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UAE Edition
Grade 10 Advanced
Student Edition

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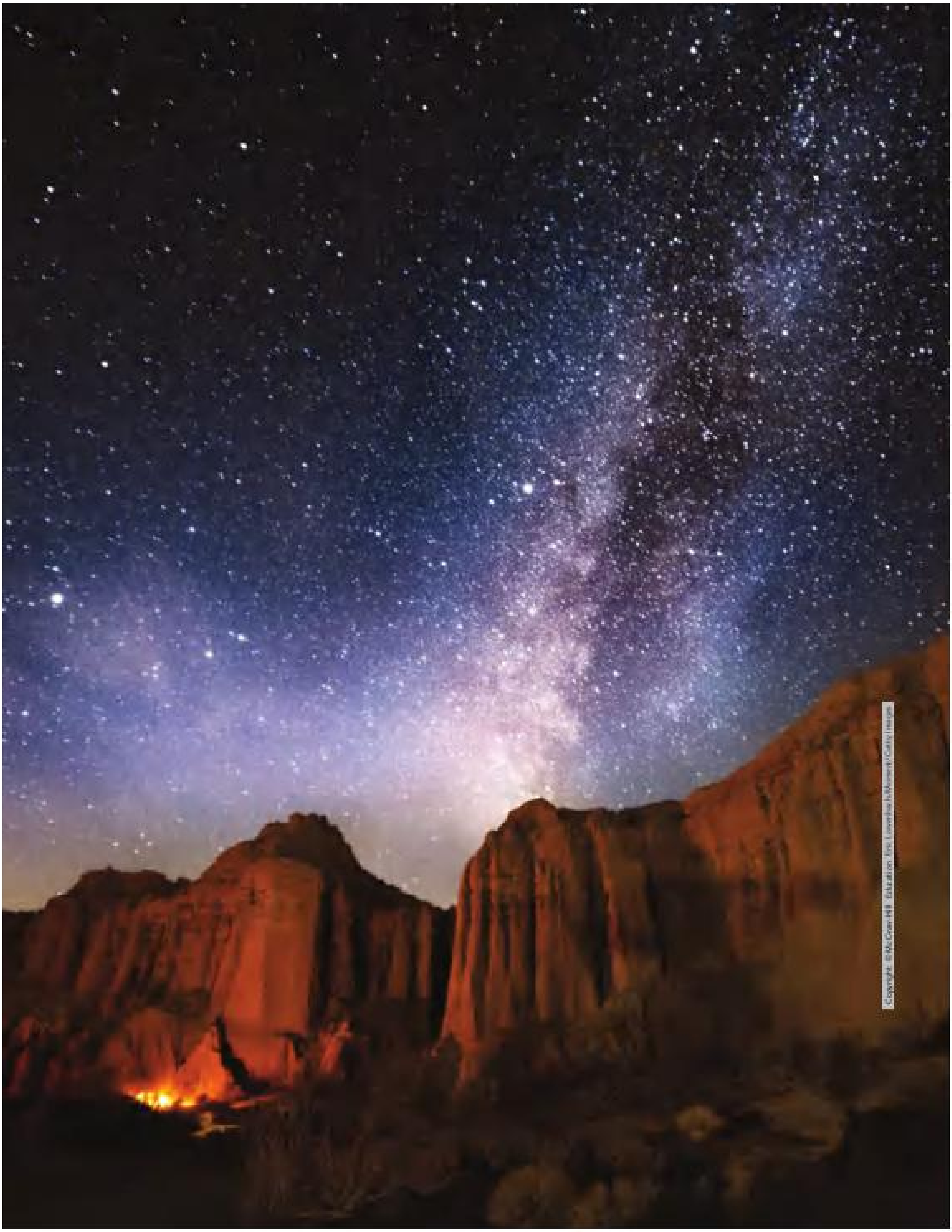


ELECTRONS IN ATOMS

ENCOUNTER THE PHENOMENON

How do we know what stars are made of?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

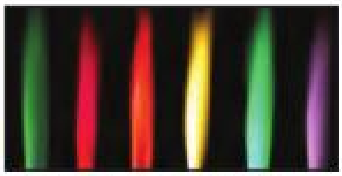
CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about how we know what stars are made of.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.



LESSON 1: Explore & Explain: Failures of the Wave Model

LESSON 2: Explore & Explain: Bohr's Atomic Model

LESSON 1
LIGHT AND QUANTIZED ENERGY

FOCUS QUESTION

What is light made of?

The Atom and Unanswered Questions

After discovering three subatomic particles in the early 1900s, scientists continued their quest to understand atomic structure and the arrangement of electrons within atoms.

Rutherford proposed that all of an atom's positive charge and virtually all of its mass are concentrated in a nucleus that is surrounded by fast-moving electrons. The model did not explain how the atom's electrons are arranged in the space around the nucleus. Nor did it address the question of why the negatively charged electrons are not pulled into the atom's positively charged nucleus. Rutherford's nuclear model did not begin to account for the differences and similarities in chemical behavior among the various elements.

For example, consider the elements lithium, sodium, and potassium, which are found in different periods on the periodic table but have similar chemical behaviors. All three elements appear metallic in nature, and their atoms react vigorously with water to liberate hydrogen gas. In fact, as shown in **Figure 1**, both sodium and potassium react so violently that the hydrogen gas can ignite and even explode.

Figure 1 Different elements can have similar reactions with water.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practice: Wave Characteristics
HS-PS4-1. Use mathematical representations to support a claim regarding relationships among the frequency, wavelength, and speed of waves traveling in various media.

CCC **Identify Crosscutting Concepts**
Create a table of the crosscutting concepts and fill in examples you find as you read.

In the early 1900s, scientists began to unravel the puzzle of chemical behavior. They observed that certain elements emitted visible light when heated in a flame. Analysis of the emitted light revealed that an element's chemical behavior is related to the arrangement of the electrons in its atoms. To understand this relationship and the nature of atomic structure, it will be helpful to first understand the nature of light.

The Wave Nature of Light

Visible light is a type of **electromagnetic radiation**—a form of energy that exhibits wavelike behavior as it travels through space. It can be modeled as a wave of changing electric and magnetic fields. Other examples of electromagnetic radiation include microwaves, X rays, and television and radio waves.

Characteristics of waves

All waves can be described by several characteristics, a few of which might be familiar to you from everyday experience. You might have seen concentric waves when dropping an object into water, as shown in **Figure 2a**.

The **wavelength** (represented by λ, the Greek letter lambda) is the shortest distance between equivalent points on a continuous wave. For example, in **Figure 2b**, the wavelength is measured from crest to crest or from trough to trough. Wavelength is usually expressed in meters, centimeters, or nanometers (1 nm = 1×10^{-9} m).

The **frequency** (represented by ν, the Greek letter nu) is the number of waves that pass a given point per second. One hertz (Hz), the SI unit of frequency, equals one wave per second. In calculations, frequency is expressed with units of waves per second, (1/s) or (s^{-1}); the term waves is understood. A particular frequency can be expressed in the following ways: 652 Hz = 652 waves/second = 652/s = 652 s^{-1}.

The **amplitude** of a wave is the wave's height from the origin to a crest, or from the origin to a trough, as illustrated in **Figure 2b.** Wavelength and frequency do not affect the amplitude of a wave.

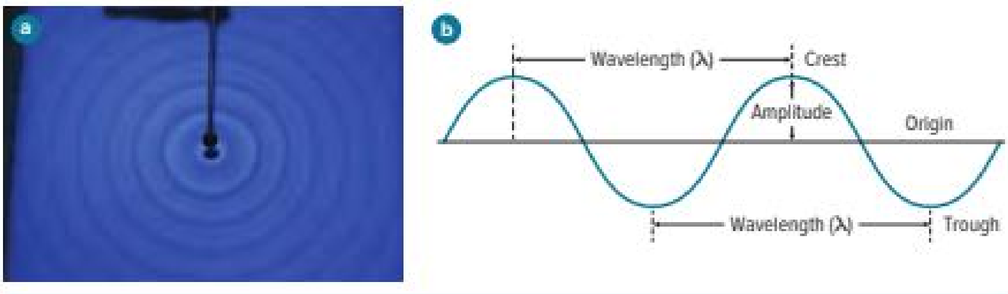


Figure 2 a. The concentric waves in the water show the characteristic properties of all waves. **b.** Amplitude, wavelength, and frequency are the main characteristics of waves.

Identify *a crest, a trough, and one wavelength in the photo.*

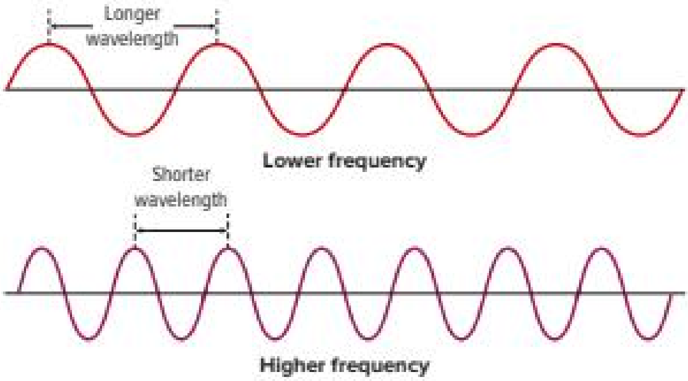


Figure 3 These waves illustrate the relationship between wavelength and frequency. As frequency increases, wavelength decreases.

Infer *Does frequency or wavelength affect amplitude?*

All electromagnetic waves, including visible light, travel at a speed of 3.00×10^8 m/s in a vacuum. Because the speed of light is such an important and universal value, it is given its own symbol, c. The speed of light is the product of its wavelength (λ) and its frequency (ν).

Electromagnetic Wave Relationship

$$c = \lambda\nu$$

c is the speed of light in a vacuum.
λ is the wavelength.
ν is the frequency.

The speed of light in a vacuum is equal to the product of the wavelength and the frequency.

Although the speed of all electromagnetic waves in a vacuum is the same, waves can have different wavelengths and frequencies. As you can see from the equation above, wavelength and frequency are inversely related; in other words, as one quantity increases, the other decreases. To better understand this relationship, examine the two waves illustrated in **Figure 3**. Although both waves travel at the speed of light, you can see that the red wave has a longer wavelength and lower frequency than the violet wave.

Electromagnetic spectrum

Sunlight, which is one example of white light, contains a nearly continuous range of wavelengths and frequencies. White light passing through a prism separates into a continuous spectrum of colors similar to the spectrum in **Figure 4**. These are the colors of the visible spectrum. The spectrum is called continuous because each point of it corresponds to a unique wavelength and frequency.

Figure 4 When white light passes through a prism, it is separated into a continuous spectrum of its different components—red, orange, yellow, green, blue, indigo, and violet light.

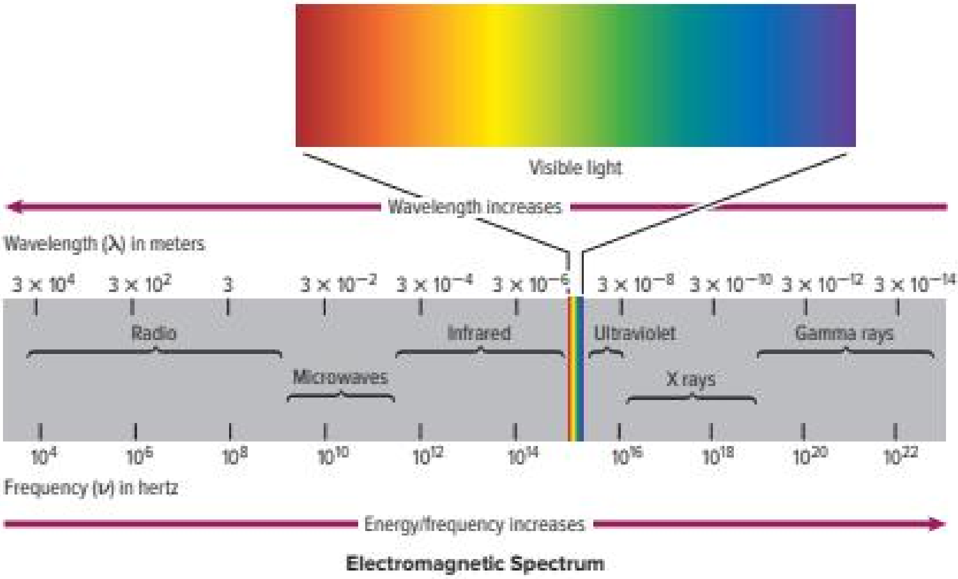


Figure 5 The electromagnetic spectrum covers a wide range of frequencies. The visible-light section of the spectrum is very narrow. As frequency and energy increase, wavelength decreases.

The visible spectrum of light, shown in **Figure 4**, comprises only a small portion of the complete electromagnetic spectrum. The complete electromagnetic spectrum is illustrated in **Figure 5**. The **electromagnetic spectrum,** also called the EM spectrum, includes all forms of electromagnetic radiation, with the only differences in the types of radiation being their frequencies and wavelengths.

Note in **Figure 4** that the bend varies with the wavelengths as they pass through the prism, resulting in the sequence of the colors red, orange, yellow, green, blue, indigo, and violet. In examining the energy of the radiation shown in **Figure 5**, note that energy increases with increasing frequency. Thus, looking back at **Figure 3**, the violet light, with its greater frequency, has more energy than the red light. This relationship between frequency and energy will be explained in the next lesson. The wavelength and frequency of a wave are related to one another by the speed of travel of the wave, which depends on the type of wave and the medium through which it is passing. For light waves, you can use the formula $c = \lambda\nu$ to calculate the wavelength or frequency of any wave.

PHYSICS Connection Electromagnetic radiation from diverse origins constantly bombards us. In addition to the radiation from the Sun, technology such as radio and TV signals, phone relay stations, lightbulbs, medical X-ray equipment, and particle accelerators also produce radiation. Natural sources on Earth, such as lightning, natural radioactivity, and even the glow of fireflies, also contribute. Our knowledge of the universe is based on electromagnetic radiation emitted by distant objects and detected with instruments on Earth.

Get It?

Explain how wavelength and frequency of a wave are related.

EXAMPLE Problem 1

CALCULATING WAVELENGTH OF AN EM WAVE Microwaves are used to cook food and transmit information. What is the wavelength of a microwave that has a frequency of 3.44×10^9 Hz?

1 ANALYZE THE PROBLEM

You are given the frequency of a microwave. You also know that because microwaves are part of the electromagnetic spectrum, their speeds, frequencies, and wavelengths are related by the formula $c = \lambda\nu$. The value of c is a known constant. First, solve the equation for wavelength, then substitute the known values and solve.

Known	Unknown
$\nu = 3.44 \times 10^9 \text{Hz}$	$\lambda = ?\ \text{m}$
$c = 3.00 \times 10^8 \text{ m/s}$	

2 SOLVE FOR THE UNKNOWN

Solve the equation relating the speed, frequency, and wavelength of an electromagnetic wave for wavelength (λ).

$c = \lambda\nu$ — State the electromagnetic wave relationship.

$\lambda = c/\nu$ — Solve for λ.

$\lambda = \dfrac{3.00 \times 10^8 \text{ m/s}}{3.44 \times 10^9 \text{ Hz}}$ — Substitute $c = 3.00 \times 10^8$ m/s and $\nu = 3.44 \times 10^9$ Hz.

Note that hertz is equivalent to 1/s or s^{-1}.

$\lambda = \dfrac{3.00 \times 10^8 \text{ m}/\cancel{s}}{3.44 \times 10^9 \cancel{s^{-1}}}$ — Divide numbers and units.

$\lambda = 8.72 \times 10^{-2} \text{m}$

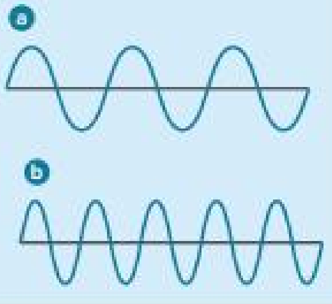
3 EVALUATE THE ANSWER

The answer is correctly expressed in a unit of wavelength (m). Both of the known values in the problem are expressed with three significant figures, so the answer should have three significant figures, which it does. The value for the wavelength is within the wavelength range for microwaves shown in **Figure 5.**

PRACTICE Problems

ADDITIONAL PRACTICE

1. Objects get their colors from reflecting only certain wavelengths when hit with white light. Light reflected from a green leaf is found to have a wavelength of 4.90×10^{-7} m. What is the frequency of the light?
2. When light or longer wavelength electromagnetic radiation is absorbed in matter, it is generally converted into thermal energy (heat). Shorter wavelength electromagnetic radiation (ultraviolet, X-rays, gamma rays) can ionize atoms and cause damage to living cells. X-rays can penetrate body tissues and are widely used to diagnose and treat disorders of internal body structures. What is the frequency of an X-ray with a wavelength of 1.15×10^{-10} m?
3. After careful analysis, an electromagnetic wave is found to have a frequency of 7.8×10^6 Hz. What is the speed of the wave?
4. **CHALLENGE** While an FM radio station broadcasts at a frequency of 94.7 MHz, an AM station broadcasts at a frequency of 820 kHz. What are the wavelengths of the two broadcasts? Which of the drawings on the right corresponds to the FM station? To the AM station?



The Particle Nature of Light

While considering light as a wave explains much of its everyday behavior, it fails to adequately describe important aspects of light's interactions with matter. The wave model of light cannot explain why heated objects emit only certain frequencies of light at a given temperature, or why some metals emit electrons when light of a specific frequency shines on them. Scientists realized that a new model or a revision of the wave model of light was needed to address these phenomena.

The quantum concept

When objects are heated, they emit glowing light. **Figure 6** illustrates this phenomenon with iron. A piece of iron appears dark gray at room temperature, glows red when heated sufficiently, and turns orange, then bluish in color at even higher temperatures. As the iron gets hotter, it has more energy and emits different colors of light. These different colors correspond to different frequencies and wavelengths.

The wave model could not explain the emission of these different wavelengths. In 1900, German physicist Max Planck (1858–1947) began searching for an explanation of this phenomenon. His study led him to a startling conclusion: matter can gain or lose energy only in small, specific amounts called quanta. A **quantum** is the minimum amount of energy that can be gained or lost by an atom.

Planck proposed that the energy emitted by hot objects was quantized. He also showed that there is a direct relationship between the energy of a quantum and the frequency of emitted radiation.

Figure 6 The wavelength of the light emitted by heated metal, such as the iron above, depends on the temperature. At room temperature, iron is gray. When heated, it first turns red, then glowing orange.

Identify *the color of the piece of iron with the greatest kinetic energy.*

Energy of a Quantum

$$E_{\text{quantum}} = h\nu$$

E_{quantum} represents energy.
h is Planck's constant.
ν represents frequency.

The energy of a quantum is given by the product of Planck's constant and the frequency.

Planck's constant, h, has a value of 6.626×10^{-34} J·s, where J is the symbol for joule, the SI unit of energy. The equation shows that the energy of radiation increases as the radiation's frequency, ν, increases. According to Planck's theory, for a given frequency, ν, matter can emit or absorb energy only in whole-number multiples of $h\nu$; that is, $1h\nu$, $2h\nu$, $3h\nu$, and so on.

A useful analogy is that of a child building a wall with wooden blocks. The child can add to or take away from the wall only in increments of whole numbers of blocks. Similarly, matter can have only certain amounts of energy—quantities of energy between these values do not exist.

Planck and other physicists of the time thought the concept of quantized energy was revolutionary, and some found it disturbing. Prior experience had led scientists to think that energy could be absorbed and emitted in continually varying quantities, with no minimum limit to the amount. For example, think about heating a cup of water in a microwave oven. It seems that you can add any amount of thermal energy to the water by regulating the power and duration of the microwaves. Instead, the water's temperature increases in infinitesimal steps as its molecules absorb quanta of energy. Because these steps are so small, the temperature seems to rise in a continuous, rather than a stepwise, manner.

Real-World Chemistry

The Photoelectric Effect

SOLAR ENERGY is sometimes used to power road signs. Photovoltaic cells use the photoelectric effect to convert the energy of light into electric energy.

The photoelectric effect

Scientists also knew that the wave model of light could not explain a phenomenon called the photoelectric effect. In the **photoelectric effect,** electrons, called photoelectrons, are emitted from a metal's surface when light at or above a certain frequency shines on the surface, as shown in **Figure 7**.

The wave model predicts that given enough time, even low-energy, low-frequency light would accumulate and supply enough energy to eject photoelectrons from a metal. In reality, a metal will not eject photoelectrons below a specific frequency of incident light. For example, no matter how intensely or how long it shines, light with a frequency less than 1.14×10^{15} Hz does not eject photoelectrons from silver. But even dim light with a frequency equal to or greater than 1.14×10^{15} Hz ejects photoelectrons from silver.

Get It?

Describe the photoelectric effect.

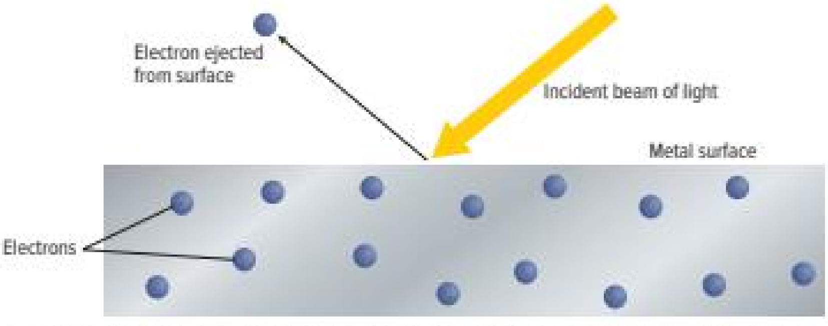


Figure 7 The photoelectric effect occurs when light of a certain frequency strikes a metal surface and ejects electrons. When the intensity of the light increases, the number of electrons ejected increases. When the frequency (energy) of the light increases, the energy of the ejected electrons increases.

Light's dual nature

To explain the photoelectric effect, Albert Einstein proposed in 1905 that light has a dual nature. A beam of light has wavelike and particlelike properties. It can be thought of, or modeled, as a beam of bundles of energy called photons. A **photon** is a massless particle that carries a quantum of energy. Extending Planck's idea of quantized energy, Einstein calculated that a photon's energy depends on its frequency.

Energy of a Photon

$$E_{photon} = h\nu$$

E_{photon} represents energy.
h is Planck's constant.
ν represents frequency.

The energy of a photon is given by the product of Planck's constant and the frequency.

Einstein also proposed that the energy of a photon must have a certain threshold value to cause the ejection of a photoelectron from the surface of the metal. Thus, even small numbers of photons with energy above the threshold value will cause the photoelectric effect. Einstein won the Nobel Prize in Physics in 1921 for this work.

EXAMPLE Problem 2

CALCULATE THE ENERGY OF A PHOTON Every object gets its color by reflecting a certain portion of incident light. The color is determined by the wavelength of the reflected photons, thus by their energy. What is the energy of a photon from the violet portion of the Sun's light if it has a frequency of $7.230 \times 10^{14}\ s^{-1}$?

1 ANALYZE THE PROBLEM

Known	Unknown
$\nu = 7.230 \times 10^{14}\ s^{-1}$	$E_{photon} = ?\ J$
$h = 6.626 \times 10^{-34}\ J{\cdot}s$	

2 SOLVE FOR THE UNKNOWN

$E_{photon} = h\nu$ — State the equation for the energy of a photon.

$E_{photon} = (6.626 \times 10^{-34}\ J{\cdot}s)(7.230 \times 10^{14}\ s^{-1})$ — Substitute $h = 6.626 \times 10^{-34}\ J{\cdot}s$ and $\nu = 7.230 \times 10^{14}\ s^{-1}$.

$E_{photon} = 4.791 \times 10^{-19}\ J$ — Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

As expected, the energy of a single photon of light is extremely small. The unit is joules, an energy unit, and there are four significant figures.

PRACTICE Problems

ADDITIONAL PRACTICE

5. Calculate the energy possessed by a single photon of each of the following types of electromagnetic radiation.
   a. $6.32 \times 10^{20}\ s^{-1}$ **b.** 9.50×10^{13} Hz **c.** $1.05 \times 10^{16}\ s^{-1}$
6. The blue color in some fireworks occurs when copper(I) chloride is heated to approximately 1500 K and emits blue light of wavelength 4.50×10^{2} nm. How much energy does one photon of this light carry?
7. **CHALLENGE** The microwaves used to heat food have a wavelength of 0.125 m. What is the energy of one photon of the microwave radiation?

Atomic Emission Spectra

Have you ever wondered how light is produced in the glowing tubes of neon signs? This is another phenomenon that cannot be explained by the wave model of light. The light of a neon sign is produced by passing electricity through a tube filled with neon gas. Neon atoms in the tube absorb energy and become excited. These excited atoms return to their stable state by emitting light to release that energy. If the light emitted by the neon is passed through a glass prism, neon's atomic emission spectrum is produced.

The **atomic emission spectrum** of an element is the set of frequencies of the electromagnetic waves emitted by atoms of the element. **Figure 8** shows the purple-pink glow produced by excited hydrogen atoms and the visible portion of hydrogen's emission spectrum responsible for producing the glow. Note that an atomic emission spectrum is not a continuous spectrum. Rather, it consists of several individual lines of color corresponding to the frequencies of radiation emitted by the atoms.

Each element's atomic emission spectrum is unique and can be used to identify an element. For example, when a platinum wire is dipped into a strontium nitrate solution and then held in a burner flame, the strontium atoms emit a characteristic red color.

Get It?

Explain how an emission spectrum is produced.

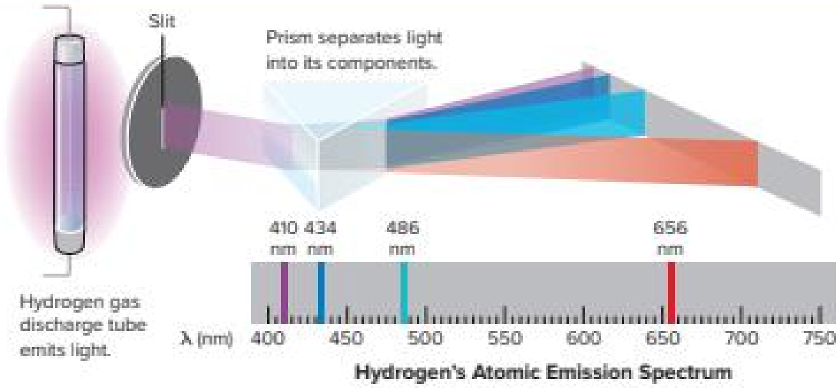


Figure 8 The purple light emitted by hydrogen can be separated into its different components using a prism. Hydrogen has an atomic emission spectrum that comprises four lines of different wavelengths.
Determine *Which line has the highest energy?*

STEM CAREER Connection

Astrochemist

Do you like chemistry, planetary science, chemical biology, physics, astronomy, and computational science? A career in astrochemisty may be the career for you. Astrochemists use telescopes, satellites, and space vehicles to collect spectroscopic data and analyze it. In this career, knowledge from several scientific disciplines is used to analyze and model the data collected.

ACADEMIC VOCABULARY

phenomenon
an observable fact or event *During rainstorms, electric currents often pass from the sky to Earth—a phenomenon called lightning.*

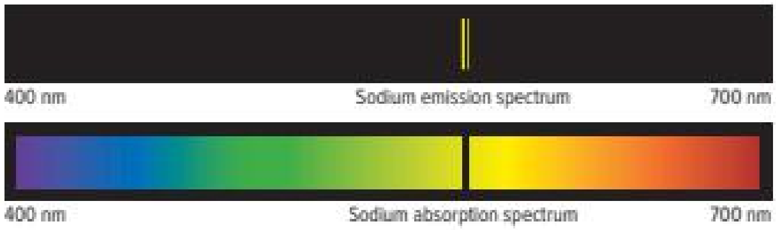


Figure 9 When excited sodium atoms return to a less excited state, they emit light at certain frequencies, producing an emission spectrum. When a continuous spectrum of light passes through sodium gas, atoms in the gas absorb light at those same frequencies, producing an absorption spectrum with dark spectral lines.

ASTRONOMY Connection Astronomers use atomic spectra to determine the composition of the outer layers of stars. When a continuous spectrum of light from within a star passes through the outer layers of the star, atoms in the outer layers absorb light at certain frequencies, producing an absorption spectrum. The lines in the absorption spectrum reveal what elements are in the outer layers of the star because the frequencies absorbed in an element's absorption spectrum are the same as those emitted in the element's emission spectrum, as shown for sodium in **Figure 9.**

Check Your Progress

Summary

- All waves are defined by their wavelengths, frequencies, amplitudes, and speeds.
- In a vacuum, all electromagnetic waves travel at the speed of light.
- All electromagnetic waves have both wave and particle properties.
- Matter emits and absorbs energy in quanta.
- White light produces a continuous spectrum. An element's emission spectrum consists of a series of lines of individual colors.

Demonstrate Understanding

8. **Describe** the relationship between changing electric and magnetic fields and particles.
9. **Compare and contrast** continuous spectrum and emission spectrum.
10. **Discuss** the way in which Einstein utilized Planck's quantum concept to explain the photoelectric effect.
11. **Calculate** Heating 235 g of water from 22.6°C to 94.4°C in a microwave oven requires 7.06×10^4 J of energy. If the microwave frequency is $2.88 \times 10^{10}\ s^{-1}$, how many quanta are required to supply the 7.06×10^4 J?
12. **Interpret Scientific Illustrations** Use **Figure 5** and your knowledge of electromagnetic radiation to match the numbered items with the lettered items. The numbered items may be used more than once or not at all.

a. longest wavelength	1. gamma rays
b. highest frequency	2. ultraviolet light
c. greatest energy	3. radio waves

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LESSON 2
QUANTUM THEORY AND THE ATOM

FOCUS QUESTION

Why does every element produce a unique atomic emission spectrum?

Bohr's Model of the Atom

The dual wave-particle model of light accounted for several previously unexplainable phenomena, but scientists still did not understand the relationships among atomic structure, electrons, and atomic emission spectra. Recall that hydrogen's atomic emission spectrum is discontinuous; that is, it is made up of only certain frequencies of light. Why are the atomic emission spectra of elements discontinuous rather than continuous? Niels Bohr, a Danish physicist working in Rutherford's laboratory in 1913, proposed a quantum model for the hydrogen atom that seemed to answer this question. Bohr's model also correctly predicted the frequencies of the lines in hydrogen's atomic emission spectrum.

Energy states of hydrogen

Building on Planck's and Einstein's concepts of quantized energy, Bohr proposed that the hydrogen atom has only certain allowable energy states, as illustrated in **Figure 10**. The lowest allowable energy state of an atom is called its **ground state.** When an atom gains energy, it is said to be in an excited state.

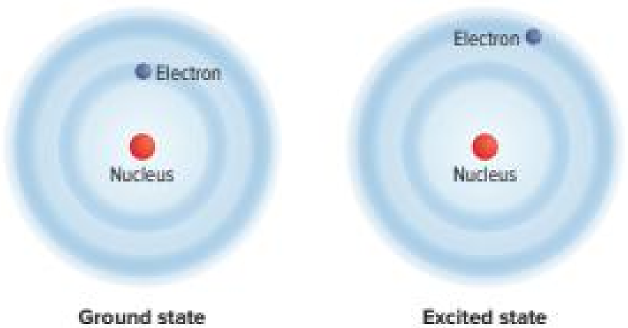


Figure 10 The figure shows an atom that has one electron. Note that the illustration is not to scale. In its ground state, the electron is associated with the lowest energy level. When the atom is in an excited state, the electron is associated with a higher energy level.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Laboratory: The Photoelectric Effect
Use mathematical and computational thinking to observe patterns in the stability and instability of physical systems.

Inquiry into Chemistry: Design Atomic Models
Plan and carry out an investigation to create a model of the charged subsections of an atom.

Table 1 Bohr's Description of the Hydrogen Atom

Bohr's Atomic Orbit	Quantum Number	Orbit Radius (nm)	Corresponding Atomic Energy Level	Relative Energy
First	$n = 1$	0.0529	1	E_1
Second	$n = 2$	0.212	2	$E_2 = 4E_1$
Third	$n = 3$	0.476	3	$E_3 = 9E_1$
Fourth	$n = 4$	0.846	4	$E_4 = 16E_1$
Fifth	$n = 5$	1.32	5	$E_5 = 25E_1$
Sixth	$n = 6$	1.90	6	$E_6 = 36E_1$
Seventh	$n = 7$	2.59	7	$E_7 = 49E_1$

Bohr suggested that the electron in a hydrogen atom moves around the nucleus in only certain allowed circular orbits. The smaller the electron's orbit, the lower the atom's energy state, or energy level. Conversely, the larger the electron's orbit, the higher the atom's energy state, or energy level. Bohr assigned a number, n, called a **quantum number,** to each orbit. He also calculated the radius of each orbit. **Table 1** shows data for the first seven energy levels of a hydrogen atom according to Bohr's model.

The hydrogen line spectrum

Bohr suggested that a hydrogen atom is in the ground state when its single electron is in the $n = 1$ orbit, also called the first energy level. In the ground state, the atom does not radiate energy. When energy is added from an outside source, the electron moves to a higher-energy orbit, putting the atom in an excited state. When the atom is in an excited state, the electron can drop from the higher-energy orbit to a lower-energy orbit, as shown in **Figure 11.**

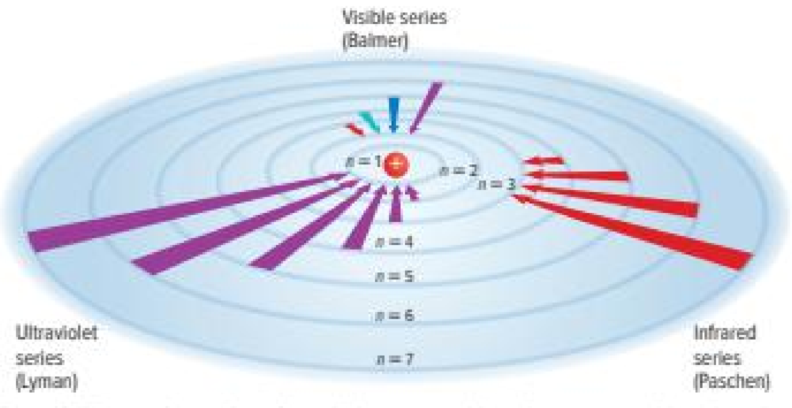


Figure 11 When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon is emitted. The ultraviolet (Lyman), visible (Balmer), and infrared (Paschen) series correspond to electrons dropping to $n = 1$, $n = 2$, and $n = 3$, respectively.

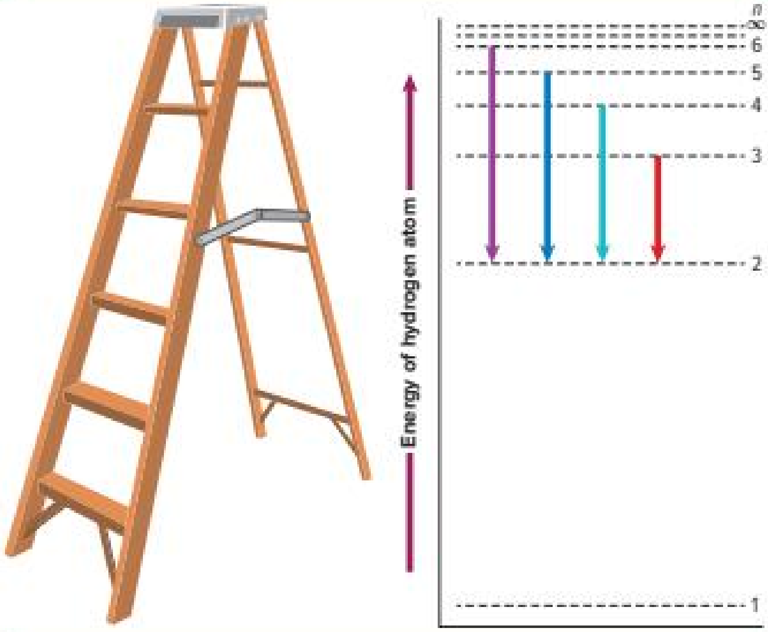


Figure 12 Only certain energy levels are allowed. The energy levels are similar to the rungs of a ladder. The four visible lines correspond to electrons dropping from a higher n to the orbit $n = 2$. As n increases, the hydrogen atom's energy levels are closer to each other.

As a result of this transition, the atom emits a photon corresponding to the energy difference between the two levels.

$$\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}} = E_{\text{photon}} = h\nu$$

Because only certain atomic energies are possible, only certain frequencies of electromagnetic radiation can be emitted.

You might compare hydrogen's atomic energy states to rungs on a ladder, as shown in **Figure 12**. A person can climb up or down the ladder only from rung to rung. Similarly, the hydrogen atom's electron can move only from one allowable orbit to another, and therefore, can emit or absorb only certain amounts of energy, corresponding to the energy difference between the two orbits. Unlike rungs on a ladder, however, the hydrogen atom's energy levels are not evenly spaced.

Figure 12 also illustrates the four electron transitions that account for visible lines in hydrogen's atomic emission spectrum, shown in **Figure 8**. Electron transitions from higher-energy orbits to the second orbit account for all of hydrogen's visible lines, which form the Balmer series. Other electron transitions have been measured that are not visible, such as the Lyman series (ultraviolet), in which electrons drop into the $n = 1$ orbit, and the Paschen series (infrared), in which electrons drop into the $n = 3$ orbit.

Get It?

Explain why different colors of light result from electron behavior in the atom.

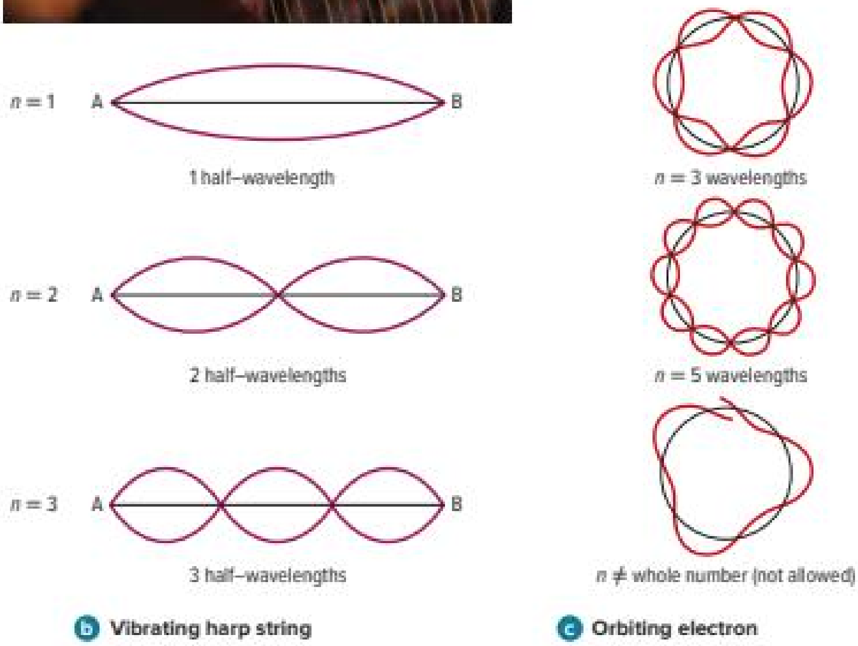
The limits of Bohr's model

Bohr's model explained hydrogen's observed spectral lines. However, the model failed to explain the spectrum of any other element. Moreover, Bohr's model did not fully account for the chemical behavior of atoms. In fact, although Bohr's idea of quantized energy levels laid the groundwork for atomic models to come, later experiments demonstrated that the Bohr model was fundamentally incorrect. The movements of electrons in atoms are not completely understood even now; however, substantial evidence indicates that electrons do not move around the nucleus in circular orbits.

The Quantum Mechanical Model of the Atom

Scientists in the mid-1920s, convinced that the Bohr atomic model was incorrect, formulated new and innovative explanations of how electrons are arranged in atoms. In 1924, a French graduate student in physics named Louis de Broglie (1892–1987) proposed a new idea, shown in **Figure 13** and discussed on the following page.

Figure 13 **a.** The string on the harp vibrates between two fixed endpoints. **b.** The vibrations of a string between the two fixed endpoints labeled A and B are limited to multiples of half-wavelengths. **c.** Electrons on circular orbits can only have whole numbers of wavelengths.



Electrons as waves

De Broglie had been thinking that Bohr's quantized electron orbits had characteristics similar to those of waves. For example, as **Figures 13a** and **13b** show, only multiples of half-wavelengths are possible on a plucked harp string because the string is fixed at both ends. Similarly, de Broglie saw that only whole numbers of wavelengths are allowed in a circular orbit of fixed radius, as shown in **Figure 13c**.

De Broglie also reflected on the fact that light—at one time thought to be strictly a wave phenomenon—has both wave and particle characteristics. These thoughts led de Broglie to pose a new question: If waves can have particlelike behavior, could the opposite also be true? That is, can particles of matter, including electrons, behave like waves?

De Broglie knew that if an electron has wavelike motion and is restricted to circular orbits of fixed radius, only certain wavelengths, frequencies, and energies are possible. Developing his idea, de Broglie derived the following equation, called the **de Broglie equation.**

Particle Electromagnetic-Wave Relationship

$$\lambda = \frac{h}{m\nu}$$

λ represents wavelength.
h is Planck's constant.
m represents mass of the particle.
ν represents velocity.

The wavelength of a particle is the ratio of Planck's constant and the product of the particle's mass and its velocity.

The de Broglie equation predicts that all moving particles have wave characteristics. Note that the equation includes Planck's constant. Planck's constant is an exceedingly small number, 6.626×10^{-34} J·s, which helps explain why it is difficult or impossible to observe the wave characteristics of objects at the scale of everyday experience. For example, an automobile moving at 25 m/s and having a mass of 910 kg has a wavelength of 2.9×10^{-38} m, far too small to be seen or detected. By comparison, an electron moving at the same speed has the easily measured wavelength of 2.9×10^{-5} m. Subsequent experiments have proven that electrons and other moving particles do indeed have wave characteristics.

Get It?

Identify which variables in the de Broglie equation represent wavelike properties.

The Heisenberg uncertainty principle

Step by step, scientists such as Rutherford, Bohr, and de Broglie had been unraveling the mysteries of the atom. However, a conclusion reached by the German theoretical physicist Werner Heisenberg (1901–1976) proved to have profound implications for atomic models.

Heisenberg showed that it is impossible to take any measurement of an object without disturbing the object. Imagine trying to locate a hovering, helium-filled balloon in a darkened room. If you wave your hand about, you can locate the balloon's position when you touch it. However, when you touch the balloon, you transfer energy to it and change its position. You could also detect the balloon's position by turning on a flashlight. Using this method, photons of light reflected from the balloon would reach your eyes and reveal the balloon's location. Because the balloon is a macroscopic object, the effect of the rebounding photons on its position is very small and not observable.

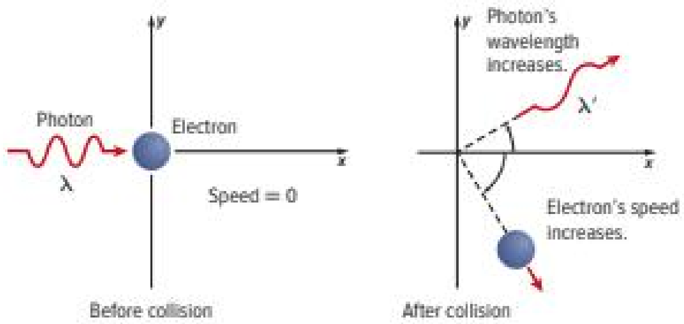


Figure 14 When a photon interacts with an electron at rest, both the velocity and the position of the electron are modified. This illustrates the Heisenberg uncertainty principle. It is impossible to know at the same time the position and the velocity of a particle.

Explain *Why has the photon's energy changed?*

Now imagine trying to determine an electron's location by "bumping" it with a high-energy photon. Because such a photon has about the same energy as an electron, the interaction between the two particles changes both the wavelength of the photon and the position and velocity of the electron, as shown in **Figure 14**. In other words, the act of observing the electron produces a significant, unavoidable uncertainty in the position and motion of the electron. Heisenberg's analysis of interactions, such as those between photons and electrons, led him to his historic conclusion. The **Heisenberg uncertainty principle** states that it is fundamentally impossible to know precisely both the velocity and position of a particle at the same time.

Although scientists of the time found Heisenberg's principle difficult to accept, it has been proven to describe the fundamental limitations of what can be observed. The interaction of a photon with a macroscopic object such as a helium-filled balloon has so little effect on the balloon that the uncertainty in its position is too small to measure. But that is not the case with an electron moving at 6×10^6 m/s near an atomic nucleus. The uncertainty of the electron's position is at least 10^{-9} m, about 10 times greater than the diameter of the entire atom.

The Heisenberg uncertainty principle also means that it is impossible to assign fixed paths for electrons like the circular orbits in Bohr's model. The only quantity that can be known is the probability for an electron to occupy a certain region around the nucleus.

Get It?

Identify the only quantity of an electron's orbit that can be determined.

CCC CROSSCUTTING CONCEPTS

Cause and Effect What empirical evidence did scientists have that supports the claim that electrons have both particle and wave properties?

The Schrödinger wave equation

In 1926, Austrian physicist Erwin Schrödinger (1887–1961) furthered the wave-particle theory proposed by de Broglie. Schrödinger derived an equation that treated the hydrogen atom's electron as a wave. Schrödinger's new model for the hydrogen atom seemed to apply equally well to atoms of other elements--an area in which Bohr's model failed. The atomic model in which electrons are treated as waves is called the wave mechanical model of the atom or the **quantum mechanical model of the atom.** Like Bohr's model, the quantum mechanical model limits an electron's energy to certain values. However, unlike Bohr's model, the quantum mechanical model makes no attempt to describe the electron's path around the nucleus.

Get It?

Compare and contrast Bohr's model and the quantum mechanical model.

Electron's probable location

The Schrödinger wave equation is too complex to be considered here. However, each solution to the equation is known as a wave function, which is related to the probability of finding the electron within a particular volume of space around the nucleus. The wave function predicts a three-dimensional region around the nucleus, called an **atomic orbital,** which describes the electron's probable location. An atomic orbital is like a fuzzy cloud in which the density at a given point is proportional to the probability of finding the electron at that point.

Figure 15a illustrates the probability map that describes the electron in the atom's lowest energy state. The probability map can be thought of as a time-exposure photograph of the electron moving around the nucleus, in which each dot represents the electron's location at an instant in time. The high density of dots near the nucleus indicates the electron's most probable location. However, it is also possible that the electron might be found at a considerable distance from the nucleus.

Get It?

Describe where electrons are located in an atom.

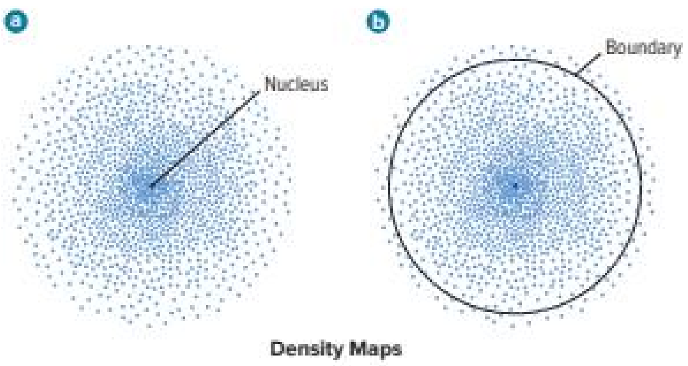


Figure 15 The density map represents the probability of finding an electron at a given position around the nucleus. **a.** The higher density of points near the nucleus shows that the electron is more likely to be found close to the nucleus. **b.** At any given time, there is a 90% probability of finding the electron within the circular region shown. This surface is sometimes chosen to represent the boundary of the atom. In this illustration, the circle corresponds to a projection of the 3-dimensional sphere that contains the electrons.

Hydrogen's Atomic Orbitals

Because the boundary of an atomic orbital is fuzzy, the orbital does not have an exact defined size. To overcome the inherent uncertainty about the electron's location, chemists arbitrarily draw an orbital's surface to contain 90% of the electron's total probability distribution. This means that the probability of finding the electron within the boundary is 0.9 and the probability of finding it outside the boundary is 0.1. In other words, it is more likely to find the electron close to the nucleus and within the volume defined by the boundary, than to find it outside the volume. The circle shown in **Figure 15b** encloses 90% of the lowest-energy orbital of hydrogen.

Principal quantum number

Recall that the Bohr atomic model assigns quantum numbers to electron orbits. Similarly, the quantum mechanical model assigns four quantum numbers to atomic orbitals. The first one is the **principal quantum number** (n) and indicates the relative size and energy of atomic orbitals. As n increases, the orbital becomes larger, the electron spends more time farther from the nucleus, and the atom's energy increases. Therefore, n specifies the atom's major energy levels. Each major energy level is called a **principal energy level.** An atom's lowest principal energy level is assigned a principal quantum number of 1. When the hydrogen atom's single electron occupies an orbital with $n = 1$, the atom is in its ground state. Up to 7 energy levels have been detected for the hydrogen atom, giving n values ranging from 1 to 7.

Energy sublevels

Principal energy levels contain **energy sublevels.** Principal energy level 1 consists of a single sublevel, principal energy level 2 consists of two sublevels, principal energy level 3 consists of three sublevels, and so on. To better understand the relationship between the atom's energy levels and sublevels, picture the seats in a wedge-shaped section of a theater, as shown in **Figure 16**. As you move away from the stage, the rows become higher and contain more seats. Similarly, the number of energy sublevels in a principal energy level increases as n increases.

Get It?

Explain the relationship between energy levels and sublevels.

Figure 16 Energy levels can be thought of as rows of seats in a theater. The rows that are higher up and farther from the stage contain more seats. Similarly, energy levels related to orbitals farther from the nucleus contain more sublevels.

Shapes of orbitals

Sublevels are labeled *s, p, d,* or f according to the shapes of the atom's orbitals. All s orbitals are spherical, and all p orbitals are dumbbell-shaped; however, not all d or f orbitals have the same shape. Each orbital can contain, at most, two electrons. The single sublevel in principal energy level 1 corresponds to a spherical orbital called the 1s orbital. The two sublevels in principal energy level 2 are designated 2s and 2p. The 2s sublevel corresponds to the 2s orbital, which is spherical like the 1s orbital but larger in size, as shown in **Figure 17a**. The 2p sublevel corresponds to three dumbbell-shaped p orbitals designated $2p_x$, $2p_y$, and $2p_z$. The subscripts *x, y,* and *z* merely designate the orientations of p orbitals along the *x, y,* and *z* coordinate axes, as shown in **Figure 17b**. Each of the p orbitals related to an energy sublevel has the same energy.

Get It?

Describe the shapes of *s* and *p* orbitals.

Principal energy level 3 consists of three sublevels designated 3s, 3p, and 3d. Each d sublevel relates to five orbitals of equal energy. Four of the d orbitals have identical shapes but different orientations along the *x, y,* and *z* coordinate axes. However, the fifth orbital, d_{z^2}, is shaped and oriented differently than the other four. The shapes and orientations of the five d orbitals are illustrated in **Figure 17c**. The fourth principal energy level ($n = 4$) contains a fourth sublevel, called the 4f sublevel, which relates to seven f orbitals of equal energy. The f orbitals have complex, multilobed shapes.

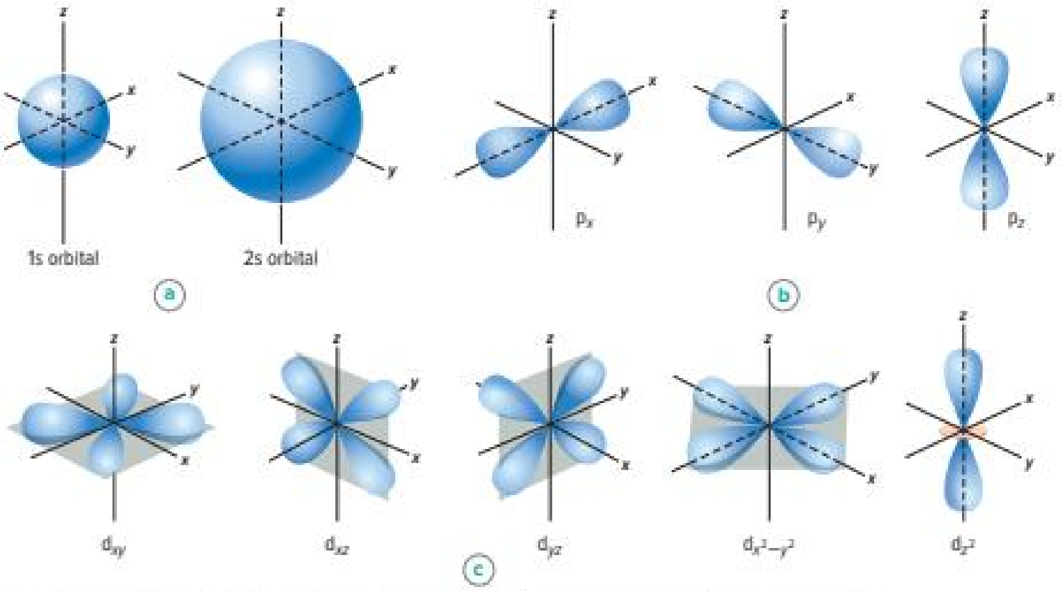


Figure 17 The shapes of atomic orbitals describe the probable distribution of electrons in energy sublevels.
a. All s orbitals are spherical, and their size increases with increasing principal quantum number.
b. The three p orbitals are dumbbell-shaped and are oriented along the three perpendicular *x, y,* and *z* axes.
c. Four of the five d orbitals have the same shape but lie in different planes. The d_{z^2} orbital has its own unique shape.

Table 2 Hydrogen's First Four Principal Energy Levels

Principal Quantum Number (n)	Sublevels (Types of Orbitals) Present	Number of Orbitals Related to Sublevel	Total Number of Orbitals Related to Principal Energy Level (n^2)
1	s	1	1
2	s p	1 3	4
3	s p d	1 3 5	9
4	s p d f	1 3 5 7	16

Hydrogen's first four principal energy levels, sublevels, and related atomic orbitals are summarized in **Table 2**. Note that the number of orbitals related to each sublevel is always an odd number, and that the maximum number of orbitals related to each principal energy level equals n^2.

At any given time, the electron in a hydrogen atom can occupy just one orbital. You can think of the other orbitals as unoccupied spaces—spaces available should the atom's energy increase or decrease.

Check Your Progress

Summary

- Bohr's atomic model attributes hydrogen's emission spectrum to electrons dropping from higher-energy to lower-energy orbits.
- The de Broglie equation relates a particle's wavelength to its mass, its velocity, and Planck's constant.
- The quantum mechanical model assumes that electrons have wave properties.
- Electrons occupy three-dimensional regions of space called atomic orbitals.

Demonstrate Understanding

13. **Explain** the reason, according to Bohr's atomic model, why atomic emission spectra contain only certain frequencies of light.
14. **Differentiate** between the wavelength of visible light and the wavelength of a moving soccer ball.
15. **Explain** why the location of an electron in an atom is uncertain using the Heisenberg uncertainty principle. How is the location of electrons in atoms defined?
16. **Compare and contrast** Bohr's model and the quantum mechanical model of the atom.
17. **Enumerate** the sublevels contained in the hydrogen atom's first four energy levels. What orbitals are related to each s sublevel and each p sublevel?
18. **Calculate** Use the information in **Table 1** to calculate how many times larger the hydrogen atom's seventh Bohr radius is than its first Bohr radius.

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LESSON 3

ELECTRON CONFIGURATION

FOCUS QUESTION

How are electrons arranged in atoms?

Ground-State Electron Configuration

The arrangement of electrons in an atom is called the atom's **electron configuration.** Because low-energy systems are more stable than high-energy systems, electrons in an atom tend to assume the arrangement that gives the atom the lowest energy possible. The most stable, lowest-energy arrangement of the electrons is called the element's ground-state electron configuration.

Three rules, or principles—the aufbau principle, the Pauli exclusion principle, and Hund's rule—define how electrons can be arranged in an atom's orbitals.

The aufbau principle

The **aufbau principle** states that each electron occupies the lowest energy orbital available. Therefore, your first step in determining an element's ground-state electron configuration is learning the sequence of atomic orbitals from lowest energy to highest energy. This sequence, known as an aufbau diagram, is shown in **Figure 18**. In the diagram, each box represents an atomic orbital.

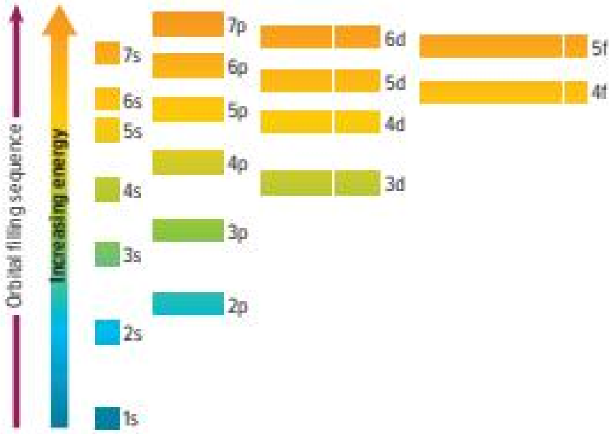


Figure 18 The aufbau diagram shows the energy of each sublevel relative to the energy of other sublevels.

Determine *Which sublevel has the greater energy, 4d or 5p?*

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Virtual Investigation: Electron Configuration

Obtain, evaluate, and communicate information on the patterns present in the periodic table that translate into patterns of electron states.

Laboratory: Electron Charge-to-Mass Ratio

Analyze and interpret data to determine the proportion of charge to mass for an electron.

Table 3 Features of the Aufbau Diagram

Feature	Example
All orbitals related to an energy sublevel are of equal energy.	All three 2p orbitals are of equal energy.
In a multi-electron atom, the energy sublevels within a principal energy level have different energies.	The three 2p orbitals are of higher energy than the 2s orbital.
In order of increasing energy, the sequence of energy sublevels within a principal energy level is s, p, d, and f.	If $n = 4$, then the sequence of energy sublevels is 4s, 4p, 4d, and 4f.
Orbitals related to energy sublevels within one principal energy level can overlap orbitals related to energy sublevels within another principal level.	The orbital related to the atom's 4s sublevel has a lower energy than the five orbitals related to the 3d sublevel.

Table 3 summarizes several features of the aufbau diagram. Although the aufbau principle describes the sequence in which orbitals are filled with electrons, it is important to know that atoms are not built up electron by electron.

The Pauli exclusion principle

Every electron has an associated spin, similar to the way a top spins on its point. Like a top, an electron is able to spin in only one of two directions. The **Pauli exclusion principle,** proposed by Austrian physicist Wolfgang Pauli (1900–1958), states that a maximum of two electrons can occupy a single atomic orbital, but only if the electrons have opposite spins.

Electrons in orbitals can be represented by arrows in boxes. An arrow pointing up $\boxed{\uparrow}$ represents the electron spinning in one direction, and an arrow pointing down $\boxed{\downarrow}$ represents the electron spinning in the opposite direction. An empty box $\boxed{}$ represents an unoccupied orbital, a box containing a single up arrow $\boxed{\uparrow}$ represents an orbital with one electron, and a box containing both up and down arrows $\boxed{\uparrow\downarrow}$ represents a filled orbital containing a pair of electrons with opposite spins.

Hund's rule

The fact that negatively charged electrons repel each other affects the distribution of electrons in equal-energy orbitals. **Hund's rule** states that single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. For example, the boxes below show the sequence in which six electrons occupy the three 2p orbitals. One electron enters each of the orbitals before a second electron enters any of the orbitals.

1. $\boxed{\uparrow}\boxed{}\boxed{}$ 2. $\boxed{\uparrow}\boxed{\uparrow}\boxed{}$ 3. $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

4. $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$ 5. $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$ 6. $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$

Get It?

State the three rules that define how electrons are arranged in atoms.

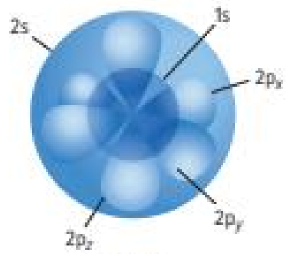
Electron Arrangement

You can represent an atom's electron configuration using one of two convenient methods: orbital diagrams or electron configuration notation.

Orbital diagrams

As mentioned earlier, electrons in orbitals can be represented by arrows in boxes. Each box is labeled with the principal quantum number and sublevel associated with the orbital. For example, the orbital diagram for a ground-state carbon atom, shown below, contains two electrons in the 1s orbital, two electrons in the 2s orbital, and one electron in two of three separate 2p orbitals. The third 2p orbital remains unoccupied.

[↑↓] [↑↓] [↑][↑][ ]
1s 2s 2p



Neon atom

[↑↓] [↑↓] [↑↓][↑↓][↑↓]
1s 2s 2p

Figure 19 The 1s, 2s, and 2p orbitals of a neon atom overlap.
Determine *how many electrons a neon atom has.*

Electron configuration notation

The electron configuration notation designates the principal energy level and energy sublevel associated with each of the atom's orbitals and includes a superscript representing the number of electrons in the orbital. For example, the electron configuration notation of a ground-state carbon atom is written $1s^2 2s^2 2p^2$.

Orbital diagrams and electron configuration notations for the elements in periods one and two of the periodic table are shown in **Table 4. Figure 19** illustrates how the 1s, 2s, $2p_x$, $2p_y$, and $2p_z$ orbitals illustrated earlier in **Figure 17** overlap in the neon atom.

Table 4 Electron Configurations and Orbital Diagrams for Elements 1–10

Element	Atomic Number	Orbital Diagram 1s 2s $2p_x$ $2p_y$ $2p_z$	Electron Configuration Notation
Hydrogen	1	[↑]	$1s^1$
Helium	2	[↑↓]	$1s^2$
Lithium	3	[↑↓] [↑]	$1s^2 2s^1$
Beryllium	4	[↑↓] [↑↓]	$1s^2 2s^2$
Boron	5	[↑↓] [↑↓] [↑]	$1s^2 2s^2 2p^1$
Carbon	6	[↑↓] [↑↓] [↑] [↑]	$1s^2 2s^2 2p^2$
Nitrogen	7	[↑↓] [↑↓] [↑] [↑] [↑]	$1s^2 2s^2 2p^3$
Oxygen	8	[↑↓] [↑↓] [↑↓] [↑] [↑]	$1s^2 2s^2 2p^4$
Fluorine	9	[↑↓] [↑↓] [↑↓] [↑↓] [↑]	$1s^2 2s^2 2p^5$
Neon	10	[↑↓] [↑↓] [↑↓] [↑↓] [↑↓]	$1s^2 2s^2 2p^6$

Note that the electron configuration notation does not usually show the orbital distributions of electrons related to a sublevel. It is understood that a designation such as nitrogen's $2p^3$ represents the orbital occupancy $2p_x^{\,1}2p_y^{\,1}2p_z^{\,1}$.

For sodium, the first ten electrons occupy 1s, 2s, and 2p orbitals. Then, according to the aufbau sequence, the eleventh electron occupies the 3s orbital. The electron configuration notation and orbital diagram for sodium are written as follows.

$$1s^22s^22p^63s^1$$

Noble-gas notation Noble gases are the elements in the last column of the periodic table. Their outermost energy levels are full, and they are unusually stable. Noble-gas notation uses bracketed symbols to represent the electron configurations of noble gases. For example, [He] represents the electron configuration for helium, $1s^2$, and [Ne] represents the electron configuration for neon, $1s^22s^22p^6$.

Compare the electron configuration for neon with sodium's configuration above. Note that the inner-level configuration for sodium is identical to the electron configuration for neon. Using noble-gas notation, sodium's electron configuration can be shortened to the form $[Ne]3s^1$. The electron configuration for an element can be represented using the noble-gas notation for the noble gas in the previous period and the electron configuration for the additional orbitals being filled. The complete and abbreviated (using noble-gas notation) electron configurations of the period 3 elements are shown in **Table 5.**

Table 5 Electron Configurations for Elements 11–18

Element	Atomic Number	Complete Electron Configuration	Electron Configuration Using Noble Gas
Sodium	11	$1s^22s^22p^63s^1$	$[Ne]3s^1$
Magnesium	12	$1s^22s^22p^63s^2$	$[Ne]3s^2$
Aluminum	13	$1s^22s^22p^63s^23p^1$	$[Ne]3s^23p^1$
Silicon	14	$1s^22s^22p^63s^23p^2$	$[Ne]3s^23p^2$
Phosphorus	15	$1s^22s^22p^63s^23p^3$	$[Ne]3s^23p^3$
Sulfur	16	$1s^22s^22p^63s^23p^4$	$[Ne]3s^23p^4$
Chlorine	17	$1s^22s^22p^63s^23p^5$	$[Ne]3s^23p^5$
Argon	18	$1s^22s^22p^63s^23p^6$	$[Ne]3s^23p^6$ or [Ar]

Get It?

Explain how to write the noble-gas notation for an element. What is the noble-gas notation for calcium?

SCIENCE USAGE V. COMMON USAGE

period

Science usage: a horizontal row of elements in the current periodic table

There are seven periods in the current periodic table.

Common usage: an interval of time determined by some recurring phenomenon

The period of Earth's orbit is one year.

WORD ORIGIN

aufbau

comes from the German word *aufbauen*, which means to *build up* or *arrange*

Exceptions to predicted configurations

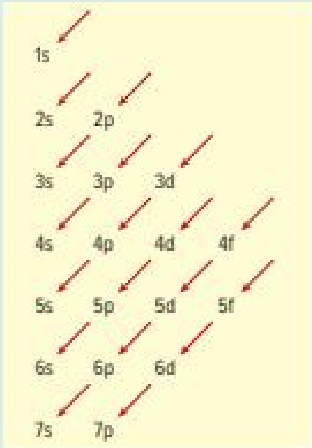
You can use the aufbau diagram to write correct ground-state electron configurations for all elements up to and including vanadium, atomic number 23. However, if you were to proceed in this manner, your configurations for chromium, $[Ar]4s^2 3d^4$, and copper, $[Ar]4s^2 3d^9$, would be incorrect. The correct configurations for these two elements are $[Ar]4s^1 3d^5$ for chromium and $[Ar]4s^1 3d^{10}$ for copper. The electron configurations for these two elements, as well as those of several other elements, illustrate the increased stability of half-filled and filled sets of s and d orbitals.

PROBLEM-SOLVING STRATEGY

Filling Atomic Orbitals

By drawing a sublevel diagram and following the arrows, you can write the ground-state electron configuration for any element.

1. Sketch the sublevel diagram on a blank piece of paper.
2. Determine the number of electrons in one atom of the element for which you are writing the electron configuration. The number of electrons in a neutral atom equals the element's atomic number.
3. Starting with 1s, write the aufbau sequence of atomic orbitals by following the diagonal arrows from the top of the sublevel diagram to the bottom. When you complete one line of arrows, move to the right, to the beginning of the next line of arrows. As you proceed, add superscripts indicating the numbers of electrons in each set of atomic orbitals. Continue only until you have sufficient atomic orbitals to accommodate the total number of electrons in one atom of the element.
4. Apply noble-gas notation.



The sublevel diagram shows the order in which the orbitals are usually filled.

Apply the Strategy

Write the ground-state electron configuration for zirconium.

PRACTICE Problems

ADDITIONAL PRACTICE

19. Write ground-state electron configurations for the following elements.
    - a. bromine (Br)
    - b. strontium (Sr)
    - c. antimony (Sb)
    - d. rhenium (Re)
    - e. terbium (Tb)
    - f. titanium (Ti)
20. A chlorine atom in its ground state has a total of seven electrons in orbitals related to the atom's third energy level. How many of the seven electrons occupy p orbitals? How many of the 17 electrons in a chlorine atom occupy p orbitals?
21. When a sulfur atom reacts with other atoms, electrons in the atom's third energy level are involved. How many such electrons does a sulfur atom have?
22. An element has the ground-state electron configuration $[Kr]5s^2 4d^{10} 5p^1$. It is part of some semiconductors and used in various alloys. What element is it?
23. **CHALLENGE** In its ground state, an atom of an element has two electrons in all orbitals related to the atom's highest energy level for which $n = 6$. Using noble-gas notation, write the electron configuration for this element, and identify the element.

Valence Electrons

Only certain electrons, called valence electrons, determine the chemical properties of an element. **Valence electrons** are defined as electrons in the atom's outermost orbitals–generally those orbitals associated with the atom's highest principal energy level. For example, a sulfur atom contains 16 electrons, only six of which occupy the outermost 3s and 3p orbitals, as shown by sulfur's electron configuration, $[Ne]3s^23p^4$. Sulfur has six valence electrons. Similarly, although a cesium atom has 55 electrons, it has just one valence electron, the 6s electron shown in cesium's electron configuration, $[Xe]6s^1$.

Get It?
Cite Evidence How do the properties of electrons influence the properties of elements?

Electron-dot structures

Because valence electrons are involved in forming chemical bonds, chemists often represent them visually using a simple shorthand method, called electron-dot structure. An atom's **electron-dot structure** consists of the element's symbol, which represents the atomic nucleus and inner-level electrons, surrounded by dots representing all of the atom's valence electrons. In writing an atom's electron-dot structure, dots representing valence electrons are placed one at a time on the four sides of the symbol (they may be placed in any sequence) and then paired up until all are shown. **Table 6** shows examples for the second period.

Table 6 Electron Configurations and Dot Structures

Element	Atomic Number	Electron Configuration	Electron-Dot Structure
Lithium	3	$1s^22s^1$	$\text{Li}\cdot$
Beryllium	4	$1s^22s^2$	$\cdot\text{Be}\cdot$
Boron	5	$1s^22s^22p^1$	$\cdot\dot{\text{B}}\cdot$
Carbon	6	$1s^22s^22p^2$	$\cdot\underset{\cdot}{\dot{\text{C}}}\cdot$
Nitrogen	7	$1s^22s^22p^3$	$\cdot\underset{\cdot}{\ddot{\text{N}}}\cdot$
Oxygen	8	$1s^22s^22p^4$	$:\underset{\cdot}{\ddot{\text{O}}}\cdot$
Fluorine	9	$1s^22s^22p^5$	$:\underset{\cdot\cdot}{\ddot{\text{F}}}\cdot$
Neon	10	$1s^22s^22p^6$	$:\underset{\cdot\cdot}{\ddot{\text{Ne}}}:$

EXAMPLE Problem 3

ELECTRON-DOT STRUCTURES Some toothpastes contain stannous fluoride, a compound of tin and fluorine. What is tin's electron-dot structure?

1 ANALYZE THE PROBLEM

Consult the periodic table to determine the total number of electrons in a tin atom. Write out tin's electron configuration, and determine its number of valence electrons. Then use the rules for electron-dot structures to draw the electron-dot structure for tin.

2 SOLVE FOR THE UNKNOWN

Tin has an atomic number of 50. Thus, a tin atom has 50 electrons.

$[Kr]5s^24d^{10}5p^2$ Write out tin's electron configuration using noble-gas notation. The closest noble gas is Kr.

The two 5s and the two 5p electrons (the electrons in the orbitals related to the atom's highest principal energy level) represent tin's four valence electrons. Draw the four valence electrons around tin's chemical symbol (Sn) to show tin's electron-dot structure. $\cdot\underset{\cdot}{\dot{\text{Sn}}}\cdot$

EXAMPLE Problem 3 (continued)

3 EVALUATE THE ANSWER

The correct symbol for tin (Sn) has been used, and the rules for drawing electron-dot structures have been correctly applied.

PRACTICE Problems ADDITIONAL PRACTICE

24. Draw electron-dot structures for atoms of the following elements.

a. magnesium **b.** thallium **c.** xenon

25. An atom of an element has a total of 13 electrons. What is the element, and how many electrons are shown in its electron-dot structure?

26. CHALLENGE This element exists in the solid state at room temperature and at normal atmospheric pressure and is found in emerald gemstones. It is known to be one of the following elements: carbon, germanium, sulfur, cesium, beryllium, or argon. Identify the element based on the electron-dot structure at right.

·X·

Check Your Progress

Summary

- The arrangement of electrons in an atom is called the atom's electron configuration.
- Electron configurations are defined by the aufbau principle, the Pauli exclusion principle, and Hund's rule.
- An element's valence electrons determine the chemical properties of the element.
- Electron configurations can be represented using orbital diagrams, electron configuration notation, and electron-dot structures.

Demonstrate Understanding

27. **Apply** the Pauli exclusion principle, the aufbau principle, and Hund's rule to write the electron configuration and draw the orbital diagram for each of the following elements.

a. silicon b. fluorine c. calcium d. krypton

28. **Define** *valence electron.*

29. **Illustrate** and describe the sequence in which ten electrons occupy the five orbitals related to an atom's d sublevel.

30. **Extend** the aufbau sequence through an element that has not yet been identified, but whose atoms would completely fill 7p orbitals. How many electrons would such an atom have? Write its electron configuration using noble-gas notation for the previous noble gas, radon.

31. **Interpret Scientific Illustrations** Which is the correct electron-dot structure for an atom of selenium? Explain.

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SCIENTIFIC BREAKTHROUGHS

Batteries of the Future: Super Charged!

Batteries are fundamental to modern technology. Since the invention of the lithium ion battery, researchers have been looking for a better energy source for cars, smart phones, and computers.

Lithium ion batteries are used in electronics.

The New Wave of Battery Power

In most batteries available in devices today, the electrolyte is liquid. However, some researchers are working to develop safer batteries with water-based, air, or solid electrolytes.

The U.S. Army Research Laboratory, in collaboration with the University of Maryland, is shaking things up with new technology that uses a saltwater electrolyte. The researchers say this eliminates the fire and explosion risk associated with some non-aqueous lithium ion batteries, which is especially a concern for military personnel in combat situations. The technology needs to be perfected before it is made commercially available, but so far, it is the first battery of its kind to reach the 4.0 volt mark essential for many electronics.

Lithium-air batteries are another promising possibility for the future. It is projected they could provide three times as much power for a given weight compared with lithium-ion batteries. Researchers across the country are working to determine which electrolyte material, such as lithium iodide, will be most efficient in these batteries to make them cheaper and more powerful.

Solid state batteries use polymer electrolytes to eliminate the liquid electrolyte entirely. This creates a safer, fire-resistant, more powerful, and rechargeable energy source. Solid state batteries are an important advancement to the electric automotive industry, which is currently limited by the range of the best lithium ion batteries.

Given the pervasiveness of battery-powered electronics, new faster, stronger, and safer options could be the catalyst for the next big breakthrough in battery technology.

MAKE AND DEFEND A CLAIM

Research one type of battery described in this feature. Write a report summarizing why you think this battery will or will not be successful in the marketplace.

STUDY GUIDE

GO ONLINE to study with your Science Notebook

Lesson 1 LIGHT AND QUANTIZED ENERGY

- All waves are defined by their wavelengths, frequencies, amplitudes, and speeds.

$$c = \lambda\nu$$

- In a vacuum, all electromagnetic waves travel at the speed of light.
- All electromagnetic waves have both wave and particle properties.
- Matter emits and absorbs energy in quanta.

$$E_{quantum} = h\nu$$

- White light produces a continuous spectrum. An element's emission spectrum consists of a series of lines of individual colors.

- electromagnetic radiation
- wavelength
- frequency
- amplitude
- electromagnetic spectrum
- quantum
- Planck's constant
- photoelectric effect
- photon
- atomic emission spectrum

Lesson 2 QUANTUM THEORY AND THE ATOM

- Bohr's atomic model attributes hydrogen's emission spectrum to electrons dropping from higher-energy to lower-energy orbits.

$$\Delta E = E_{higher\text{-}energy\ orbit} - E_{lower\text{-}energy\ orbit} = E_{photon} = h\nu$$

- The de Broglie equation relates a particle's wavelength to its mass, its velocity, and Planck's constant.

$$\lambda = h/m\nu$$

- The quantum mechanical model assumes that electrons have wave properties.
- Electrons occupy three-dimensional regions of space called atomic orbitals.

- ground state
- quantum number
- de Broglie equation
- Heisenberg uncertainty principle
- quantum mechanical model of the atom
- atomic orbital
- principal quantum number
- principal energy level
- energy sublevel

Lesson 3 ELECTRON CONFIGURATION

- The arrangement of electrons in an atom is called the atom's electron configuration.
- Electron configurations are defined by the aufbau principle, the Pauli exclusion principle, and Hund's rule.
- An element's valence electrons determine the chemical properties of the element.
- Electron configurations can be represented using orbital diagrams, electron configuration notation, and electron-dot structures.

- electron configuration
- aufbau principle
- Pauli exclusion principle
- Hund's rule
- valence electron
- electron-dot structure

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

How do we know what stars are made of?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

SEP Data Analysis Lab

What electron transitions account for the Balmer series?

Hydrogen's emission spectrum comprises three series of lines. Some wavelengths are ultraviolet (Lyman series) and infrared (Paschen series). Visible wavelengths comprise the Balmer series. The Bohr atomic model attributes these spectral lines to transitions from higher-energy states with electron orbits in which $n = n_i$ to lower-energy states with smaller electron orbits in which $n = n_f$.

CER Analyze and Interpret Data

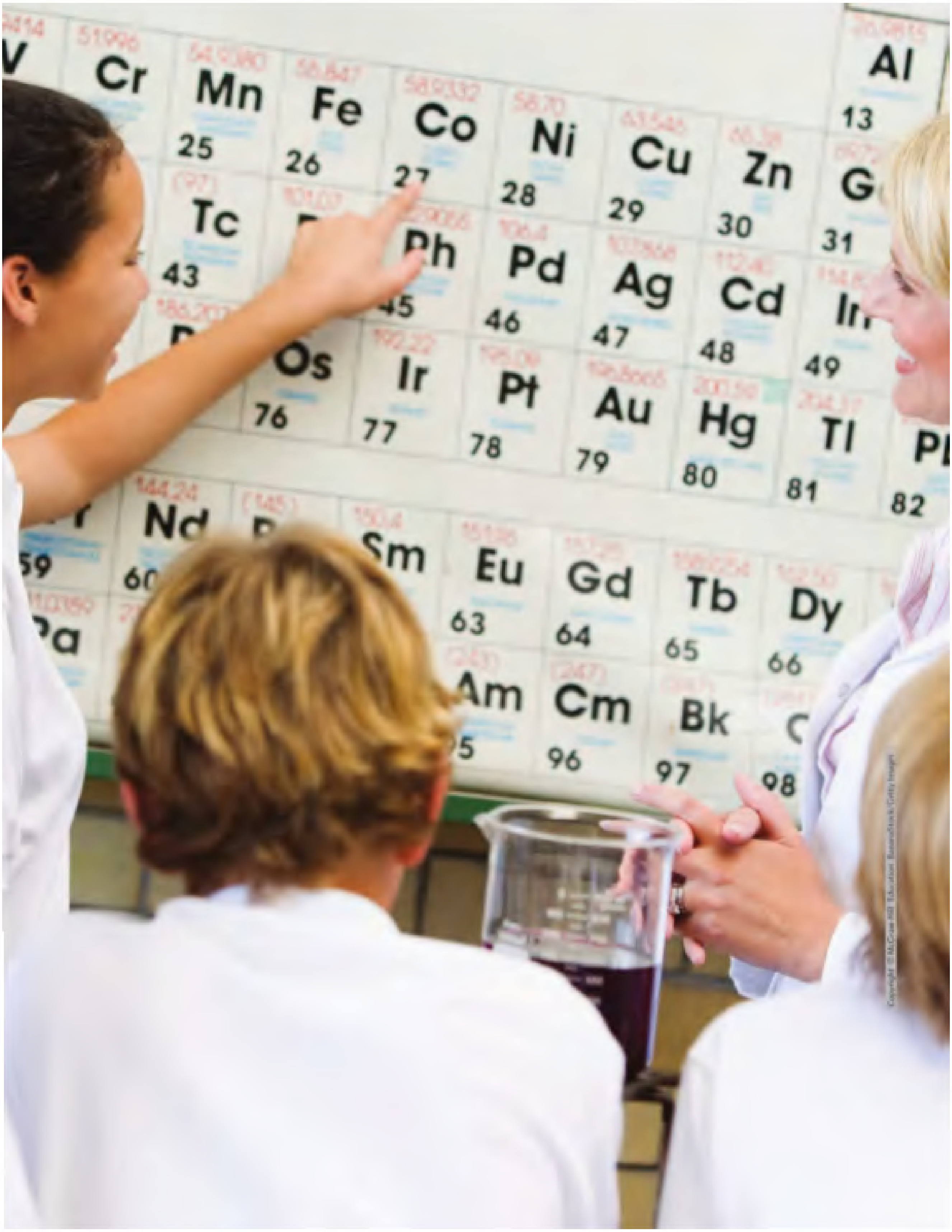
Some hydrogen balmer lines are designated H_α (6562 Å), H_β (4861 Å), H_γ (4340 Å), and H_δ (4101 Å). Each wavelength (λ) is related to an electron transition within a hydrogen atom by the following equation, in which $1.09678 \times 10^7\ m^{-1}$ is known as the Rydberg constant.

$$\frac{1}{\lambda} = 1.09678 \times 10^7 \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)\ m^{-1}$$

For hydrogen's Balmer series, electron orbit transitions occur from larger orbits to the $n = 2$ orbit; that is, $n_f = 2$.

CER Analyze and Interpret Data

1. **Calculate the wavelengths for the following electron orbit transitions.**
   a. $n_i = 3$; $n_f = 2$ **c.** $n_i = 5$; $n_f = 2$
   b. $n_i = 4$; $n_f = 2$ **d.** $n_i = 6$; $n_f = 2$
2. **Claim, Evidence, Reasoning** Relate the Balmer-series wavelengths you calculated in Question 1 to those determined experimentally. Allowing for experimental error and calculation uncertainty, do the wavelengths match? Explain your answer. One angstrom (Å) equals 10^{-10} m.
3. **Apply** the formula $E = hc/\lambda$ to determine the energy per quantum for each of the orbit transitions in Question 1.


Cr
Mn 25
Fe 26
Co 27
Ni 28
Cu 29
Zn 30
Al 13
31
Tc 43
Pd 46
Ag 47
Cd 48
49
Os 76
Ir 77
Pt 78
Au 79
Hg 80
Tl 81
82
Nd 60
Sm
Eu 63
Gd 64
Tb 65
Dy 66
59
Am
Cm 96
Bk 97
98

THE PERIODIC TABLE AND PERIODIC LAW

ENCOUNTER THE PHENOMENON

What can we learn from the periodic table?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

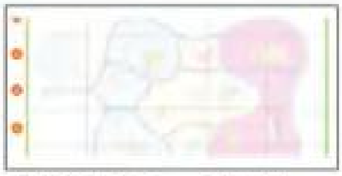
CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about what we can learn from the periodic table.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.



LESSON 1: Explore & Explain: Groups and Periods

LESSON 2: Explore & Explain: Electron Configuration and the Periodic Table

LESSON 1

DEVELOPMENT OF THE MODERN PERIODIC TABLE

FOCUS QUESTION

How are elements organized in the periodic table?

Development of the Periodic Table

In the late 1700s, French scientist Antoine Lavoisier (1743–1794) compiled a list of all elements that were known at the time. The list, shown in **Table 1**, contained 33 elements organized in four categories. Many of these elements, such as silver, gold, carbon, and oxygen, have been known since prehistoric times.

The 1800s brought a large increase in the number of known elements. The advent of electricity, which was used to break down compounds into their components, and the development of the spectrometer, which was used to identify the newly isolated elements, played major roles in the advancement of chemistry. The industrial revolution of the mid-1800s also played a major role, which led to the development of many new chemistry-based industries, such as the manufacture of petrochemicals, soaps, dyes, and fertilizers. By 1870, there were over 60 known elements.

Along with the discovery of new elements came volumes of new scientific data related to the elements and their compounds. Chemists of the time were overwhelmed with learning the properties of so many new elements and compounds. What chemists needed was a tool for organizing the many facts associated with the elements.

Table 1 Lavoisier's Table of Simple Substances (Old English Names)

Gases	light, heat, dephlogisticated air, phlogisticated gas, inflammable air
Metals	antimony, silver, arsenic, bismuth, cobalt, copper, tin, iron, manganese, mercury, molybdena, nickel, gold, platina, lead, tungsten, zinc
Nonmetals	sulphur, phosphorus, pure charcoal, radical muriatique*, radical fluorique*, radical boracique*
Earths	chalk, magnesia, barote, clay, siliceous earth

*no English name

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Virtual Investigation: Periodic Properties

Analyze and interpret the data in the periodic table of elements for patterns of organization and the properties of matter.

ChemLAB: Investigate Descriptive Chemistry

Analyze and interpret data to determine patterns of properties in representative elements.

Organizing the elements

A significant step toward developing a tool for organizing the elements and the large amount of data about their properties came in 1860, when chemists agreed upon a method for accurately determining the atomic masses of the elements. Until this time, different chemists used different mass values in their work, making the results of one chemist's work hard to reproduce by another.

With newly agreed-upon atomic masses for the elements, the search for relationships between atomic mass and elemental properties, and a way to organize the elements, began in earnest.

John Newlands

In 1864, English chemist John Newlands (1837–1898) proposed an organizational scheme for the elements. He noticed that when the elements were arranged by increasing atomic mass, their properties repeated every eighth element. A pattern such as this is called periodic because it repeats in a specific manner. Newlands named the periodic relationship that he observed in chemical properties the law of octaves, after the musical octave in which notes repeat every eighth tone.

Figure 1 shows how Newlands organized 14 of the elements known in the mid-1860s. Acceptance of the law of octaves was hampered because the law did not work for all of the known elements. Also, the use of the word octave was harshly criticized by fellow scientists, who thought that the musical analogy was unscientific. While his law was not generally accepted, the passage of a few years would show that Newlands was basically correct: the properties of elements do repeat in a periodic way.

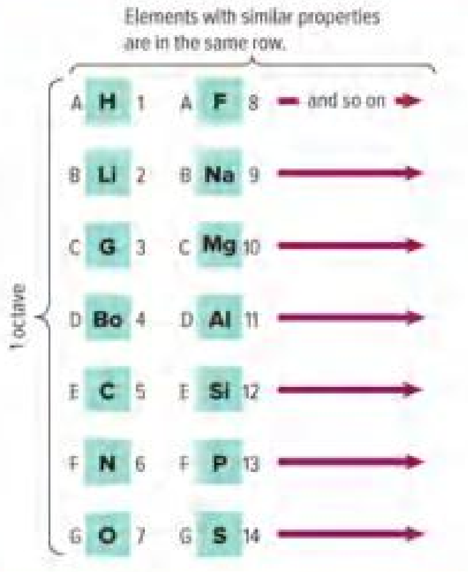


Figure 1 John Newlands noticed that the properties of elements repeated every eighth element in the same way musical notes repeat every eighth note and form octaves.

Meyer and Mendeleev

In 1869, German chemist Lothar Meyer (1830–1895) and Russian chemist Dmitri Mendeleev (1834–1907) each demonstrated a connection between atomic mass and the properties of elements. Mendeleev, however, is generally given more credit than Meyer because he published his organizational scheme first.

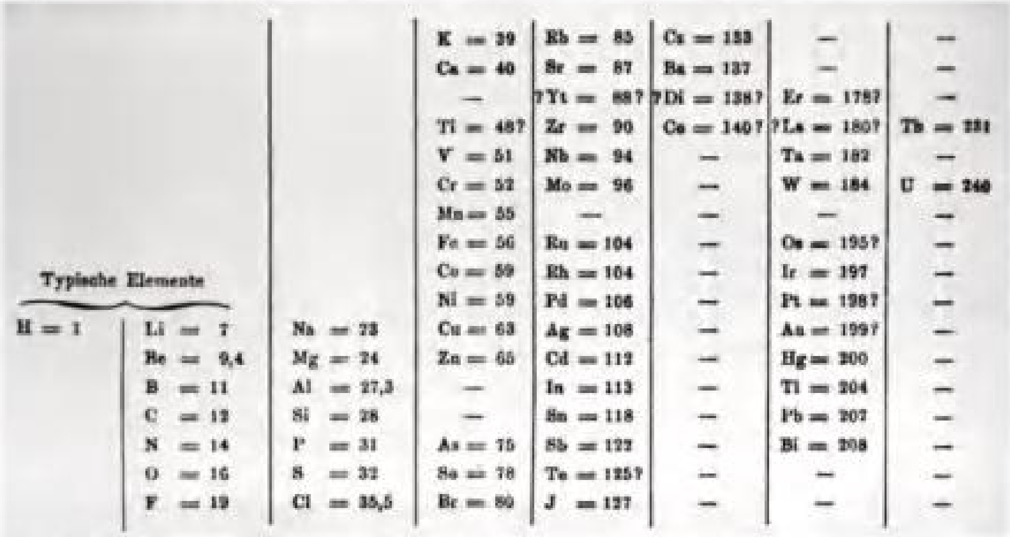
Like Newlands several years earlier, Mendeleev noticed that when the elements were ordered by increasing atomic mass, there was a periodic pattern in their properties.

Get It?

Describe the pattern that both Newlands and Mendeleev noticed about the properties of the elements.

By arranging the elements in order of increasing atomic mass into columns with similar properties, Mendeleev organized the elements into a periodic table.

Mendeleev's table, shown in **Figure 2**, became widely accepted because he predicted the existence and properties of undiscovered elements that were later found. Mendeleev left blank spaces in the table where he thought the undiscovered elements should go. By noting trends in the properties of known elements, he was able to predict the properties of the yet-to-be-discovered elements scandium, gallium, and germanium.



			K = 39	Rb = 85	Cs = 133	—	—
			Ca = 40	Sr = 87	Ba = 137	—	—
			—	?Yt = 88?	?Di = 138?	Er = 178?	—
			Ti = 48?	Zr = 90	Ce = 140?	?La = 180?	Th = 231
			V = 51	Nb = 94	—	Ta = 182	—
			Cr = 52	Mo = 96	—	W = 184	U = 240
			Mn = 55	—	—	—	—
			Fe = 56	Ru = 104	—	Os = 195?	—
Typische Elemente			Co = 59	Rh = 104	—	Ir = 197	—
			Ni = 59	Pd = 106	—	Pt = 198?	—
H = 1	Li = 7	Na = 23	Cu = 63	Ag = 108	—	Au = 199?	—
	Be = 9,4	Mg = 24	Zn = 65	Cd = 112	—	Hg = 200	—
	B = 11	Al = 27,3	—	In = 113	—	Tl = 204	—
	C = 12	Si = 28	—	Sn = 118	—	Pb = 207	—
	N = 14	P = 31	As = 75	Sb = 122	—	Bi = 208	—
	O = 16	S = 32	Se = 78	Te = 125?	—	—	—
	F = 19	Cl = 35,5	Br = 80	J = 127	—	—	—

Figure 2 In the first version of his table, published in 1869, Mendeleev arranged elements with similar chemical properties horizontally. He left empty spaces for elements that were not yet discovered.

Moseley

Mendeleev's table, however, was not completely correct. After several new elements were discovered and the atomic masses of the known elements were more accurately determined, it became apparent that several elements in his table were not in the correct order. Arranging the elements by mass resulted in several elements being placed in groups of elements with differing properties. The reason for this problem was determined in 1913 by English chemist Henry Moseley (1887–1915). Moseley discovered that atoms of each element contain a unique number of protons in their nuclei—the number of protons being equal to the atom's atomic number. By arranging the elements in order of increasing atomic number, the problems with the order of the elements in the periodic table were solved and a clear periodic pattern of properties resulted.

The statement that there is a periodic repetition of chemical and physical properties of the elements when they are arranged by increasing atomic number is called the **periodic law.**

Get It?

Compare and contrast the ways in which Mendeleev and Moseley organized the elements.

Table 2 summarizes the contributions of Newlands, Meyer, Mendeleev, and Moseley to the development of the periodic table.

The periodic table brought order to seemingly unrelated facts and became a significant tool for chemists. It is a useful reference for understanding and predicting the properties of elements and for organizing knowledge of atomic structure.

Table 2 Contributions to the Classification of Elements

John Newlands (1837–1898)
• arranged elements by increasing atomic mass • noticed the repetition of properties every eighth element • created the law of octaves
Lothar Meyer (1830–1895)
• demonstrated a connection between atomic mass and elements' properties • arranged the elements in order of increasing atomic mass
Dmitri Mendeleev (1834–1907)
• demonstrated a connection between atomic mass and elements' properties • arranged the elements in order of increasing atomic mass • predicted the existence and properties of undiscovered elements
Henry Moseley (1887–1915)
• discovered that atoms contain a unique number of protons called the atomic number • arranged elements in order of increasing atomic number, which resulted in a periodic pattern of properties

The Modern Periodic Table

The modern periodic table consists of boxes, each containing an element name, symbol, atomic number, and atomic mass. A typical box from the table is shown in **Figure 3**.

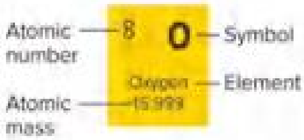


Figure 3 A typical box from the periodic table contains the element's name, its chemical symbol, its atomic number, and its atomic mass.

The table orders elements horizontally by the number of protons in an atom's nucleus, and places those with similar chemical properties in columns. The columns are called **groups** or families. The rows are called **periods.**

The periodic table is shown in **Figure 4** on the next page and on the inside back cover of your textbook. Becoming familiar with the periodic table will help you understand how the properties of different elements relate to one another.

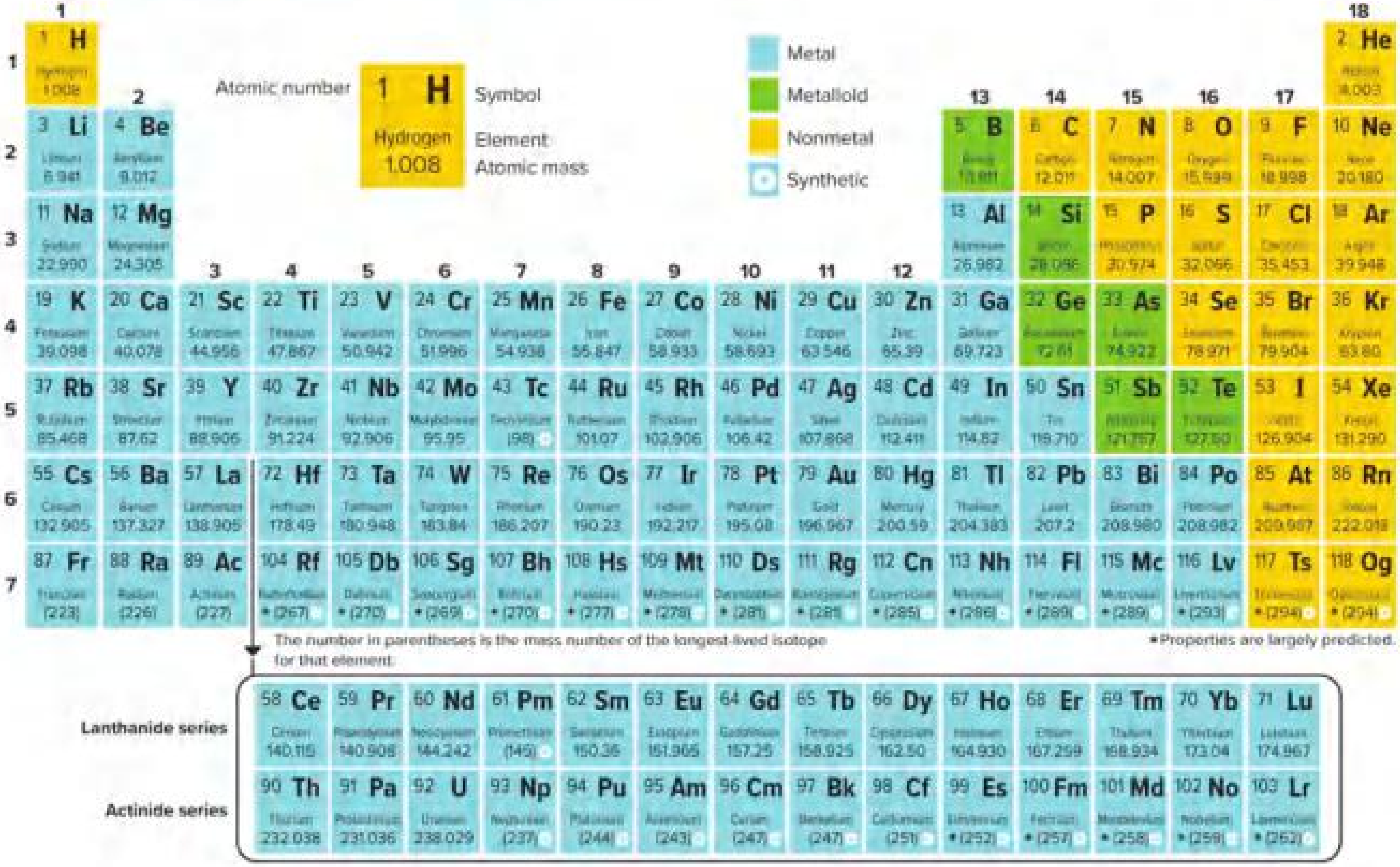
CCC CROSSCUTTING CONCEPTS

Patterns Different patterns can be observed in the periodic table. The patterns organize and can predict the properties of elements. Compare and contrast the periods and groups of the table, shown in **Figure 4,** based on their atomic number and atomic mass. Create a graphic organizer or other simple visual that will help you and your classmates remember the patterns.

WORD ORIGIN

periodic
comes from the Greek word *periodos*, meaning *way around, circuit*

Figure 4 **Periodic Table of the Elements**



Groups and periods

Beginning with hydrogen in period 1, there are a total of seven periods. Each group is numbered 1 through 18. For example, period 4 contains potassium and calcium. Oxygen is in group 16. The elements in groups 1, 2, and 13 to 18 possess a wide range of chemical and physical properties. For this reason, they are often referred to as the main group, or **representative elements.** The elements in groups 3 to 12 are referred to as the **transition elements.** Elements are also classified as metals, nonmetals, and metalloids.

Metals

Elements that are generally shiny when smooth and clean, solid at room temperature, and good conductors of heat and electricity are called **metals.** Most metals are also malleable and ductile, meaning that they can be pounded into thin sheets and drawn into wires, respectively, as shown in **Figure 5**.

Most representative elements and all transition elements are metals. If you look at boron (B) in column 13, you will see a heavy stairstep line that zigzags down to astatine (At) at the bottom of group 17. This stairstep line is a visual divider between the metals and the nonmetals on the table. In the periodic table shown in **Figure 4** metals are represented by the blue boxes.

Alkali Metals Except for hydrogen, all of the elements on the left side of the table are metals. The group 1 elements (except for hydrogen) are known as the **alkali metals.** Because they are so reactive, alkali metals usually exist as compounds with other elements. Two familiar alkali metals are sodium (Na), one of the components of salt, and lithium (Li), often used in batteries.

Figure 5 Copper, like most metals, is ductile and conducts electricity well. For these reasons copper is used for electrical wiring.

SCIENCE USAGE V. COMMON USAGE

conductor

Science usage: a substance or body capable of transmitting electricity, heat, or sound

Copper is a good conductor of heat.

Common usage: a person who conducts an orchestra, chorus, or other group of musical performers

The new conductor helped the orchestra perform at its best.

Alkaline Earth Metals The **alkaline earth metals** are in group 2. They are also highly reactive. Calcium (Ca) and magnesium (Mg), two minerals important for your health, are examples of alkaline earth metals. Because magnesium is strong and relatively light, it is used in applications in which strength and low mass are important, as shown in **Figure 6**.

Figure 6 Because magnesium is light and strong, it is often used in the production of safety devices such as these caribiners used by climbers.

Transition and Inner Transition Metals The transition elements are divided into **transition metals** and **inner transition metals.** The two sets of inner transition metals, known as the **lanthanide series** and **actinide series,** are located along the bottom of the periodic table. The rest of the elements in groups 3 to 12 make up the transition metals. Elements from the lanthanide series are used extensively as phosphors, substances that emit light when struck by electrons. Because it is strong and light, the transition metal titanium is used to make frames for bicycles and eyeglasses.

Nonmetals

BIOLOGY Connection Nonmetals occupy the upper-right side of the periodic table. They are represented by the yellow boxes in **Figure 4**. **Nonmetals** are elements that are generally gases or brittle, dull-looking solids. They are poor conductors of heat and electricity. The only nonmetal that is a liquid at room temperature is bromine (Br). The most abundant element in the human body is the nonmetal oxygen, which constitutes 65% of the body mass.

Group 17 comprises highly reactive elements that are known as **halogens.** Like the group 1 and group 2 elements, the halogens are often part of compounds. Compounds made with the halogen fluorine (F) are commonly added to toothpaste and drinking water to prevent tooth decay. The extremely unreactive group 18 elements are commonly called the **noble gases.** They are used in applications where their unreactivity is an advantage, such as in lasers, a variety of light bulbs, and neon signs.

Metalloids

The elements in the green boxes bordering the stairstep line in **Figure 4** are called metalloids, or semimetals.

Metalloids have physical and chemical properties of both metals and nonmetals. Silicon (Si) and germanium (Ge) are two important metalloids used extensively in computer chips and solar cells. Silicon is also used to make prosthetics or in lifelike applications, as shown in **Figure 7**.

This introduction to the periodic table touches only the surface of its durable explanatory power. You can refer to the Elements Handbook at the end of your textbook to learn more about the elements and their various groups.

Figure 7 Scientists developing submarine technology created this robot that looks and swims like a real fish. Its body is made of a silicon resin that softens in water.

Check Your Progress

Summary

- The elements were first organized by increasing atomic mass, which led to inconsistencies. Later, they were organized by increasing atomic number.
- The periodic law states that when the elements are arranged by increasing atomic number, there is a periodic repetition of their chemical and physical properties.
- The periodic table organizes the elements into periods (rows) and groups or families (columns); elements with similar properties are in the same group.
- Elements are classified as metals, nonmetals, or metalloids.

Demonstrate Understanding

1. **Describe** the development of the periodic table.
2. **Sketch** a simplified version of the periodic table, and indicate the location of metals, nonmetals, and metalloids.
3. **Describe** the general characteristics of metals, nonmetals, and metalloids.
4. **Identify** each of the following as a representative element or a transition element.
   a. lithium (Li)
   b. platinum (Pt)
   c. promethium (Pm)
   d. carbon (C)
5. **Compare** For each of the given elements, list two other elements with similar chemical properties.
   a. iodine (I) b. barium (Ba) c. iron (Fe)
6. **Compare** According to the periodic table, which two elements have an atomic mass less than twice their atomic number?
7. **Interpret Data** A company plans to make an electronic device. They need to use an element that has chemical behavior similar to that of silicon (Si) and lead (Pb). The element must have an atomic mass greater than that of sulfur (S) but less than that of cadmium (Cd). Use the periodic table to predict which element the company will use.

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LESSON 2

CLASSIFICATION OF THE ELEMENTS

FOCUS QUESTION

Why do elements in the same group have similar properties?

Organizing the Elements by Electron Configuration

As you learned previously, electron configuration determines the chemical properties of an element. Writing out electron configurations using the aufbau diagram can be tedious. Fortunately, you can determine an atom's electron configuration and its number of valence (outermost) electrons from its position on the periodic table. The repeating patterns of the table reflect patterns of outer electron states. The electron configurations for some of the group 1 elements are listed in **Table 3**. All four configurations have a single electron in their outermost orbitals.

Valence electrons

Recall that electrons in the highest principal energy level of an atom are called valence electrons. Each of the group 1 elements has one valence electron. The group 1 elements have similar chemical properties because they have the same number of valence electrons. This is one of the most important relationships in chemistry: atoms in the same group have similar chemical properties because they have the same number of valence electrons. Each group 1 element has a valence electron configuration of s^1. Each group 2 element has a valence electron configuration of s^2. Groups 1, 2, and 13 to 18 all have their own valence electron configurations.

Table 3 Electron Configuration for the Group 1 Elements

Period 1	hydrogen	$1s^1$	$1s^1$
Period 2	lithium	$1s2s^1$	$[He]2s^1$
Period 3	sodium	$1s^22s^22p^63s^1$	$[Ne]3s^1$
Period 4	potassium	$1s^22s^22p^63s^23p64s^1$	$[Ar]4s^1$

3D THINKING DCI CCC SEP

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

? **Revisit the Encounter the Phenomenon Question**

What information from this lesson can help you answer the module question?

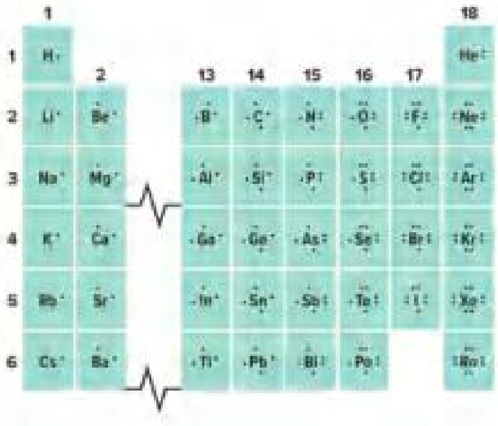


Figure 8 The figure shows the electron-dot structure of most representative elements.
Observe *How does the number of valence electrons vary within a group?*

Valence electrons and period

The energy level of an element's valence electrons indicates the period on the periodic table in which it is found. For example, lithium's valence electron is in the second energy level and lithium is found in period 2. Now look at gallium, with its electron configuration of $[Ar]4s^2 3d^{10} 4p^1$. Gallium's valence electrons are in the fourth energy level, and gallium is found in the fourth period.

Valence electrons of the representative elements

Elements in group 1 have one valence electron; group 2 elements have two valence electrons. Group 13 elements have three valence electrons, group 14 elements have four, and so on. The noble gases in group 18 each have eight valence electrons, with the exception of helium, which has only two valence electrons. **Figure 8** shows how the electron-dot structures you studied previously illustrate the connection between group number and number of valence electrons. Notice that the number of valence electrons for the elements in groups 13 to 18 is ten less than their group number.

The s-, p-, d-, and f-Block Elements

The periodic table has columns and rows of varying sizes. The reason behind the table's odd shape becomes clear if it is divided into sections, or blocks, representing the atom's energy sublevel being filled with valence electrons. Because there are four different energy sublevels (s, p, d, and f), the periodic table is divided into four distinct blocks, as shown in **Figure 9** on the next page.

s-Block elements

The s-block consists of groups 1 and 2, and the element helium. Group 1 elements have partially filled s orbitals containing one valence electron and electron configurations ending in s^1. Group 2 elements have completely filled s orbitals containing two valence electrons and electron configurations ending in s^2. Because s orbitals hold two electrons at most, the s-block spans two groups.

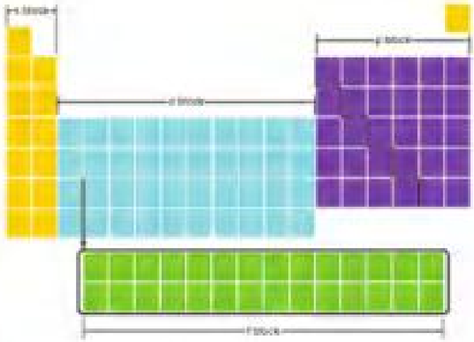


Figure 9 The periodic table is divided into four blocks—s, p, d, and f.

Analyze *What is the relationship between the maximum number of electrons an energy sublevel can hold and the number of columns in that block on the diagram?*

p-Block elements

After the s sublevel is filled, the valence electrons next occupy the p sublevel. The p-block is comprised of groups 13 through 18 and contains elements with filled or partially filled p orbitals. There are no p-block elements in period 1 because the p sublevel does not exist for the first principal energy level (n = 1). The first p-block element is boron (B), which is in the second period. The p-block spans six groups because the three p orbitals can hold a maximum of six electrons.

The group 18 elements, which are called the noble gases, are unique members of the p-block. The atoms of these elements are so stable that they undergo virtually no chemical reactions. The electron configurations of the first four noble gas elements are shown in **Table 4**. Here, both the s and p orbitals corresponding to the period's principal energy level are completely filled. This arrangement of electrons results in an unusually stable atomic structure. Together, the s- and p-blocks comprise the representative elements.

d-Block elements

The d-block contains the transition metals and is the largest of the blocks. With some exceptions, d-block elements are characterized by a filled outermost s orbital of energy level n, and filled or partially filled d orbitals of energy level $n - 1$.

Table 4 **Noble Gas Electron Configuration**

Period	Principal Energy Level	Element	Electron Configuration
1	$n = 1$	helium	$1s^2$
2	$n = 2$	neon	$[He]2s2^2p^6$
3	$n = 3$	argon	$[Ne]3s^23p^6$
4	$n = 4$	krypton	$[Ar]4s^23d^{10}4p^6$

ACADEMIC VOCABULARY

structure

something made up of more-or-less interdependent elements or parts

Many scientists were involved in the discovery of the structure of the atom.

As you move across a period, electrons fill the d orbitals. For example, scandium (Sc), the first d-block element, has an electron configuration of $[Ar]4s^23d^1$. Titanium (Ti), the next element on the table, has an electron configuration of $[Ar]4s^23d^2$. Note that titanium's filled outermost s orbital has an energy level of $n = 4$, while the d orbital, which is partially filled, has an energy level of $n = 3$.

As you learned previously, the aufbau principle states that the 4s orbital has a lower energy level than the 3d orbital. Therefore, the 4s orbital is filled before the 3d orbital. The five d orbitals can hold a total of 10 electrons; thus, the d-block spans 10 groups on the periodic table.

f-Block elements

The f-block contains the inner transition metals. Its elements are characterized by a filled, or partially filled outermost s orbital, and filled or partially filled 4f and 5f orbitals.

The electrons of the f sublevel do not fill their orbitals in a predictable manner. Because there are seven f orbitals holding up to a maximum of 14 electrons, the f-block spans 14 columns of the periodic table.

EXAMPLE Problem 1

ELECTRON CONFIGURATION AND THE PERIODIC TABLE Strontium, which is used to produce red fireworks, has an electron configuration of $[Kr]5s^2$. Without using the periodic table, determine the group, period, and block of strontium.

1 ANALYZE THE PROBLEM

You are given the electron configuration of strontium.

Known	Unknown
Electron configuration = $[Kr]5s^2$	Group = ?
	Period = ?
	Block = ?

2 SOLVE FOR THE UNKNOWN

The s^2 indicates that strontium's valence electrons fill the s sublevel. Thus, strontium is in **group 2** of the **s-block**.

For representative elements, the number of valence electrons can indicate the group number.

The 5 in $5s^2$ indicates that strontium is in **period 5**.

The number of the highest energy level indicates the period number.

3 EVALUATE THE ANSWER

The relationships between electron configuration and position on the periodic table have been correctly applied.

STEM CAREER Connection

Physical Chemist

Does the idea of using computers and sophisticated laboratory instruments to model, simulate, and analyze experimental results appeal to you? Are you someone who enjoys developing new theories? Physical chemists are interested in how matter behaves at the molecular and atomic level. They have a strong interest and background in chemistry, physics, and math.

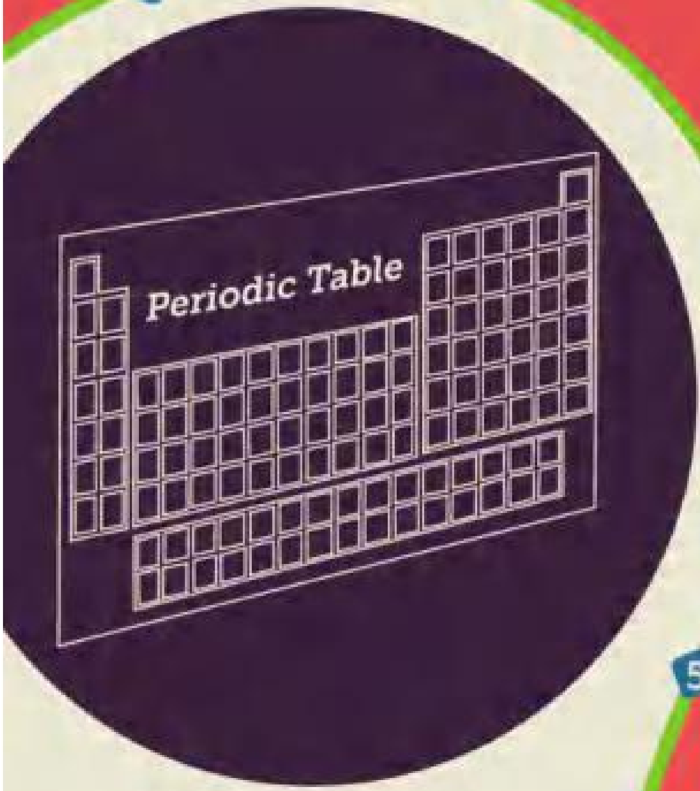


Figure 10

History of the Periodic Table

The modern periodic table is the result of the work of many scientists over the centuries who studied elements and discovered periodic patterns in their properties.

1 **1789** Antoine Lavoisier defines the chemical element, develops a list of all known elements, and distinguishes between metals and nonmetals.

2 **1814** An academic journal publishes Jöns Jacob Berzelius's paper proposing letter symbols for the known elements.

3 **1869** Lothar Meyer and Dmitri Mendeleev independently develop tables based on element characteristics and predict the properties of unknown elements.

4 **1894-1900** The noble gases—argon, helium, krypton, neon, xenon, and radon—become a new group in the periodic table.

5 **1913** Henry Moseley determines the atomic number of known elements and establishes that element properties vary periodically with atomic number.

6 **1940** Synthesized elements with an atomic number larger than 92 become part of a new block of the periodic table called the actinides.

7 **1969** Researchers at the University of California, Berkeley synthesize the first element heavier than the actinides. It has a half-life of 4.7 seconds and is named rutherfordium.

8 **2016** The syntheses of elements 113, 115, 117, and 118 are confirmed. This completes the seventh period of the periodic table.

PRACTICE Problems ADDITIONAL PRACTICE

8. Without using the periodic table, determine the group, period, and block of an atom with the following electron configurations.
   a. $[Ne]3s^2$ b. $[He]2s^2$ c. $[Kr]5s^2 4d^{10} 5p^5$
9. What are the symbols for the elements with the following valence electron configurations?
   a. s^2d^1 b. s^2p^3 c. s^2p^6
10. **CHALLENGE** Write the electron configuration of the following elements.
    a. the group 2 element in the fourth period
    b. the group 12 element in the fourth period
    c. the noble gas in the fifth period
    d. the group 16 element in the second period

The development of the periodic table took many years, but like all scientific knowledge, it is open to change. As new elements are synthetized, identified, and named, and as new data about elements arise from experimentation, the periodic table is updated.

Refer to **Figure 10** on the previous page to learn more about the history of the periodic table and the work of the many scientists who contributed to its development. The periodic table is an essential tool in understanding and exploring chemistry and you will use it throughout your study of the subject.

Check Your Progress

Summary

- The periodic table has four blocks (s, p, d, f).
- Elements within a group have similar chemical properties.
- The group number for elements in groups 1 and 2 equals the element's number of valence electrons.
- The energy level of an atom's valence electrons equals its period number.

Demonstrate Understanding

11. **Explain** what determines the blocks in the periodic table.
12. **Determine** in which block of the periodic table are the elements having the following valence electron configurations.
    a. s^2p^4 c. s^2d^1
    b. s^1 d. s^2p^1
13. **Infer** Xenon, a nonreactive gas used in strobe lights, is a poor conductor of heat and electricity. Would you expect xenon to be a metal, a nonmetal, or a metalloid? Where would you expect it to be on the periodic table? Explain.
14. **Explain** why elements within a group have similar chemical properties.
15. **Model** Make a simplified sketch of the periodic table, and label the s-, p-, d-, and f-blocks.

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

LESSON 3

PERIODIC TRENDS

FOCUS QUESTION

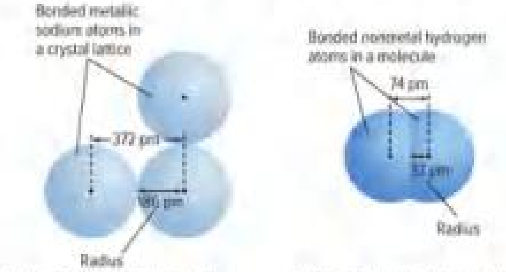
How can you use the periodic table to predict an element's properties?

Atomic Radius

Many properties of the elements tend to change in a predictable way, known as a trend, as you move across a period or down a group. Atomic size is one such periodic trend. The sizes of atoms are influenced by electron configuration.

Recall that the electron cloud surrounding a nucleus does not have a clearly defined edge. The outer limit of an electron cloud is defined as the spherical surface within which there is a 90% probability of finding an electron. However, this surface does not exist in a physical way, as the outer surface of a golf ball does. Atomic size is defined by how closely an atom lies to a neighboring atom. Because the nature of the neighboring atom can vary from one substance to another, the size of the atom itself also tends to vary somewhat from substance to substance.

For metals such as sodium, the atomic radius is defined as half the distance between adjacent nuclei in a crystal of the element, as shown in **Figure 11**. For elements that commonly occur as molecules, such as many nonmetals, the atomic radius is defined as half the distance between nuclei of identical atoms that are chemically bonded together. The atomic radius of a nonmetal diatomic hydrogen molecule (H_2) is shown in **Figure 11**.



The radius of a metal atom is one-half the distance between two adjacent atoms in the crystal.

The radius of a nonmetal atom is often determined from a molecule of two identical atoms.

Figure 11 Atomic radii depend on the type of bonds that atoms form.

3D THINKING DCI CCC SEP

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practice: Electron Patterns of Atoms

HS-PS1-1. Use the periodic table as a model to predict the relative properties of elements based on patterns of electrons in the outermost energy level of atoms.

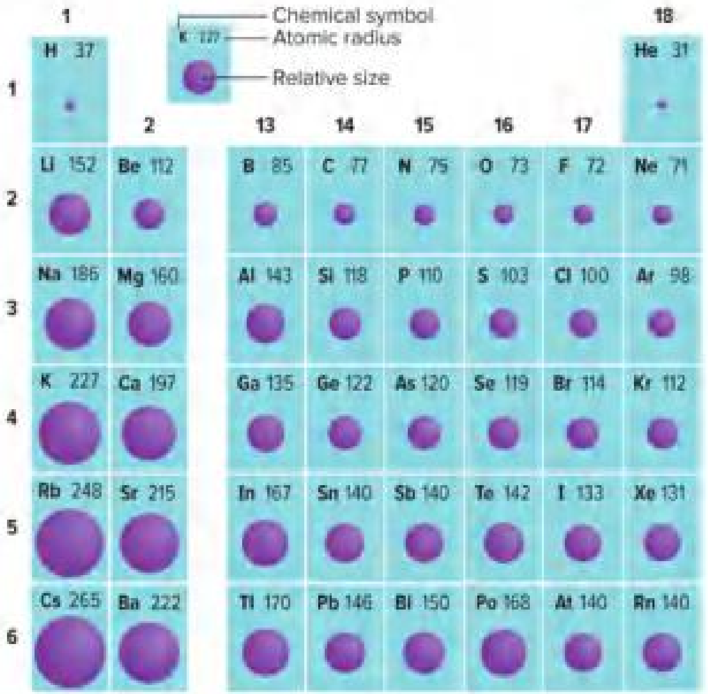


Figure 12 The atomic radii of the representative elements, given in picometers (10^{-12} m), vary as you move from left to right within a period and down a group.

Infer *why the atomic radii increase as you move down a group.*

Trends within periods

In general, there is a decrease in atomic radii as you move from left to right across a period. This trend is illustrated in **Figure 12**. It is caused by the increasing positive charge in the nucleus and the fact that the principal energy level within a period remains the same. Each successive element has one additional proton and electron, and each additional electron is added to orbitals corresponding to the same principal energy level. Moving across a period, no additional electrons come between the valence electrons and the nucleus. Thus, the valence electrons are not shielded from the increased nuclear charge, which pulls the outermost electrons closer to the nucleus.

Get It?

Discuss how the fact that the principal energy level remains the same within a period explains the decrease in the atomic radii across a period.

Trends within groups

Atomic radii generally increase as you move down a group on the periodic table. The nuclear charge increases, and electrons are added to orbitals corresponding to successively higher principal energy levels. However, the increased nuclear charge does not pull the outer electrons toward the nucleus to make the atom smaller as you might expect. Why does the increased nuclear charge not make the atom smaller?

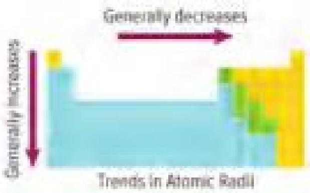


Figure 13 Atomic radii generally decrease from left to right in a period and generally increase as you move down a group.

Moving down a group, the outermost orbital increases in size along with the increasing principal energy level; thus, the atom becomes larger. The larger orbital means that the outer electrons are farther from the nucleus. This increased distance offsets the pull of the increased nuclear charge. Also, as additional orbitals between the nucleus and the outer electrons are occupied, these electrons shield the outer electrons from the nucleus. **Figure 13** summarizes the group and period trends.

EXAMPLE Problem 2

ELECTRON CONFIGURATION AND THE PERIODIC TABLE Which has the largest atomic radius: carbon (C), fluorine (F), beryllium (Be), or lithium (Li)? Answer without referring to **Figure 12**. Explain your answer in terms of trends in atomic radii.

1 ANALYZE THE PROBLEM

You are given four elements. First, determine the groups and periods the elements occupy. Then apply the general trends in atomic radii to determine which has the largest atomic radius.

2 SOLVE FOR THE UNKNOWN

From the periodic table, all the elements are found to be in period 2. — Determine the periods.

Ordering the elements from left-to-right across the period yields: Li, Be, C, and F.

The first element in period 2, lithium, has the largest radius. — Apply the trend of decreasing radii across a period.

3 EVALUATE THE ANSWER

The period trend in atomic radii has been correctly applied. Checking radii values in **Figure 12** verifies the answer.

PRACTICE Problems

ADDITIONAL PRACTICE

Answer the following questions using your knowledge of group and period trends in atomic radii. Do not use the atomic radii values in Figure 12 to answer the questions.

16. Which has the largest atomic radius: magnesium (Mg), silicon (Si), sulfur (S), or sodium (Na)? The smallest?

17. The figure on the right shows helium, krypton, and radon. Which one is krypton? How can you tell?

18. Can you determine which of two unknown elements has the larger radius if the only known information is that the atomic number of one of the elements is 20 greater than the other? Explain.

19. CHALLENGE Determine which element in each pair has the largest atomic radius:

a. the element in period 2, group 1; or the element in period 3, group 18

b. the element in period 5, group 2; or the element in period 3, group 16

c. the element in period 3, group 14; or the element in period 6, group 15

d. the element in period 4, group 18; or the element in period 2, group 16

Ionic Radius

Atoms can gain or lose one or more electrons to form ions. Because electrons are negatively charged, atoms that gain or lose electrons acquire a net charge. Thus, an **ion** is an atom or a bonded group of atoms that has a positive or negative charge.

You will learn more about ions later, but for now, consider how the formation of an ion affects the size of an atom.

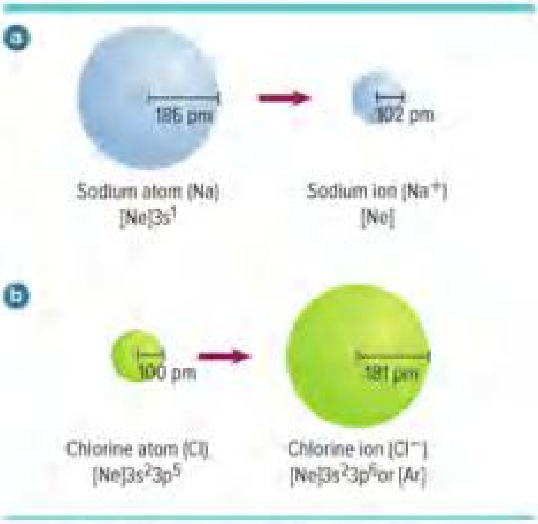


Figure 14 The size of atoms varies greatly when they form ions.

a. Positive ions are smaller than the neutral atoms from which they form.

b. Negative ions are larger than the neutral atoms from which they form.

Losing electrons

When atoms lose electrons and form positively charged ions, they always become smaller. The reason is twofold. The electron lost from the atom will almost always be a valence electron. The loss of a valence electron can leave a completely empty outer orbital, which results in a smaller radius. Furthermore, the electrostatic repulsion between the now-fewer number of remaining electrons decreases. As a result, they experience a greater nuclear charge allowing these remaining electrons to be pulled closer to the positively charged nucleus.

Figure 14a illustrates how the radius of sodium decreases when sodium atoms form positive ions. The outer orbital of the sodium atom is unoccupied in the sodium ion, so the sodium ion is much smaller than the sodium atom.

Gaining electrons

When atoms gain electrons and form negatively charged ions, they become larger. The addition of an electron to an atom increases the electrostatic repulsion between the atom's outer electrons, forcing them to move farther apart. The increased distance between the outer electrons results in a larger radius.

Figure 14b shows how the radius of chlorine increases when chlorine atoms form negative ions. Adding an electron to a chlorine atom increases the electrostatic repulsion among its valence electrons. The increased repulsion causes the electrons to move farther apart and results in the radius of a chloride ion being almost twice as large as that of a chlorine atom.

Get It?
Explain why a lithium ion is smaller than a lithium atom.

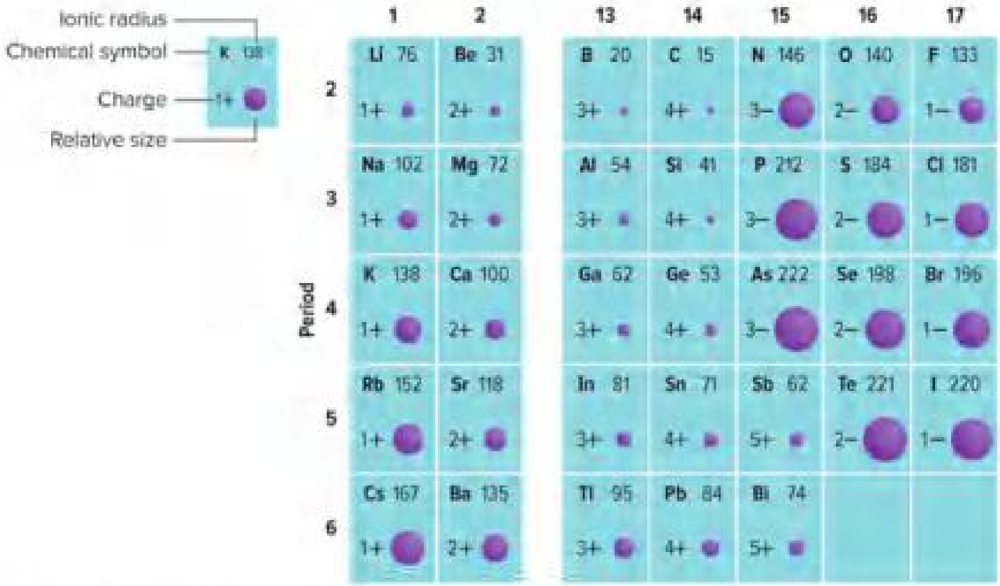


Figure 15 The ionic radii of most of the representative elements are shown in picometers (10^{-12} m).
Explain *why the ionic radii increase for both positive and negative ions as you move down a group.*

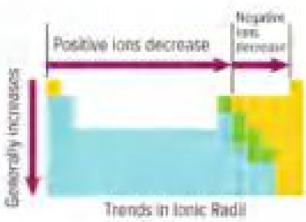
Trends within periods

The ionic radii of most of the representative elements are shown in **Figure 15**. Note that elements on the left side of the table form smaller positive ions, and elements on the right side of the table form larger negative ions.

In general, as you move from left to right across a period, the size of the positive ions gradually decreases. Then, beginning in group 15 or 16, the size of the much-larger negative ions also gradually decreases.

Trends within groups

As you move down a group, an ion's outer electrons are in orbitals corresponding to higher principal energy levels, resulting in a gradual increase in ionic size. Thus, the ionic radii of both positive and negative ions increase as you move down a group. The group and period trends in ionic radii are summarized in **Figure 16**.



Get It?

Explain why calcium has a greater atomic radius than magnesium.

Figure 16 The diagram summarizes the general trends in ionic radii.

Ionization Energy

To form a positive ion, an electron must be removed from a neutral atom. This requires energy. The energy is needed to overcome the attraction between the positive charge of the nucleus and the negative charge of the electron.

Ionization energy is defined as the energy required to remove an electron from a gaseous atom. For example, 8.64×10^{-19} J is required to remove an electron from a gaseous lithium atom. The energy required to remove the first outermost electron from an atom is called the first ionization energy. The first ionization energy of lithium equals 8.64×10^{-19} J. The loss of the electron results in the formation of a Li^{+} ion. The first ionization energies of the elements in periods 1 through 5 are plotted on the graph in **Figure 17**.

Get It?
Define *ionization energy.*

Think of ionization energy as an indication of how strongly an atom's nucleus holds onto its valence electrons. A high ionization energy value indicates the atom has a strong hold on its electrons. Atoms with large ionization energy values are less likely to form positive ions. Likewise, a low ionization energy value indicates an atom loses an outer electron easily. Such atoms are likely to form positive ions. Lithium's low ionization energy, for example, is important for its use in lithium-ion computer backup batteries, where the ability to lose electrons easily makes a battery that can quickly provide a large amount of electrical power.

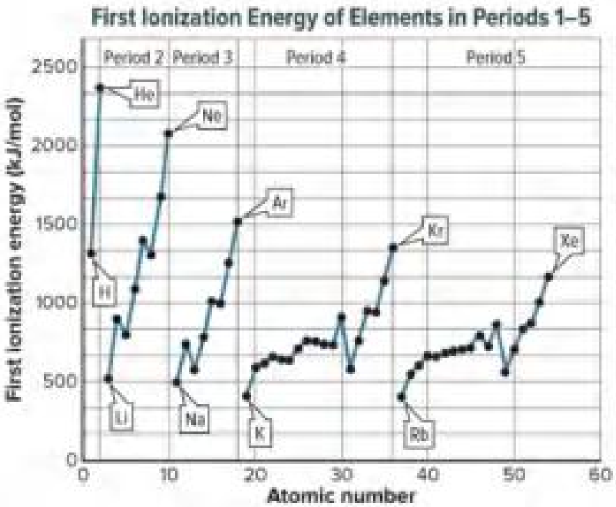


Figure 17 The first ionization energies for elements in periods 1 through 5 are shown as a function of the atomic number.

Describe *how ionization energy and atomic number are related as shown on this scatter plot.*

Each set of connected points on the graph in **Figure 17** represents the elements in a period. The group 1 metals have low ionization energies. Thus, group 1 metals (Li, Na, K, Rb) are likely to form positive ions. The group 18 elements (He, Ne, Ar, Kr, Xe) have high ionization energies and are unlikely to form ions. The stable electron configuration of gases of group 18 greatly limits their reactivity.

Removing more than one electron

After removing the first electron from an atom, it is possible to remove additional electrons. The amount of energy required to remove a second electron from a 1+ ion is called the second ionization energy, the amount of energy required to remove a third electron from a 2+ ion is called the third ionization energy, and so on. **Table 5** lists the first through ninth ionization energies for elements in period 2.

Reading across **Table 5** from left to right, you will see that the energy required for each successive ionization always increases. However, the increase in energy does not occur smoothly. Note that for each element there is an ionization for which the required energy increases dramatically. For example, the second ionization energy of lithium (7300 kJ/mol) is much greater than its first ionization energy (520 kJ/mol). This means that a lithium atom is likely to lose its first valence electron but extremely unlikely to lose its second.

Get It?

Infer how many electrons carbon is likely to lose.

If you examine the **Table 5**, you will notice that the ionization at which the large increase in energy occurs is related to the atom's number of valence electrons. The element lithium has one valence electron and the increase occurs after the first ionization energy. Lithium easily forms the common lithium 1+ ion but is unlikely to form a lithium 2+ ion. The increase in ionization energy shows that atoms hold onto their inner core electrons much more strongly than they hold onto their valence (outermost) electrons.

Table 5 Successive Ionization Energies for the Period 2 Elements

Element	Valence Electrons	Ionization Energy (kJ/mol)*								
		1st	2nd	3rd	4th	5th	6th	7th	8th	9th
Li	1	520	7300	11,810						
Be	2	900	1760	14,850	21,010					
B	3	800	2430	3660	25,020	32,820				
C	4	1090	2350	4620	6220	37,830	47,280			
N	5	1400	2860	4580	7480	9440	53,270	64,360		
O	6	1310	3390	5300	7470	10,980	13,330	71,870	84,080	
F	7	1680	3370	6050	8410	11,020	15,160	17,870	92,040	106,430
Ne	8	2080	3950	6120	9370	12,180	15,240	20,000	23,070	115,380

*mol is an abbreviation for mole, a quantity of matter

Trends within periods

As shown in **Figure 17** and by the values in **Table 5**, first ionization energies generally increase as you move from left to right across a period. The increased nuclear charge of each successive element produces an increased hold on the valence electrons.

Figure 18 Ionization energies generally increase from left to right in a period and generally decrease as you move down a group.

Trends within groups

First ionization energies generally decrease as you move down a group. This decrease in energy occurs because atomic size increases as you move down the group. Less energy is required to remove the valence electrons farther from the nucleus. **Figure 18** summarizes the group and period trends in first ionization energies.

Octet rule

When a sodium atom loses its single valence electron to form a 1+ sodium ion, its electron configuration changes as shown below.

Sodium atom $1s^2 2s^2 2p^6 3s^1$ Sodium ion $1s^2 2s^2 2p^6$

Note that the sodium ion has the same electron configuration as neon ($1s^2 2s^2 2p^6$), a noble gas. This observation leads to one of the most important principles in chemistry, the **octet rule.** The octet rule states that atoms tend to gain, lose, or share electrons in order to acquire a full set of eight valence electrons. This reinforces what you learned earlier, that the electron configuration of filled s and p orbitals of the same energy level (consisting of eight valence electrons) is unusually stable. Note that the first-period elements are an exception to the rule, as they are complete with only two valence electrons.

Real-World Chemistry

Ionization Energy

SCUBA DIVING The increased pressure that scuba divers experience far below the water's surface can cause too much oxygen to enter their blood, which would result in confusion and nausea. To avoid this, divers sometimes use a gas mixture called heliox—oxygen diluted with helium. Helium's high ionization energy ensures that it will not react chemically in the bloodstream.

Electronegativity

The **electronegativity** of an element indicates the relative ability of its atoms to attract electrons in a chemical bond. As shown in **Figure 19**, on the next page, electronegativity generally decreases as you move down a group. **Figure 19** also indicates that electronegativity generally increases as you move from left to right across a period. Fluorine is the most electronegative element, with a value of 3.98, meaning it attracts electrons more strongly than any other element in a chemical bond. Cesium and francium are the least electronegative elements, with values of 0.79 and 0.70, repectively. In a chemical bond, the atom with the greater electronegativity more strongly attracts the bond's electrons. Note that because the noble gases form very few compounds, they do not have electronegativity values.

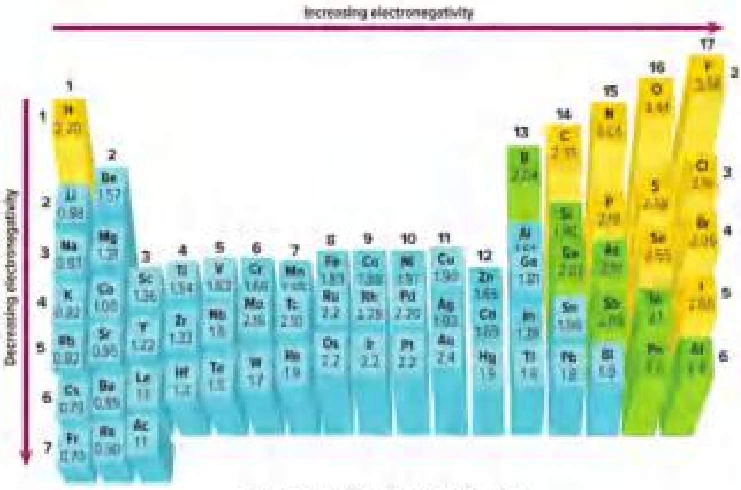


Figure 19 The electronegativity values for most of the elements are shown. The values are given in Paulings, a unit named after American scientist Linus Pauling (1901–1994).

Infer *why electronegativity values are not listed for the noble gases.*

Check Your Progress

Summary

- Atomic and ionic radii decrease from left to right across a period, and increase as you move down a group.
- Ionization energies generally increase from left to right across a period, and decrease down a group.
- The octet rule states that atoms gain, lose, or share electrons to acquire a full set of eight valence electrons.
- Electronegativity generally increases from left to right across a period, and decreases down a group.

Demonstrate Understanding

20. **Explain** how the period and group trends in atomic radii are related to electron configuration.
21. **Indicate** whether fluorine or bromine has a larger value for each of the following properties.
    a. electronegativity
    b. ionic radius
    c. atomic radius
    d. ionization energy
22. **Explain** why it takes more energy to remove the second electron from a lithium atom than it does to remove the fourth electron from a carbon atom.
23. **Calculate** Determine the differences in electronegativity, ionic radius, atomic radius, and first ionization energy for oxygen and beryllium.
24. **Make and Use Graphs** Graph the atomic radii of the representative elements in periods 2, 3, and 4 versus their atomic numbers. Connect the points of elements in each period, so that there are three separate curves on the graph. Summarize the trends in atomic radii shown on your graph. Explain.

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

NATURE OF SCIENCE

The Evolving Periodic Table

Chemists have used the periodic table since its development in the late 1860s, but it has evolved over the years, and it is still evolving today.

The periodic table has evolved since it was developed in the 1800s, but it still remains an important fixture in chemistry classrooms today.

The Adaptable Periodic Table

The design of the periodic table that we use today was developed in the 19th century, before the discovery of all of the naturally occurring elements, including the noble gases, and before the synthesis of elements. Initially, the elements were organized on the periodic table by atomic mass. This caused some inconsistencies. Once the atomic number was used to align the elements into rows and properties were used to organize the elements into columns, the modern periodic table was born.

The modern periodic table demonstrates the elegance of the nature of science. The early periodic table evolved to incorporate new information. For example, a new column was added when the noble gases were discovered. Period 7 is now completely filled after the recently discovered elements 113, 115, 117, and 118 were added.

Newly synthesized elements must now be added to another row. If and when these new elements are discovered, period 8 will be added to the periodic table.

When the periodic table was being developed chemists did not understand why these groups of elements had similar properties, but they recognized the periodic trend in these properties. Today, students and chemists understand that elements in a group have similar properties and the same number of valence electrons. Our understanding of the atom evolved along with our understanding of what elements in a group on the periodic table have in common in terms of valence electrons, reactivity, and properties.

While the nature of atoms and elements were not understood at the time of the periodic table's development, the original table was well designed. As new information was discovered about atoms and elements, the periodic table evolved and incorporated the new information resulting in a table that is just as useful today as it was when it was first developed.

APPLY SCIENTIFIC PRINCIPLES AND EVIDENCE

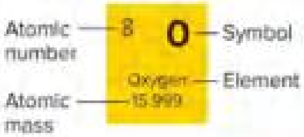
Explain how a discovery about an element or elements was incorporated into the modern periodic table.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 DEVELOPMENT OF THE MODERN PERIODIC TABLE

- The elements were first organized by increasing atomic mass, which led to inconsistencies. Later, they were organized by increasing atomic number.
- The periodic law states that when the elements are arranged by increasing atomic number, there is a periodic repetition of their chemical and physical properties.
- The periodic table organizes the elements into periods (rows) and groups or families (columns); elements with similar properties are in the same group.
- Elements are classified as metals, nonmetals, or metalloids.



- periodic law
- group
- period
- representative element
- transition element
- metal
- alkali metal
- alkaline earth metal
- transition metal
- inner transition metal
- lanthanide series
- actinide series
- nonmetal
- halogen
- noble gas
- metalloid

Lesson 2 CLASSIFICATION OF THE ELEMENTS

- The periodic table has four blocks (s, p, d, f).
- Elements within a group have similar chemical properties.
- The group number for elements in groups 1 and 2 equals the element's number of valence electrons.
- The energy level of an atom's valence electrons equals its period number.

Lesson 3 PERIODIC TRENDS

- Atomic and ionic radii decrease from left to right across a period, and increase as you move down a group.
- Ionization energies generally increase from left to right across a period, and decrease down a group.
- The octet rule states that atoms gain, lose, or share electrons to acquire a full set of eight valence electrons.
- Electronegativity generally increases from left to right across a period, and decreases down a group.

- ion
- ionization energy
- octet rule
- electronegativity

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

What can we learn from the periodic table?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will apply your evidence from this module and complete your project.

GO FURTHER

SEP Data Analysis Lab

Can you predict the properties of an element?

Francium was discovered in 1939, but its existence was predicted by Mendeleev in the 1870s. It is the least stable of the first 101 elements, with a half-life of just 22 minutes for its most stable isotope. Use the properties of other alkali metals, shown in the table, to predict some of francium's properties.

CER Analyze and Interpret Data

Use the given information about the known properties of the alkali metals to devise a method for predicting the corresponding properties of francium.

1. **Claim, Evidence, Reasoning** Devise an approach that clearly displays the trends for each of the properties given in the table and allows you to extrapolate a value for francium. Use the periodic law as a guide.
2. **Claim, Evidence, Reasoning** Predict whether francium is a solid, a liquid, or a gas. How can you support your prediction?
3. **Infer** which column of data presents the greatest possible error in making a prediction. Explain.
4. **Determine** why producing 1 million francium atoms per second is not enough to make measurements, such as density or melting point.

Alkali Metals Data

Element	Melting Point (°C)	Boiling Point (°C)	Radius (pm)
Lithium	180.5	1342	152
Sodium	97.8	883	186
Potassium	63.4	759	227
Rubidium	39.3	688	248
Cesium	28.4	671	265
Francium	?	?	?

IONIC COMPOUNDS AND METALS

ENCOUNTER THE PHENOMENON

Why do some crystals form cubes?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

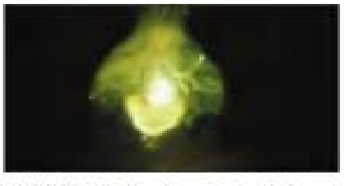
CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about why some crystals form cubes.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.



LESSON 2: Explore & Explain: Ionic Bond Formation

LESSON 2: Explore & Explain: Properties of Ionic Compounds

LESSON 1

ION FORMATION

FOCUS QUESTION

Why do elements form compounds?

Valence Electrons and Chemical Bonds

Imagine going on a scuba dive, diving below the ocean's surface and observing the awe-inspiring world below. You might explore the colorful and exotic organisms teeming around a coral reef, such as the one shown in **Figure 1.** The reef is formed from a compound called calcium carbonate, which is just one of thousands of compounds found on Earth. How do so many compounds form from the relatively few elements known to exist? The answer to this question involves the electron structure of atoms and the nature of the forces between atoms.

In previous chapters, you learned that elements within a group on the periodic table have similar properties. Many of these properties depend on the number of valence electrons the atom has. These valence electrons are involved in the formation of chemical bonds between two atoms. A **chemical bond** is the force that holds two atoms together.

Figure 1 As carbon dioxide dissolves in ocean water, carbonate ions are produced. Coral polyps capture these carbonate ions, producing crystals of calcium carbonate, which they secrete as an exoskeleton. Over time, the coral reef forms. A coral reef is a complex habitat that supports coral, algae, mollusks, echinoderms, and a variety of fishes.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Review the News

Obtain information from a current news story about ion formation. Evaluate your source and communicate your findings to your class.

Chemical bonds can form by the attraction between the positive nucleus of one atom and the negative electrons of another atom, or by the attraction between positive ions and negative ions. This chapter discusses chemical bonds formed by ions, atoms that have acquired a positive or negative charge. You will learn about bonds that form from the sharing of electrons in a later chapter.

Valence electrons

Recall that an electron-dot structure is a type of diagram used to keep track of valence electrons. Electron-dot structures are especially helpful when used to illustrate the formation of chemical bonds. **Table 1** shows several examples of electron-dot structures. For example, carbon, with an electron configuration of $1s^22s^22p^2$, has four valence electrons in the second energy level. These valence electrons are represented by the four dots around the symbol C in the table.

Table 1 Electron-Dot Structures

Group	1	2	13	14	15	16	17	18
Diagram	Li•	•Be•	•Ḃ•	•Ċ•	•N̈•	•Ö:	:F̈:	:N̈e:

Also, recall that ionization energy refers to how easily an atom loses an electron and that electron affinity indicates how much attraction an atom has for electrons. Noble gases, which have high ionization energies and low electron affinities, show a general lack of chemical reactivity. Other elements on the periodic table react with each other, forming numerous compounds. The difference in reactivity is directly related to the valence electrons.

The difference in reactivity involves the octet—the stable arrangement of eight valence electrons in the outer energy level. Unreactive noble gases have electron configurations that have a full outermost energy level. This level is filled with two electrons for helium ($1s^2$) and eight electrons for the other noble gases (ns^2np^6). Elements tend to react to acquire the stable electron structure of a noble gas.

Positive Ion Formation

A positive ion forms when an atom loses one or more valence electrons in order to attain a noble gas configuration. A positively charged ion is called a **cation.** To understand the formation of a positive ion, compare the electron configurations of the noble gas neon (atomic number 10) and the alkali metal sodium (atomic number 11).

Neon atom (Ne) $1s^22s^22p^6$

Sodium atom (Na) $1s^22s^22p^63s^1$

Note that the sodium atom has one 3s valence electron; it differs from the noble gas neon by that single valence electron. When sodium loses this outer valence electron, the resulting electron configuration is identical to that of neon.

Figure 2 shows how a sodium atom loses its valence electron to become a sodium cation. By losing an electron, the sodium atom acquires the stable outer electron configuration of neon. It is important to understand that although sodium now has the electron configuration of neon, it is not neon. It is a sodium ion with a single positive charge. The 11 protons that establish the character of sodium still remain within its nucleus.

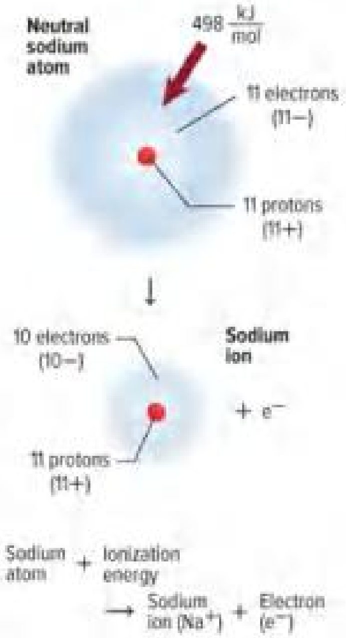


Figure 2 In the formation of a positive ion, a neutral atom loses one or more valence electrons. The atom is neutral because it contains equal numbers of protons and electrons; the ion, however, contains more protons than electrons and has a positive charge.

Analyze *Does the removal of an electron from a neutral atom require energy or release energy?*

Get It?

Identify the number of electrons in the outermost energy level that are associated with maximum stability.

Metal ions

Metal atoms are reactive because they lose valence electrons easily. The group 1 and 2 metals are the most reactive metals on the periodic table. For example, potassium and magnesium, group 1 and 2 elements, respectively, form K^+ and Mg^{2+} ions. Some group 13 atoms also form ions. The ions formed by metal atoms in groups 1, 2, and 13 are summarized in **Table 2**.

Transition metal ions

Recall that, in general, transition metals have an outer energy level of ns^2. Going from left to right across a period, atoms of each element fill an inner d sublevel. When forming positive ions, transition metals commonly lose their two valence electrons, forming 2+ ions. However, it is also possible for d electrons to be lost. Thus, transition metals also commonly form ions of 3+ or greater, depending on the number of d electrons in the electron structure. It is difficult to predict the number of electrons that will be lost. For example, iron (Fe) forms both Fe^{2+} and Fe^{3+} ions. A useful rule of thumb for these metals is that they form ions with a 2+ or a 3+ charge.

Get It?

Explain in your own words why transition metals can form ions with 2+ or 3+ charges.

Table 2 Group 1, 2, and 13 Ions

Group	Configuration	Charge of Ion Formed
1	[noble gas]ns^1	1+ when the s^1 electron is lost
2	[noble gas]ns^2	2+ when the s^2 electrons are lost
13	[noble gas]ns^2np^1	3+ when the s^2p^1 electrons are lost

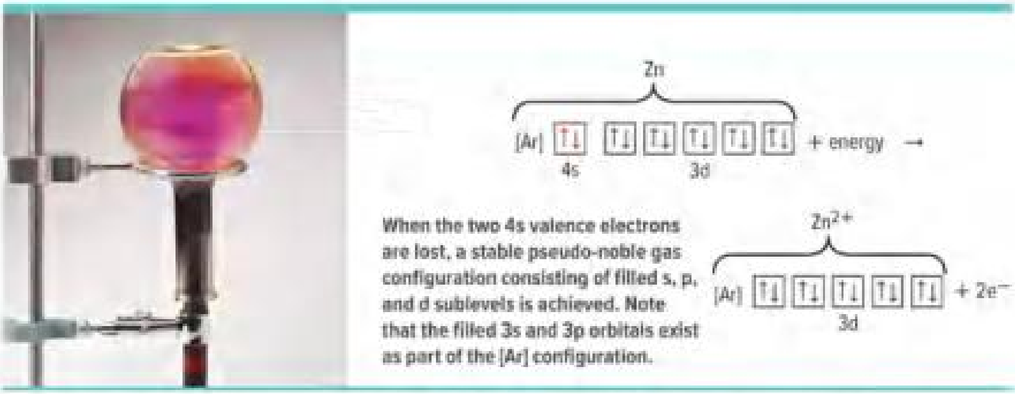


Figure 3 When zinc reacts with iodine, the heat of the reaction causes solid iodine to sublimate into a purple vapor. At the bottom of the tube, ZnI_2 is formed containing Zn^{2+} ions with a pseudo-noble gas configuration.

Pseudo-noble gas configurations

Although the formation of an octet is the most stable electron configuration, other electron configurations can also provide some stability. For example, elements in groups 11–14 lose electrons to form an outer energy level containing full s, p, and d sublevels. These relatively stable electron arrangements are referred to as pseudo-noble gas configurations. The zinc atom has the electron configuration of $1s^22s^22p^63s^23p^64s^23d^{10}$, in **Figure 3.** When forming an ion, the zinc atom loses the two 4s electrons in the outer energy level, and the stable configuration of $1s^22s^22p^63s^23p^63d^{10}$ results in a pseudo-noble gas configuration.

Negative Ion Formation

Nonmetals, which are located on the right side of the periodic table, easily gain electrons to attain a stable outer electron configuration. Examine **Figure 4.** To attain a noble-gas configuration, chlorine gains one electron, forming an ion with a 1− charge. After gaining the electron, the chloride ion has the electron configuration of an argon atom.

Chlorine atom (Cl)	$1s^22s^22p^63s^23p^5$
Argon atom (Ar)	$1s^22s^22p^63s^23p^6$
Chloride ion (Cl^-)	$1s^22s^22p^63s^23p^6$

An **anion** is a negatively charged ion. To designate an anion, the ending *-ide* is added to the root name of the element. Thus, a chlorine atom becomes a chloride anion.

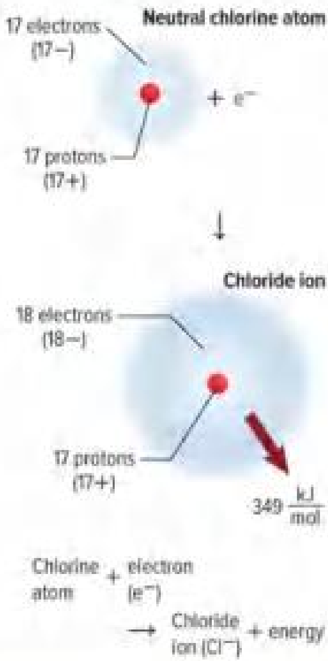


Figure 4 During the formation of the negative chloride ion, a neutral atom gains one electron. The process releases 349 kJ/mol of energy.

Compare *How do the energy changes accompanying positive ion and negative ion formation compare?*

Table 3 Group 15–17 Ions

Group	Configuration	Charge of Ion Formed
15	[noble gas]ns^2np^3	3– when three electrons are gained
16	[noble gas]ns^2np^4	2– when two electrons are gained
17	[noble gas]ns^2np^5	1– when one electron is gained

Nonmetal ions

As shown in **Table 3**, nonmetals gain the number of electrons that, when added to their valence electrons, equals 8. For example, consider phosphorus, with five valence electrons. To form a stable octet, the atom gains three electrons and forms a phosphide ion with a 3– charge. Likewise, oxygen, with six valence electrons, gains two electrons and forms an oxide ion with a 2– charge.

Some nonmetals can lose or gain other numbers of electrons to form an octet. For example, in addition to gaining three electrons, phosphorus can lose five. However, in general, group 15 elements gain three electrons, group 16 elements gain two, and group 17 elements gain one to achieve an octet.

Check Your Progress

Summary

- A chemical bond is the force that holds two atoms together.
- Some atoms form ions to gain stability. A stable configuration involves a complete outer energy level, usually consisting of eight valence electrons.
- Ions are formed by the loss or gain of valence electrons.
- The number of protons remains unchanged during ion formation.

Demonstrate Understanding

1. **Relate** the properties of atoms, their position in the periodic table, and their number of valence electrons to their chemical reactivity.
2. **Describe** two different causes of the force of attraction in a chemical bond.
3. **Apply** Why are all of the elements in group 18 relatively unreactive, whereas those in group 17 are very reactive?
4. **Summarize** ionic bond formation by correctly pairing these terms: cation, anion, electron gain, and electron loss.
5. **Apply** Write out the electron configuration for each atom. Then, predict the change that must occur in each to achieve a noble-gas configuration.

   a. nitrogen b. sulfur c. barium d. lithium
6. **Model** Draw models to represent the formation of the positive calcium ion and the negative bromide ion.

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

LESSON 2

IONIC BONDS AND IONIC COMPOUNDS

FOCUS QUESTION

How are the ions in ionic compounds arranged?

Formation of an Ionic Bond

What do the reactions shown in **Figure 5** have in common? In both cases, elements react with each other to form a compound. **Figure 5a** shows the reaction between the elements sodium and chlorine. During this reaction, a sodium atom transfers its valence electron to a chlorine atom and becomes a positive ion. The chlorine atom accepts the electron into its outer energy level and becomes a negative ion. The oppositely charged ions attract each other, forming the compound sodium chloride. The electrostatic force that holds oppositely charged particles together in an ionic compound is referred to as an **ionic bond.** Compounds that contain ionic bonds are **ionic compounds.** If ionic bonds occur between metals and the nonmetal oxygen, oxides form. Most other ionic compounds are called salts.

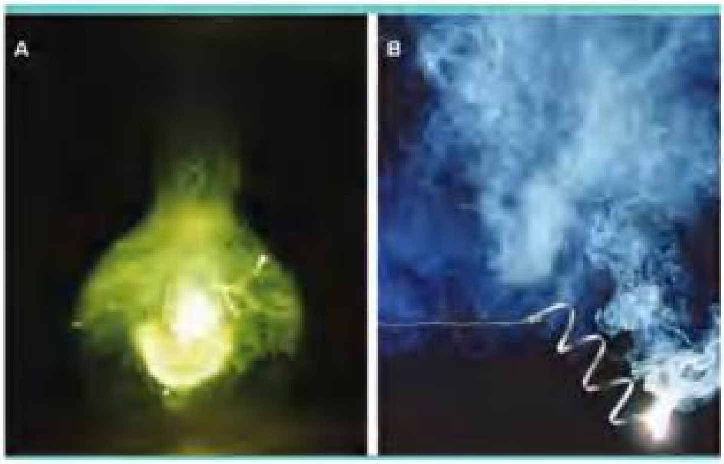


Figure 5 Each of these chemical reactions produces an ionic compound while releasing a large amount of energy. **a.** The reaction between elemental sodium and chlorine gas produces a white crystalline solid. **b.** When a ribbon of magnesium metal burns in air, it forms the ionic compound magnesium oxide.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

ChemLAB: Synthesize an Ionic Compound
Carry out an investigation to determine if a compound has ionic bonds based on physical properties.

Probeware Lab: Conductivity
Obtain, evaluate, and communicate information to discover the function of an electric current to determine the properties of a dissolved solid.

Binary ionic compounds

Thousands of compounds contain ionic bonds. Many ionic compounds are binary, which means that they contain only two different elements. Binary ionic compounds contain a metallic cation and a nonmetallic anion. Sodium chloride (NaCl) is a binary compound because it contains two different elements, sodium and chlorine. Magnesium oxide (MgO), the reaction product shown in **Figure 5b** on the previous page, is also a binary ionic compound.

Table 4 summarizes several ways in which the formation of an ionic compound such as sodium chloride can be represented.

Table 4 **Formation of Sodium Chloride**

Chemical Equation

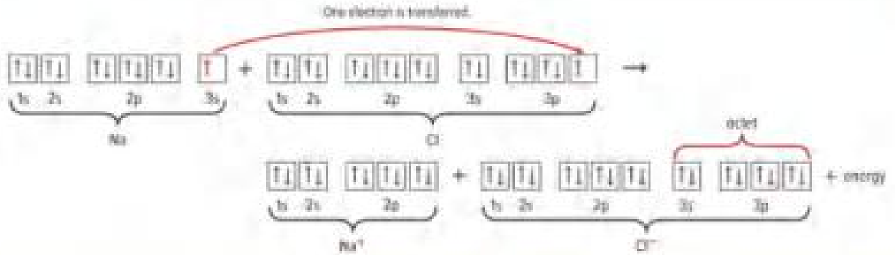
$$Na + Cl \rightarrow Na^{+} + Cl^{-} + \text{energy}$$

Electron Configurations

One electron is transferred.

$$\underbrace{[Ne]3s^1}_{Na} + \underbrace{[Ne]3s^2 3p^5}_{Cl} \rightarrow \underbrace{[Ne]}_{Na^+} + \underbrace{[Ar]}_{Cl^-} + \text{energy}$$

Orbital Notation

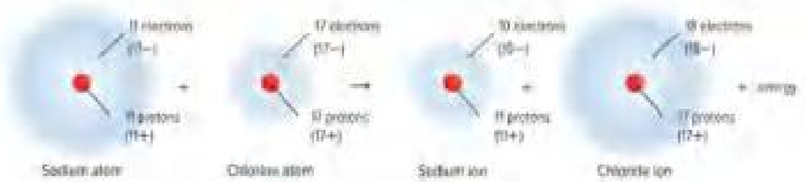


Electron-Dot Structures

One electron is transferred.

$$Na\cdot + \cdot\ddot{\underset{..}{Cl}}: \rightarrow [Na]^{+} + [:\ddot{\underset{..}{Cl}}:]^{-} + \text{energy}$$

Atomic Models



What role does ionic charge play in the formation of ionic compounds? To answer this question, consider how calcium fluoride forms. Calcium has the electron configuration $[Ar]4s^2$, and needs to lose two electrons to attain the stable configuration of argon. Fluorine has the configuration $[He]2s^22p^5$, and must gain one electron to attain the stable configuration of neon. Because the number of electrons lost and gained must be equal, two fluorine atoms are needed to accept the two electrons lost from the calcium atom. As you can see, the overall charge of one unit of calcium fluoride (CaF_2) is zero.

$$1\ \cancel{\text{Ca ion}}\left(\frac{2+}{\cancel{\text{Ca ion}}}\right)+2\ \cancel{\text{F ions}}\left(\frac{1-}{\cancel{\text{F ion}}}\right)= (1)\ (2+) + (2)(1-) = 0$$

Next, consider aluminum oxide, the whitish coating that forms on aluminum chairs. To acquire a noble-gas configuration, each aluminum atom loses three electrons and each oxygen atom gains two electrons. Thus, three oxygen atoms are needed to accept the six electrons lost by two aluminum atoms. The neutral compound formed is aluminum oxide (Al_2O_3).

$$2\ \cancel{\text{Al ions}}\left(\frac{3+}{\cancel{\text{Al ion}}}\right)+ 3\ \cancel{\text{O ions}}\left(\frac{2-}{\cancel{\text{O ion}}}\right) = 2(3+) + 3(2-) = 0$$

Real-World Chemistry

Ionic Compounds

MINERAL SUPPLEMENTS To function properly, your body requires a daily intake of many different minerals. To ensure they are getting what they need, many people take a daily multivitamin and a mineral supplement. The minerals in these supplements come from a variety of ionic compounds. In fact, the majority of minerals found in mineral supplements come from ground-up rocks.

PRACTICE Problems ADDITIONAL PRACTICE

Explain how an ionic compound forms from these elements.

7. sodium and nitrogen
8. lithium and oxygen
9. strontium and fluorine
10. aluminum and sulfur
11. **CHALLENGE** Explain how elements in the two groups shown on the periodic table at the right combine to form an ionic compound.

Properties of Ionic Compounds

The chemical bonds in a compound determine many of its properties and applications. For ionic compounds, electrical forces in the ionic bonds produce unique physical structures, unlike those of other compounds. The physical structures of ionic compounds also contribute to their bulk physical properties.

Physical structure

In an ionic compound, large numbers of positive ions and negative ions exist together in a ratio determined by the number of electrons transferred from the metal atom to the nonmetal atom. These ions are packed into a regular repeating pattern that balances the forces of attraction and repulsion between the ions.

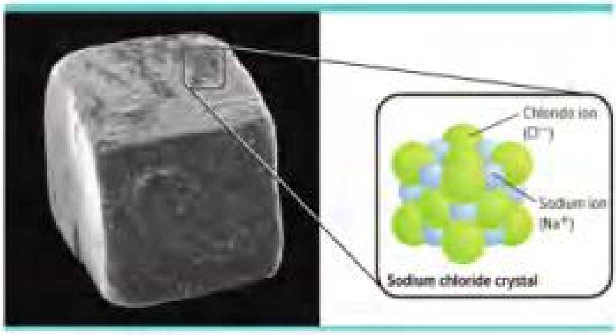


Figure 6 When viewed with a scanning electron microscope, the cubic shape of a sodium chloride crystal is visible. The structure of the crystal is highly ordered.

Compare *the shape of the crystal at the atomic scale to the shape of the crystal at the bulk scale.*

Examine the pattern of the ions in the sodium chloride crystal shown in **Figure 6**. Note the highly organized nature of an ionic crystal—the consistent spacing of the ions and the uniform pattern formed by them. Although the ion sizes are not the same, each sodium ion in the crystal is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. What shape would you expect a large crystal of this compound to be? As shown in **Figure 6**, the one-to-one ratio of sodium and chloride ions produces a highly ordered cubic crystal. As in all ionic compounds, in NaCl, no single unit consisting of only one sodium ion and one chloride ion is formed. Instead, large numbers of sodium ions and chloride ions exist together. If you can, obtain a magnifying lens and use it to examine some crystals of table salt (NaCl). What is the shape of these small salt crystals?

Get It?

Explain what determines the ratio of positive ions to negative ions in an ionic mineral crystal.

The strong attractions among the positive ions and the negative ions in an ionic compound result in the formation of a **crystal lattice.** A crystal lattice is a three-dimensional geometric arrangement of particles. In a crystal lattice, each positive ion is surrounded by negative ions, and each negative ion is surrounded by positive ions. Ionic crystals vary in shape due to the sizes and relative numbers of the ions bonded, as shown by the minerals in **Figure 7**.

Figure 7 Aragonite ($CaCO_3$), barite ($BaSO_4$), and beryl ($Be_3Al_2Si_6O_{18}$) are examples of minerals that are ionic compounds. The ions that form them are bonded together in a crystal lattice. Differences in ion size and charge result in different ionic crystal shapes, a topic that will be discussed later.

EARTH SCIENCE Connection The minerals in **Figure 7**, on the previous page, are just a few of the types studied by mineralogists, scientists who study minerals. They make use of several classification schemes to organize the thousands of known minerals. Color, crystal structure, hardness, chemical, magnetic, and electric properties, and numerous other characteristics are used to classify minerals. The types of anions minerals contain can also be used to identify them. For example, more than one-third of all known minerals are silicates, which are minerals that contain an anion that is a combination of silicon and oxygen. Halides contain fluoride, chloride, bromide, or iodide ions. Other mineral classes include boron-containing anions known as borates and carbon-oxygen containing anions known as carbonates.

Get It?

Identify In **Figure 7**, which mineral is a silicate, and which mineral is a carbonate? Can you classify any of the minerals in the video about crystal formation in caves?

Physical properties

Melting point, boiling point, and hardness are bulk physical properties of matter that are determined by the strength of electrical forces between particles that make up the matter. Because ionic bonds are relatively strong, ionic crystals require a large amount of energy to be broken apart. Thus, ionic crystals have high melting points and high boiling points, as shown in **Table 5**. Many crystals, including gemstones, have brilliant colors. These colors are due to the presence of transition metals in the crystal lattices.

Table 5 Melting and Boiling Points of Some Ionic Compounds

Compound	Melting Point (°C)	Boiling Point (°C)
NaI	660	1304
KBr	734	1435
NaBr	747	1390
$CaCl_2$	782	>1600
NaCl	801	1413
MgO	2852	3600

Ionic crystals are also hard, rigid, brittle solids due to the strong attraction between electric charges that holds the ions in place. When an external force is applied to the crystal—a force strong enough to overcome the attractive forces holding the ions in position within the crystal—the crystal cracks or breaks apart, as shown in **Figure 8**. The crystal breaks apart because the applied force repositions the like-charged ions next to each other; the resulting repulsion between electric forces breaks apart the crystal.

Undisturbed ionic crystal
Before the force is applied, the crystal has a uniform pattern of ions.

Applied force realigns particles.
If the applied force is strong enough, it pushes the ions out of alignment.

Forces of repulsion break crystal apart.
A repulsive force created by nearby like-charged ions breaks apart the crystal.

Figure 8 Strong attractive forces hold the ions in place until a force strong enough to overcome the attraction is applied.

Another property—the ability of a material to conduct electricity—depends on the availability of freely moving charged particles. Ions are charged particles, so whether they are free to move determines whether an ionic compound conducts electricity. In the solid state, the ions in an ionic compound are locked into fixed positions by strong attractive forces. As a result, ionic solids do not conduct electricity.

The situation changes dramatically, however, when an ionic solid melts to become a liquid or is dissolved in solution. The ions—previously locked in position—are now free to move and conduct an electric current. Both ionic compounds in solution and in the liquid state are excellent conductors of electricity. An ionic compound whose aqueous solution conducts an electric current is called an **electrolyte.**

Energy and the Ionic Bond

During every chemical reaction, energy is either absorbed or released. If energy is absorbed during a chemical reaction, the reaction is endothermic. If energy is released, it is exothermic. The formation of ionic compounds from positive ions and negative ions is always exothermic. The attraction of the positive ion for the negative ions close to it forms a more stable system that is lower in energy than the individual ions. If the amount of energy released during bond formation is reabsorbed, the bonds holding the positive ions and negative ions together will break apart.

Lattice energy

Because the ions in an ionic compound are arranged in a crystal lattice, the energy required to separate 1 mol of the ions of an ionic compound is referred to as the **lattice energy.** The strength of the electrical forces holding ions in place is reflected by the lattice energy. The greater the lattice energy, the stronger the force of attraction.

Lattice energy is directly related to the size of the ions bonded. Smaller ions form compounds with more closely spaced ionic charges. Because the electrostatic force of attraction between opposite charges increases as the distance between the charges decreases, smaller ions produce stronger attractions and greater lattice energies. For example, the lattice energy of a lithium compound is greater than that of a potassium compound with the same anion because a lithium ion is smaller than a potassium ion.

Get It?

Explain the relationship between lattice energy and the size of the ions in an ionic compound.

SCIENCE USAGE v. COMMON USAGE

conduct

Science usage: to transmit light, heat, sound, or electricity

The material did not conduct electricity well.

Common usage: to guide or lead

It was the manager's job to conduct the training session.

The value of lattice energy is also affected by the charge of the ion. The ionic bond formed from the attraction of ions with larger positive or negative charges generally has a greater lattice energy. The lattice energy of MgO is almost four times greater than that of NaF because the charge of the ions in MgO is greater than the charge of the ions in NaF. The lattice energy of $SrCl_2$ is between the lattice energies of MgO and NaF because $SrCl_2$ contains ions with both higher and lower charges.

Table 6 Lattice Energies of Some Ionic Compounds

Compound	Lattice Energy (kJ/mol)	Compound	Lattice Energy (kJ/mol)
KI	632	KF	808
KBr	671	AgCl	910
RbF	774	NaF	910
NaI	682	LiF	1030
NaBr	732	$SrCl_2$	2142
NaCl	769	MgO	3795

Table 6 shows the lattice energies of some ionic compounds. Examine the lattice energies of RbF and KF. Because K+ has a smaller ionic radius than Rb+, KF has a greater lattice energy than RbF. This confirms that lattice energy is related to ion size. Notice the lattice energies of $SrCl_2$ and AgCl. How do they show the relationship between lattice energy and the charge of the ions involved?

Check Your Progress

Summary

- Ionic compounds contain ionic bonds formed by the attraction of oppositely charged ions.
- Ions in an ionic compound are arranged in a repeating pattern known as a crystal lattice.
- Ionic compound properties are related to ionic bond strength.
- Ionic compounds conduct an electric current in the liquid phase and in aqueous solution.
- Lattice energy is the energy needed to remove 1 mol of ions from its lattice.

Demonstrate Understanding

12. **Explain** how an ionic compound made up of charged particles can be electrically neutral.
13. **Describe** the endothermic and exothermic energy changes associated with ionic bond formation, and relate these to stability.
14. **Identify** three physical properties of ionic compounds that are associated with ionic bonds, and relate them to bond strength.
15. **Explain** how ions form bonds, and describe the structure of the resulting compound.
16. **Relate** lattice energy to ionic-bond strength.
17. **Apply** Use electron configurations, orbital notation, and electron-dot structures to represent the formation of an ionic compound from the metal strontium and the nonmetal chlorine.
18. **Design** a concept map that shows the relationships among ionic bond strength, physical properties of ionic compounds, lattice energy, and stability.

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LESSON 3

NAMES AND FORMULAS FOR IONIC COMPOUNDS

FOCUS QUESTION

What are the names and formulas of ionic compounds?

Formulas for Ionic Compounds

Because chemists around the world need to be able to communicate with one another, they have developed a set of rules for naming compounds. Using this standardized naming system, you can write a chemical formula from a compound's name and name a compound given its chemical formula.

Recall that an ionic compound is made up of ions arranged in a repeating pattern. The chemical formula for an ionic compound, called a **formula unit**, represents the simplest ratio of the ions involved. For example, the formula unit of magnesium chloride is $MgCl_2$ because the magnesium and chloride ions exist in a 1:2 ratio.

The overall charge of a formula unit is zero because the formula unit represents the entire crystal, which is electrically neutral. The formula unit for $MgCl_2$ contains one Mg^{2+} ion and two Cl^- ions, for a total charge of zero.

Monatomic ions

Binary ionic compounds are composed of positively charged monatomic ions of a metal and negatively charged monatomic ions of a nonmetal. A **monatomic ion** is a one-atom ion, such as Mg^{2+} or Br^-. **Table 7** indicates the charges of common monatomic ions according to their location on the periodic table. What is the formula for the beryllium ion? The iodide ion? The nitride ion?

Table 7 Common Monatomic Ions

Group	Atoms that Commonly Form Ions	Charge of Ions
1	H, Li, Na, K, Rb, Cs	1+
2	Be, Mg, Ca, Sr, Ba	2+
15	N, P, As	3−
16	O, S, Se, Te	2−
17	F, Cl, Br, I	1−

Transition metals, which are in groups 3 through 12, and metals in groups 13 and 14 are not included in **Table 7** because of the variance in ionic charges of atoms in the groups. Most transition metals and metals in groups 13 and 14 can form several different positive ions.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Review the News

Obtain information from a current news story about recent additions to the periodic table. Evaluate your source and communicate your findings to your class.

Oxidation numbers

The oxidation number of a monatomic ion, also called its oxidation state, is equal to the net charge of the ion. As shown in **Table 8**, most transition metals and group 13 and 14 metals have more than one possible ionic charge. Note that the ionic charges given in the table are the most common ones, but not the only ones possible.

The oxidation number of an element in an ionic compound equals the number of electrons transferred from the atom to form the ion. For example, a sodium atom transfers one electron to a chlorine atom to form sodium chloride. This results in Na^+ and Cl^-. Thus, the oxidation number of sodium in the compound is 1+ because one electron was transferred from the sodium atom. Because an electron is transferred to the chlorine atom, its oxidation number is 1−.

Table 8 Monatomic Metal Ions

Group	Common Ions
3	Sc^{3+}, Y^{3+}, La^{3+}
4	Ti^{2+}, Ti^{3+}
5	V^{2+}, V^{3+}
6	Cr^{2+}, Cr^{3+}
7	Mn^{2+}, Mn^{3+}, Tc^{2+}
8	Fe^{2+}, Fe^{3+}
9	Co^{2+}, Co^{3+}
10	Ni^{2+}, Pd^{2+}, Pt^{2+}, Pt^{4+}
11	Cu^{+}, Cu^{2+}, Ag^{+}, Au^{+}, Au^{3+}
12	Zn^{2+}, Cd^{2+}, Hg_2^{2+}, Hg^{2+}
13	Al^{3+}, Ga^{2+}, Ga^{3+}, In^{+}, In^{2+}, In^{3+}, Tl^{+}, Tl^{3+}
14	Sn^{2+}, Sn^{4+}, Pb^{2+}, Pb^{4+}

Formulas for binary ionic compounds

In the chemical formula for any ionic compound, the symbol of the cation is always written first, followed by the symbol of the anion. Subscripts, which are small numbers to the lower right of a symbol, represent the number of ions of each element in an ionic compound. If no subscript is written, it is assumed to be one. You can use oxidation numbers to write formulas for ionic compounds. Recall that ionic compounds have no charge. If you add the oxidation number of each ion multiplied by the number of these ions in a formula unit, the total must be zero.

Suppose you need to determine the formula for one formula unit of the compound that contains sodium and fluoride ions. Start by writing the symbol and charge for each ion: Na^+ and F^-. The ratio of ions in a formula unit of the compound must show that the number of electrons lost by the metal equals the number of electrons gained by the nonmetal. This occurs when one sodium atom transfers one electron to the fluorine atom; the formula unit is NaF.

Get It?

Relate the charge of an ion to its oxidation number.

STEM CAREER Connection

Pharmacist

Would you like using chemistry to help people treat disease? Pharmacists prepare and dispense medications to patients and provide expertise on the safe and proper use of medications. They may work in retail pharmacies or in hospitals. Pharmacists also conduct health screenings, give immunizations such as flu shots, and advise patients on general health topics such as diet, exercise, and managing stress.

ACADEMIC VOCABULARY

transfer

to cause to pass from one to another

Carlos had to transfer to a new school when his parents moved to a new neighborhood.

EXAMPLE Problem 1

FORMULA FOR AN IONIC COMPOUND Determine the formula for the ionic compound formed from potassium and oxygen.

1 ANALYZE THE PROBLEM

You are given that potassium and oxygen ions form an ionic compound, the formula for the compound is the unknown. First, write out the symbol and oxidation number for each ion involved in the compound. Potassium, from group 1, forms 1+ ions, and oxygen, from group 16, forms 2− ions.

K^{+} O^{2-}

Because the charges are not the same, you need to determine the subscripts to use to indicate the ratio of positive ions to negative ions.

2 SOLVE FOR THE UNKNOWN

A potassium atom loses one electron, while an oxygen atom gains two electrons. If combined in a one-to-one ratio, the number of electrons lost by potassium will not balance the number of electrons gained by oxygen. Thus, two potassium ions are needed for each oxide ion. The formula is **K_2O**.

3 EVALUATE THE ANSWER

The overall charge of the compound is zero.

$$2\ \cancel{\text{K ions}}\left(\frac{1+}{\cancel{\text{K ion}}}\right) + 1\ \cancel{\text{O ion}}\left(\frac{2-}{\cancel{\text{O ion}}}\right) = 2(1+) + 1(2-) = 0$$

EXAMPLE Problem 2

FORMULA FOR AN IONIC COMPOUND Determine the formula for the compound formed from aluminum ions and sulfide ions.

1 ANALYZE THE PROBLEM

You are given that aluminum and sulfur form an ionic compound; the formula for the ionic compound is the unknown. First, determine the charges of each ion. Aluminum, from group 13, forms 3+ ions, and sulfur, from group 16, forms 2− ions.

Al^{3+} S^{2-}

Each aluminum atom loses three electrons, while each sulfur atom gains two electrons. The number of electrons lost must equal the number of electrons gained.

2 SOLVE FOR THE UNKNOWN

The smallest number that can be divided evenly by both 2 and 3 is 6. Therefore, six electrons are transferred. Three sulfur atoms accept the six electrons lost by two aluminum atoms. The correct formula, **Al_2S_3**, shows two aluminum ions bonded to three sulfide ions.

3 EVALUATE THE ANSWER

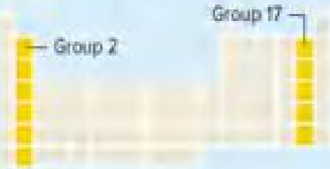
The overall charge of one formula unit of this compound is zero.

$$2\ \cancel{\text{Al ions}}\left(\frac{3+}{\cancel{\text{Al ion}}}\right) + 3\ \cancel{\text{S ions}}\left(\frac{2-}{\cancel{\text{S ion}}}\right) = 2(3+) + 3(2-) = 0$$

PRACTICE Problems ADDITIONAL PRACTICE

Write formulas for the ionic compounds formed by the following ions.
Use units as a guide to your solutions.

19. potassium and iodide
20. magnesium and chloride
21. aluminum and bromide
22. cesium and nitride
23. CHALLENGE Write the general formula for the ionic compound formed by elements from the two groups shown on the periodic table at the right.



Formulas for polyatomic ionic compounds

Many ionic compounds contain **polyatomic ions,** which are ions made up of more than one atom. **Table 9** and **Figure 9** list some common polyatomic ions. Also see **Table R-5** in the Student Resources. A polyatomic ion acts as an individual ion in a compound and its charge applies to the entire group of atoms. Thus, the formula for a polyatomic compound follows the same rules used for a binary compound.

Because a polyatomic ion exists as a unit, never change subscripts of the atoms within the ion. If more than one polyatomic ion is needed, place parentheses around the ion and write the appropriate subscript outside the parentheses. For example, consider the compound formed from the ammonium ion (NH_4^+) and the oxide ion (O^{2-}). To balance the charges, the compound must have two ammonium ions for each oxide ion. To add a subscript to ammonium, enclose it in parentheses, then add the subscript. The correct formula is $(NH_4)_2O$.

Table 9 Common Polyatomic Ions

Ion	Name	Ion	Name
NH_4^+	ammonium	IO_4^-	periodate
NO_2^-	nitrite	$C_2H_3O_2^-$	acetate
NO_3^-	nitrate	$H_2PO_4^-$	dihydrogen phosphate
OH^-	hydroxide	CO_3^{2-}	carbonate
CN^-	cyanide	SO_3^{2-}	sulfite
MnO_4^-	permanganate	SO_4^{2-}	sulfate
HCO_3^-	hydrogen carbonate	$S_2O_3^{2-}$	thiosulfate
ClO^-	hypochlorite	O_2^{2-}	peroxide
ClO_2^-	chlorite	CrO_4^{2-}	chromate
ClO_3^-	chlorate	$Cr_2O_7^{2-}$	dichromate
ClO_4^-	perchlorate	HPO_4^{2-}	hydrogen phosphate
BrO_3^-	bromate	PO_4^{3-}	phosphate
IO_3^-	iodate	AsO_4^{3-}	arsenate

Figure 9 Ammonium and phosphate ions are polyatomic; that is, they are made up of more than one atom. Each polyatomic ion, however, acts as a single unit and has one charge.

Identify *What are the charges of the ammonium ion and phosphate ion, respectively?*

EXAMPLE Problem 3

FORMULA FOR A POLYATOMIC IONIC COMPOUND A compound formed by calcium ions and phosphate ions is often used in fertilizers. Write the compound's formula.

1 ANALYZE THE PROBLEM

You know that calcium and phosphate ions form an ionic compound; the formula for the compound is the unknown. First, write each ion along with its charge. Calcium, from group 2, forms 2+ ions, and the polyatomic phosphate acts as a single unit with a 3− charge.

$$Ca^{2+} \qquad PO_4^{3-}$$

Each calcium atom loses two electrons, while each polyatomic phosphate group gains three electrons. The number of electrons lost must equal the number of electrons gained.

2 SOLVE FOR THE UNKNOWN

The smallest number evenly divisible by both charges is 6. Thus, a total of six electrons are transferred. The negative charge from two phosphate ions equals the positive charge from three calcium ions. In the formula, place the polyatomic ion in parentheses and add a subscript to the outside. The correct formula for the compound is **$Ca_3(PO_4)_2$**.

3 EVALUATE THE ANSWER

The overall charge of one formula unit of calcium phosphate is zero.

$$3\ \cancel{\text{Ca ions}}\left(\frac{2+}{\cancel{\text{Ca ion}}}\right) + 2\ \cancel{PO_4\text{ ions}}\left(\frac{3-}{\cancel{PO_4\text{ ion}}}\right) = 3(2+) + 2(3-) = 0$$

PRACTICE Problems ADDITIONAL PRACTICE

Write formulas for ionic compounds composed of the following ions.
Use units as a guide to your solutions.

24. sodium and nitrate

25. calcium and chlorate

26. aluminum and carbonate

27. **CHALLENGE** Write the formula for an ionic compound formed by ions from a group 2 element and polyatomic ions composed of only carbon and oxygen.

Names for Ions and Ionic Compounds

Scientists use a systematic approach when naming ionic compounds. Because ionic compounds have both cations and anions, the naming system accounts for both of these ions.

Naming an oxyanion

An **oxyanion** is a polyatomic ion composed of an element, usually a nonmetal, bonded to one or more oxygen atoms. More than one oxyanion exists for some nonmetals, such as nitrogen and sulfur. These ions are easily named using the rules in **Table 10**.

Table 10 Oxyanion Naming Conventions for Sulfur and Nitrogen

• Identify the ion with the greatest number of oxygen atoms. This ion is named using the root of the nonmetal and the suffix *-ate*.				
• Identify the ion with fewer oxygen atoms. This ion is named using the root of the nonmetal and the suffix *-ite*.				
Examples:	NO_3^- nitr*ate*	NO_2^- nitr*ite*	SO_4^{2-} sulf*ate*	SO_3^{2-} sulf*ite*

As shown in **Table 11**, chlorine forms four oxyanions that are named according to the number of oxygen atoms present. Names of similar oxyanions formed by other halogens follow the rules used for chlorine. For example, bromine forms the bromate ion (BrO_3^-), and iodine forms the periodate ion (IO_4^-) and the iodate ion (IO_3^-).

Table 11 Oxyanion Naming Conventions for Chlorine

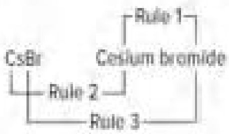
• The oxyanion with the greatest number of oxygen atoms is named using the prefix *per-*, the root of the nonmetal, and the suffix *-ate*.
• The oxyanion with one fewer oxygen atom is named using the root of the nonmetal and the suffix *-ate*.
• The oxyanion with two fewer oxygen atoms is named using the root of the nonmetal and the suffix *-ite*.
• The oxyanion with three fewer oxygen atoms is named using the prefix hypo-, the root of the nonmetal, and the suffix *-ite*.
Examples: ClO_4^- perchlorate ClO_3^- chlorate ClO_2^- chlorite ClO^- hypochlorite

Naming ionic compounds

Chemical nomenclature is a systematic way of naming compounds. Now that you are familiar with chemical formulas, you can use the following five rules to name ionic compounds.

1. Name the cation followed by the anion. Remember that the cation is always written first in the formula.
2. For monatomic cations, use the element name.
3. For monatomic anions, use the root of the element name plus the suffix *-ide*.

   Example:



4. To distinguish between multiple oxidation numbers of the same element, the name of the chemical formula must indicate the oxidation number of the cation. The oxidation number is written as a Roman numeral in parentheses after the name of the cation.

   Note: This rule applies to the transition metals and metals on the right side of the periodic table, which often have more than one oxidation number. See **Table 8** earlier in this lesson. It does not apply to group 1 and group 2 cations, as they have only one oxidation number.

   Examples: Fe^{2+} and O_2^- ions form FeO, known as iron(II) oxide.
   Fe^{3+} and O_2^- ions form Fe_2O_3, known as iron(III) oxide.

5. When the compound contains a polyatomic ion, simply use the name of the polyatomic ion in place of the anion or cation.

   Examples: The name for NaOH is sodium hydroxide.
   The name for $(NH_4)_2S$ is ammonium sulfide.

PRACTICE Problems ADDITIONAL PRACTICE

Interpret the formula representations and name these compounds.

28. NaBr

29. $CaCl_2$

30. KOH

31. $Cu(NO_3)_2$

32. Ag_2CrO_4

33. CHALLENGE The ionic compound NH_4ClO_4 is a key reactant used in solid rocket boosters, such as those that powered the Space Shuttle into orbit. Name this compound.

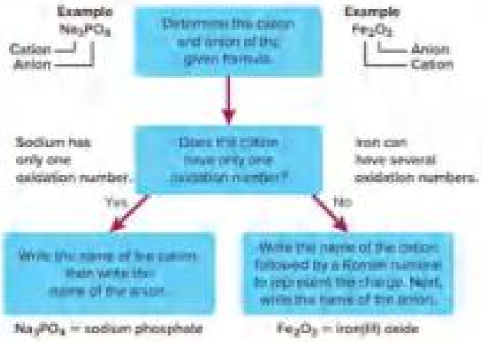
PROBLEM-SOLVING STRATEGY

Naming Ionic Compounds

Naming ionic compounds is easy if you follow this naming-convention flowchart.

Apply the Strategy

Name the compounds KOH and Ag_2CrO_4 using this flowchart.



Naming ionic compounds is important in communicating the cation and anion in a crystalline solid or aqueous solution. The Problem-Solving Strategy reviews the steps used in naming ionic compounds if the formula is known. How might you change the diagram to help you write the formulas for ionic compounds if you know their names?

Check Your Progress

Summary

- A formula unit gives the ratio of cations to anions in the ionic compound.
- A monatomic ion is formed from one atom. The charge of a monatomic ion is equal to its oxidation number.
- Roman numerals indicate the oxidation number of cations having multiple possible oxidation states.
- Polyatomic ions consist of more than one atom and act as a single unit.
- To indicate more than one polyatomic ion in a chemical formula, place parentheses around the polyatomic ion and use a subscript.

Demonstrate Understanding

34. **State** the order in which the ions associated with a compound composed of potassium and bromine would be written in the chemical formula and the compound name.
35. **Determine** the difference between a monatomic ion and a polyatomic ion, and give an example of each.
36. **Apply** Ion X has a charge of 2+, and ion Y has a charge of 1−. Write the formula unit of the compound formed from the ions.
37. **State** the name and formula for the compound formed from Mg and Cl.
38. **Write** the name and formula for the compound formed from sodium ions and nitrite ions.
39. **Analyze** What subscripts would you most likely use if the following substances formed an ionic compound?
    a. an alkali metal and a halogen
    b. an alkali metal and a nonmetal from group 16
    c. an alkaline earth metal and a halogen
    d. an alkaline earth metal and a nonmetal from group 16

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LESSON 4

METALLIC BONDS AND THE PROPERTIES OF METALS

FOCUS QUESTION

How do metals and ionic compounds compare and contrast?

Metallic Bonds

Although metals are not ionic, they share several properties with ionic compounds. The bonding in both metals and ionic compounds is based on the attraction of particles with unlike charges. In the solid state, metals often form lattices similar to ionic crystal lattices. In such a lattice, 8 to 12 other metal atoms closely surround each metal atom.

A sea of electrons

In a metallic lattice, metal atoms do not share their valence electrons with neighboring atoms, nor do they lose their valence electrons. Instead, the outer energy levels of the metal atoms overlap, as shown in **Figure 10**. This unique arrangement is described by the **electron sea model,** which proposes that all the metal atoms in a metallic solid contribute their valence electrons to form a "sea" of electrons that surrounds the metal cations in the lattice. The electrons present in the outer energy levels of the bonding metallic atoms are not held by any specific atom and can move easily from one atom to the next. Because they are free to move, they are often referred to as **delocalized electrons.** When the atom's outer electrons move freely throughout the solid, a metallic cation is formed. Each such ion is bonded to the lattice by the sea of valence electrons. A **metallic bond** is the attraction of a metallic cation for delocalized electrons.

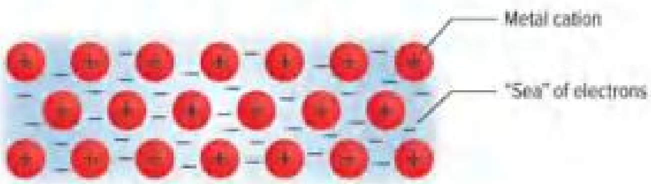

Figure 10 The valence electrons in metals (shown as a **blue cloud** of minus signs) are evenly distributed among the metallic cations (**shown in red**). Attractions between positive cations and the negative "sea" hold the metal atoms together in a lattice.

Explain *Why are electrons in metals known as delocalized electrons?*

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science and Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Modeling Electrostatic Forces—Ionic and Metallic Bonding
HS-PS2-4. Use mathematical representations of Newton's Law of Gravitation and Coulomb's Law to describe and predict the gravitational and electrostatic forces between objects.

Quick Investigation: Observe Properties
Construct an explanation of the structures and malleable properties of steel.

Properties of metals

The physical properties of metals at the bulk scale can be explained by metallic bonding. These properties provide evidence of the strength of metallic bonds.

Table 12 Melting and Boiling Points

Element	Melting Point (°C)	Boiling Point (°C)
Lithium	180	1342
Tin	232	2602
Aluminum	660	2519
Barium	727	1897
Silver	962	2162
Copper	1085	2562

Melting and boiling points The melting points of metals vary greatly. Mercury is a liquid at room temperature, which makes it useful in scientific instruments such as thermometers and barometers. On the other hand, tungsten has a melting point of 3422°C. Lightbulb filaments are usually made from tungsten, as are certain spacecraft parts. In general, metals have moderately high melting points and high boiling points, as shown in **Table 12**. The melting points are less extreme than the boiling points because the cations and electrons are mobile in a metal, so it does not take much energy for them to be able to move past each other. However, during boiling, atoms must be completely separated from the lattice, which requires much more energy.

Thermal conductivity and electrical conductivity The movement of mobile electrons around positive metallic cations makes metals good conductors. The delocalized electrons move heat from one place to another much more quickly than the electrons in a material that does not contain mobile electrons. Mobile electrons easily move as part of an electric current when an electric potential is applied to a metal. These same delocalized electrons interact with light, absorbing and releasing photons, thereby creating the property of luster in metals.

Malleability, ductility, and durability Metals are malleable, which means they can be hammered into sheets, and they are ductile, which means they can be drawn into wire. **Figure 11** shows how the mobile particles involved in metallic bonding can be pushed or pulled past each other. Metals are generally durable. Although metallic cations are mobile in a metal, they are strongly attracted to the electrons surrounding them and are not easily removed from the metal.

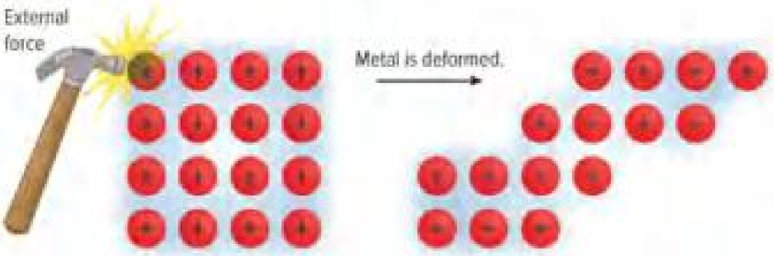

Figure 11 An applied force causes metal ions to move through delocalized electrons, making metals malleable and ductile.

Describe *in your own words what happens to metal ions when a metal is struck with a hammer.*

Hardness and strength The mobile electrons in transition metals consist not only of the two outer s electrons but also of the inner d electrons. As the number of delocalized electrons increases, so do the properties of hardness and strength. For example, strong metallic bonds are found in transition metals such as chromium, iron, and nickel, whereas alkali metals are considered soft because they have only one delocalized electron, ns^1.

Get It?

Contrast the behavior of metals and ionic compounds when each is struck by a hammer.

Metal Alloys

Due to the nature of metallic bonds, it is relatively easy to introduce other elements into the metallic crystal, forming an alloy. An **alloy** is a mixture of elements that has metallic properties. Because of their unique blend of properties, alloys have a wide range of commercial applications. Stainless steel, brass, and cast iron are a few of the many useful alloys. **Table 13** lists some commercially important alloys and their uses.

You likely have items made of alloys in your home, but you might not have realized it. Magnets often are used to attach decorative items or mementos to refrigerator doors. Candlesticks made of brass are common home accessories. Yellow gold jewelry is made of an alloy of copper and gold. Silver gold jewelry often is made of an alloy of gold with silver or palladium. Stainless steel often is used to make kitchen sinks, appliance doors and exposed sides, and tableware.

Table 13 Commercial Alloys

Common Name	Composition	Uses
Alnico	Fe 50%, Al 20%, Ni 20%, Co 10%	magnets
Brass	Cu 67–90%, Zn 10–33%	plumbing, hardware, lighting
Bronze	Cu 70–95%, Zn 1–25%, Sn 1–18%	bearings, bells, medals
Cast iron	Fe 96–97%, C 3–4%	casting
Gold, 10-carat	Au 42%, Ag 12–20%, Cu 37.46%	jewelry
Lead shot	Pb 99.8%, As 0.2%	shotgun shells
Pewter	Sn 70–95%, Sb 5–15%, Pb 0–15%	tableware
Stainless steel	Fe 73–79%, Cr 14–18%, Ni 7–9%	instruments, sinks
Sterling silver	Ag 92.5%, Cu 7.5%	tableware, jewelry

CCC CROSSCUTTING CONCEPTS

Structure and Function Make a list of the properties of metals discussed in this lesson. For each item on the list, explain how that property could be seen as evidence of the structure of metals at the atomic scale.

WORD ORIGIN

alloy

comes from the Latin word *alligare*, which means *to bind*

Figure 12 Bicycle frames are sometimes made of 3/2.5 titanium alloy, an alloy of titanium containing 3% aluminum and 2.5% vanadium.

Properties of alloys

The properties of alloys differ somewhat from the properties of the elements they contain. For example, steel is iron mixed with at least one other element. Some properties of iron are present, but steel has additional properties, such as increased strength. Alloys are classified into two basic types, substitutional alloys and interstitial alloys.

Substitutional alloys In a substitutional alloy, some of the atoms in the original metal are replaced by other metals of similar atomic size. Sterling silver is an example of a substitutional alloy. In sterling silver, copper atoms replace some of the silver atoms in the metallic crystal. The resulting solid has properties of both silver and copper.

Interstitial alloys An interstitial alloy, such as the titanium alloy shown in **Figure 12,** is formed when the small holes (interstices) in a metallic crystal are filled with smaller atoms. The best-known interstitial alloy is carbon steel, in which holes in an iron crystal are filled with carbon atoms. Iron alone is relatively soft and malleable, but the added carbon makes the solid harder, stronger, and less ductile.

Check Your Progress

Summary

- A metallic bond forms when metal cations attract freely moving, delocalized valence electrons.
- In the electron sea model, electrons move through the metallic crystal and are not held by any particular atom.
- The electron sea model explains the physical properties of metallic solids.
- Metal alloys are formed when a metal is mixed with one or more other elements.

Demonstrate Understanding

40. **Contrast** the structures of ionic compounds and metals.
41. **Explain** how the conductivity of electricity and the high boiling points of metals are explained by metallic bonding.
42. **Contrast** the cause of the attraction in ionic bonds and metallic bonds.
43. **Summarize** alloy types by correctly pairing these terms and phrases: substitutional, interstitial, replaced, and filled in.
44. **Design** an experiment that could be used to distinguish between a metallic solid and an ionic solid. Include at least two different methods for comparing the solids. Explain your reasoning.
45. **Model** Draw a model to represent the physical property of metals known as ductility, or the ability to be drawn into a wire. Base your drawing on the electron sea model shown in **Figure 10.**

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ENGINEERING & TECHNOLOGY

From Salty to Fresh

Desalination is the process of removing dissolved salts from seawater. Desalination technologies are becoming more important as other sources of fresh water are depleted.

Desalination plants such as this one can produce millions of liters of fresh water from seawater every day.

Taking the Salt out of Saltwater

Only 1 percent of all water on Earth is accessible fresh water. Most of Earth's water is in oceans. As fresh water sources are depleted, looking to the oceans as a source of water seems an obvious solution. The challenge for engineers is to develop desalination processes that are efficient, economical, and environmentally friendly.

Thermal distillation is a technique that has been used for thousands of years. In this process, heat is used to evaporate seawater. The water becomes a vapor and the other substances in seawater are left behind as a residue. The evaporated water is collected and condensed by cooling. The condensed water contains no salts or other dissolved impurities and is safe to drink.

A variety of technologies use thermal distillation, ranging from small solar stills that produce fresh water for individuals to huge distillation plants that produce fresh water for cities.

The main problem with thermal distillation is that it requires a lot of energy. Since the mid-1900s, desalination research has focused on another process called reverse osmosis, which uses much less energy than thermal distillation. In reverse osmosis, water is subjected to high pressure to cause it to pass through a semi-permeable membrane. The membrane allows water molecules but not ions to pass through, resulting in salt-free, drinkable water on one side of the membrane, and concentrated seawater on the other.

Reverse osmosis can be used in small devices and large desalination plants. The technology has drawbacks, though, including potential harm to marine life when the concentrated seawater is released back to the ocean. Engineers and scientists continue to improve desalination technologies to ensure a stable, safe source of drinking water for the future.

DEVELOP A MODEL TO ILLUSTRATE

Make diagrams that compare desalination using thermal distillation and reverse osmosis. Be sure to include labels and captions.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 ION FORMATION

- A chemical bond is the force that holds two atoms together.
- Some atoms form ions to gain stability. This stable configuration involves a complete outer energy level, usually consisting of eight valence electrons.
- Ions are formed by the loss or gain of valence electrons.
- The number of protons remains unchanged during ion formation.

- chemical bond
- cation
- anion

Lesson 2 IONIC BONDS AND IONIC COMPOUNDS

- Ionic compounds contain ionic bonds formed by the attraction of oppositely charged ions.
- Ions in an ionic compound are arranged in a repeating pattern known as a crystal lattice.
- Ionic compound properties are related to ionic bond strength.
- Ionic compounds conduct an electric current in the liquid phase and in aqueous solution.
- Lattice energy is the energy needed to remove 1 mol of ions from its lattice.

- ionic bond
- ionic compound
- crystal lattice
- electrolyte
- lattice energy

Lesson 3 NAMES AND FORMULAS FOR IONIC COMPOUNDS

- A formula unit gives the ratio of cations to anions in the ionic compound.
- A monatomic ion is formed from one atom. The charge of a monatomic ion is equal to its oxidation number.
- Roman numerals indicate the oxidation number of cations having multiple possible oxidation states.
- Polyatomic ions consist of more than one atom and act as a single unit.
- To indicate more than one polyatomic ion in a chemical formula, place parentheses around the polyatomic ion and use a subscript.

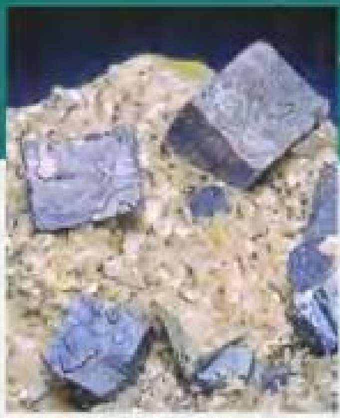
- formula unit
- monatomic ion
- polyatomic ion
- oxyanion

Lesson 2 METALLIC BONDS AND THE PROPERTIES OF METALS

- A metallic bond forms when metal cations attract freely moving, delocalized valence electrons.
- In the electron sea model, electrons move through the metallic crystal and are not held by any particular atom.
- The electron sea model explains the physical properties of metallic solids.
- Metal alloys are formed when a metal is mixed with one or more other elements.

- electron sea model
- delocalized electron
- metallic bond
- alloy

THREE-DIMENSIONAL THINKING
Module Wrap-Up



REVISIT THE PHENOMENON

Why do some crystals form cubes?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT
Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

Based on Real Data*

SEP Data Analysis Lab

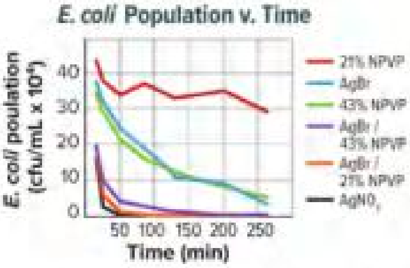
Can embedding nanoparticles of silver into a polymer give the polymer antimicrobial properties?

Researchers tested the antimicrobial properties of a new composite material—the polymer poly(4-vinyl-N-hexylpyridinium bromide), known as NPVP, which attracts cations. It is known that silver ions from silver bromide and silver nitrate exhibit antimicrobial activity. Silver bromide was embedded into the NPVP polymer. Scientists tested the antimicrobial properties of the composite material. Their results, illustrated in the graph, show the growth of *E. coli* bacteria over a period of approximately four hours. Each line represents the *E. coli* population in response to the introduction of a particular substance.

CER Analyze and Interpret Data

1. **Claim** Does the addition of silver bromide (AgBr) ions to NPVP improve the antimicrobial properties of the composite?
2. **Evidence, Reasoning** Does a composite polymer containing NPVP and silver bromide show antimicrobial properties? Explain your answer.

Data and Observations



*Data obtained from: Sambhy, V., et al. Published on the Web 7/7/2006. Silver Bromide Nanoparticle/Polymer Composites. Journal of the American Chemical Society.

COVALENT BONDING

ENCOUNTER THE PHENOMENON

Why does water expand when it freezes?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

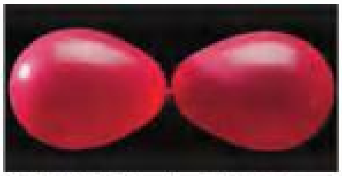
CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about why water expands when it freezes.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.



LESSON 4: Explore & Explain: Hybridization and Molecular Shape

LESSON 5: Explore & Explain: Intermolecular Forces and Properties of Covalent Compounds

LESSON 1
THE COVALENT BOND

FOCUS QUESTION

How do atoms bond in covalent molecules?

Why do atoms bond?

To understand why new compounds form, recall what you know about elements that do not tend to form new compounds–the noble gases. You learned that all noble gases have stable electron arrangements. This stable arrangement consists of a full outer energy level and has lower potential energy than other electron arrangements. Because of their stable configurations, noble gases seldom form compounds. Other elements frequently form compounds, such as the hydrogen and oxygen that form the water shown in **Figure 1**.

Gaining stability

The stability of an atom, ion, or compound is related to its energy, with lower energy states being more stable. In ionic bonds, metals and nonmetals gain stability by transferring electrons to form ions. The resulting ions have stable noble-gas electron configurations. In this module, you will learn that valence electron sharing is another way atoms can acquire the stable electron configuration of noble gases, resulting in stable molecules with less energy than the same set of atoms separated.

Figure 1 Each water droplet is made up of water molecules. Each water molecule is made up of two hydrogen atoms and one oxygen atom that have bonded by sharing electrons. The shapes of the drops are due to intermolecular forces acting on the water molecules.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Electron States and Simple Chemical Reactions
HS-PS1-2. Construct and revise an explanation for the outcome of a simple chemical reaction based on the outermost electron states of atoms, trends in the periodic table, and knowledge of the patterns of chemical properties.

Revisit the Encounter the Phenomenon Question
What information from this lesson can help you answer the module question?

What is a covalent bond?

You just read that atoms can share electrons to form stable electron configurations. How does this occur? Are there different ways in which electrons can be shared? How are the properties of these compounds different from those formed by ions? Read on to answer these questions.

Shared electrons

Atoms in nonionic compounds share electrons. The chemical bond that results from sharing valence electrons is a **covalent bond.** A **molecule** is formed when two or more atoms bond covalently. In a covalent bond, the shared electrons are considered to be part of the outer energy levels of both atoms involved. Covalent bonding generally can occur between elements that are near each other on the periodic table. The majority of covalent bonds form between atoms of nonmetallic elements.

Covalent bond formation

Diatomic molecules, such as hydrogen (H_2), nitrogen (N_2), oxygen (O_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2), form when two atoms of each element share electrons. They exist this way because the two-atom molecules are more stable than the individual atoms.

Consider fluorine, which has an electron configuration of $1s^22s^22p^5$. Each fluorine atom has seven valence electrons and needs another electron to form an octet. As two fluorine atoms approach each other, several forces act, as shown in **Figure 2**. Two repulsive forces act on the atoms, one from each atom's like-charged electrons and one from each atom's like-charged protons. A force of attraction also acts, as one atom's protons attract the other atom's electrons. As the fluorine atoms move closer, the attraction of the protons in each nucleus for the other atom's electrons increases until a point of maximum net attraction is achieved. At that point, the two atoms bond covalently, and a molecule forms. Now, each atom has a completed octet because they are each sharing a pair of electrons. If the two nuclei move closer, the repulsion forces increase and exceed the attractive forces.

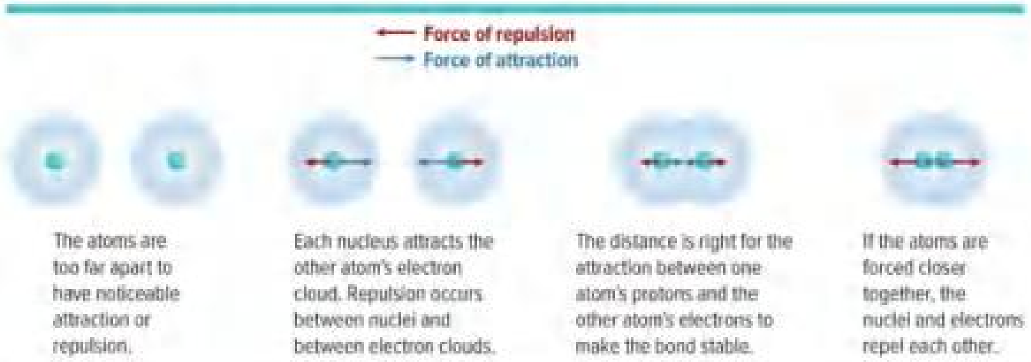


Figure 2 The arrows in this diagram show the net forces of attraction and repulsion acting on two fluorine atoms as they move toward each other. The overall force between two atoms is the result of electron-electron repulsion, nucleus-nucleus repulsion, and nucleus-electron attraction. At the position of maximum net attraction, a covalent bond forms.

Relate *How is the stability of the bond related to the forces acting on the atoms?*

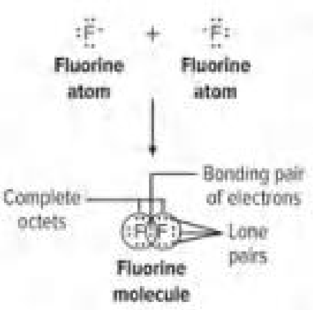


Figure 3 Two fluorine atoms share a pair of electrons to form a covalent bond. Note that the shared electron pair gives each atom a complete octet.

Infer *how the energy of the fluorine atoms changes when they form a covalent bond.*

The most stable arrangement of atoms in a covalent bond exists at some optimal distance between nuclei. At this point, the net attraction is greater than the net repulsion. Fluorine exists as a diatomic molecule because the sharing of one pair of electrons gives each fluorine atom a stable noble-gas configuration. As shown in **Figure 3**, each fluorine atom in the fluorine molecule has one pair of electrons that are covalently bonded (shared) and three pairs of electrons that are unbonded (not shared). Unbonded pairs are also known as lone pairs.

Single Covalent Bonds

When only one pair of electrons is shared, such as in a hydrogen molecule, it is a single covalent bond. The shared electron pair is often referred to as the bonding pair. For a hydrogen molecule, shown in **Figure 4,** each covalently bonded atom equally attracts the pair of shared electrons. Thus, the two shared electrons belong to each atom simultaneously, which gives each hydrogen atom the noble-gas configuration of helium ($1s^2$) and lower energy. The hydrogen molecule is more stable than either hydrogen atom is by itself.

Recall that electron-dot diagrams can be used to show valence electrons of atoms. In a **Lewis structure,** they represent the arrangement of electrons in a molecule. A line or a pair of vertical dots between the symbols of elements represents a single covalent bond in a Lewis structure. For example, a hydrogen molecule is written as H–H or H:H.

Group 17 and single bonds

The halogens—the group 17 elements, such as fluorine—have seven valence electrons. To form an octet, one more electron is needed. Therefore, atoms of group 17 elements form single covalent bonds with atoms of other nonmetals, such as carbon. You have already read that the atoms of some group 17 elements form covalent bonds with identical atoms. For example, fluorine exists as F_2, and chlorine exists as Cl_2. Because atoms from group 17 need only one more electron to form an octet they are very unstable and reactive.

Figure 4 When two hydrogen atoms share a pair of electrons, each hydrogen atom is stable because it has a full outer energy level.

Group 16 and single bonds

An atom of a group 16 element can share two electrons and can form two covalent bonds. Oxygen is a group 16 element with an electron configuration of $1s^2 2s^2 2p^4$. Water is composed of two hydrogen atoms and one oxygen atom. Each hydrogen atom has the noble-gas configuration of helium when it shares one electron with oxygen. Oxygen, in turn, has the noble-gas configuration of neon when it shares one electron with each hydrogen atom. **Figure 5a** shows the Lewis structure for a molecule of water. Notice that the oxygen atom has two single covalent bonds and two unshared pairs of electrons.

Group 15 and single bonds

Group 15 elements form three covalent bonds with atoms of nonmetals. Nitrogen is a group 15 element with the electron configuration of $1s^2 2s^2 2p^3$. Ammonia (NH_3) has three single covalent bonds. Three nitrogen electrons bond with the three hydrogen atoms leaving one pair of unshared electrons on the nitrogen atom. **Figure 5b** shows the Lewis structure for an ammonia molecule. Nitrogen also forms similar compounds with atoms of group 17 elements, such as nitrogen trifluoride (NF_3), nitrogen trichloride (NCl_3), and nitrogen tribromide (NBr_3). Each atom of these group 17 elements and the nitrogen atom share an electron pair.

Group 14 and single bonds

Atoms of group 14 elements form four covalent bonds. A methane molecule (CH_4) forms when one carbon atom bonds with four hydrogen atoms. Carbon, a group 14 element, has an electron configuration of $1s^2 2s^2 2p^2$. With four valence electrons, carbon needs four more electrons for a noble gas configuration. Therefore, when carbon bonds with other atoms, it forms four bonds. Because a hydrogen atom, a group 1 element, has one valence electron, it takes four hydrogen atoms to provide the four electrons needed by a carbon atom. The Lewis structure for methane is shown in **Figure 5c**. Carbon also forms single covalent bonds with other nonmetal atoms, including those in group 17.

Get It?
Describe how a Lewis structure shows a single covalent bond.

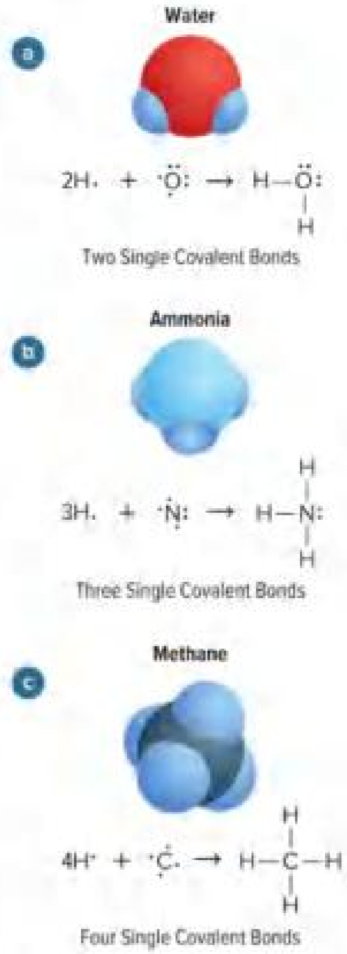


Figure 5 These chemical equations show how atoms share electrons and become stable. As shown by the Lewis structure for each molecule, all atoms in each molecule achieve a full outer energy level.

Describe *For the central atom in each molecule, describe how the octet rule is met.*

Figure 6 The frosted-looking portions of this globe were chemically etched using hydrogen fluoride (HF), a weak acid. Hydrogen fluoride reacts with silica, the major component of glass, and forms gaseous silicon tetrafluoride (SiF_4) and water.

EXAMPLE Problem 1

LEWIS STRUCTURE OF A MOLECULE The pattern on the glass shown in **Figure 6** was made by chemically etching its surface with hydrogen fluoride (HF). Draw the Lewis structure for a molecule of hydrogen fluoride.

1 ANALYZE THE PROBLEM

You are given the information that hydrogen and fluorine form the molecule hydrogen fluoride. An atom of hydrogen, a group 1 element, has only one valence electron. It can bond with any nonmetal atom when they share one pair of electrons. An atom of fluorine, a group 17 element, needs one electron to complete its octet. Therefore, a single covalent bond forms when atoms of hydrogen and fluorine bond.

2 SOLVE FOR THE UNKNOWN

To draw a Lewis structure, first draw the electron-dot diagram for each of the atoms. Then, rewrite the chemical symbols and draw a line between them to show the shared pair of electrons. Finally, add dots to show the unshared electron pairs.

3 EVALUATE THE ANSWER

Each atom in the new molecule now has a noble-gas configuration and is stable.

PRACTICE Problems

ADDITIONAL PRACTICE

Draw the Lewis structure for each molecule.

1. PH_3
2. H_2S
3. HCl
4. CCl_4
5. SiH_4
6. **CHALLENGE** Draw a generic Lewis structure for a molecule formed between atoms of group 1 and group 16 elements.

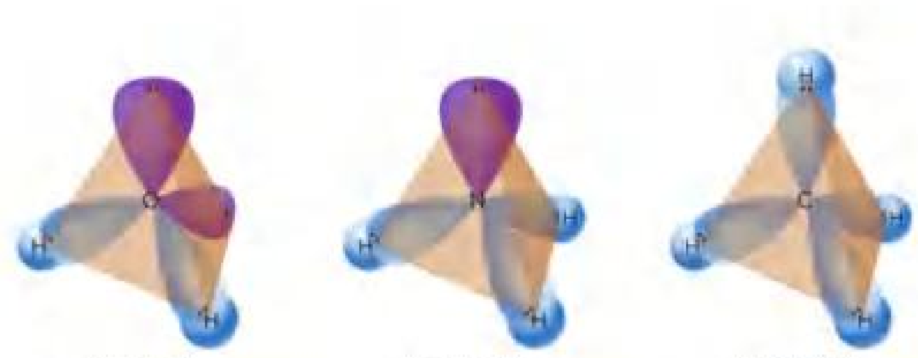


Figure 7 Sigma bonds formed in each of these molecules when the atomic orbital of each hydrogen atom overlapped end-to-end with the orbital of the central atom.

Interpret *Identify the number of sigma bonds in each molecule.*

The sigma bond

Single covalent bonds are also called **sigma bonds,** represented by the Greek letter sigma (σ). A sigma bond occurs when the pair of shared electrons is in an area centered between the two atoms. When two atoms share electrons, their valence atomic orbitals overlap end-to-end, concentrating the electrons in a bonding orbital between the two atoms. A bonding orbital is a localized region where bonding electrons will most likely be found. Sigma bonds can form when an s orbital overlaps with another s orbital or a p orbital, or two p orbitals overlap end-to-end. Water (H_2O), ammonia (NH_3), and methane (CH_4) have sigma bonds, as shown in **Figure 7**.

Get It?

List the orbitals that can form sigma bonds in a covalent compound, such as the DNA molecule shown in the photograph at the beginning of the module.

Multiple Covalent Bonds

In some molecules, atoms have noble-gas configurations when they share more than one pair of electrons with one or more atoms. Sharing multiple pairs of electrons forms multiple covalent bonds. A double covalent bond and a triple covalent bond are examples of multiple bonds. Carbon, nitrogen, oxygen, and sulfur atoms often form multiple bonds with other nonmetals. How do you know if two atoms will form a multiple bond? In general, the number of valence electrons needed to form an octet equals the number of covalent bonds that can form.

ACADEMIC VOCABULARY

overlap

to occupy the same area in part

The two driveways overlap at the street forming a common entrance.

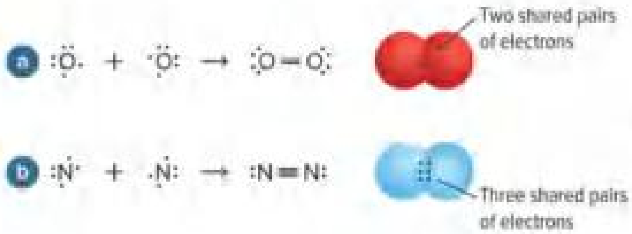


Figure 8 Multiple covalent bonds form when two atoms share more than one pair of electrons.

a. Two oxygen atoms form a double bond.

b. A triple bond forms between two nitrogen atoms.

Double bonds

A double covalent bond forms when two pairs of electrons are shared between two atoms. For example, atoms of the element oxygen only exist as diatomic molecules. Each oxygen atom has six valence electrons and must obtain two additional electrons for a noble-gas configuration, as shown in **Figure 8a**. A double covalent bond forms when each oxygen atom shares two electrons; a total of two pairs of electrons are shared between the two atoms.

Triple bonds

A triple covalent bond forms when three pairs of electrons are shared between two atoms. Diatomic nitrogen (N_2) molecules contain a triple covalent bond. Each nitrogen atom shares three electron pairs, forming a triple bond with the other nitrogen atom as shown in **Figure 8b**.

The pi bond

A multiple covalent bond consists of one sigma bond and at least one pi bond. A **pi bond,** represented by the Greek letter pi (π), forms when parallel orbitals overlap and share electrons. The shared electron pair of a pi bond occupies the space above and below the line that represents where the two atoms are joined together.

It is important to note that molecules having multiple covalent bonds contain both sigma and pi bonds. A double covalent bond, as shown in **Figure 9**, consists of one pi bond and one sigma bond. A triple covalent bond consists of two pi bonds and one sigma bond.

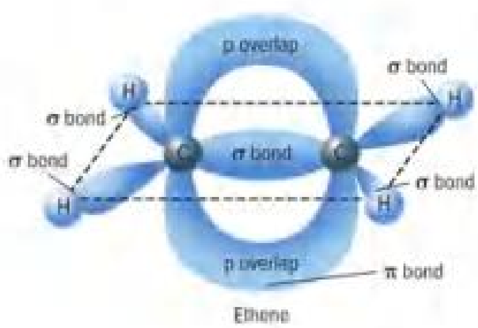


Figure 9 Notice how the multiple bond between the two carbon atoms in ethene (C_2H_4) consists of a sigma bond and a pi bond. The sigma bond is formed by the end-to-end overlap of orbitals directly between the two carbon atoms. The carbon atoms are close enough that the side-by-side p orbitals overlap and form the pi bond. This results in a doughnut-shaped cloud around the sigma bond.

The Strength of Covalent Bonds

Recall that a covalent bond involves attractive and repulsive forces. In a molecule, nuclei and electrons attract each other, but nuclei repel other nuclei, and electrons repel other electrons. When this balance of forces is upset, a covalent bond can be broken. Because covalent bonds differ in strength, some bonds break more easily than others. Several factors influence the strength of covalent bonds.

Bond length

The strength of a covalent bond depends on the distance between the bonded nuclei. The distance between the two bonded nuclei at the position of maximum attraction is called bond length, as shown in **Figure 10**. It is determined by the sizes of the two bonding atoms and how many electron pairs they share. Bond lengths for molecules of fluorine (F_2), oxygen (O_2), and nitrogen (N_2) are listed in **Table 1**. Notice that as the number of shared electron pairs increases, the bond length decreases.

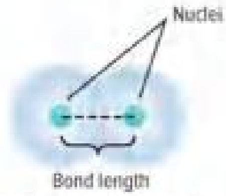


Figure 10 Bond length is the distance from the center of one nucleus to the center of the other nucleus of two bonded atoms.

Bond length and bond strength are also related: the shorter the bond length, the stronger the bond. Therefore, a single bond, such as that in F_2, is weaker than a double bond, such as that in O_2. Likewise, the double bond in O_2 is weaker than the triple bond in N_2.

Table 1 Covalent Bond Type, Bond Length, and Bond-Dissociation Energy

Molecule	Bond Type	Bond Length	Bond-Dissociation Energy
F_2	single covalent	1.43×10^{-10} m	159 kJ/mol
O_2	double covalent	1.21×10^{-10} m	498 kJ/mol
N_2	triple covalent	1.10×10^{-10} m	945 kJ/mol

Bonds and energy

An energy change occurs when a bond between atoms in a stable molecule forms or breaks. The amount of energy needed to break apart the molecule must be at least the amount released during its formation. This relationship between the amount of energy released during bond formation and the amount of energy needed to break covalent bonds is an example of how energy flows through chemical systems. The amount of energy required to break a specific covalent bond is called bond-dissociation energy and is always a positive value. The bond-dissociation energies for the covalent bonds in molecules of fluorine, oxygen, and nitrogen are listed in **Table 1**.

Bond-dissociation energy also indicates the strength of a chemical bond because of the inverse relationship between bond energy and bond length. As indicated in **Table 1**, the smaller the bond length is, the greater the bond-dissociation energy. In addition, the sum of the bond-dissociation energy values for all of the bonds in a molecule is the amount of chemical potential energy in a molecule of that compound.

Get It?

Compare the flow of energy in forming a molecule with the flow of energy when the molecule breaks apart.

The total energy change of a chemical reaction is determined from the energy of the bonds broken and formed. An **endothermic reaction** occurs when a greater amount of energy is required to break the existing bonds in the reactants than is released when the new bonds form in the products. An **exothermic reaction** occurs when more energy is released during product bond formation than is required to break bonds in the reactants. **Figure 11** illustrates a common exothermic reaction. You will study exothermic and endothermic reactions in much greater detail when you study the energy changes in chemical reactions.

Figure 11 Breaking the C–C bonds in charcoal and the O–O bonds in the oxygen in air requires an input of energy. Energy is released as heat and light when bonds form, producing CO_2. Thus, the burning of charcoal is an exothermic reaction.

Check Your Progress

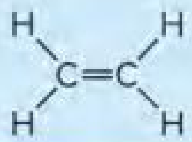
Summary

- Covalent bonds form when atoms share one or more pairs of electrons.
- Sharing one pair, two pairs, and three pairs of electrons forms single, double, and triple covalent bonds, respectively.
- Orbitals overlap directly in sigma bonds. Parallel orbitals overlap in pi bonds. A single covalent bond is a sigma bond but multiple covalent bonds are made of both sigma and pi bonds.
- Bond-dissociation energy is needed to break a covalent bond.

Demonstrate Understanding

7. **Describe** how the energy of a molecule compares to the same set of atoms separated.
8. **Describe** how the stability of an atom relates to the octet rule and the formation of covalent bonds.
9. **Illustrate** the formation of single, double, and triple covalent bonds using Lewis structures.
10. **Compare and contrast** ionic bonds and covalent bonds.
11. **Contrast** sigma bonds and pi bonds.
12. **Apply** Create a graph using the bond-dissociation energy data and the bond-length data in **Table 1**. Describe the relationship between bond length and bond-dissociation energy.
13. **Compare** the bond-dissociation energies needed to break the bonds in the structures below. Then explain how these values relate to the amount of energy needed to form the compounds.

a. H−C≡C−H

b. 

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LESSON 2

NAMING MOLECULES

FOCUS QUESTION

How do you name molecules?

Naming Binary Molecular Compounds

Many molecular compounds have common names, but they also have scientific names that reveal their composition. To write the formulas and names of molecules, you will use processes similar to those described for ionic compounds.

Start with a binary molecular compound. Note that a binary molecular compound is composed only of two nonmetal atoms—not metal atoms or ions. An example is dinitrogen monoxide (N_2O), a gaseous anesthetic that is more commonly known as nitrous oxide or laughing gas. The naming of N_2O is explained in the following rules.

1. The first element in the formula is always named first, using the entire element name. **N is the symbol for *nitrogen*.**
2. The second element in the formula is named using its root and adding the suffix *-ide*. **O is the symbol for oxygen so the second word is *oxide*.**
3. Prefixes are used to indicate the number of atoms of each element that are present in the compound. **Table 2** lists the most common prefixes used. **There are two atoms of nitrogen and one atom of oxygen, so the first word is *dinitrogen* and the second word is *monoxide*.**

Table 2 Prefixes in Covalent Compounds

Number of Atoms	Prefix	Number of Atoms	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Review the News

Obtain information from a current news story about naming molecules. Evaluate your source and communicate your findings to your class.

There are exceptions to using the prefixes shown in **Table 2**. The first element in the compound name never uses the *mono-* prefix. For example, CO is carbon monoxide, not monocarbon monoxide. Also, if using a prefix results in two consecutive vowels, one of the vowels is usually dropped to avoid an awkward pronunciation. For example, notice that the oxygen atom in CO is called monoxide, not monooxide.

EXAMPLE Problem 2

NAMING BINARY MOLECULAR COMPOUNDS Name the compound P_2O_5, which is used as a drying and dehydrating agent.

1 ANALYZE THE PROBLEM

You are given the formula for a compound. The formula contains the elements and the number of atoms of each element in one molecule of the compound. Because only two different elements are present and both are nonmetals, the compound can be named using the rules for naming binary molecular compounds.

2 SOLVE FOR THE UNKNOWN

First, name the elements involved in the compound.

phosphorus	The first element, represented by P, is phosphorus.
oxide	The second element, represented by O, is oxygen. Add the suffix *-ide* to the root of oxygen, *ox-*.
phosphorus oxide	Combine the names.

Now modify the names to indicate the number of atoms present in a molecule.

diphosphorus pentoxide	From the formula P_2O_5, you know that two phosphorus atoms and five oxygen atoms make up a molecule of the compound. From Table 2, you know that *di-* is the prefix for two and *penta-* is the prefix for five. The *a* in *penta-* is not used because oxide begins with a vowel.

3 EVALUATE THE ANSWER

The name diphosphorus pentoxide shows that a molecule of the compound contains two phosphorus atoms and five oxygen atoms, which agrees with the compound's chemical formula, P_2O_5.

PRACTICE Problems

ADDITIONAL PRACTICE

Name each of the binary covalent compounds listed below.

14. CO_2

15. SO_2

16. NF_3

17. CCl_4

18. CHALLENGE What is the formula for diarsenic trioxide?

Common names for some molecular compounds

Have you ever enjoyed an icy, cold glass of dihydrogen monoxide on a hot day? You probably have but you most likely called it by its common name, water. Recall that many ionic compounds have common names in addition to their scientific ones. For example, baking soda is sodium hydrogen carbonate and common table salt is sodium chloride.

Many binary molecular compounds, such as nitrous oxide and water, were discovered and given common names long before the present-day naming system was developed. Other binary covalent compounds that are generally known by their common names rather than their scientific names are ammonia (NH_3), hydrazine (N_2H_4), and nitric oxide (NO).

Get It?

Apply What are the scientific names for ammonia, hydrazine, and nitric oxide?

Naming Acids

Water solutions, also called aqueous solutions, of some molecules are acidic and are named as acids. Acids are important compounds with specific properties. If a compound produces hydrogen ions (H^+) in solution, it is an acid. For example, HCl produces H^+ in solution and is an acid. Two common types of acids exist—binary acids and oxyacids.

Naming binary acids

A binary acid contains hydrogen and one other element. The naming of the common binary acid known as hydrochloric acid is explained in the following rules.

1. The first word has the prefix *hydro-* to name the hydrogen part of the compound. The rest of the first word consists of a form of the root of the second element plus the suffix *-ic*. **HCl (hydrogen and chlorine) becomes *hydrochloric*.**
2. The second word is always *acid*. **Thus, HCl in a water solution is called *hydrochloric acid*.**

Although the term *binary* indicates exactly two elements, a few acids that contain more than two elements are named according to the rules for naming binary acids.

If no oxygen is present in the formula for the acidic compound, the acid is named in the same way as a binary acid, except that the root of the second part of the name is the root of the polyatomic ion that the acid contains. For example, HCN, which is composed of hydrogen and the cyanide ion, is called *hydrocyanic acid* in solution.

Naming oxyacids

An acid that contains both a hydrogen atom and an oxyanion is referred to as an **oxyacid**. Recall that an oxyanion is a polyatomic ion containing one or more oxygen atoms. The following rules explain how to name nitric acid, an important oxyacid that has a chemical formula of HNO_3.

Table 3 Naming Oxyacids

Compound	Oxyanion	Acid Suffix	Acid Name
$HClO_3$	chlorate	-ic	chloric acid
$HClO_2$	chlorite	-ous	chlorous acid
HNO_3	nitrate	-ic	nitric acid
HNO_2	nitrite	-ous	nitrous acid

1. First, identify the oxyanion present. The first word of an oxyacid's name consists of the root of the oxyanion and the prefix *per-* or *hypo-* if it is part of the oxyanion's name. The first word of the oxyacid's name also has a suffix that depends on the oxyanion's suffix. If the oxyanion's name ends with the suffix *-ate*, replace it with the suffix *-ic*. If the name of the oxyanion ends with the suffix *-ite*, replace it with the suffix *-ous*. **NO_3^-, the nitrate ion, becomes *nitric*.**
2. The second word of the name is always *acid*. **HNO_3 (hydrogen and the nitrate ion) becomes *nitric acid*.**

Table 3 shows how the names of several oxyacids follow these rules. Notice that the hydrogen in an oxyacid is not part of the name.

You have learned that naming covalent compounds follows different sets of rules depending on the composition of the compound. **Table 4** summarizes the formulas and names of several covalent compounds. Note that an acid, whether a binary acid or an oxyacid, can have a common name in addition to its compound name.

PRACTICE Problems ADDITIONAL PRACTICE

Name the following acids. Assume each compound is dissolved in water.

19. HI

20. $HClO_4$

21. $HClO_3$

22. H_2SO_4

23. H_2S

24. CHALLENGE What is the formula for periodic acid?

Table 4 Formulas and Names of Some Covalent Compounds

Formula	Common Name	Molecular Compound Name
H_2O	water	dihydrogen monoxide
NH_3	ammonia	nitrogen trihydride
N_2H_4	hydrazine	dinitrogen tetrahydride
HCl	muriatic acid	hydrochloric acid
$C_9H_8O_4$	aspirin	2-(acetyloxy)benzoic acid

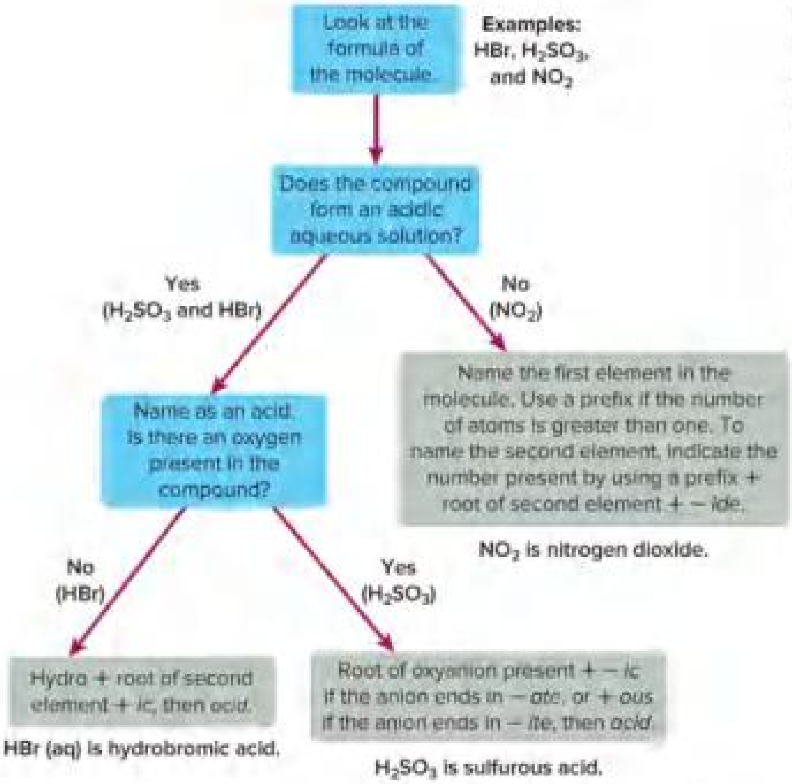

Figure 12 Use this flowchart to name molecular compounds when their formulas are known.

Apply *Which compound in the figure is an oxyacid? Which is a binary acid?*

The flowchart in **Figure 12** can help you determine the name of a molecular covalent compound. To use the chart, start at the top and work downward by reading the text contained in the colored boxes. Apply the text at each step of the flowchart to the formula of the compound that you wish to name.

Writing Formulas from Names

The name of a molecular compound reveals its composition and is important in communicating the nature of the compound. Given the name of any binary molecule, you should be able to write the correct chemical formula. The prefixes used in a name indicate the exact number of each atom present in the molecule and determine the subscripts used in the formula. If you are having trouble writing formulas from the names for binary compounds, you might want to review the naming rules listed on the pages at the beginning of this lesson.

Get It?

Interpret How do you determine the correct subscripts to use in order to write the chemical formula of a binary molecular compound when given the compound's name?

The formula for an acid can also be derived from the name. It is helpful to remember that all binary acids contain hydrogen and one other element. Also, keep in mind that the first part of the name of a binary acid will always use the prefix *hydro-*.

For oxyacids—acids containing oxyanions—you will need to know the names of the common oxyanions. If you need to review the formulas and names of oxyanions, see **Table 9** in the previous module.

PRACTICE Problems ADDITIONAL PRACTICE

Identify the formula for each compound.

25. silver chloride

26. dihydrogen monoxide

27. chlorine trifluoride

28. diphosphorus trioxide

29. disulfur decafluoride

30. CHALLENGE What is the formula for carbonic acid?

Check Your Progress

Summary

- Names of covalent molecular compounds include prefixes for the number of each atom present. The final letter of the prefix is dropped if the element name begins with a vowel.
- Molecules that produce H^+ in solution are acids. Binary acids contain hydrogen and one other element. Oxyacids contain hydrogen and an oxyanion.

Demonstrate Understanding

31. **Summarize** the rules for naming binary molecular compounds.
32. **Define** a binary molecular compound.
33. **Describe** the difference between a binary acid and an oxyacid.
34. **Apply** Using the system of rules for naming binary molecular compounds, describe how you would name the molecule N_2O_4.
35. **Apply** Write the molecular formula for each of these compounds: *iodic acid, disulfur trioxide, dinitrogen monoxide*, and *hydrofluoric acid*.
36. **State** the molecular formula for each compound listed below.
    a. dinitrogen trioxide
    b. nitrogen monoxide
    c. hydrochloric acid
    d. chloric acid
    e. sulfuric acid
    f. sulfurous acid

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LESSON 3

MOLECULAR STRUCTURES

FOCUS QUESTION

How are electrons shared in covalent molecules?

Structural Formulas

You have already studied the structure of ionic compounds—substances formed from ionic bonds. Covalent molecules described in this module have structures that are different from those of ionic compounds. When studying the molecular structures of covalent compounds, various models are used as representations of the molecules.

The molecular formula, which shows the element symbols and numerical subscripts, tells you the type and number of each atom in a molecule. As shown in **Figure 13**, there are several different models that can be used to represent a molecule. Note that in the ball-and-stick and space-filling molecular models, atoms of each specific element are represented by spheres of a representative color, as shown in **Table R-1** in the Student Resources. These colors are used for identifying the atoms if the chemical symbol of the element is not present.

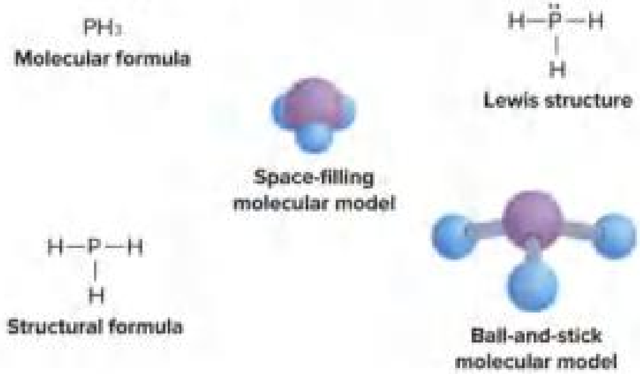


Figure 13 All of these models can be used to show the relative locations of atoms and electrons in the phosphorus trihydride (phosphine) molecule.

Compare and contrast *the types of information contained in each model.*

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science and Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Revisit the Encounter the Phenomenon Question

What information from this lesson can help you answer the module question?

One of the most useful molecular models is the **structural formula,** which uses letter symbols and bonds to show relative positions of atoms. You can predict the structural formula for many molecules by drawing the Lewis structure. You have already seen some simple examples of Lewis structures, but more involved structures are needed to help you determine the shapes of molecules.

Lewis structures

Although it is fairly easy to draw Lewis structures for most compounds formed by nonmetals, it is a good idea to follow a regular procedure. Whenever you need to draw a Lewis structure, follow the steps outlined in this Problem-Solving Strategy.

PROBLEM-SOLVING STRATEGY

Drawing Lewis Structures

1. Predict the location of certain atoms. The atom that has the least attraction for shared electrons will be the central atom in the molecule. This element is usually the one closer to the left side of the periodic table. The central atom is located in the center of the molecule; all other atoms become terminal atoms.

   Hydrogen is always a terminal, or end, atom. Because it can share only one pair of electrons, hydrogen can be connected to only one other atom.
2. Determine the number of electrons available for bonding.

   This number is equal to the total number of valence electrons in the atoms that make up the molecule.
3. Determine the number of bonding pairs.

   To do this, divide the number of electrons available for bonding by two.
4. Place the bonding pairs.

   Place one bonding pair (single bond) between the central atom and each of the terminal atoms.
5. Determine the number of electron pairs remaining.

   To do this, subtract the number of pairs used in Step 4 from the total number of bonding pairs determined in Step 3. These remaining pairs include lone pairs as well as pairs used in double and triple bonds. Place lone pairs around each terminal atom (except H atoms) bonded to the central atom to satisfy the octet rule. Any remaining pairs will be assigned to the central atom.
6. Determine whether the central atom satisfies the octet rule.

   Is the central atom surrounded by four electron pairs? If not, it does not satisfy the octet rule. To satisfy the octet rule, convert one or two of the lone pairs on the terminal atoms into a double bond or a triple bond between the terminal atom and the central atom. These pairs are still associated with the terminal atom as well as with the central atom. Remember that carbon, nitrogen, oxygen, and sulfur often form double and triple bonds.

Apply the Strategy

Study Example Problems 3 through 5 to see how the steps in the Problem-Solving Strategy are applied.

EXAMPLE Problem 3

LEWIS STRUCTURE FOR A COVALENT COMPOUND WITH SINGLE BONDS Ammonia is a raw material used in the manufacture of many products, including fertilizers, cleaning products, and explosives. Draw the Lewis structure for ammonia (NH_3).

1 ANALYZE THE PROBLEM

Ammonia molecules consist of one nitrogen atom and three hydrogen atoms. Because hydrogen must be a terminal atom, nitrogen is the central atom.

2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons available for bonding.

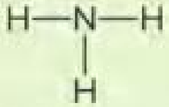
$$1 \text{ N atom} \times \frac{5 \text{ valence electrons}}{1 \text{ N atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{1 \text{ H atom}} = 8 \text{ valence electrons}$$

There are 8 valence electrons available for bonding.

$$\frac{8 \text{ electrons}}{2 \text{ electrons/pair}} = 4 \text{ pairs}$$

Determine the total number of bonding pairs. To do this, divide the number of available electrons by two.

Four pairs of electrons are available for bonding.



Place a bonding pair (a single bond) between the central nitrogen atom and each terminal hydrogen atom.

Determine the number of bonding pairs remaining.

4 pairs total − 3 pairs used
= 1 pair available

Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

The remaining pair—a lone pair—must be added to either the terminal atoms or the central atom. Because hydrogen atoms can have only one bond, they have no lone pairs.

Place the remaining lone pair on the central nitrogen atom.

3 EVALUATE THE ANSWER

Each hydrogen atom shares one pair of electrons, as required, and the central nitrogen atom shares three pairs of electrons and has one lone pair, providing a stable octet.

PRACTICE Problems

ADDITIONAL PRACTICE

37. Draw the Lewis structure for BH_3.
38. **CHALLENGE** A nitrogen trifluoride molecule contains numerous lone pairs. Draw its Lewis structure.

CCC CROSSCUTTING CONCEPTS

System and System Models Refer to the five common molecular models, shown on the first page of this lesson. With a partner, choose 2 compounds listed in question number 36 in the lesson 2 Review It!. Individually, without sharing your work, use 5 index cards to draw each model type to represent each compound, using a total of 10 index cards. Share your cards with your partner and compare your models. If they differ, work together to reach an agreed configuration.

EXAMPLE Problem 4

LEWIS STRUCTURE FOR A COVALENT COMPOUND WITH MULTIPLE BONDS Carbon dioxide is a product of all cellular respiration. Draw the Lewis structure for carbon dioxide (CO_2).

1 ANALYZE THE PROBLEM

The carbon dioxide molecule consists of one carbon atom and two oxygen atoms. Because carbon has less attraction for shared electrons, carbon is the central atom, and the two oxygen atoms are terminal.

2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons available for bonding.

$$1 \text{ C atom} \times \frac{4 \text{ valence electrons}}{1 \text{ C atom}} + 2 \text{ O atoms} \times \frac{6 \text{ valence electrons}}{1 \text{ O atom}} = 16 \text{ valence electrons}$$

There are 16 valence electrons available for bonding.

$$\frac{16 \text{ electrons}}{2 \text{ electrons/pair}} = 8 \text{ pairs}$$

Determine the total number of bonding pairs by dividing the number of available electrons by two.

Eight pairs of electrons are available for bonding.

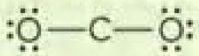
O—C—O

Place a bonding pair (a single bond) between the central carbon atom and each terminal oxygen atom.

Determine the number of electron pairs remaining.

8 pairs total − 2 pairs used
= 6 pairs available

Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.



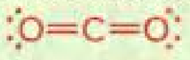
Add three lone pairs to each terminal oxygen atom.

Determine the number of electron pairs remaining.

6 pairs available − 6 pairs used
= 0 pairs available

Subtract the lone pairs from the pairs available.

Examine the incomplete structure above (showing the placement of the lone pairs). Note that the carbon atom does not have an octet and that there are no more electron pairs available. To give the carbon atom an octet, the molecule must form double bonds.



Use a lone pair from each O atom to form a double bond with the C atom.

3 EVALUATE THE ANSWER

Both carbon and oxygen now have an octet, which satisfies the octet rule.

PRACTICE Problems ADDITIONAL PRACTICE

39. Draw the Lewis structure for ethylene, C_2H_4.

40. CHALLENGE A molecule of carbon disulfide contains both lone pairs and multiple-covalent bonds. Draw its Lewis structure.

Lewis structures for polyatomic ions

To find the total number of electrons available for bonding in a polyatomic ion, first find the number available in the atoms present in the ion. Then, add the ion charge if the ion is negative, or subtract the ion charge if the ion is positive. Compared to the number of valence electrons present in the atoms that make up the ion, more electrons are present if the ion is negatively charged and fewer are present if the ion is positive.

EXAMPLE Problem 5

LEWIS STRUCTURE FOR A POLYATOMIC ION Draw the correct Lewis structure for the polyatomic ion phosphate (PO_4^{3-}).

1 ANALYZE THE PROBLEM

You are given that the phosphate ion consists of one phosphorus atom and four oxygen atoms and has a charge of 3−. Because phosphorus has less attraction for shared electrons than oxygen, phosphorus is the central atom and the four oxygen atoms are terminal atoms.

2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons available for bonding.

$$1\text{ P atom} \times \frac{5\text{ valence electrons}}{1\text{ P atom}} + 4\text{ O atoms} \times \frac{6\text{ valence electrons}}{1\text{ O atom}}$$

$$+\ 3\text{ electrons from the negative charge} = 32\text{ valence electrons}$$

$$\frac{32\text{ electrons}}{2\text{ electrons/pair}} = 16\text{ pairs}$$

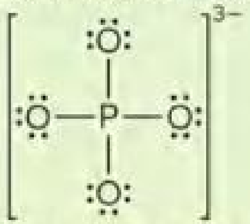
Determine the total number of bonding pairs.

Draw single bonds from each terminal oxygen atom to the central phosphorus atom.

16 pairs total − 4 pairs used = 12 pairs available

Subtract the number of pairs used from the total number of pairs of electrons available.

Add three lone pairs to each terminal oxygen atom. 12 pairs available − 12 lone pairs used = 0



Subtracting the lone pairs used from the pairs available verifies that there are no electron pairs available for the phosphorus atom. The Lewis structure for the phosphate ion is shown.

3 EVALUATE THE ANSWER

All of the atoms have an octet, and the group has a net charge of 3−.

PRACTICE Problems

ADDITIONAL PRACTICE

41. Draw the Lewis structure for the NH_4^+ ion.

42. CHALLENGE The ClO_4^- ion contains numerous lone pairs. Draw its Lewis structure.

Resonance Structures

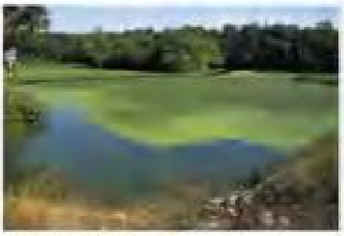
Using the same sequence of atoms, it is possible to have more than one correct Lewis structure when a molecule or polyatomic ion has both a double bond and a single bond. Consider the polyatomic ion nitrate (NO_3^-), shown in **Figure 14a**. Three equivalent structures can be used to represent the nitrate ion.

Resonance is a condition that occurs when more than one valid Lewis structure can be written for a molecule or ion. The two or more correct Lewis structures that represent a single molecule or ion are referred to as resonance structures. Resonance structures differ only in the position of the electron pairs, never the atom positions. The location of the lone pairs and bonding pairs differs in resonance structures. The molecule O_3 and the polyatomic ions NO_3^-, NO_2^-, SO_3^{2-}, and CO_3^{2-} all exhibit resonance.

It is important to note that each molecule or ion that exhibits resonance behaves as if it has only one structure. Refer to **Figure 14b**. Experimentally measured bond lengths show that the bonds are identical to each other. They are shorter than single bonds but longer than double bonds. The actual bond length is an average of the bonds in the resonance structures.

Real-World Chemistry

Phosphorus and Nitrogen



ALGAL BLOOMS Phosphorus and nitrogen are nutrients required for algae growth. Both can enter lakes and streams from discharges of sewage and industrial waste, and in fertilizer runoff. If these substances build up in a body of water, a rapid growth of algae, known as an algal bloom, can occur, forming a thick layer of green slime over the water's surface. When the algae use up the supply of nutrients, they die and decompose. This process reduces the amount of dissolved oxygen in the water that is available to other aquatic organisms.

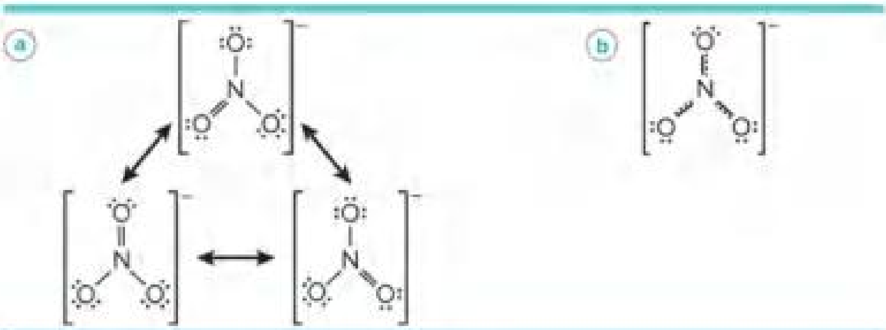


Figure 14 The nitrate ion (NO_3^-) exhibits resonance.

a. These resonance structures differ only in the location of the double bond. The locations of the nitrogen and oxygen atoms stay the same.

b. The actual nitrate ion is like an average of the three resonance structures in **a**. The dotted lines indicate possible locations of the double bond.

PRACTICE Problems ADDITIONAL PRACTICE

Draw the Lewis resonance structures for the following molecules.

43. NO_2^- **44.** SO_2 **45.** O_3

46. CHALLENGE Draw the Lewis resonance structure for the ion SO_3^{2-}.

Exceptions to the Octet Rule

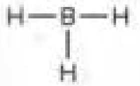
Generally, atoms attain an octet when they bond with other atoms. Some molecules and ions, however, do not obey the octet rule. There are several reasons for these exceptions.

Figure 15 The central nitrogen atom in this NO_2 molecule does not satisfy the octet rule; the nitrogen atom has only seven electrons in its outer energy level.

Odd number of valence electrons

First, a small group of molecules might have an odd number of valence electrons and be unable to form an octet around each atom. For example, NO_2 has five valence electrons from nitrogen and 12 from oxygen, totaling 17, which cannot form an exact number of electron pairs. See **Figure 15**. ClO_2 and NO are other examples of molecules with odd numbers of valence electrons.

Suboctets and coordinate covalent bonds



Another exception to the octet rule is due to a few compounds that form suboctets—stable configurations with fewer than eight electrons present around an atom. This group is relatively rare, and BH_3 is an example. Boron, a group 13 metalloid, forms three covalent bonds with other nonmetallic atoms.The boron atom shares only six electrons—too few to form an octet. Such compounds tend to be reactive and can share an entire pair of electrons donated by another atom.

A **coordinate covalent bond** forms when one atom donates both of the electrons to be shared with an atom or ion that needs two electrons to form a stable electron arrangement with lower potential energy. Refer to **Figure 16**. Atoms or ions with lone pairs often form coordinate covalent bonds with atoms or ions that need two more electrons.

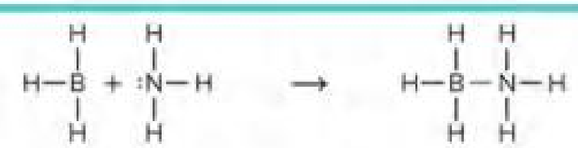


Figure 16 In this reaction between boron trihydride (BH_3) and ammonia (NH_3), the nitrogen atom donates both electrons that are shared by boron and nitrogen, forming a coordinate covalent bond.

Interpret *Does the coordinate covalent bond in the product molecule satisfy the octet rule?*

Figure 17 Prior to the reaction of PCl_3 and Cl_2, every reactant atom follows the octet rule. After the reaction, the product, PCl_5, has an expanded octet containing ten electrons.

Expanded octets

The third group of compounds that does not follow the octet rule has central atoms that contain more than eight valence electrons. This electron arrangement is referred to as an expanded octet. An expanded octet can be explained by considering the d orbitals that occur in the energy levels of elements in period three or higher. An example of an expanded octet, shown in **Figure 17**, is the bond formation in the molecule PCl_5. Five bonds are formed with ten electrons shared in one s orbital, three p orbitals, and one d orbital. Another example is the molecule SF_6, which has six bonds sharing 12 electrons in an s orbital, three p orbitals, and two d orbitals. When you draw the Lewis structures for these compounds, either extra lone pairs are added to the central atom or more than four bonding atoms are present in the molecule.

Get It?

Summarize three reasons why some molecules do not conform to the octet rule.

EXAMPLE Problem 6

LEWIS STRUCTURE: EXCEPTION TO THE OCTET RULE Xenon is a noble gas that will form a few compounds with nonmetals that strongly attract electrons. Draw the correct Lewis structure for xenon tetrafluoride (XeF_4).

1 ANALYZE THE PROBLEM

You are given that a molecule of xenon tetrafluoride consists of one xenon atom and four fluorine atoms. Xenon has less attraction for electrons, so it is the central atom.

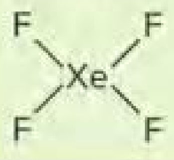
2 SOLVE FOR THE UNKNOWN

First, find the total number of valence electrons.

$$1 \text{ Xe atom} \times \frac{8 \text{ valence electrons}}{1 \text{ Xe atom}} + 4 \text{ F atoms} \times \frac{7 \text{ valence electrons}}{1 \text{ F atom}} = 36 \text{ valence electrons}$$

$$\frac{36 \text{ electrons}}{2 \text{ electrons/pair}} = 18 \text{ pairs}$$

Determine the total number of bonding pairs.



Use four bonding pairs to bond the four F atoms to the central Xe atom.

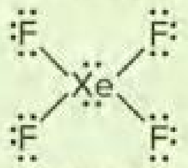
EXAMPLE Problem 6 (continued)

18 pairs available − 4 pairs used = 14 pairs available

Determine the number of remaining pairs.

$14 \text{ pairs} - 4 \text{ F atoms} \times \frac{3 \text{ pairs}}{1 \text{ F atom}} = 2 \text{ pairs unused}$

Add three pairs to each F atom to obtain an octet. Determine how many pairs remain.



Place the two remaining pairs on the central Xe atom.

3 EVALUATE THE ANSWER

This structure gives xenon 12 total electrons, an expanded octet. Xenon compounds, such as the XeF_4 shown here, are toxic because they are highly reactive.

PRACTICE Problems

ADDITIONAL PRACTICE

Draw the expanded octet Lewis structure for each molecule.

47. ClF_3

48. PCl_5

49. CHALLENGE Draw the Lewis structure for the molecule formed when six fluorine atoms and one sulfur atom bond covalently.

Check Your Progress

Summary

- Different models can be used to represent molecules.
- Resonance occurs when more than one valid Lewis structure exists for the same molecule.
- Exceptions to the octet rule occur in some molecules.

Demonstrate Understanding

50. **Describe** the information contained in a structural formula.
51. **State** the steps used to draw Lewis structures in your own words.
52. **Summarize** exceptions to the octet rule by correctly pairing these molecules and phrases: odd number of valence electrons, PCl_5, ClO_2, BH_3, expanded octet, less than an octet.
53. **Evaluate** A classmate states that a binary compound having only sigma bonds displays resonance. Could the classmate's statement be true?
54. **Draw** the resonance structures for the dinitrogen oxide (N_2O) molecule.
55. **Draw** the Lewis structures for CN^-, SiF_4, HCO_3^-, and, AsF_6^-.

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LESSON 4

MOLECULAR SHAPES

FOCUS QUESTION

What shapes do molecules form?

VSEPR Model

Once a Lewis structure is drawn, the molecular geometry, or shape, of a molecule can be determined using the **V**alence **S**hell **E**lectron **P**air **R**epulsion model, or **VSEPR model.** This model is based on an arrangement that minimizes the repulsion of shared and unshared electron pairs around the central atom. Recall that the attraction and repulsion between electric charges at the atomic scale explain the structure, properties, and transformations of matter.

To understand the VSEPR model better, imagine balloons that are inflated to similar sizes and tied together, as shown in **Figure 18**. Each balloon represents an electron-dense region. When a set of balloons is connected at a central point, which represents a central atom, the balloons naturally form a shape that minimizes interactions between the balloons.

Figure 18 Electron pairs in a molecule are located as far apart as they can be, just as these balloons are arranged. Two pairs form a linear shape. Three pairs form a trigonal planar shape. Four pairs form a tetrahedral shape.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Modeling Electrostatic Forces—Covalent Bonding
HS-PS2-4. Use mathematical representations of Newton's Law of Gravitation and Coulomb's Law to describe and predict the gravitational and electrostatic forces between objects.

Virtual Investigation: Molecular Shapes
Develop and use models to discover the structure and function of atomic electric charges.

Bond angle

The electron pairs in a molecule repel one another in a similar way. These forces cause the atoms in a molecule to be positioned at fixed angles relative to one another. The angle formed by two terminal atoms and the central atom is a bond angle. Bond angles predicted by VSEPR are supported by experimental evidence.

Unshared pairs of electrons are also important in determining the shape of the molecule. These electrons occupy a slightly larger orbital than shared electrons. Therefore, shared bonding orbitals are pushed together by unshared pairs.

BIOLOGY Connection The shape of a molecule determines many of its physical and chemical properties. For example, scientists discovered that the shape of food molecules is important to our sense of taste. The surface of your tongue is covered with taste buds, each of which contains from 50 to 100 taste receptor cells. Taste receptor cells can detect five distinct tastes—sweet, bitter, salty, sour, and MSG (monosodium glutamate).

The shapes of food molecules are determined by their chemical structures. When a molecule enters a taste bud, it must have the correct shape for the nerve in each receptor cell to respond and send a message to the brain. The brain then interprets the message as a certain taste. When such molecules bind to sweet receptors, they are sensed as sweet. The greater the number of food molecules that fit a sweet receptor cell, the sweeter the food tastes. Sugars and artificial sweeteners are not the only sweet molecules. Some proteins found in fruits are also sweet molecules. Some common molecular shapes are illustrated in **Table 5** and **Table 6** on the following pages.

Hybridization

Sometimes atomic orbitals fuse to form a new hybridized orbital. **Hybridization** occurs when two things are combined and the result has characteristics of both. For example, during chemical bonding, if electrons come from 2 different atomic orbitals, such as a p or an s, they must rearrange or combine into a hybrid orbital with the same shape and energy level. To understand hybrid orbitals, consider the bonding involved in the methane molecule (CH_4). The hybrid orbitals in a carbon atom are shown in **Figure 19** on the next page in blue. Note, as shown in the electron notation, that although carbon initially has only 2 electrons in its p orbital, a 1 s electron is promoted from the s to the p orbital so that a total of 4 unpaired electrons can be shared. These 4 unpaired electrons are rearranged into four hybrid sp^3 orbitals that can now be used to bond with 4 hydrogen atoms.

WORD ORIGIN

trigonal planar

comes from the Latin words *trigonum*, which means *triangular, and plan-*, which means *flat*

CCC CROSSCUTTING CONCEPTS

Structure and Function Designing new systems requires examining properties of materials, structures of components, and connections of components to solve a problem. Write a blog post for a science web site that applies this statement to developing robots with artificial taste receptors that mimic human ones.

The hybrid orbitals in a carbon atom are shown in **Figure 19**. The hybrid orbital is called an sp^3 orbital because the four hybrid orbitals form from one s orbital and three p orbitals.

The number of atomic orbitals that mix and form the hybrid orbital equals the total number of pairs of electrons, as shown in **Table 5** and **Table 6**. In addition, the number of hybrid orbitals formed equals the number of atomic orbitals mixed. For example, $AlCl_3$ has a total of three pairs of electrons and VSEPR predicts a trigonal planar molecular shape. This shape results when one s and two p orbitals on the central atom, Al, mix and form three identical sp^2 hybrid orbitals.

Lone pairs also occupy hybrid orbitals. Compare the hybrid orbitals of $BeCl_2$ in **Table 5** and H_2O in **Table 6**. Both compounds contain three atoms. Why does an H_2O molecule contain sp^3 orbitals? There are two lone pairs on the central oxygen atom in H_2O. Therefore, there must be four hybrid orbitals—two for bonding and two for the lone pairs.

Recall that multiple covalent bonds consist of one sigma bond and one or more pi bonds. Only the two electrons in the sigma bond occupy hybrid orbitals such as sp and sp^2. The remaining unhybridized p orbitals overlap to form pi bonds. It is important to note that single, double, and triple covalent bonds contain only one hybrid orbital. Thus, CO_2, with two double bonds, forms sp hybrid orbitals.

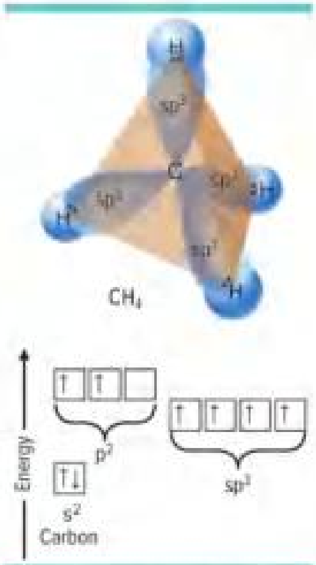


Figure 19 Notice that the hybrid orbitals have an intermediate amount of potential energy when compared with the energy of the original s and p orbitals. According to VSEPR theory, a tetrahedral shape minimizes repulsion between the hybrid orbitals in a CH_4 molecule.

Identify *How many faces does the tetrahedral shape formed by the sp^3 orbitals have?*

Get It?
Describe how repulsion between electric charges at the atomic scale explains the structure of a methane molecule.

Table 5 Molecular Shapes: 2 or 3 Total Pairs

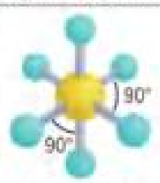
Molecule	Total Pairs	Shared Pairs	Lone Pairs	Hybrid Orbitals	Molecular Shape*
$BeCl_2$	2	2	0	sp	 Linear
$AlCl_3$	3	3	0	sp^2	 Trigonal planar

The $BeCl_2$ molecule contains only two pairs of electrons shared with the central Be atom. These bonding electrons have the maximum separation, a bond angle of 180°, and the molecular shape is linear.

The three bonding electron pairs in $AlCl_3$ have maximum separation in a trigonal planar shape with 120° bond angles.

*Balls represent atoms, sticks represent bonds, and lobes represent lone pairs of electrons.

Table 6 **Molecular Shapes: 4, 5, or 6 Total Pairs**

Molecule	Total Pairs	Shared Pairs	Lone Pairs	Hybrid Orbitals	Molecular Shape*
CH_4	4	4	0	sp^3	 Tetrahedral
NH_3	4	3	1	sp^3	 Trigonal pyramidal
H_2O	4	2	2	sp^3	 Bent
$NbBr_5$	5	5	0	sp^3d	 Trigonal bipyramidal
SF_6	6	6	0	sp^3d^2	 Octahedral

When the central atom in a molecule has four pairs of bonding electrons, as CH_4 does, the shape is tetrahedral. The bond angles are 109.5°.

NH_3 has three single covalent bonds and one lone pair. The lone pair takes up a greater amount of space than the shared pairs. There is stronger repulsion between the lone pair and the bonding pairs than between two bonding pairs. The resulting geometry is trigonal pyramidal, with 107.3° bond angles.

Water has two covalent bonds and two lone pairs. Repulsion between the lone pairs causes the angle to be 104.5°, less than both tetrahedral and trigonal pyramid. As a result, water molecules have a bent shape.

The $NbBr_5$ molecule has five pairs of bonding electrons. The trigonal bipyramidal shape minimizes the repulsion of these shared electron pairs.

As with $NbBr_5$, SF_6 has no unshared electron pairs on the central atom. However, six shared pairs arranged about the central atom result in an octahedral shape.

*Balls represent atoms, sticks represent bonds, and lobes represent lone pairs of electrons.

EXAMPLE Problem 7

FIND THE SHAPE OF A MOLECULE Phosphorus trihydride, a colorless gas, is produced when organic materials, such as fish flesh, rot. What is the shape of a phosphorus trihydride molecule? Predict the bond angle and identify hybrid orbitals.

1 ANALYZE THE PROBLEM

A phosphorus trihydride molecule has three hydrogen atoms bonded to a central phosphorus atom.

EXAMPLE Problem 7 (continued)

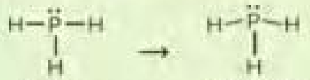
2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons and the number of electron pairs.

$$1\text{ P atom} \times \frac{5\text{ valence electrons}}{1\text{ P atom}} + 3\text{ H atoms} \times \frac{1\text{ valence electron}}{1\text{ H atom}} = 8\text{ valence electrons}$$

$$\frac{8\text{ electrons}}{2\text{ electrons/pair}} = 4\text{ pairs}$$

Determine the total number of bonding pairs.



Lewis structure → **Molecular shape**

Draw the Lewis structure, using one pair of electrons to bond each H atom to the central P atom and assigning the lone pair to the P atom.

The molecular shape is trigonal pyramidal with a predicted 107° bond angle and sp^3 hybrid orbitals.

3 EVALUATE THE ANSWER

All electron pairs are used and each atom has a stable electron configuration.

PRACTICE Problems ADDITIONAL PRACTICE

Determine the molecular shape, bond angle, and hybrid orbitals for each molecule.

56. BF_3 **57.** OCl_2 **58.** BeF_2 **59.** CF_4

60. CHALLENGE For a NH_4^+ ion, identify its molecular shape, bond angle, and hybrid orbitals.

Check Your Progress

Summary

- VSEPR model theory states that electron pairs repel each other and determine both the shape of and bond angles in a molecule.
- Hybridization explains the observed shapes of molecules by the presence of equivalent hybrid orbitals.

Demonstrate Understanding

61. **Summarize** how the VSEPR model helps explain how electric charges affect bonding and molecular shape in covalent compounds.
62. **Define** the term *bond angle.*
63. **Apply** Use the term *hybridization* to describe the bonds in a methane molecule.
64. **Compare** the size of an orbital that has a shared electron pair with one that has a lone pair.
65. **Identify** the type of hybrid orbitals present and bond angles for a molecule with a tetrahedral shape.
66. **Compare** the molecular shapes and hybrid orbitals of PF_3 and PF_5 molecules. Explain why their shapes differ.
67. **List** in a table the Lewis structure, molecular shape, bond angle, and hybrid orbitals for molecules of CS_2, CH_2O, H_2Se, CCl_2F_2, and NCl_3.

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LESSON 5

ELECTRONEGATIVITY AND POLARITY

FOCUS QUESTION

How does molecular shape affect the way that covalent compounds are held together?

Electronegativity and Bond Character

The type of bond formed during a reaction is related to each atom's attraction for electrons. The version of the periodic table of the elements shown in **Figure 20** lists electronegativity values. Recall that electronegativity indicates the relative ability of an atom to attract electrons for bonding. Because noble gases do not generally form compounds, the noble gases are not shown on this table.

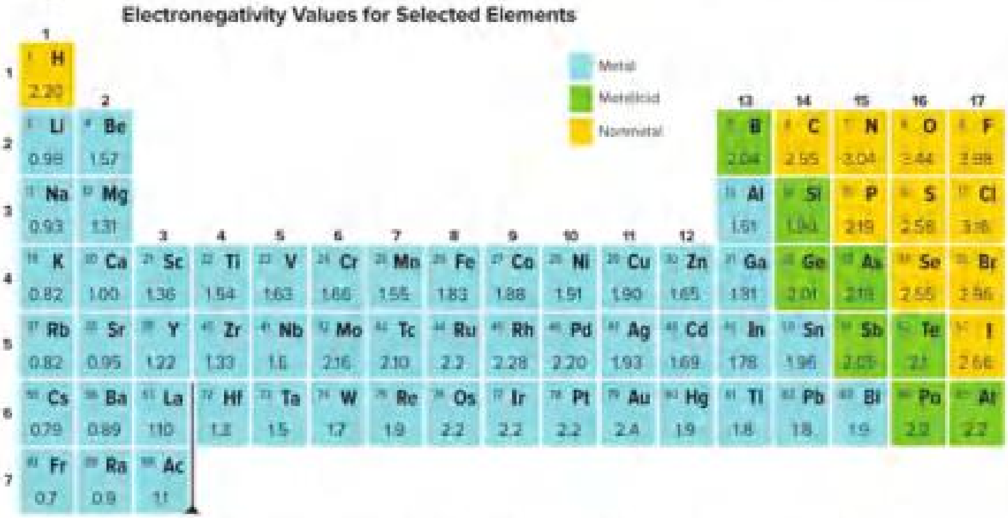


Figure 20 Electronegativity values are derived by comparing an atom's attraction for shared electrons to that of a fluorine atom's attraction for shared electrons. Note, the electonegativity values for the lanthanide and actinide series, which are not shown, range from 1.12 to 1.7.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Communicate Information about Contact and Noncontact Forces

HS-PS2-6. Communicate scientific and technical information about why the molecular-level structure is important in the functioning of designed materials.

Revisit the Encounter the Phenomenon Question

What information from this lesson can help you answer the module question?

Table 7 EN Difference Between Atoms within a Compound and Bond Character

Electronegativity Difference	Bond Character
> 1.7	mostly ionic
$0.4 - 1.7$	polar covalent
< 0.4	mostly covalent
0	nonpolar covalent

Electronegativity

The scale of electronegativities—shown in **Figure 20**—allows chemists to evaluate the affinity of specific atoms in a compound for electrons. Electonegativity is a measure of the tendency of an atom to accept an electron. Excluding noble gases, electronegativity increases with increasing atomic number within a period and decreases with increasing atomic number within a group. This shows that some atoms are more capable of attracting electrons than others. The explanation for why this occurs becomes clear when you consider the number of valence electrons an atom has. As you move from left to right on the table, the number of electrons needed to complete the octet rule becomes less, thereby increasing an atom's affinity for electrons in order to reach a more stable state.

Note that flourine has the greatest electronegativity value (3.98), while francium has the least (0.7). Although xenon is a noble gas, it can sometimes form bonds with highly electonegative atoms, such as flourine.

Get It?

Explain how electric charges at the atomic scale relate to the electronegativity of an atom.

Bond character

A chemical bond between atoms of different elements is never completely ionic or covalent. The character of a bond depends on how strongly each of the bonded atoms attracts electrons. As shown in **Table 7**, the character and type of a chemical bond can be predicted using the electronegativity difference between atoms that bond within a compound. Electrons in bonds between identical atoms have an electronegativity difference of zero—meaning that the electrons are equally shared between the two atoms. This type of bond is considered nonpolar covalent, or a pure covalent bond.

On the other hand, because different elements have different electronegativities, the electron pairs in a covalent bond between different atoms are not shared equally. Unequal sharing results in a **polar covalent bond.** When there is a large difference in the electronegativity between bonded atoms, an electron is transferred from one atom to the other, which results in bonding that is primarily ionic.

Bonding is not often clearly ionic or covalent. An electronegativity difference of 1.70 is considered 50 percent covalent and 50 percent ionic. As the difference in electronegativity increases, the bond becomes more ionic in character. As shown in **Table 7**, ionic bonds form when the electronegativity difference is greater than 1.70. However, this cutoff is sometimes inconsistent with experimental observations of two nonmetals bonding together.

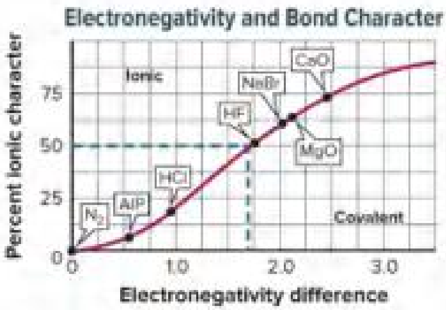


Figure 21 This graph shows that the difference in electronegativity between bonding atoms determines the percent ionic character of the bond. Above 50% ionic character, bonds are mostly ionic.

Figure 21 summarizes the range of chemical bonding between two atoms. What percent ionic character is a bond between two atoms that have an electronegativity difference of 2.00? Where would LiBr be plotted on the graph?

Get It?

Describe how electronegativity relates to polar covalent bonding.

Polar Covalent Bonds

As you just learned, polar covalent bonds form because not all atoms that share electrons attract them equally. A polar covalent bond is similar to a tug-of-war in which the two teams are not of equal strength. Although both sides share the rope, the stronger team pulls more of the rope toward its side. When a polar bond forms, the shared electron pair or pairs are pulled toward one of the atoms. Thus, the electrons spend more time around that atom than the other atom. This results in partial charges at the ends of the bond.

The Greek letter delta (δ) is used to represent a partial charge. In a polar covalent bond, δ^- represents a partial negative charge and δ^+ represents a partial positive charge. As shown in **Figure 22**, δ^- and δ^+ can be added to a molecular model to indicate the polarity of the covalent bond. The more-electronegative atom is at the partially negative end, while the less-electronegative atom is at the partially positive end. The resulting polar bond often is referred to as a dipole (two poles).

Electronegativity	Cl = 3.16
Electronegativity	H = 2.20
Difference	= 0.96

Figure 22 Chlorine's electronegativity is higher than that of hydrogen. Therefore, in a molecule containing hydrogen and chlorine, the shared pair of electrons is with the chlorine atom more often than it is with the hydrogen atom. Symbols are used to indicate the partial charge at each end of the molecule from this unequal sharing of electrons.

Molecular polarity

Covalently bonded molecules are either polar or nonpolar; which type depends on the location and nature of the covalent bonds in the molecule. A distinguishing feature of nonpolar molecules is that they are not attracted by an electric field. Polar molecules, however, are attracted by an electric field. Because polar molecules are dipoles with partially charged ends, they have an uneven electron density. This results in the tendency of polar molecules to align with an electric field.

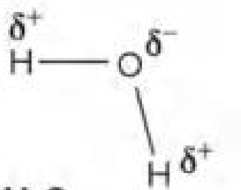
Polarity and molecular shape

You can learn why some molecules are polar and some are not by comparing water (H_2O) and carbon tetrachloride (CCl_4) molecules. Both molecules have polar covalent bonds. According to **Figure 20** earlier in the lesson, the electronegativity difference between a hydrogen atom and an oxygen atom is 1.24. The electronegativity difference between a chlorine atom and a carbon atom is 0.61. Although these electronegativity differences vary, a H–O bond and a C–Cl bond are considered to be polar covalent.

$$\overset{\delta^+}{\mathrm{H}}\!-\!\overset{\delta^-}{\mathrm{O}} \qquad \overset{\delta^+}{\mathrm{C}}\!-\!\overset{\delta^-}{\mathrm{Cl}}$$

According to their molecular formulas, both molecules have more than one polar covalent bond. However, only the water molecule is polar. Why might one molecule with polar covalent bonds be polar, while a second molecule with polar covalent bonds is nonpolar? The answer lies in the shapes of the molecules.

The shape of an H_2O molecule, as determined by VSEPR, is bent because the central oxygen atom has lone pairs of electrons, as shown in **Figure 23**. Because the polar H–O bonds are asymmetric in a water molecule, the molecule has a definite positive end and a definite negative end. Thus, it is polar.



H_2O

Figure 23 The bent shape of a water molecule makes it polar.

STEM CAREER Connection

Surface Chemist

Surface chemistry involves the study of chemical reactions at surfaces and interfaces, usually between two different solid, liquid, or gas phases. Surface chemists are interested in developing new surface treatments, using materials such as polymer films or nanoparticles to solve industrial problems. A bachelor's degree, master's degree, or Ph.D. is needed, as an understanding of chemical synthesis is required.

CCC CROSSCUTTING CONCEPTS

Patterns Use graph paper to construct the graph in **Figure 21**. Calculate the electronegativity differences for the following: PbCl, CH, CO, NaCl, LiI, SiO. Plot the values on the curve to determine how they are bonded. What evidence can you cite to explain the pattern in electronegativity when moving left to right on the periodic table?

A CCl_4 molecule is tetrahedral, and therefore, symmetrical, as shown in **Figure 24**. The electric charge measured at any distance from its center is identical to the charge measured at the same distance to the opposite side. The average center of the negative charge is located on the carbon atom. The positive center is also located on the carbon atom. Because the partial charges are balanced, CCl_4 is a nonpolar molecule. Note that symmetric molecules are usually nonpolar, and molecules that are asymmetric are polar as long as the bond type is polar.

Is the molecule of ammonia (NH_3), shown in **Figure 25**, polar? It has a central nitrogen atom and three terminal hydrogen atoms. Its shape is trigonal pyramidal because of the lone pair of electrons present on the nitrogen atom. Using **Figure 20**, you can find that the electronegativity difference of hydrogen and nitrogen is 0.84, making each N–H bond polar covalent. The charge distribution is unequal because the molecule is asymmetric. Thus, the molecule is polar.

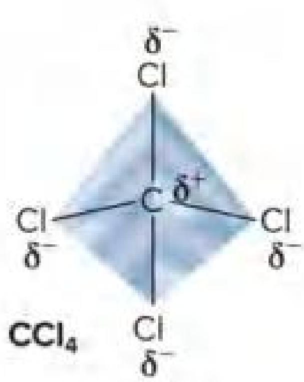


Figure 24 The symmetry of a CCl_4 molecule results in an equal distribution of charge, and the molecule is nonpolar.

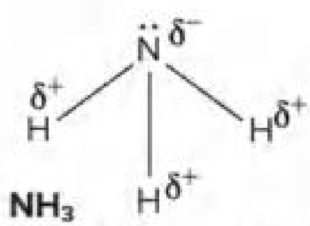


Figure 25 The asymmetric shape of an ammonia molecule results in an unequal charge distribution, and the molecule is polar.

Properties of Covalent Compounds

Table salt, an ionic solid, and table sugar, a covalent solid, are similar in appearance. However, these compounds behave differently when heated. Salt does not melt, but sugar melts at a relatively low temperature. Does the type of bonding in a compound affect its properties?

Intermolecular forces

Differences in properties of atoms are a result of differences in both attractive and repulsive forces. In a covalent compound, the covalent bonds between atoms in molecules are strong, but the attraction forces between molecules are relatively weak. These weak attraction forces are known as intermolecular forces, or van der Waals forces. Intermolecular forces vary in strength but are weaker than the bonds that join atoms in a molecule or those that join ions in an ionic compound.

Get It?

Distinguish covalent bonds from van der Waals forces.

Van der Waals forces, unlike ionic or covalent bonds, are distance-dependent, meaning that as the distance between interacting molecules increases, the van der Waals forces quickly vanish and are no longer present.

Figure 26 Symmetric covalent molecules, such as oil and most petroleum products, are nonpolar. Asymmetric molecules, such as water, are usually polar. As shown in this photo, polar and nonpolar substances usually do not mix.

Infer *Will water alone clean oil from a fabric?*

There are different types of intermolecular forces. Between nonpolar molecules, the force is weak and is called a dispersion force, or induced dipole. The force between oppositely charged ends of two polar molecules is called a dipole-dipole force. The more polar the molecule, the stronger the dipole-dipole force. The third force, a hydrogen bond, is especially strong. It forms between the hydrogen end of one dipole and a fluorine, oxygen, or nitrogen atom on another dipole. You will study intermolecular forces in more detail when you study states of matter.

Get It?

Compare dispersion forces, dipole-dipole forces, and hydrogen bonds.

Solubility of polar molecules

The physical property known as solubility is the ability of a substance to dissolve in another substance. The bond type and the shape of the molecules present determine solubility. Polar molecules and ionic compounds are usually soluble in polar substances, but nonpolar molecules dissolve only in nonpolar substances, as shown in **Figure 26**.

Forces and properties

The properties of covalent molecular compounds are related to the relatively weak intermolecular forces holding the molecules together. These weak forces result in the relatively low melting and boiling points of molecular substances compared with those of ionic substances. That is why, when heated moderately, sugar melts but salt does not.

Weak intermolecular forces also explain why many molecular substances exist as gases or vaporize readily at room temperature. Oxygen (O_2), carbon dioxide (CO_2), and hydrogen sulfide (H_2S) are examples of covalent compounds that are gases at room temperature. Because the hardness of a substance depends on the intermolecular forces between individual molecules, many covalent molecules are relatively soft solids. Paraffin, found in candles and other products, is a common example of a covalent solid.

In the solid phase, molecules align to form a crystal lattice. This molecular lattice is similar to that of an ionic solid, but with less attraction between particles. The structure of the lattice is affected by molecular shape and the type of intermolecular force. Most molecular information has been determined by studying molecular solids.

Covalent Network Solids

There are some solids, often called covalent network solids, that are composed only of atoms interconnected by a network of covalent bonds. Quartz and diamond are two common examples of network solids.

In contrast to molecular solids, network solids are typically brittle, nonconductors of heat or electricity, and extremely hard. Analyzing the structure of a diamond explains some of its properties. In a diamond, each carbon atom is bonded to four other carbon atoms. This tetrahedral arrangement, which is shown in **Figure 27**, forms a strongly bonded crystal system that is extremely hard and has a very high melting point.

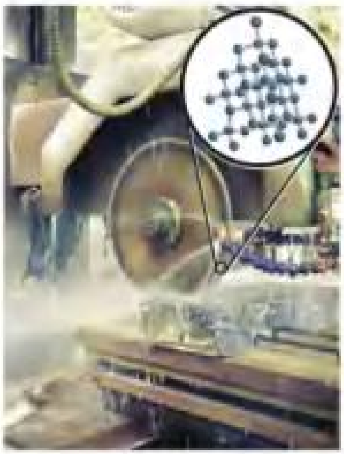


Figure 27 Network solids are often used in cutting tools because of their extreme hardness. Here, a diamond-tipped saw blade cuts through stone.

Check Your Progress

Summary

- Attraction and repulsion of electrical charges at the atomic scale explain the electronegativity, bonding, and polar characteristics of compounds.
- The electronegativity difference between atoms in a compound determines the character of a bond between atoms.
- Polar bonds occur when electrons are not shared equally, forming a dipole.
- The spatial arrangement of polar bonds in a molecule determines the overall polarity of a molecule.
- Molecules attract each other by weak intermolecular forces. In a covalent network solid, each atom is covalently bonded to many other atoms.

Demonstrate Understanding

68. **Summarize** how electonegativity relates to bonding and what pattern of electronegativity is observed in elements as you move from left to right on the periodic table.
69. **Describe** a polar covalent bond.
70. **Describe** a polar molecule.
71. **List** three properties of a covalent compound in the solid phase.
72. **Categorize** bond types using electronegativity differences.
73. **Generalize** Describe the general characteristics of covalent network solids.
74. **Predict** the type of bond that will form between the following pair of atoms:
    a. H and S b. C and H c. Na and S
75. **Identify** each molecule as polar or nonpolar: SCl_2, CS_2, and CF_4.
76. **Determine** whether a compound made of hydrogen and sulfur atoms is polar or nonpolar.
77. **Draw** the Lewis structures for the molecules SF_4 and SF_6. Analyze each structure to determine whether the molecule is polar or nonpolar.

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SCIENCE & SOCIETY

Plastics: The Good, The Bad, and the Ugly

Plastic's properties make it a very versatile material. Its chemical structure makes it strong, lightweight, and easy to produce. However, because plastic doesn't decompose very easily, it's causing some very worrying effects on the environment.

Large patches of plastic waste and microplastics float in Earth's oceans.

Past, Present, and Future

The first completely synthetic plastic, Bakelite, was invented in 1907. By the 1950s, plastic could be found in almost every room in an American house, from plastic-wrapped baked goods in the kitchen to plastic-coated children's toys in the nursery.

Plastic's properties are due to its structure. Plastic is formed when carbon monomers covalently bond in long chains. The resulting material is easily moldable and durable. However, the arrangement of bonds in synthetic plastic polymers makes them difficult to be broken by bacteria and other decomposers. The result is that most plastics aren't biodegradable.

Many plastics are recyclable, but the processes used to recycle plastics are expensive. As a result, much of the world's plastic doesn't make it into the recycling bin, and instead finds its way into the environment, carried by runoff into the ocean.

The problem of plastic pollution isn't limited to unsightly bags and bottles lying on the beach. Once in the environment, plastic can be broken down by sunlight into smaller pieces called microplastics. After they find their way into an ocean ecosystem, microplastics have long-reaching effects on marine life.

Thanks to education and advocacy groups, the public is paying attention to this problem. Some communities have recycling processes that are easier for the public to use, and many are using less plastic. Some scientists are concentrating on creating plant-based plastics that are more biodegradable, while others are studying bacteria and fungi that are able to use plastic for energy. The key to solving the problem with plastics will be to find a balance between making plastics work for both people and the environment.

EVALUATE A DESIGN SOLUTION

Research one way in which scientists are trying to solve the problem of plastic in the environment. Create an infographic or other visual display about how this solution might work in your community.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 THE COVALENT BOND

- Covalent bonds form when atoms share one or more pairs of electrons.
- Sharing one pair, two pairs, and three pairs of electrons forms single, double, and triple covalent bonds, respectively.
- Orbitals overlap directly in sigma bonds. Parallel orbitals overlap in pi bonds. A single covalent bond is a sigma bond, but multiple covalent bonds are made of both sigma and pi bonds.
- Bond dissociation energy is needed to break a covalent bond.

- covalent bond
- molecule
- Lewis structure
- sigma bond
- pi bond
- endothermic reaction
- exothermic reaction

Lesson 2 NAMING MOLECULES

- Names of covalent molecular compounds include prefixes for the number of each atom present. The final letter of the prefix is dropped if the element name begins with a vowel.
- Molecules that produce H^+ in solution are acids. Binary acids contain hydrogen and one other element. Oxyacids contain hydrogen and an oxyanion.

- oxyacid

Lesson 3 MOLECULAR STRUCTURES

- Different models can be used to represent molecules.
- Resonance occurs when more than one valid Lewis structure exists for the same molecule.
- Exceptions to the octet rule occur in some molecules.

- structural formula
- resonance
- coordinate covalent bond

Lesson 4 MOLECULAR SHAPES

- VSEPR model theory states that electron pairs repel each other and determine both the shape of and bond angles in a molecule.
- Hybridization explains the observed shapes of molecules by the presence of equivalent hybrid orbitals.

- VSEPR model
- hybridization

Lesson 5 ELECTRONEGATIVITY AND POLARITY

- The electronegativity difference determines the character of a bond between atoms.
- Polar bonds occur when electrons are not shared equally, forming a dipole.
- The spatial arrangement of polar bonds in a molecule determines the overall polarity of a molecule.
- Molecules attract each other by weak intermolecular forces. In a covalent network solid, each atom is covalently bonded to many other atoms.

- polar covalent bond

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

Why does water expand when it freezes?

CER Claim, Evidence, Reasoning

Explain your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

Based on Real Data*

SEP Data Analysis Lab

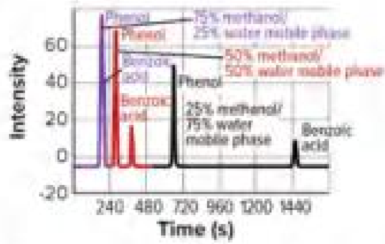
How does the polarity of the mobile phase affect chromatograms?

High-performance liquid chromatography, or HPLC, is used by analytical chemists to separate mixtures of solutes. During HPLC, components that are strongly attracted to the extracting solvent are retained longer by the moving phase and tend to appear early on a chromatograph. Several scientists performed HPLC using a methanol-water mixture as the extracting solvent to separate a phenol-benzoic acid mixture. Their results are shown in the graph. The peak areas on the chromatograph indicate the amount of each component present in the mixture.

CER Analyze and Interpret Data

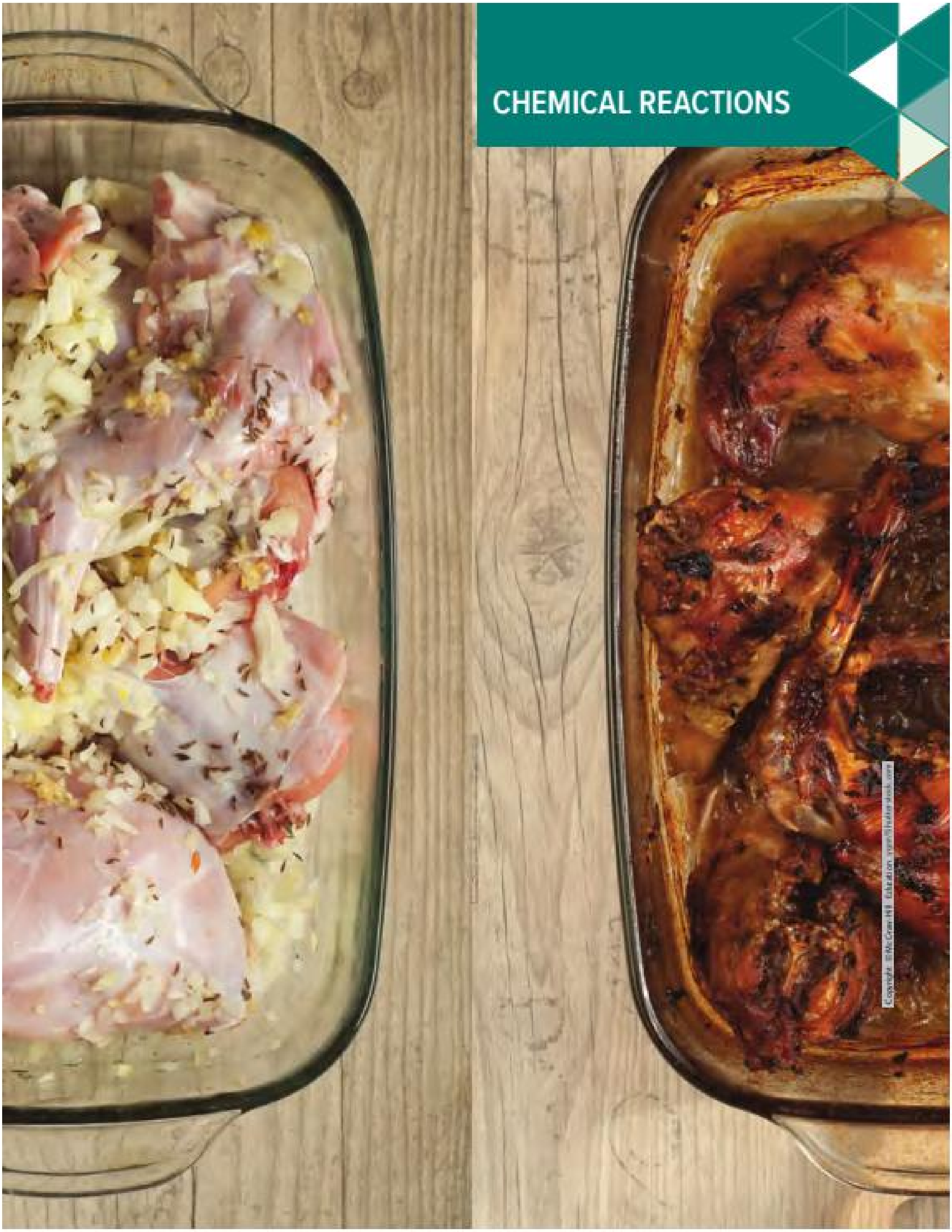
1. **Explain** the different retention times shown on the chromatograms.
2. **Infer** from the graph the component, phenol or benzoic acid, that is in excess. Explain your answer.
3. **Infer** which component of the mixture has more polar molecules.
4. **Determine** the most effective composition of the mobile phase (of those tested) for separating phenol from benzoic acid. Explain.

Data and Observations



*Data obtained from: Joseph, Seema M. and Palasota, John A. 2001. The combined effects of pH and percent methanol on the HPLC separation of benzoic acid and phenol. Journal of Chemical Education 78:1381.

CHEMICAL REACTIONS



CHEMICAL REACTIONS

ENCOUNTER THE PHENOMENON

What happens to food when you cook it?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

CER Claim, Evidence, Reasoning

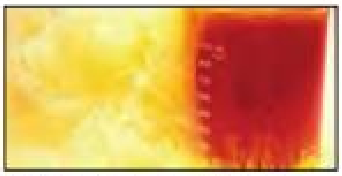
Make Your Claim Use your CER chart to make a claim about what happens to food when you cook it.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.

LESSON 1: Explore & Explain: Evidence of Chemical Reactions



LESSON 1: Explore & Explain: Representing Chemical Reactions

LESSON 1

REACTIONS AND EQUATIONS

FOCUS QUESTION

How are chemical reactions modeled?

Chemical Reactions

Do you know that the foods you eat, the fibers in your clothes, and the plastic in your CDs have something in common? Foods, fibers, and plastics are produced when the atoms in substances are rearranged to form different substances. Rearrangments of atoms are happening all the time. Atoms are rearranged during the cooking of food, inside the bodies of living things, in vehicle engines, in factories, and in the atmosphere.

The process by which the atoms of one or more substances are rearranged to form different substances is called a **chemical reaction.** A chemical reaction is another name for a chemical change, which you read about previously.

Chemical reactions affect every part of your life. They break down your food, producing the energy you need to live. Chemical reactions in the engines of cars and buses provide the energy to power the vehicles. They produce natural fibers, such as cotton and wool, in plants and animals. In factories, they produce synthetic fibers such as nylon, which is shown in **Figure 1.** Nylon is used in many familiar products, including carpeting, clothing, sports equipment, and tires.

Figure 1 When adipoyl chloride in dichloromethane reacts with hexanediamine, nylon is formed.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Inquiry into Chemistry: Solve It: Mystery of the Moonlight Ride
Obtain information on a water decomposition reaction and describe its structure and function as it relates to fuel.

Virtual Investigation: Balancing Chemical Equations
Use mathematics and computational thinking to balance chemical equations.

Figure 2 Each of these photos illustrates evidence of a chemical reaction.

Evidence of a chemical reaction

Although some chemical reactions are hard to detect, many reactions provide physical evidence that they have occurred. A temperature change can indicate a chemical reaction. Many reactions, such as those that occur during the burning of wood, release energy in the form of heat and light. Other chemical reactions absorb heat.

In addition to a temperature change, color change can indicate that a chemical reaction has occurred. For example, you might have noticed that the color of some nails that are left outside changes from silver to orange-brown in a short time. The color change is evidence that a chemical reaction occurred between the iron in the nail and the oxygen in air. Odor, gas bubbles, and the formation of a solid are other indications of chemical change. Each of the photographs in **Figure 2** shows evidence of a chemical reaction.

Get It?

Cite Evidence Describe any evidence of a chemical reaction you note in the photo at the beginning of this module.

Representing Chemical Reactions

Chemists use statements called equations to represent chemical reactions. Equations show a reaction's **reactants,** which are the starting substances, and **products,** which are the substances formed during the reaction. Chemical equations do not express numerical equalities as mathematical equations do because during chemical reactions the reactants are used up as the products form. Instead, the equations used by chemists show the direction in which the reaction progresses. Therefore, an arrow rather than an equal sign is used to separate the reactants from the products. You read the arrow as react to produce or yield. The reactants are written to the left of the arrow, and the products are written to the right of the arrow. When there are two or more reactants, or when there are two or more products, a plus sign separates each reactant or each product. These elements of equation notation are shown below.

$$\text{Reactant 1} + \text{Reactant 2} \rightarrow \text{Product 1} + \text{Product 2}$$

Symbols for equations

You have already seen that the plus symbol is used to separate two or more reactants or products on each side of an equation, and that an arrow is used to separate the reactants from the products. When a reaction is reversible, a two-way arrow is used to show that the reaction can proceed in either direction.

As well, symbols are used to show the physical states of the reactants and products. Reactants and products can exist as solids, liquids, and gases. When they are dissolved in water, they are said to be aqueous. It is important to show the physical states of a reaction's reactants and products in an equation because the physical states provide clues about how the reaction occurs. Some basic symbols used in equations are shown in **Table 1**.

Table 1
Symbols Used in Equations

Symbol	Purpose
+	separates two or more reactants or products
→	separates reactants from products
⇌	separates reactants from products and indicates a reversible reaction
(s)	identifies a solid state
(l)	identifies a liquid state
(g)	identifies a gaseous state
(aq)	identifies a water solution

Word equations

You can use statements called word equations to indicate the reactants and products of chemical reactions. The word equation below describes the reaction between aluminum (Al) and bromine (Br), which is shown in **Figure 3**. Aluminum is a solid, and bromine is a liquid. The brownish-red cloud in the photograph is excess bromine. The reaction's product, which is solid particles of aluminum bromide ($AlBr_3$), settles on the bottom of the beaker.

$$\text{Reactant 1} + \text{Reactant 2} \rightarrow \text{Product 1}$$

$$\text{aluminum(s)} + \text{bromine(l)} \rightarrow \text{aluminum bromide(s)}$$

The word equation reads, "Aluminum and bromine react to produce aluminum bromide."

Figure 3 Science, like all other disciplines, has a specialized language that allows specific information to be communicated in a uniform manner. This reaction between aluminum and bromine can be described by a word equation, a skeleton equation, or a balanced chemical equation.

Skeleton equations

Although word equations help to describe chemical reactions, they lack important information. A skeleton equation uses chemical formulas rather than words to identify the reactants and the products. For example, the skeleton equation for the reaction between aluminum and bromine uses the formulas for aluminum, bromine, and aluminum bromide in place of words.

$$Al(s) + Br_2(l) \rightarrow AlBr_3(s)$$

How would you write the skeleton equation that describes the reaction between carbon and sulfur to form carbon disulfide? Carbon and sulfur are solids. First, write the chemical formulas for each of the reactants to the left of the arrow. Then, separate the reactants with a plus sign. Add symbols to each formula to indicate the physical states of the compounds.

$$C(s) + S(s) \rightarrow$$

Finally, write the chemical formula for the product, liquid carbon disulfide, to the right of the arrow and indicate its physical state. The result is the skeleton equation for the reaction.

$$C(s) + S(s) \rightarrow CS_2(l)$$

This skeleton equation tells us that carbon in the solid state reacts with sulfur in the solid state to produce carbon disulfide in the liquid state. The skeleton equation does not, however, indicate the amounts of each compound involved in the reaction.

Get It?

Compare What information do skeleton equations provide that word equations lack?

PRACTICE Problems

ADDITIONAL PRACTICE

Write skeleton equations for the following word equations.

1. Hydrogen and bromine gases react to yield hydrogen bromide.

   hydrogen(g) + bromine(g) → hydrogen bromide(g)

2. When carbon monoxide and oxygen react, carbon dioxide forms.

   carbon monoxide(g) + oxygen(g) → carbon dioxide(g)

3. **CHALLENGE** Write the word equation and the skeleton equation for the following reaction: when heated, solid potassium chlorate yields solid potassium chloride and oxygen gas.

ACADEMIC VOCABULARY

formula

an expression using chemical symbols to represent a substance

The chemical formula for water is H_2O.

CCC CROSSCUTTING CONCEPTS

Energy and Matter The total amount of matter is conserved in a chemical reaction. Write a paragraph that argues that a skeleton equation does not provide enough information to show this to be true. Include evidence to support your argument.

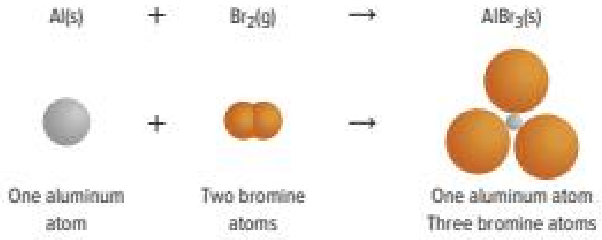


Figure 4 The information conveyed by skeleton equations is limited. In this case, the skeleton equation is correct, but it does not show the exact number of atoms that interact. Refer to **Table R-1** in the Student Resources for a key to atom color conventions.

Chemical equations

Like word equations, skeleton equations lack some information about reactions. Recall that the law of conservation of mass states that in a chemical change, matter is neither created nor destroyed. Chemical equations must show that atoms are conserved during a reaction. Skeleton equations lack that information.

Look at **Figure 4**. The skeleton equation for the reaction between aluminum and bromine shows that one aluminum atom and two bromine atoms react to produce a substance containing one aluminum atom and three bromine atoms. Was a bromine atom created in the reaction? Atoms are not created in chemical reactions, so more information is needed. To accurately represent a chemical reaction using an equation, the equation must show equal numbers of atoms of each reactant and each product on both sides of the reaction arrow. Such an equation is called a balanced chemical equation. A **chemical equation** is a statement that uses chemical formulas to show the identities and relative amounts of the substances involved in a chemical reaction.

Balancing Chemical Equations

The balanced equation for the reaction between aluminum and bromine is shown in **Figure 5**. To balance an equation, you must find the correct coefficients for the chemical formulas in the skeleton equation. A **coefficient** in a chemical equation is the number written in front of a reactant or product. Coefficients are usually whole numbers and are not usually written if the value is one. The coefficients in a balanced equation describe the lowest whole number ratio of the amounts of all of the reactants and products. **Table 2** shows the steps for balancing a chemical equation.

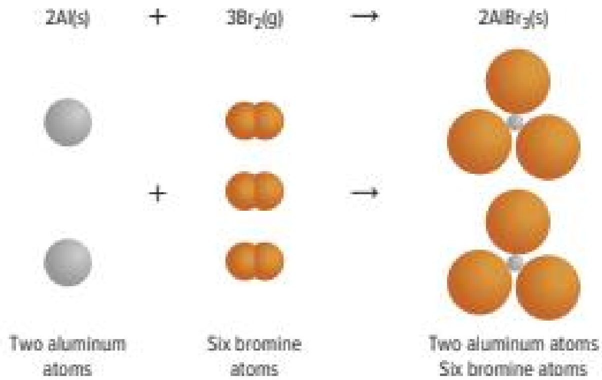


Figure 5 In a balanced chemical equation, the number of particles on the reactant side of the equation equals the number of particles on the product side of the equation. In this case, two aluminum atoms and six bromine atoms are needed on both sides of the equation.

Conservation of mass A fundamental concept of chemistry is the law of conservation of mass. All chemical reactions obey the law that matter, including atoms, is neither created nor destroyed. Therefore, it is important that chemical equations are balanced to show that reactions obey this law. The fact that atoms are conserved is used to describe and balance chemical equations, as shown in **Table 2**.

Table 2 Steps for Balancing Equations

Step	Process	Example
1	*Write the skeleton equation for the reaction.* Make sure that the chemical formulas correctly represent the substances. An arrow separates the reactants from the products, and a plus sign separates multiple reactants and products. Show the physical states of all reactants and products.	$H_2(g)$ + $Cl_2(g)$ → $HCl(g)$  Two hydrogen atoms + Two chlorine atoms → One hydrogen atom One chlorine atom
2	*Count the atoms of the elements in the reactants.* If a reaction involves identical polyatomic ions in the reactants and products, count each polyatomic ion as a single element. This reaction does not involve any polyatomic ions. Two atoms of hydrogen and two atoms of chlorine are reacting.	H_2 + Cl_2 → 2 atoms H 2 atoms Cl
3	*Count the atoms of the elements in the products.* One atom of hydrogen and one atom of chlorine are produced.	HCl 1 atom H + 1 atom Cl
4	*Change the coefficients to make the number of atoms of each element equal on both sides of the equation, showing that atoms are conserved.* Never change a subscript in a chemical formula to balance an equation because doing so changes the identity of the substance.	H_2 + Cl_2 → 2HCl 2 atoms H, 2 atoms Cl, 2 atoms H + 2 atoms Cl  Two hydrogen atoms + Two chlorine atoms → Two hydrogen atoms Two chlorine atoms
5	*Write the coefficients in their lowest possible ratio.* The coefficients should be the smallest possible whole numbers. The ratio 1 hydrogen to 1 chlorine to 2 hydrogen chloride (1:1:2) is the lowest-possible ratio because the coefficients cannot be reduced further and still remain whole numbers.	$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ 1:1:2 1 H_2 to 1 Cl_2 to 2 HCl
6	*Check your work.* Make sure that the chemical formulas are written correctly. Then, check that the number of atoms of each element is equal on both sides of the equation.	H_2 + Cl_2 → 2HCl 2 atoms H, 2 atoms Cl, 2 atoms H + 2 atoms Cl There are two hydrogen atoms and two chlorine atoms on both sides of the equation.

EXAMPLE Problem 1

WRITING A BALANCED CHEMICAL EQUATION Write the balanced chemical equation for the reaction in which aqueous sodium hydroxide and aqueous calcium bromide react to produce solid calcium hydroxide and aqueous sodium bromide.

1 ANALYZE THE PROBLEM

You are given the reactants and products in a chemical reaction. Start with a skeleton equation, and use the steps given in **Table 2** for balancing chemical equations.

2 SOLVE FOR THE UNKNOWN

Write the skeleton equation for the chemical reaction. Be sure to put the reactants on the left side of the arrow and the products on the right. Separate the substances with plus signs, and indicate their physical states.

$$NaOH(aq) + CaBr_2(aq) \rightarrow Ca(OH)_2(s) + NaBr(aq)$$

1 Na, 1 O, 1 H, 1 Ca, 2 Br	Count the atoms of each element in the reactants.
1 Na, 2 O, 2 H, 1 Ca, 1 Br	Count the atoms of each element in the products.
$2NaOH + CaBr_2 \rightarrow Ca(OH)_2 + NaBr$	Insert the coefficient 2 in front of NaOH to balance the hydroxide ions.
$2NaOH + CaBr_2 \rightarrow Ca(OH)_2 + 2NaBr$	Insert the coefficient 2 in front of NaBr to balance the Na and Br atoms.
The ratio of the coefficients is 2:1:1:2.	Write the coefficients in their lowest-possible ratio.
Reactants: 2 Na, 2 OH, 1 Ca, 2 Br Products: 2 Na, 2 OH, 1 Ca, 2 Br	Check to make sure that the number of atoms of each element is equal on both sides of the equation.

3 EVALUATE THE ANSWER

The chemical formulas for all substances are written correctly.

The number of atoms of each element is equal on both sides of the equation. The coefficients are written in the lowest possible ratio. The balanced chemical equation for the reaction is

$$2NaOH(aq) + CaBr_2(aq) \rightarrow Ca(OH)_2(s) + 2NaBr(aq)$$

Real-World Chemistry

Calcium Hydroxide

REEF AQUARIUMS An aqueous solution of calcium hydroxide is used in reef aquariums to provide calcium for animals such as snails and corals. Calcium hydroxide reacts with the carbon dioxide in the water to produce calcium and bicarbonate ions. Reef animals use the calcium to grow shells and strong skeletal systems.

PRACTICE Problems

ADDITIONAL PRACTICE

Write chemical equations for each of the following reactions.

4. In water, iron(III) chloride reacts with sodium hydroxide, producing solid iron(III) hydroxide and sodium chloride.
5. Liquid carbon disulfide reacts with oxygen gas, producing carbon dioxide gas and sulfur dioxide gas.
6. **CHALLENGE** A piece of zinc metal is added to a solution of dihydrogen sulfate. This reaction produces a gas and a solution of zinc sulfate.

Balancing Chemical Equations

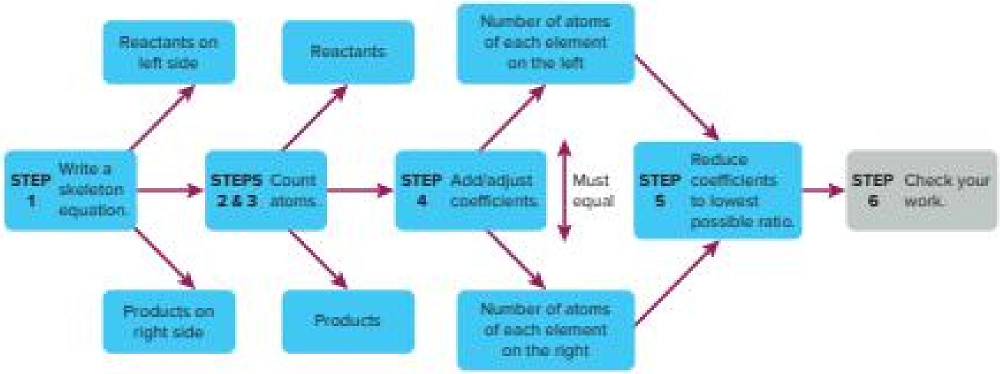


Figure 6 Use this flowchart to help you master the skill of balancing equations.

Figure 6 summarizes the steps for balancing equations. Most chemical equations can be balanced by the process you learned in this lesson.

Check Your Progress

Summary

- Some physical changes are evidence that indicate a chemical reaction has occurred.
- Word equations and skeleton equations provide important information about a chemical reaction.
- A chemical equation gives the identities and relative amounts of the reactants and products that are involved in a chemical reaction.
- Balancing an equation involves adjusting the coefficients until the number of atoms of each element is equal on both sides of the equation.

Demonstrate Understanding

7. **Explain** why it is important that a chemical equation be balanced in terms of the fact that atoms are conserved.
8. **List** three types of physical evidence that indicate a chemical reaction has occurred.
9. **Compare and contrast** a skeleton equation and a chemical equation.
10. **Explain** why it is important to reduce coefficients in a balanced equation to the lowest-possible whole-number ratio.
11. **Analyze** When balancing a chemical equation, can you adjust the subscript in a formula? Explain.
12. **Assess** Is the following equation balanced? If not, correct the coefficients to balance the equation.
    $2K_2CrO_4(aq) + Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + PbCrO_4(s)$
13. **Evaluate** Aqueous phosphoric acid and aqueous calcium hydroxide react to form solid calcium phosphate and water. Write a balanced chemical equation for this reaction.

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

LESSON 2

CLASSIFYING CHEMICAL REACTIONS

FOCUS QUESTION

What are the different types of chemical reactions?

Types of Chemical Reactions

The fact that atoms are conserved, together with knowledge of the chemical properties of the elements involved, can be used to describe and predict chemical reactions. Chemists classify chemical reactions into several categories. Knowing the categories of chemical reactions can help you remember and understand them. It can also help you recognize patterns and predict the products of many reactions.

Chemists distinguish among four reaction types: synthesis, combustion, decomposition, and replacement reactions. By analyzing and comparing the reactants and products of a variety of chemical reactions, you will notice patterns that will help you to classify them. Note, however, that some reactions fit into more than one of these types.

Synthesis Reactions

In **Figure 7**, two sodium atoms react with a molecule of chlorine to produce sodium chloride. This reaction is a **synthesis reaction**—a chemical reaction in which two or more substances (A and B) react to produce a single product (AB).

$$A + B \rightarrow AB$$

When two elements react, the reaction is always a synthesis reaction.

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

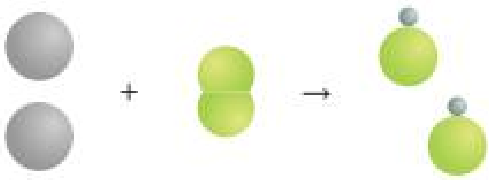


Figure 7 In this synthesis reaction, two elements, sodium and chlorine, react to produce one compound, sodium chloride.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

ChemLAB: Develop an Activity Series
Obtain and evaluate information on the effects of metal reactivity to predict chemical reactions.

Laboratory: Single-Replacement Reactions
Obtain and evaluate information on the effects of a single replacement reaction to understand the conservation of atoms.

Two compounds can also combine to form one compound. For example, the reaction between calcium oxide (CaO) and water (H_2O) to form calcium hydroxide ($Ca(OH)_2$) is a synthesis reaction.

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

Another type of synthesis reaction involves a reaction between a compound and an element, as happens when sulfur dioxide gas (SO_2) reacts with oxygen gas (O_2) to form sulfur trioxide (SO_3).

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

Combustion Reactions

The synthesis reaction between sulfur dioxide and oxygen can also be classified as a **combustion reaction.** In a combustion reaction, such as the one shown in **Figure 8**, oxygen combines with a substance and releases energy in the form of heat and light. Oxygen can combine in this way with many different substances, making combustion reactions common.

A combustion reaction occurs between hydrogen and oxygen when hydrogen is heated, as illustrated in **Figure 9**. Water is formed during the reaction, and a large amount of energy is released. Another important combustion reaction occurs when coal is burned to produce energy. Coal is called a fossil fuel because it contains the remains of plants that lived long ago. It is composed primarily of the element carbon. Coal-burning power plants generate electric power in many parts of the United States. The primary reaction that occurs in these plants is between carbon and oxygen.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Figure 8 The light produced by a sparkler is the result of a combustion reaction between oxygen and different metals.

Figure 9 During a combustion reaction between oxygen and hydrogen, water is formed.

WORD ORIGIN

combustion
comes from the Latin word *comburere*, meaning *to burn*

CCC CROSSCUTTING CONCEPTS

Patterns Study the different types of chemical reactions in this lesson. Then create a poster that identifies the different patterns observed for the reaction types. What evidence supports the patterns you used in your poster?

Note that the combustion reactions just mentioned are also synthesis reactions. However, not all combustion reactions are synthesis reactions. For example, the reaction involving methane gas (CH_4) and oxygen illustrates a combustion reaction in which one substance replaces another in the formation of products.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Methane, which belongs to a group of substances called hydrocarbons, is the major component of natural gas. A fireplace that uses natural gas as fuel is shown in **Figure 10.** All hydrocarbons contain carbon and hydrogen and burn in oxygen to yield carbon dioxide and water.

Figure 10 The combustion of natural gas in this fireplace provides warmth and light.

PRACTICE Problems ADDITIONAL PRACTICE

Write chemical equations for the following reactions. Classify each reaction into as many categories as possible.

14. The solids aluminum and sulfur react to produce aluminum sulfide.

15. Water and dinitrogen pentoxide gas react to produce aqueous hydrogen nitrate.

16. The gases nitrogen dioxide and oxygen react to produce dinitrogen pentoxide gas.

17. CHALLENGE Sulfuric acid (H_2SO_4) and sodium hydroxide solutions react to produce aqueous sodium sulfate and water.

Decomposition Reactions

Some chemical reactions are essentially the opposite of synthesis reactions. These reactions are classified as **decomposition reactions.** A decomposition reaction is one in which a single compound breaks down into two or more elements or new compounds. In generic terms, decomposition reactions can be represented as follows.

$$AB \rightarrow A + B$$

Decomposition reactions often require an energy source, such as heat, light, or electricity, to occur. For example, ammonium nitrate breaks down into dinitrogen monoxide and water when the reactant is heated to a high temperature.

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(g)$$

Notice that this decomposition reaction involves one reactant compound breaking down into two product compounds. Other decomposition reactions involve one reactant compound breaking down into more than two products. The products of a decomposition reaction may be elements, compounds, or one or more of each.

Figure 11 The decomposition of sodium azide, which produces a gas, is the chemical reaction that inflates air bags.

The outcome of another decomposition reaction is shown in **Figure 11**. Automobile safety air bags inflate rapidly as sodium azide pellets decompose. A device that can provide an electric signal to start the reaction is packaged inside air bags along with the sodium azide pellets. When the device is activated, sodium azide decomposes, producing nitrogen gas that quickly inflates the air bag.

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

Notice that the decomposition of sodium azide produces sodium metal in addition to the harmless gas nitrogen. Sodium metal is highly reactive and caustic and therefore poses a safety concern when the air bag deploys. Designers overcame this problem by adding iron(III) oxide. The sodium reacts with the iron(III) oxide to produce sodium oxide, which then reacts with carbon dioxide and water vapour in the air to produce the much safer substance sodium hydrogen carbonate, also known as baking soda. These reactions all take place very quickly.

PRACTICE Problems ADDITIONAL PRACTICE

Write chemical equations for the following decomposition reactions.

18. Aluminum oxide(s) decomposes when electricity passes through it.

19. Nickel(II) hydroxide(s) decomposes to produce nickel(II) oxide(s) and water.

20. CHALLENGE Heating sodium hydrogen carbonate(s) produces sodium carbonate(aq) and water. Carbon dioxide gas is also produced.

Replacement Reactions

In contrast to synthesis, combustion, and decomposition reactions, many chemical reactions are replacement reactions and involve the replacement of an element in a compound. These replacement reactions are also known as displacement reactions. There are two types of replacement reactions: single-replacement reactions and double-replacement reactions.

Single-replacement reactions

The reaction between lithium and water is shown in **Figure 12**. The following chemical equation shows that a lithium atom replaces one of the hydrogen atoms in a water molecule.

$$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$$

A reaction in which the atoms of one element replace the atoms of another element in a compound is called a **single-replacement reaction.** The following generic equation can be used to represent single-replacement reactions such as the reaction of lithium with water to form lithium hydroxide and hydrogen.

$$A + BX \rightarrow AX + B$$

Get It?

Analyze In the reaction between lithium and water, which element replaces hydrogen in water?

Lithium + Water | Copper + Silver Nitrate

Figure 12 In a single-replacement reaction, the atoms of one element replace the atoms of another element in a compound.

Metal replaces hydrogen or another metal The reaction between lithium and water is one type of single-replacement reaction, in which a metal replaces a hydrogen atom in a water molecule. Another type of single-replacement reaction occurs when one metal replaces another metal in a compound dissolved in water. **Figure 12** shows a single-replacement reaction occurring when copper wire is placed in aqueous silver nitrate. The crystals that are accumulating on the copper bar are the silver atoms that the copper atoms replaced.

$$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$$

A metal will not always replace another metal in a compound dissolved in water because metals differ in their reactivities. Reactivity is the ability to react with another substance. An activity series of some metals is shown in **Figure 13**. This series orders metals by reactivity with other metals. Single-replacement reactions are used to determine a metal's position on the list. The most active metals are at the top of the list. The least active metals are at the bottom. Similarly, the reactivity of each halogen has been determined and listed, as shown in **Figure 13**.

Most active

METALS
Lithium
Rubidium
Potassium
Calcium
Sodium
Magnesium
Aluminum
Manganese
Zinc
Iron
Nickel
Tin
Lead
Copper
Silver
Platinum
Gold

Least active

Most active

HALOGENS
Fluorine
Chlorine
Bromine
Iodine

Least active

Figure 13 An activity series is a useful tool for determining the result of a single-replacement reaction.

You can use the activity series to predict whether or not certain reactions will occur. A specific metal can replace any metal listed below it that is in a compound. It cannot replace any metal listed above it. For example, copper atoms replace silver atoms in a solution of silver nitrate. However, if you place a silver wire in aqueous copper(II) nitrate, the silver atoms will not replace the copper. Silver is listed below copper in the activity series, so no reaction occurs. The letters NR (no reaction) are commonly used to indicate that a reaction will not occur.

$$Ag(s) + Cu(NO_3)_2(aq) \rightarrow NR$$

Nonmetal replaces nonmetal A third type of single-replacement reaction involves the replacement of a nonmetal in a compound by another nonmetal. Halogens are frequently involved in these types of reactions. Like metals, halogens exhibit different activity levels in single-replacement reactions. The reactivities of halogens, determined by single-replacement reactions, are also shown in **Figure 13**.

The most active halogen is fluorine, and the least active is iodine. A more reactive halogen replaces a less reactive halogen that is part of a compound dissolved in water. A halogen cannot replace any halogen listed above it. For example, fluorine replaces bromine in water containing dissolved sodium bromide, as shown in the following chemical equation. However, bromine does not replace fluorine in water containing dissolved sodium fluoride.

$$F_2(g) + 2NaBr(aq) \rightarrow 2NaF(aq) + Br_2(l)$$

$$Br_2(g) + 2NaF(aq) \rightarrow NR$$

Get It?
Explain how a single-replacement reaction works.

EXAMPLE Problem 2

SINGLE-REPLACEMENT REACTIONS Predict the products that will result when these reactants combine, and write a balanced chemical equation for each reaction.

a. $Fe(s) + CuSO_4(aq) \rightarrow$

b. $Br_2(l) + MgCl_2(aq) \rightarrow$

c. $Mg(s) + AlCl_3(aq) \rightarrow$

1 ANALYZE THE PROBLEM

You are given three sets of reactants. Using **Figure 13,** you must first determine if each reaction occurs. Then, if a reaction is predicted, you can determine the product(s) of the reaction. With this information you can write a skeleton equation for the reaction. Finally, you can use the steps for balancing chemical equations to write the complete balanced chemical equation.

2 SOLVE FOR THE UNKNOWN

a. Iron is listed above copper in the activity series. Therefore, the first reaction will occur because iron is more reactive than copper. In this case, iron will replace copper. The skeleton equation for this reaction is

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

This equation is balanced.

b. In the second reaction, chlorine is more reactive than bromine because bromine is listed below chlorine in the activity series. Therefore, the reaction will not occur. The skeleton equation for this situation is

$$Br(l) + MgCl_2(aq) \rightarrow NR$$

No balancing is required.

c. Magnesium is listed above aluminum in the activity series. Therefore, the third reaction will occur because magnesium is more reactive than aluminum. In this case, magnesium will replace aluminum. The skeleton equation for this reaction is

$$Mg(s) + AlCl_3(aq) \rightarrow Al(s) + MgCl_2(aq)$$

This equation is not balanced. The balanced equation is

$$3Mg(s) + 2AlCl_3(aq) \rightarrow 2Al(s) + 3MgCl_2(aq)$$

3 EVALUATE THE ANSWER

The activity series shown in **Figure 13** supports the reaction predictions. The chemical equations balance because the number of atoms of each substance is equal on both sides of the equation.

Real-World Chemistry

Single-Replacement Reactions

ZINC PLATING Tools made of steel are often covered with a layer of zinc to prevent corrosion. Zinc is more reactive than the lead in steel. During zinc plating, the zinc replaces some of the surface lead, coating the steel.

PRACTICE Problems

ADDITIONAL PRACTICE

Predict whether the following single-replacement reactions will occur. If a reaction occurs, write a balanced equation for the reaction.

21. $K(s) + ZnCl_2(aq) \rightarrow$

22. $Cl_2(g) + HF(aq) \rightarrow$

23. $Fe(s) + Na_3PO_4(aq) \rightarrow$

24. CHALLENGE $Al(s) + Pb(NO_3)_2(aq) \rightarrow$

$$AX + BY \rightarrow AY + BX$$

$$Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$$

Figure 14 The color-coding in the generic equation for a double-replacement reaction and in the equation for the reaction between calcium hydroxide and hydrochloric acid shows the anions changing places.

Double-replacement reactions

The final type of replacement reaction, which involves an exchange of ions between two compounds, is called a **double-replacement reaction.**

In the generic equation in **Figure 14**, A and B represent positively charged ions (cations), and X and Y represent negatively charged ions (anions). Notice that the anions have switched places and are now bonded to the other cations in the reaction. In other words, X replaces Y and Y replaces X—a double replacement. More simply, the positive and negative ions of two compounds switch places.

Figure 15 When aqueous sodium hydroxide is added to a solution of copper(II) chloride, the anions (OH^- and Cl^-) change places. The resulting products are sodium chloride, which remains in solution, and copper(II) hydroxide, the blue solid in the beaker.

The reaction between calcium hydroxide and hydrochloric acid is a double-replacement reaction.

$$Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$$

The ionic components of the reaction are Ca^{2+}, OH^-, H^+, and Cl^-. Knowing this, you can now see the two replacements of the reaction. The anions (OH^- and Cl^-) have changed places and are now bonded to the other cations (Ca^{2+} and H^+), as shown in **Figure 14**.

The reaction between sodium hydroxide and copper(II) chloride in solution is also a double-replacement reaction.

$$2NaOH(aq) + CuCl_2(aq) \rightarrow 2NaCl(aq) + Cu(OH)_2(s)$$

In this case, the anions (OH^- and Cl^-) changed places and bonded to the other cations (Na^+ and Ca^{2+}). **Figure 15** shows that the result of this reaction is a solid product, copper(II) hydroxide. A solid produced during a chemical reaction in a solution is called a **precipitate.**

Get It?

Compare How do single-replacement reactions and double-replacement reactions differ?

Table 3 Guidelines for Writing Double-Replacement Reactions

Step	Example
1. Write the components of the reactants in a skeleton equation.	$Al(NO_3)_3 + H_2SO_4$
2. Identify the cations and the anions in each compound.	$Al(NO_3)_3$ has Al^{3+} and NO_3^- H_2SO_4 has H^+ and SO_4^{2-}
3. Pair up each cation with the anion from the other compound.	Al^{3+} pairs with SO_4^{2-} H^+ pairs with NO_3^-
4. Write the formulas for the products using the pairs from Step 3.	$Al_2(SO_4)_3$ HNO_3
5. Write the complete equation for the double-replacement reaction.	$Al(NO_3)_3 + H_2SO_4 \rightarrow Al_2(SO_4)_3 + HNO_3$
6. Balance the equation.	$2Al(NO_3)_3 + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 6HNO_3$

Products of double-replacement reactions

One of the key characteristics of double-replacement reactions is the type of product that is formed when the reaction takes place. All double-replacement reactions produce either water, a precipitate, or a gas. Refer back to the two double-replacement reactions previously discussed in this section. The reaction between calcium hydroxide and hydrochloric acid produces water. A precipitate is produced in the reaction between sodium hydroxide and copper(II) chloride. An example of a double-replacement reaction that forms a gas is that of potassium cyanide and hydrobromic acid.

$$KCN(aq) + HBr(aq) \rightarrow KBr(aq) + HCN(g)$$

The basic steps to write double-replacement reactions are given in **Table 3**.

Get It?

Describe what happens to the anions in a double-replacement reaction.

PRACTICE Problems

ADDITIONAL PRACTICE

Write the balanced chemical equations for the following double-replacement reactions.

25. The two substances at right react to produce solid silver iodide and aqueous lithium nitrate.

26. Aqueous barium chloride and aqueous potassium carbonate react to produce solid barium carbonate and aqueous potassium chloride.

27. Aqueous sodium oxalate and aqueous lead(II) nitrate react to produce solid lead(II) oxalate and aqueous sodium nitrate.

28. CHALLENGE Acetic acid (CH_3COOH) and potassium hydroxide react to produce potassium acetate and water.

Table 4 **Predicting Products of Chemical Reactions**

Type of Reaction	Reactants	Probable Products	Generic Equation
Synthesis	• two or more substances	• one compound	$A + B \rightarrow AB$
Combustion	• a metal and oxygen • a nonmetal and oxygen • a compound and oxygen	• the oxide of the metal • the oxide of the nonmetal • two or more oxides	$A + O_2 \rightarrow AO$
Decomposition	• one compound	• two or more elements and/or compounds	$AB \rightarrow A + B$
Single-replacement	• a metal and a compound • a nonmetal and a compound	• a new compound and the replaced metal • a new compound and the replaced non-metal	$A + BX \rightarrow AX + B$
Double-replacement	• two compounds	• two different compounds, one of which is a solid, water, or a gas	$AX + BY \rightarrow AY + BX$

Table 4 summarizes the types of chemical reactions. Use this table to identify reactions and predict their products. First, write the chemical equation. Second, determine what is happening in the reaction. How many reactants are there? How many products are there? What happened to the elements and compounds in the reaction? Third, use your analysis to classify the reaction. Finally, compare the reaction to the generic equations in the table to check your answer.

Check Your Progress

Summary

- Classifying chemical reactions makes them easier to understand, remember, and recognize.
- Activity series of metals and halogens can be used to predict if single-replacement reactions will occur.

Demonstrate Understanding

29. Describe the four types of chemical reactions and their characteristics.
30. **Explain** how an activity series of metals is organized.
31. **Compare and contrast** single-replacement reactions and double-replacement reactions.
32. **Describe** the result of a double-replacement reaction.
33. Predict Use the fact that atoms are conserved, and your knowledge of the properties of the elements involved, to describe and predict the reaction most likely to occur when barium reacts with fluorine. Write the chemical equation for the reaction.
34. **Interpret Data** Could the following reaction occur? Explain your answer.

$$3Ni + 2AuBr_3 \rightarrow 3NiBr_2 + 2Au$$

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LESSON 3

REACTIONS IN AQUEOUS SOLUTIONS

FOCUS QUESTION

What is unique about reactions that take place in water?

Aqueous Solutions

You read previously that a solution is a homogeneous mixture. Many of the reactions discussed in the previous section involve substances dissolved in water. When a substance dissolves in water, a solution forms. An **aqueous solution** contains one or more substances called **solutes** dissolved in the water. In this case, water is the **solvent**—the most plentiful substance in the solution.

Molecular compounds in solution

Although water is always the solvent in aqueous solutions, there are many possible solutes. Some solutes, such as sucrose (table sugar) and ethanol (grain alcohol), are molecular compounds that exist as molecules in aqueous solutions. Other solutes are molecular compounds that form ions when they dissolve in water. For example, the molecular compound hydrogen chloride forms hydrogen ions and chloride ions when it dissolves in water, as shown in **Figure 16**.

An equation can be used to show this ionization process.

$$HCl(aq) \rightarrow H^{+}(aq) + Cl^{-}(aq)$$

Compounds such as hydrogen chloride that produce hydrogen ions in aqueous solution are acids. In fact, an aqueous solution of hydrogen chloride is usually referred to as hydrochloric acid.

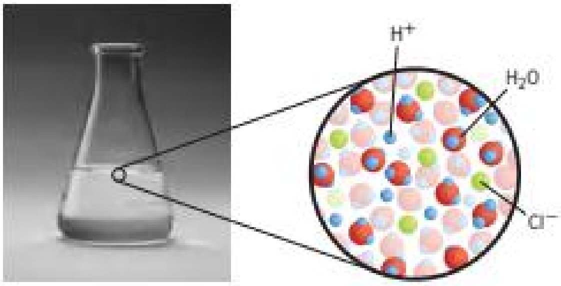

Figure 16 In water, hydrogen chloride (HCl) breaks apart into hydrogen ions (H^+) and chloride ions (Cl^-).

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Design Your Own Lab: How thick is the coating on a galvanized nail?
Plan and carry out an investigation to discover the structure and function of the coating on a galvanized nail by using your knowledge of the chemical properties involved.

Ionic compounds in solution

In addition to molecular compounds, ionic compounds might be solutes in aqueous solutions. Recall that ionic compounds consist of positive ions and negative ions held together by ionic bonds. When ionic compounds dissolve in water, their ions can separate—a process called dissociation. For example, an aqueous solution of the ionic compound sodium chloride contains Na^+ and Cl^- ions because when sodium chloride is added to water, the ions in the compound dissociate and become dispersed throughout the resulting solution.

Types of Reactions in Aqueous Solutions

When two aqueous solutions that contain ions as solutes are combined, the ions might react with one another. These reactions are always double-replacement reactions. The solvent molecules, which are all water molecules, do not usually react. Three types of products can form from the double-replacement reaction: a precipitate, water, or a gas.

Reactions that form precipitates

Some reactions that occur in aqueous solutions produce precipitates. For example, recall from Lesson 2 that when aqueous solutions of sodium hydroxide and copper(II) chloride are mixed, a double-replacement reaction occurs in which the precipitate copper(II) hydroxide forms. This reaction is shown in **Figure 17.**

$$2NaOH(aq) + CuCl_2(aq) \rightarrow 2NaCl(aq) + Cu(OH)_2(s)$$

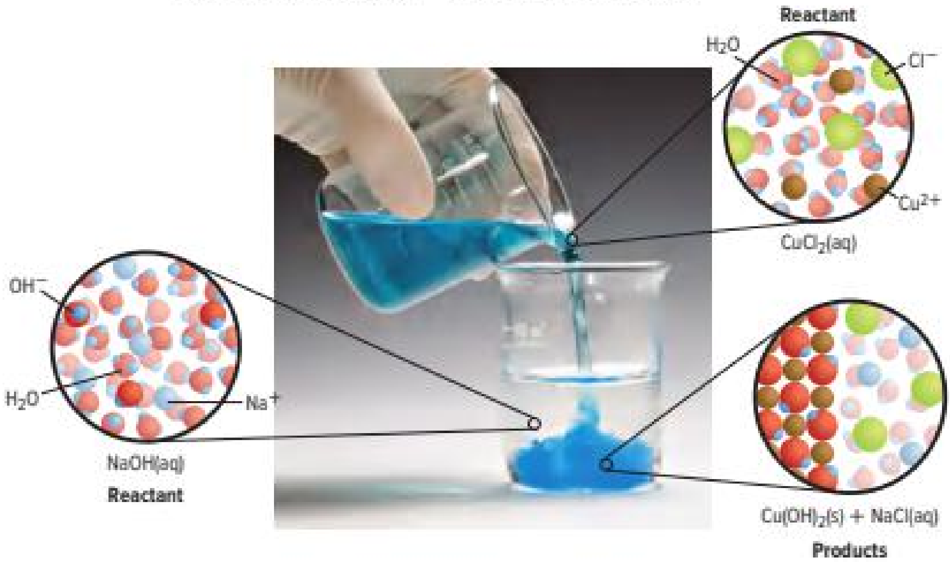

Figure 17 Like the aqueous solution of HCl in **Figure 16**, sodium hydroxide (NaOH) in an aqueous solution dissociates into sodium (Na^+) and hydroxide (OH^-) ions. Copper(II) chloride ($CuCl_2$) also dissociates into Cu^{2+} and Cl^- ions.

Interpret *What is the identity of the blue solid that is forming in the beaker?*

Note that the chemical equation does not show some details of this reaction. Sodium hydroxide and copper(II) chloride are ionic compounds. Therefore, in aqueous solutions they exist as Na^+, OH^-, Cu^{2+}, and Cl^- ions. When their solutions are combined, Cu^{2+} ions in one solution and OH^- ions in the other solution react to form the precipitate copper(II) hydroxide, $Cu(OH)_2(s)$. The Na^+ and Cl^- ions remain dissolved in the newly formed solution.

Ionic equations To show the details of reactions that involve ions in aqueous solutions, chemists use ionic equations. Ionic equations differ from chemical equations in that substances that are ions in solution are written as ions in the equation. Look again at the reaction between aqueous solutions of sodium hydroxide and copper(II) chloride. To write the ionic equation for this reaction, you must show the reactants, $NaOH(aq)$ and $CuCl_2(aq)$, and the product, $NaCl(aq)$, as ions.

$$2Na^+(aq) + 2OH^-(aq) + Cu^{2+}(aq) + 2Cl^-(aq) \rightarrow 2Na^+(aq) + 2Cl^-(aq) + Cu(OH)_2(s)$$

An ionic equation that shows all of the particles in a solution as they exist is called a **complete ionic equation.**

Note that the sodium ions and the chloride ions are both reactants and products. Because they are both reactants and products, they do not participate in the reaction. Ions that do not participate in a chemical reaction are called **spectator ions** and are not usually shown in ionic equations. Spectator ions are like spectators at a baseball game. The ions are present for the reaction but they do not affect the outcome of the reaction, just as spectators at a baseball game are present to watch the game but do not directly affect its outcome.

Net ionic equations are ionic equations that include only the particles that participate in the reaction. Net ionic equations are written from complete ionic equations by removing all spectator ions. For example, a net ionic equation is what remains after the sodium and chloride ions are crossed out of this complete ionic equation.

$$\cancel{2Na^+(aq)} + 2OH^-(aq) + Cu^{2+}(aq) + \cancel{2Cl^-(aq)} \rightarrow \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)} + Cu(OH)_2(s)$$

Only the hydroxide and copper ions are left in the net ionic equation shown below.

$$2OH^-(aq) + Cu^{2+}(aq) \rightarrow Cu(OH)_2(s)$$

Get It?

Compare How are complete ionic equations and net ionic equations different from chemical equations?

SCIENCE USAGE V. COMMON USAGE

compound

Science usage: a chemical combination of two or more different elements

Salt is a compound comprised of the elements sodium and chlorine.

Common usage: a word that consists of two or more words

Two compound words are basketball *and* textbook.

EXAMPLE Problem 3

REACTIONS THAT FORM A PRECIPITATE Write the chemical, complete ionic, and net ionic equations for the reaction between aqueous solutions of barium nitrate and sodium carbonate that forms the precipitate barium carbonate.

1 ANALYZE THE PROBLEM

You are given the word equation for the reaction between barium nitrate and sodium carbonate. You must determine the chemical formulas and relative amounts of all reactants and products to write the balanced chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation, you can write the net ionic equation. The net ionic equation will include fewer substances than the other equations.

2 SOLVE FOR THE UNKNOWN

Write the correct chemical formulas and physical states for all substances involved in the reaction.

$Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(s) + NaNO_3(aq)$

$Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(s) + 2NaNO_3(aq)$ — Balance the skeleton equation.

$Ba^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + 2Na^+(aq) + 2NO_3^-(aq)$ — Show the ions of the reactants and the products.

$Ba^{2+}(aq) + \cancel{2NO_3^-(aq)} + \cancel{2Na^+(aq)} + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + \cancel{2Na^+(aq)} + \cancel{2NO_3^-(aq)}$ — Cross out the spectator ions from the complete ionic equation.

$Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$ — Write the net ionic equation.

3 EVALUATE THE ANSWER

The net ionic equation includes fewer substances than the other equations because it shows only the reacting particles. The particles composing the solid precipitate that is the result of the reaction are no longer ions.

PRACTICE Problems ADDITIONAL PRACTICE

Write chemical, complete ionic, and net ionic equations for each of the following reactions that might produce a precipitate. Use *NR* to indicate that no reaction occurs.

35. Aqueous solutions of potassium iodide and silver nitrate are mixed, forming the precipitate silver iodide.

36. Aqueous solutions of ammonium phosphate and sodium sulfate are mixed. No precipitate forms and no gas is produced.

37. Aqueous solutions of aluminum chloride and sodium hydroxide are mixed, forming the precipitate aluminum hydroxide.

38. Aqueous solutions of lithium sulfate and calcium nitrate are mixed, forming the precipitate calcium sulfate.

39. CHALLENGE When aqueous solutions of sodium carbonate and manganese(V) chloride are mixed, a precipitate forms. The precipitate is a compound containing manganese.

Reactions that form water

Another type of double-replacement reaction that occurs in an aqueous solution produces water molecules. For example, when you mix hydrobromic acid (HBr) with a sodium hydroxide solution (NaOH), as shown in **Figure 18**, a double-replacement reaction occurs and water is formed. The chemical equation is shown below.

$$HBr(aq) + NaOH(aq) \rightarrow H_2O(l) + NaBr(aq)$$

In this case, the reactants and the product sodium bromide exist as ions in an aqueous solution. The complete ionic equation for this reaction shows these ions.

$$H^+(aq) + Br^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow$$
$$H_2O(l) + Na^+(aq) + Br^-(aq)$$

Examine the complete ionic equation. The reacting solute ions are the hydrogen ions and hydroxide ions because the sodium ions and bromine ions are spectator ions. If you cross out the spectator ions, you are left with the ions that take part in the reaction.

$$H^+(aq) + \cancel{Br^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \rightarrow$$
$$H_2O(l) + \cancel{Na^+(aq)} + \cancel{Br^-(aq)}$$

This equation is the net ionic equation for the reaction.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Get It?

Analyze In the reaction between hydrobromic acid and sodium hydroxide, why are the sodium ions and bromine ions called spectator ions?

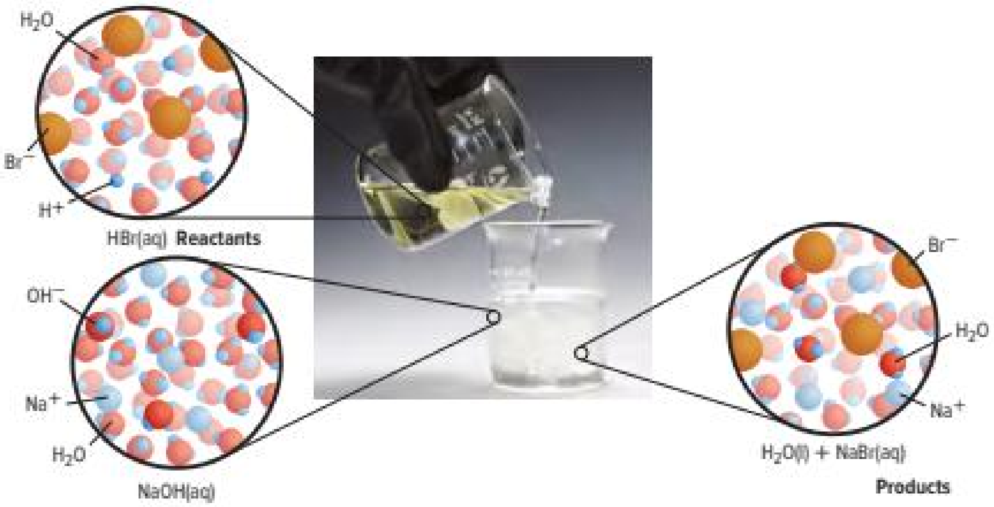

Figure 18 In water, hydrogen bromide (HBr) ionizes into H^+ and Br^- ions. Sodium hydroxide (NaOH) also dissociates into Na^+ and OH^- ions. The hydrogen ions and hydroxide ions react to form water.

Determine *Which ions are the anions in this reaction? The cations?*

After you have written the chemical, complete ionic, and net ionic equations for several double-replacement reactions that produce water, you will notice that the net ionic equation is the same for all of them. It shows hydrogen ions and hydroxide ions reacting to form water. If after writing the chemical and complete ionic equations for these types of reactions you end up with a different net ionic equation, go back and check your work.

EXAMPLE Problem 4

REACTIONS THAT FORM WATER Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous lithium hydroxide. This reaction produces water and aqueous lithium chloride.

1 ANALYZE THE PROBLEM

You are given the word equation for the reaction that occurs between hydrochloric acid and aqueous lithium hydroxide to produce water and aqueous lithium chloride. You must determine the chemical formulas for and relative amounts of all reactants and products to write the balanced chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation, you can write the net ionic equation.

2 SOLVE FOR THE UNKNOWN

Write the skeleton equation for the reaction and balance it.

$$HCl(aq) + LiOH(aq) \rightarrow H_2O(l) + LiCl(aq)$$

$$H^+(aq) + Cl^-(aq) + Li^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Li^+(aq) + Cl^-(aq)$$

Show the ions of the reactants and the products.

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Li^+(aq)} + OH^-(aq) \rightarrow H_2O(l) + \cancel{Li^+(aq)} + \cancel{Cl^-(aq)}$$

Cross out the spectator ions from the complete ionic equation.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Write the net ionic equation.

3 EVALUATE THE ANSWER

The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produces water. The particles that compose the product water are no longer ions.

PRACTICE Problems

ADDITIONAL PRACTICE

Write chemical, complete ionic, and net ionic equations for the reactions between the following substances, which produce water.

40. Mixing sulfuric acid (H_2SO_4) and aqueous potassium hydroxide produces water and aqueous potassium sulfate.

41. Mixing hydrochloric acid (HCl) and aqueous calcium hydroxide produces water and aqueous calcium chloride.

42. Mixing nitric acid (HNO_3) and aqueous ammonium hydroxide produces water and aqueous ammonium nitrate.

43. Mixing hydrosulfuric acid (H_2S) and aqueous calcium hydroxide produces water and aqueous calcium sulfide.

44. **CHALLENGE** When benzoic acid (C_6H_5COOH) and magnesium hydroxide are mixed, water and magnesium benzoate are produced.

Reactions that form gases

A third type of double-replacement reaction that occurs in aqueous solutions results in the formation of a gas. Some gases commonly produced in these reactions are carbon dioxide, hydrogen cyanide, and hydrogen sulfide. A gas-producing reaction occurs when you mix hydroiodic acid (HI) with an aqueous solution of lithium sulfide. Bubbles of hydrogen sulfide gas form in the container during the reaction. Lithium iodide is also produced.

$$2HI(aq) + Li_2S(aq) \rightarrow H_2S(g) + 2LiI(aq)$$

The reactants hydroiodic acid and lithium sulfide exist as ions in aqueous solution. Therefore, you can write an ionic equation for this reaction. The complete ionic equation includes all of the substances in the solution.

$$2H^+(aq) + 2I^-(aq) + 2Li^+(aq) + S^{2-}(aq) \rightarrow H_2S(g) + 2Li^+(aq) + 2I^-(aq)$$

Note that there are many spectator ions in the equation. When the spectator ions are crossed out, only the substances involved in the reaction remain in the equation. In this case, only hydrogen ions, sulfide ions, and hydrogen sulfide remain.

$$2H^+(aq) + \cancel{2I^-(aq)} + \cancel{2Li^+(aq)} + S^{2-}(aq) \rightarrow H_2S(g) + \cancel{2Li^+(aq)} + \cancel{2I^-(aq)}$$

This is the net ionic equation.

$$2H^+(aq) + S^{2-}(aq) \rightarrow H_2S(g)$$

Get It?

Explain why lithium ions are not shown in the above equation.

Figure 19 When vinegar and baking soda (sodium hydrogen carbonate, $NaHCO_3$) combine, the result is a vigorous bubbling that releases carbon dioxide (CO_2).

A reaction that produces carbon dioxide gas occurs in your kitchen when you mix vinegar and baking soda. Vinegar is an aqueous solution of acetic acid and water. Baking soda essentially consists of sodium hydrogen carbonate. Rapid bubbling occurs when vinegar and baking soda are combined. The bubbles are carbon dioxide gas escaping from the solution. You can see this reaction occurring in **Figure 19**.

A reaction similar to the one between vinegar and baking soda occurs when you combine any acidic solution and sodium hydrogen carbonate. In all cases, two reactions must occur almost simultaneously in the solution to produce the carbon dioxide gas. One of these is a double-replacement reaction and the other is a decomposition reaction.

For example, when you dissolve sodium hydrogen carbonate in hydrochloric acid, a gas-producing double-replacement reaction occurs. The hydrogen in the hydrochloric acid and the sodium in the sodium hydrogen carbonate replace each other.

$$HCl(aq) + NaHCO_3(aq) \rightarrow H_2CO_3(aq) + NaCl(aq)$$

Sodium chloride is an ionic compound, and its ions remain separate in the aqueous solution. However, as the carbonic acid (H_2CO_3) forms, it decomposes immediately into water and carbon dioxide.

$$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$

EXAMPLE Problem 5

REACTIONS THAT FORM GASES Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous sodium sulfide, which produces hydrogen sulfide gas.

1 ANALYZE THE PROBLEM

You are given the word equation for the reaction between hydrochloric acid (HCl) and sodium sulfide (Na_2S). You must write the skeleton equation and balance it. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions in the complete ionic equation, you can write the net ionic equation.

2 SOLVE FOR THE UNKNOWN

Write the correct skeleton equation for the reaction.

$$HCl(aq) + Na_2S(aq) \rightarrow H_2S(g) + NaCl(aq)$$

$$2HCl(aq) + Na_2S(aq) \rightarrow H_2S(g) + 2NaCl(aq)$$

Balance the skeleton equation.

$$2H^+(aq) + 2Cl^-(aq) + 2Na^+(aq) + S^{2-}(aq) \rightarrow H_2S(g) + 2Na^+(aq) + 2Cl^-(aq)$$

Show the ions of the reactants and the products.

$$2H^+(aq) + \cancel{2Cl^-(aq)} + \cancel{2Na^+(aq)} + S^{2-}(aq) \rightarrow H_2S(g) + \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)}$$

Cross out the spectator ions from the complete ionic equation.

$$2H^+(aq) + S^{2-}(aq) \rightarrow H_2S(g)$$

Write the net ionic equation in its smallest whole-number ratio.

3 EVALUATE THE ANSWER

The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produce hydrogen sulfide. The particles that compose the product are no longer ions.

PRACTICE Problems

ADDITIONAL PRACTICE

Write chemical, complete ionic, and net ionic equations for these reactions.

45. Perchloric acid ($HClO_4$) reacts with aqueous potassium carbonate, forming carbon dioxide gas and water.

46. Sulfuric acid (H_2SO_4) reacts with aqueous sodium cyanide, forming hydrogen cyanide gas and aqueous sodium sulfate.

47. Hydrobromic acid (HBr) reacts with aqueous ammonium carbonate, forming carbon dioxide gas and water.

48. Nitric acid (HNO_3) reacts with aqueous potassium rubidium sulfide, forming hydrogen sulfide gas.

49. **CHALLENGE** Aqueous potassium iodide reacts with lead nitrate in solution, forming solid lead iodide.

Double-replacement reaction

$$AX + BY \rightarrow AY + BX$$

$$HCl(aq) + NaHCO_3(aq) \rightarrow H_2CO_3(aq) + NaCl(aq)$$

$$AB \rightarrow A + B$$

$$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$

Decomposition reaction

Figure 20 When HCl is combined with $NaHCO_3$, a double-replacement reaction takes place, followed immediately by a decomposition reaction.

Overall equations

Recall that when you combine an acidic solution, such as hydrochloric acid, and sodium hydrogen carbonate, two reactions occur—a double-replacement reaction and a decomposition reaction. These reactions are shown in **Figure 20.** The two reactions can be combined and represented by one chemical equation in a process similar to adding mathematical equations. An equation that combines two reactions is called an overall equation. To write an overall equation, the reactants in the two reactions are written on the reactant side of the combined equation, and the products of the two reactions are written on the product side. Then, any substances that are on both sides of the equation are crossed out.

Reaction 1 $HCl(aq) + NaHCO_3(aq) \rightarrow H_2CO_3(aq) + NaCl(aq)$

Reaction 2 $H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$

Combined equation $HCl(aq) + NaHCO_3(aq) + \cancel{H_2CO_3(aq)} \rightarrow \cancel{H_2CO_3(aq)} + NaCl(aq) + H_2O(l) + CO_2(g)$

Overall equation $HCl(aq) + NaHCO_3(aq) \rightarrow H_2O(l) + CO_2(g) + NaCl(aq)$

In this case, the reactants in the overall equation exist as ions in aqueous solution. Therefore, a complete ionic equation can be written for the reaction.

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g) + Na^+(aq) + Cl^-(aq)$$

Note that the sodium and chloride ions are the spectator ions. When you cross them out, only the substances that take part in the reaction remain.

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g) + \cancel{Na^+(aq)} + \cancel{Cl^-(aq)}$$

The net ionic equation shows that the reaction produces water and carbon dioxide gas.

$$H^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g)$$

Get It?

Describe What is an overall equation?

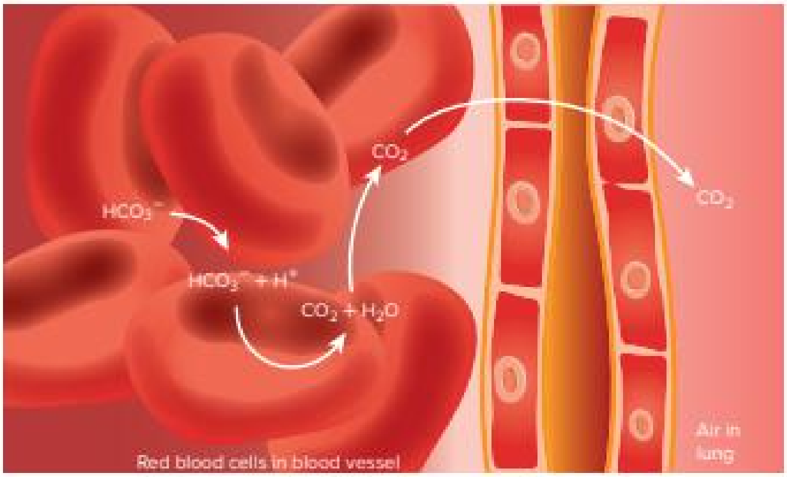

Figure 21 After a hydrogen carbonate ion (HCO_3^-) enters a red blood cell, it reacts with a hydrogen ion (H^+) to form water and carbon dioxide (CO_2). The CO_2 is exhaled from the lungs during respiration.

LIFE SCIENCE Connection The reaction between hydrogen ions and hydrogen carbonate ions to produce water and carbon dioxide is an important one in your body. This reaction is occurring in the blood vessels of your lungs as you read these words. As shown in **Figure 21**, the carbon dioxide gas produced in your cells is transported in your blood as hydrogen carbonate ions (HCO_3^-). In the blood vessels of your lungs, the HCO_3^- ions combine with H^+ ions to produce water and CO_2, which you exhale.

As you have read, the reaction between an acid and sodium hydrogen carbonate also occurs in products that are made with baking soda, which contains sodium hydrogen carbonate. Sodium hydrogen carbonate makes baked goods rise. This is partly because the sodium hydrogen carbonate reacts with an acid in the batter, such as lemon juice, to produce carbon dioxide. The gas results in bubbles that give the baked item a fluffy texture. In addition, at high temperatures, sodium hydrogen carbonate decomposes to form products that include carbon dioxide. The chemical and physical properties of sodium hydrogen carbonate mean it has many other uses. It is used as an antacid and in deodorants to absorb moisture and odors. Baking soda can be added to toothpaste to whiten teeth and freshen breath. As a paste, sodium bicarbonate can be used in cleaning and scrubbing. It is even used as a fire-suppression agent in some fire extinguishers.

Get It?

Infer Why is a mechanism needed to remove carbon dioxide gas from your cells?

STEM CAREER Connection

Hair Stylist

Most hair stylists are savvy with scissors and up to date on the latest hair trends, but did you know that they also use chemical reactions in their career? Hair stylists work with many processes, like those that permanently curl or relax hair, that involve chemical reactions. They also need to know how to handle hazardous chemicals safely.

Table 5 Reactions that Take Place in Aqueous Solutions

Type of Reaction	Description
Reactions that form precipitates	Dissolved substances are mixed. When they react, a solid is produced, visible as a white or colored cloudiness in the reaction mixture.
Reactions that form water	Evidence of the reaction may not be observable because water is colorless, odorless, and already makes up most of the solution.
Reactions that form gases	Gases such as carbon dioxide, hydrogen cyanide, and hydrogen sulfide are produced. Bubbles are produced as the reaction proceeds.

Table 5 lists the types of reactions that occur in aqueous solutions. The descriptions summarize how physical evidence such as the formation of a precipitate or the production of bubbles can help you classify a chemical reaction.

Check Your Progress

Summary

- In aqueous solutions, the solvent is always water. There are many possible solutes.
- Many molecular compounds form ions when they dissolve in water. When some ionic compounds dissolve in water, their ions separate.
- When two aqueous solutions that contain ions as solutes are combined, the ions might react with one another. The solvent molecules do not usually react.
- Reactions that occur in aqueous solutions are double-replacement reactions.

Demonstrate Understanding

50. **List** three common types of products produced by reactions that occur in aqueous solutions.
51. **Describe** solvents and solutes in aqueous solution.
52. **Distinguish** between a complete ionic equation and a net ionic equation.
53. **Write** complete ionic and net ionic equations for the reaction between sulfuric acid (H_2SO_4) and calcium carbonate ($CaCO_3$).
    $$H_2SO_4(aq) + CaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + CaSO_4(aq)$$
54. **Analyze** Complete and balance the following equation.
    $$CO_2(g) + HCl(aq) \rightarrow$$
55. **Predict** What type of product would the following reaction be most likely to produce? Explain your reasoning.
    $$Ba(OH)_2(aq) + 2HCl(aq) \rightarrow$$
56. **Formulate Equations** A reaction occurs when nitric acid (HNO_3) is mixed with an aqueous solution of potassium hydrogen carbonate. Aqueous potassium nitrate is produced. Write the chemical and net ionic equations for the reaction.

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SCIENCE & SOCIETY

How One Woman Led the FDA to Save Lives

When a drug company could not provide data to support its claims of safety for a new drug, one woman's persistence shaped the future of the Food and Drug Administration.

President John F. Kennedy and Dr. Frances Oldham Kelsey.

Drug Safety

Drug metabolism is the series of chemical reactions that breaks down a drug into compounds that are used in the body. Before a new drug is approved for human use, scientists perform research to ensure that no harmful side effects are produced as a result of these reactions. Pharmacologists – scientists who study the reactions of drugs in the body – work with the Food and Drug Administration (FDA) to evaluate data and make decisions regarding the safety and effectiveness of a proposed medication.

A Historical Example: Thalidomide

In 1960, the FDA received an application for the approval of thalidomide, a drug used to treat a variety of symptoms from nausea to sleeplessness. In Europe and other parts of the world, many doctors gave the drug to pregnant women as a treatment for morning sickness.

The application fell to pharmacologist Dr. Frances Oldham Kelsey. When she reviewed clinical studies of the drug, the lack of research to support the drug's safety concerned her. She and her colleagues were further alarmed when insistence for more data from the drug company did not result in additional evidence.

While the FDA held off approving the application, reports of devastating birth defects in babies born to mothers who had taken thalidomide surfaced around the world. By November 1961, German officials took the drug off the market, and other countries soon followed suit. By early 1962, the distributor withdrew its application for FDA approval.

Because of her persistence in the pursuit of evidence, Kelsey impacted countless lives. In 1962, President John F. Kennedy honored Dr. Kelsey with the Distinguished Federal Civilian Service award. In part because of her work, the United States Congress passed the Kefauver-Harris bill, forcing major regulatory reforms on the pharmaceutical industry.

COMMUNICATE SCIENTIFIC INFORMATION

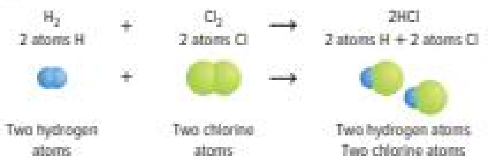
Research the metabolism of atorvastatin, a commonly prescribed medication used to lower cholesterol. Make a presentation that explains these metabolic reactions. Include any potential harmful effects or products of the reactions.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 REACTIONS AND EQUATIONS

- Some physical changes are evidence that indicate a chemical reaction has occurred.
- Word equations and skeleton equations provide important information about a chemical reaction.
- A chemical equation gives the identities and relative amounts of the reactants and products that are involved in a chemical reaction.
- Balancing an equation involves adjusting the coefficients until the number of atoms of each element is equal on both sides of the equation.



- chemical reaction
- reactant
- product
- chemical equation
- coefficient

Lesson 2 CLASSIFYING CHEMICAL REACTIONS

- Classifying chemical reactions makes them easier to understand, remember, and recognize.
- Activity series of metals and halogens can be used to predict if single-replacement reactions will occur.

- synthesis reaction
- combustion reaction
- decomposition reaction
- single-replacement reaction
- double-replacement reaction
- precipitate

Lesson 3 REACTIONS IN AQUEOUS SOLUTIONS

- In aqueous solutions, the solvent is always water. There are many possible solutes.
- Many molecular compounds form ions when they dissolve in water. When some ionic compounds dissolve in water, their ions separate.
- When two aqueous solutions that contain ions as solutes are combined, the ions might react with one another. The solvent molecules do not usually react.
- Reactions that occur in aqueous solutions are double-replacement reactions.

- aqueous solution
- solute
- solvent
- complete ionic equation
- spectator ion
- net ionic equation

THREE-DIMENSIONAL THINKING
Module Wrap-Up

REVISIT THE PHENOMENON

What happens to food when you cook it?

CER Claim, Evidence, Reasoning

Explain your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your **STEM** unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

SEP Data Analysis Lab

How can you explain the reactivities of halogens?

The location of all the halogens in group 17 in the periodic table tells you that halogens have common characteristics. Indeed, halogens are all nonmetals and have seven electrons in their outermost orbitals. However, each halogen also has its own characteristics, such as the ability to react with other substances. Examine the data table. It includes data about the atomic radii, ionization energies, and electronegativities of the halogens.

Data and Observations

Properties of Halogens

Halogen	Atomic Radius (pm)	Ionization Energy (kJ/mol)	Electro-negativity
Fluorine	72	1681	3.98
Chlorine	100	1251	3.16
Bromine	114	1140	2.96
Iodine	133	1008	2.66
Astatine	140	920	2.2

CER Analyze and Interpret Data

1. **Make graphs** Use the information in the data table to make three line graphs.
2. **Claim** Describe any periodic trends that you identify in the data.
3. **Evidence, Reasoning** Relate any periodic trends that you identify among the halogens to the activity series of the halogens.
4. **Claim, Evidence, Reasoning** Predict the location of the element astatine in the activity series of halogens. Explain.

THE MOLE

THE MOLE

ENCOUNTER THE PHENOMENON

How is counting pennies like counting atoms?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

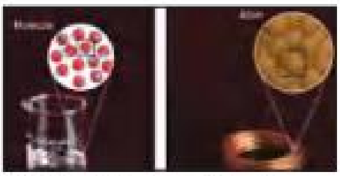
CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about how counting pennies is like counting atoms.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move **through the module.**

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.



LESSON 1: Explore & Explain: Defining the Mole

LESSON 2: Explore & Explain: Converting Between Mass and Moles

LESSON 1

MEASURING MATTER

FOCUS QUESTION

Why is it useful to group large numbers of things?

Counting Particles

If you were buying a bouquet of roses for a special occasion, you probably would not ask for 12 or 24; you would ask for one or two dozen. Similarly, you might buy a pair of gloves, a ream of paper, or a gross of pencils. Each of the units shown in **Figure 1**—a pair, a dozen, a gross, and a ream—represents a specific number of items. These units make counting objects easier. It is easier to buy and sell paper by the ream—500 sheets—than by the individual sheet.

Each of the counting units shown in **Figure 1** is appropriate for certain kinds of objects, depending primarily on their size and function. But regardless of the object—gloves, eggs, pencils, or paper—the number that the unit represents is always constant.

Chemists also need a convenient method for accurately counting the number of atoms, molecules, or formula units in a sample of a substance. However, in chemistry, systems can only be quantified indirectly as atoms are so small and there are so many of them in even the smallest sample that it is impossible to count them directly. Because of this, chemists created a counting unit called the mole. In the Launch Lab, you probably found that a mole of any object is an enormous number of items.

Matt Meadows/McGraw-Hill Education

Figure 1 Different units are used to count different types of objects. A pair is two objects, a dozen is 12, a gross is 144, and a ream is 500. What other counting units are you familiar with?

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Small-Scale Lab: Determining Avogadro's Number

Obtain and evaluate information to determine the function of Avogadro's number.

Revisit the Encounter the Phenomenon Question

What information from this lesson can help you answer the module question?

The mole

The **mole,** abbreviated mol, is the SI base unit used to measure the amount of a substance. A mole is defined as the number of carbon atoms in exactly 12 g of pure carbon-12. Through years of experimentation, it has been established that a mole of anything contains $6.022140857 \times 10^{23}$ representative particles. A representative particle is any kind of particle, such as an atom, a molecule, a formula unit, an electron, or an ion. If you write out this number, it looks like this.

602,214,085,700,000,000,000,000

The number $6.022140857 \times 10^{23}$ is called **Avogadro's number,** in honor of the Italian scientist Amedeo Avogadro, who, in 1811, proposed that equal volumes of gas at the same temperature and pressure contain the same number of particles. In this book, Avogadro's number is rounded to three significant figures, 6.02×10^{23}.

Avogadro's number would not be convenient for measuring a quantity of marbles. Avogadro's number of marbles with a five-eighth inch diameter would cover the surface of Earth to a depth of more than 14 km! It is more convenient to use the mole to measure amounts of substances. **Figure 2** shows one-mole quantities of water, copper, and salt. Each one has a different representative particle: molecules of water, atoms of copper, and formula units of NaCl.

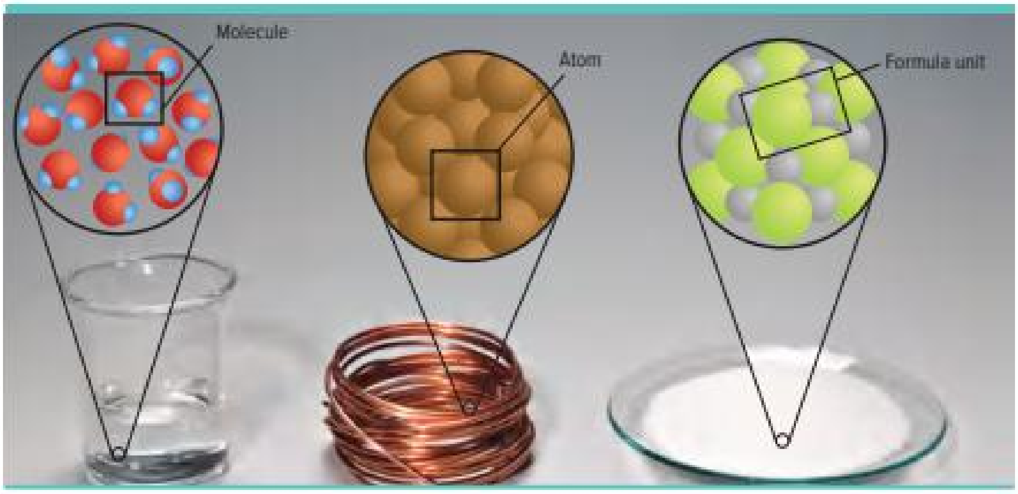


Figure 2 The amount of each different substance is 6.02×10^{23} or 1 mol, of representative particles. The representative particle for each substance is shown in a box. Refer to **Table R-1** in the Student Resources for a key to atom color conventions.

SCIENCE USAGE V. COMMON USAGE

mole

Science usage: an SI base unit used to measure the quantity of matter
The chemist measured out a mole of the compound.

Common usage: a small burrowing animal. *The damage to the lawn was caused by a mole.*

CCC CROSSCUTTING CONCEPTS

Scale, Proportion, and Quantity Some systems can only be studied indirectly as they are too small, too large, too fast, or too slow to observe directly. Make a poster that compares and contrasts Avogadro's number to quantify atoms with methods used to quantify microorganisms (turbidity) and methods used to quantify astronomical distances (light-years or parsecs). Can you think of a third example?

Converting Between Moles and Particles

Figure 3 A key to using dimensional analysis is correctly identifying the mathematical relationship between the units you are converting. The relationship shown here, 12 roses = 1 dozen roses, can be used to write two conversion factors.

Suppose you buy three-and-one-half dozen roses and want to know how many roses you have. Recall what you have learned about conversion factors. You can multiply the known quantity (3.5 dozen roses) by a conversion factor to express the quantity in the units you want (number of roses). First, identify the mathematical relationship that relates the given unit with the desired unit. **Figure 3** shows the relationship.

Relationship: 1 dozen roses = 12 roses

By dividing each side of the equality by the other side, you can write two conversion factors from the relationship.

Conversion factors: $\frac{12 \text{ roses}}{1 \text{ dozen roses}}$ and $\frac{1 \text{ dozen roses}}{12 \text{ roses}}$

Then choose the conversion factor that, when multiplied by the known quantity, results in the desired unit. When set up correctly, all units cancel except those required for the answer.

Conversion: $3.5 \cancel{\text{dozen roses}} \times \frac{12 \text{ roses}}{1 \cancel{\text{dozen roses}}} = 42 \text{ roses}$

Get It?

Determine how many pennies are in the stack on the left in the photograph at the beginning of this module. Create a conversion factor to calculate how many pennies are in the photo, knowing that there are 12 stacks of pennies.

Moles to particles

Now suppose you want to determine how many particles of sucrose are in 3.50 mol of sucrose. The relationship between moles and representative particles is given by Avogadro's number, 1 mol of representative particles = 6.02×10^{23} representative particles.

Using this relationship, there are two different conversion factors that relate representative particles and moles.

$$\frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mol}}$$

$$\frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ representative particles}}$$

Use the following conversion factor to find the number of particles in a given number of moles.

$$\text{number of moles} \times \frac{6.02 \times 10^{23} \text{ representative particles}}{1 \text{ mol}}$$

$$= \text{number of representative particles}$$

As shown in **Figure 4**, the representative particle of sucrose is a molecule. To obtain the number of sucrose molecules contained in 3.50 mol of sucrose, you need to use Avogadro's number as a conversion factor.

$$3.50\ \cancel{\text{mol sucrose}} \times \frac{6.02 \times 10^{23}\ \text{molecules sucrose}}{1\ \cancel{\text{mol sucrose}}}$$

$$= 2.11 \times 10^{24}\ \text{molecules sucrose}$$

There are 2.11×10^{24} molecules of sucrose in 3.50 mol of sucrose.

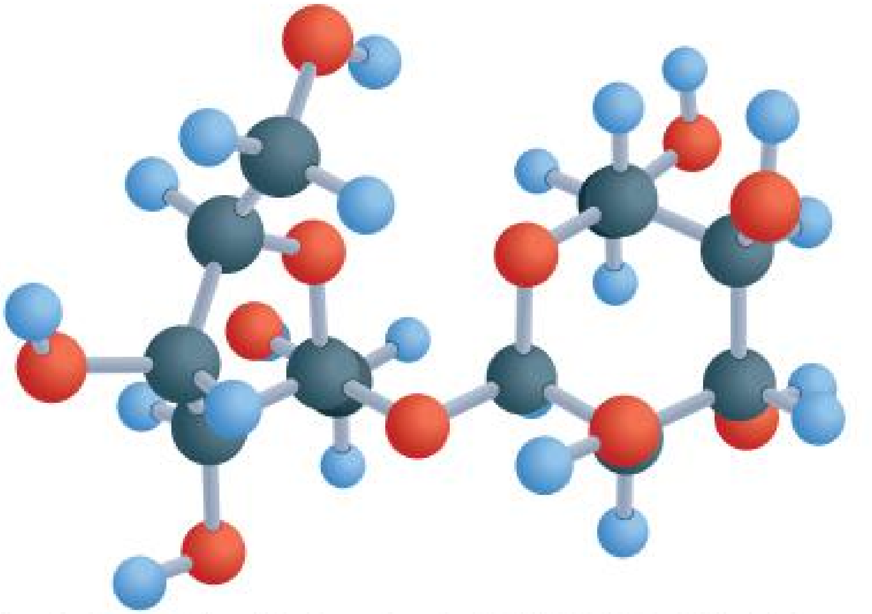


Figure 4 The representative particle of sucrose is a molecule. The ball-and-stick model shows that a molecule of sucrose is a single unit made up of carbon, hydrogen, and oxygen.

Analyze *Use the ball-and-stick model of sucrose to write the chemical formula for sucrose.*

PRACTICE Problems ADDITIONAL PRACTICE

1. Zinc (Zn) is used to form a corrosion-inhibiting surface on galvanized steel. Determine the number of Zn atoms in 2.50 mol of Zn.
2. Calculate the number of molecules in 11.5 mol of water (H_2O).
3. Silver nitrate ($AgNO_3$) is used to make several different silver halides used in photographic films. How many formula units of $AgNO_3$ are there in 3.25 mol of $AgNO_3$?
4. **CHALLENGE** Calculate the number of oxygen atoms in 5.00 mol of oxygen molecules. Oxygen is a diatomic molecule, O_2.

Particles to moles

Now suppose you want to find out how many moles are represented by a certain number of representative particles. To do this, you can use the inverse of Avogadro's number as a conversion factor.

$$\text{number of representative particles} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ representative particles}}$$

$$= \text{number of moles}$$

For example, if instead of knowing how many moles of sucrose you have, suppose you knew that a sample contained 2.11×10^{24} molecules of sucrose. To convert this number of molecules of sucrose to moles of sucrose, you need a conversion factor that has moles in the numerator and molecules in the denominator.

$$2.11 \times 10^{24} \cancel{\text{molecules sucrose}} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \cancel{\text{molecules sucrose}}}$$

$$= 3.50 \text{ mol sucrose}$$

Thus, 2.11×10^{24} molecules of sucrose is 3.50 mol of sucrose.

You can convert between moles and number of representative particles by multiplying the known quantity by the proper conversion factor. Example Problem 1 further illustrates the conversion process.

Get It?

List the two conversion factors that can be written from Avogadro's number.

EXAMPLE Problem 1

PARTICLES-TO-MOLES CONVERSION Zinc (Zn) is used as a corrosion-resistant coating on iron and steel. It is also an essential trace element in your diet. Calculate the number of moles of zinc that contain 4.50×10^{24} atoms.

1 ANALYZE THE PROBLEM

You are given the number of atoms of zinc and must find the equivalent number of moles. If you compare 4.50×10^{24} atoms Zn with 6.02×10^{23}, the number of atoms in 1 mol, you can predict that the answer shoud be less than 10 mol.

Known

number of atoms = 4.50×10^{24} atoms Zn

1 mol Zn = 6.02×10^{23} atoms Zn

Unknown

moles Zn = ? mol

2 SOLVE FOR THE UNKNOWN

Use a conversion factor—the inverse of Avogadro's number—that relates moles to atoms.

$$\text{number of atoms} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ atoms}} = \text{number of moles}$$

Apply the conversion factor.

$$4.50 \times 10^{24} \cancel{\text{atoms Zn}} \times \frac{1 \text{ mol Zn}}{6.02 \times 10^{23} \cancel{\text{atoms Zn}}} = 7.48 \text{ mol Zn}$$

Substitute number of Zn atoms = 4.50×10^{24}. Multiply and divide numbers and units.

EXAMPLE Problem 1 (continued)

3 EVALUATE THE ANSWER

Both the number of Zn atoms and Avogadro's number have three significant figures. Therefore, the answer is expressed correctly with three digits. The answer is less than 10 mol, as predicted, and has the correct unit, moles.

PRACTICE Problems

ADDITIONAL PRACTICE

5. How many moles contain each of the following?

a. 5.75×10^{24} atoms Al

b. 2.50×10^{20} atoms Fe

6. CHALLENGE Identify the representative particle for each formula, and convert the given number of representative particles to moles.

a. 3.75×10^{24} CO_2

b. 3.58×10^{23} $ZnCl_2$

Check Your Progress

Summary

- Some systems cannot be studied directly, as they are too small or large, or too fast or slow.
- The mole is a unit used to count particles of matter indirectly. One mole of a pure substance contains Avogadro's number of representative particles.
- Conversion factors written from Avogadro's relationship can be used to convert between moles and number of representative particles.
- Representative particles include atoms, ions, molecules, formula units, electrons, and other similar particles.
- One mole of carbon-12 atoms has a mass of exactly 12 g.

Demonstrate Understanding

7. **Explain** why chemists use the mole.
8. **State** the mathematical relationship between Avogadro's number and 1 mol.
9. **List** conversion factors used to convert between particles and moles.
10. **Explain** how a mole is similar to a dozen.
11. **Apply** How does a chemist count the number of particles in a given number of moles of a substance?
12. **Calculate** the mass in atomic mass units of 0.25 mol of carbon-12 atoms.
13. **Calculate** the number of representative particles of each substance.
    a. 11.5 mol of Ag
    b. 18.0 mol H_2O
    c. 0.150 mol NaCl
    d. 1.35×10^{-2} mol CH_4
14. **Arrange** these three samples from smallest to largest in terms of number of representative particles: 1.25×10^{25} atoms of zinc (Zn), 3.56 mol of iron (Fe), and 6.78×10^{22} molecules of glucose ($C_6H_{12}O_6$).

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LESSON 2

MASS AND THE MOLE

FOCUS QUESTION

What is the mass of a mole?

The Mass of a Mole

You would not expect a dozen limes to have the same mass as a dozen eggs. Since eggs and limes differ in size and composition, they have different masses, as shown in **Figure 5.** One-mole quantities of two different substances have different masses for the same reason—the substances have different compositions. This difference in mass occurs because carbon atoms differ from copper atoms. Thus, the mass of 6.02×10^{23} carbon atoms does not equal the mass of 6.02×10^{23} copper atoms.

Recall that each atom of carbon-12 has a mass of 12 amu. The atomic masses of all other elements are established relative to carbon-12. For example, an atom of hydrogen-1 has a mass of approximately 1 amu, one-twelfth the mass of a carbon-12 atom.

You will notice, however, that the atomic-masses on the periodic table are not exact integers. For example, you will find 12.011 amu for carbon and 1.008 amu for hydrogen. These noninteger values occur because the values are weighted averages of the masses of all the naturally occurring isotopes of each element.

Some elements have two naturally occuring isotopes, while others have many more than two. Usually, elements have a relatively high percentage of one isotope when compared to the other isotopes. This is why an element such as carbon has a mass of 12.011 amu, as the isotope carbon-12 is the most abundant isotope.

Figure 5 A dozen limes have approximately twice the mass of one dozen eggs. The difference in mass is reasonable because limes are different from eggs in composition and size.

3D THINKING | DCI Disciplinary Core Ideas | CCC Crosscutting Concepts | SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**
Create a table of the crosscutting concepts and fill in examples you find as you read.

? **Revisit the Encounter the Phenomenon Question**
What information from this lesson can help you answer the module question?

Molar mass

Figure 6 One mole of iron contains Avogadro's number of atoms and has a mass equal to its atomic mass in grams.

How does the mass of one atom relate to the mass of one mole of that atom? Recall that the mole is defined as the number of carbon-12 atoms in exactly 12 g of pure carbon-12. Thus, the mass of one mole of carbon-12 atoms is 12 g. The masses of all atoms are established relative to the mass of carbon-12. The mass in grams of one mole of any pure substance is called its **molar mass.**

The molar mass of any element is numerically equal to its atomic mass and has the units g/mol. As given on the periodic table, an atom of iron has an atomic mass of 55.845 amu. Thus, the molar mass of iron is 55.845 g/mol, and 1 mol (or 6.02×10^{23} atoms of iron) has a mass of 55.845 g. Note that by measuring 55.845 g of iron, you indirectly count out 6.02×10^{23} atoms of iron. **Figure 6** shows the relationship between molar mass and one mole of an element.

Using Molar Mass

Imagine that your class bought jelly beans in bulk to sell by the dozen at a candy sale. You soon realize that it is too much work to count out each dozen, so you instead decide to measure the jelly beans by mass. You find that the mass of 1 dozen jelly beans is 35 g. This relationship and the conversion factors that stem from it are as follows:

$$1 \text{ dozen jelly beans} = 35 \text{ g jelly beans}$$

$$\frac{35 \text{ g jelly beans}}{1 \text{ dozen jelly beans}} \text{ and } \frac{1 \text{ dozen jelly beans}}{35 \text{ g jelly beans}}$$

What mass of jelly beans should you measure if a customer wants 5 dozen jelly beans? To determine this mass, you would multiply the number of dozens of jelly beans to be sold by the correct conversion factor. Select the conversion factor with the units you are converting to in the numerator (g) and the units you are converting from in the denominator (dozen).

$$5 \cancel{\text{ dozen jelly beans}} \times \frac{35 \text{ g jelly beans}}{1 \cancel{\text{ dozen jelly beans}}} = 175 \text{ g jelly beans}$$

A quantity of 5 dozen jelly beans has a mass of 175 g.

Suppose you want instead to convert from mass to a number. For example, how many dozen jelly beans are in 105 g?

$$105 \cancel{\text{ g jelly beans}} \times \frac{1 \text{ dozen jelly beans}}{35 \cancel{\text{ g jelly beans}}} = 3 \text{ dozen jelly beans}$$

105 g of jelly beans represents 3 dozen jelly beans.

Get It?

Compare How are the jelly bean conversion factors used above similar to the molar mass of a compound?

Moles to mass

Now suppose that while working in a chemistry lab, you need 3.00 mol of copper (Cu) for a chemical reaction. How would you measure that amount? Like the 5 dozen jelly beans, the number of moles of copper can be converted to an equivalent mass and measured on a balance.

To calculate the mass of a given number of moles, simply multiply the number of moles by the molar mass.

$$\text{number of moles} \times \frac{\text{mass in grams}}{1 \text{ mole}} = \text{mass}$$

If you check the periodic table, you will find that copper, element 29, has an atomic mass of 63.546 amu. You know that the molar mass of an element (in g/mol) is equal to its atomic mass (given in amu). Thus, copper has a molar mass of 63.546 g/mol. By using the molar mass, you can convert 3.00 mol of copper to grams of copper.

$$3.00\ \cancel{\text{mol Cu}} \times \frac{63.546 \text{ g Cu}}{1\ \cancel{\text{mol Cu}}} = 191 \text{ g Cu}$$

So, as shown in **Figure 7**, you can measure the 3.00 mol of copper needed for the reaction by using a balance to measure out 191 g of copper. The reverse conversion—from mass to moles—also involves the molar mass as a conversion factor, but it is the inverse of the molar mass that is used. Can you explain why?

BIOLOGY Connection Biologists in the fields of biochemistry and molecular biology study molecules—often very large molecules, such as DNA, RNA, and proteins—that are involved in biological processes. When biologists discover a new protein, they can use a technique known as mass spectrometry to determine the molar mass of the protein. The molar mass, together with information gained through other techniques, can then be used to determine the composition of the protein. By applying this kind of analysis to separate parts of a large molecule, the structure of the whole molecule can be pieced together.

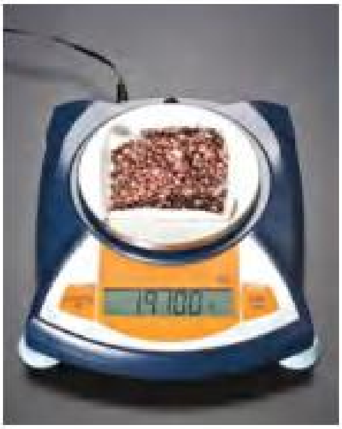

Figure 7 To measure 3.00 mol of copper, place a weighing paper on a balance, tare the balance, and then add 191 g of copper.

EXAMPLE Problem 2

MOLE-TO-MASS CONVERSION Chromium (Cr), a transition element, is a component of chrome plating. Chrome plating is used on metals and in steel alloys to control corrosion. Calculate the mass in grams of 0.0450 mol Cr.

1 ANALYZE THE PROBLEM

You are given the number of moles of chromium and must convert it to an equivalent mass using the molar mass of chromium from the periodic table. Because the sample is less than one-tenth of a mole, the answer should be less than one-tenth of the molar mass.

Known	Unknown
number of moles = 0.0450 mol Cr	mass Cr = ? g
molar mass Cr = 52.00 g/mol Cr	

EXAMPLE Problem 2 (continued)

2 SOLVE FOR THE UNKNOWN

Use a conversion factor—the molar mass—that relates grams of chromium to moles of chromium. Write the conversion factor with moles of chromium in the denominator and grams of chromium in the numerator. Substitute the known values into the equation and solve.

$\text{moles Cr} \times \frac{\text{grams Cr}}{1 \text{ mol Cr}} = \textbf{grams Cr}$ Apply the conversion factor.

$0.0450 \cancel{\text{mol Cr}} \times \frac{52.00 \text{ g Cr}}{1 \cancel{\text{mol Cr}}} = \textbf{2.34 g Cr}$ Substitute 0.450 mol for moles Cr and 52.00 g/mol for molar mass of Cr. Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The known number of moles of chromium has the smallest number of significant figures, three, so the answer is correctly stated with three digits. The answer is less than one-tenth the mass of 1 mol, as predicted, and is in grams.

Many of the values for atomic mass given in the periodic table have five significant figures. However, in Example Problem 2, the periodic table value of 51.996 g/mol Cr was rounded to 52.00 g/mol. It is generally okay to round reference values as long as you keep one more significant figure than the answer will have. In the case of Example Problem 2, 52.00 has one more significant figure than 0.0450 mol Cr, which limits the answer to three significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

15. Determine the mass in grams of each of the following.

a. 3.57 mol Al

b. 4.26 mol Si

16. CHALLENGE Convert each given quantity in scientific notation to mass in grams expressed in scientific notation.

a. 3.45×10^2 mol Co

b. 2.45×10^{-2} mol Zn

EXAMPLE Problem 3

MASS-TO-MOLE CONVERSION Calcium (Ca), the fifth most-abundant element on Earth, is always found combined with other elements because of its high reactivity. How many moles of calcium are in 525 g Ca?

1 ANALYZE THE PROBLEM

You must convert the mass of calcium to moles of calcium. The mass of calcium is more than ten times larger than the molar mass. Therefore, the answer should be greater than 10 mol.

Known	Unknown
mass = 525 g Ca	number of moles Ca = ? mol
molar mass Ca = 40.08 g/mol Ca	

EXAMPLE Problem 3 (continued)

2 SOLVE FOR THE UNKNOWN

Use a conversion factor—the inverse of molar mass—that relates moles of calcium to grams of calcium. Substitute the known values and solve.

$$\text{mass Ca} \times \frac{1 \text{ mol Ca}}{\text{grams Ca}} = \textbf{moles Ca}$$

Apply the conversion factor.

$$525 \text{ g } \cancel{\text{Ca}} \times \frac{1 \text{ mol Ca}}{40.08 \text{ g } \cancel{\text{Ca}}} = \textbf{13.1 mol Ca}$$

Substitute mass Ca = 525 g, and inverse molar mass of Ca = 1 mol/40.08 g. Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The mass of calcium has the fewest significant figures, three, so the answer is expressed correctly with three digits. As predicted, the answer is greater than 10 mol and has the expected unit.

PRACTICE Problems ADDITIONAL PRACTICE

17. Determine the number of moles in each of the following.

a. 25.5 g Ag **b.** 300.0 g S

18. CHALLENGE Convert each mass to moles. Express the answer in scientific notation.

a. 1.25×10^3 g Zn **b.** 1.00 kg Fe

Converting between mass and atoms

So far, you have learned how to convert mass to moles and moles to mass. You can go one step further and convert mass to the number of atoms. Recall the jelly beans you were selling at the candy sale. At the end of the day, you find that 550 g of jelly beans is left unsold. Without counting, can you determine how many jelly beans that is? You know that 1 dozen jelly beans has a mass of 35 g and that 1 dozen is 12 jelly beans. Thus, you can first convert the 550 g to dozens of jelly beans by using the conversion factor that relates dozens and mass.

$$550 \cancel{\text{ g jelly beans}} \times \frac{1 \text{ dozen jelly beans}}{35 \cancel{\text{ g jelly beans}}} = 16 \text{ dozen jelly beans}$$

Next, you can determine how many jelly beans are in 16 dozen by multiplying by the conversion factor that relates number of particles (jelly beans) and dozens. The conversion factor relating number of jelly beans and dozens is 12 jelly beans/dozen. Applying it yields the answer in jelly beans.

$$16 \cancel{\text{ dozen}} \times \frac{12 \text{ jelly beans}}{1 \cancel{\text{ dozen}}} = 192 \text{ jelly beans}$$

The 550 g of leftover jelly beans is equal to 192 jelly beans.

Just as you could not make a direct conversion from the mass of jelly beans to the number of jelly beans, you cannot make a direct conversion from the mass of a substance to the number of representative particles of that substance. You must first convert mass to moles by multiplying by a conversion factor that relates moles and mass.

That conversion factor is the molar mass. The number of moles must then be multiplied by a conversion factor that relates the number of representative particles to moles. For this conversion, you will use Avogadro's number.

EXAMPLE Problem 4

MASS-TO-ATOMS CONVERSION Gold (Au) is one of a group of metals called the coinage metals (copper, silver, and gold). How many atoms of gold are in a U.S. Eagle, a gold alloy bullion coin with a mass of 31.1 g Au?

1 ANALYZE THE PROBLEM

You must determine the number of atoms in a given mass of gold. Because you cannot convert directly from mass to the number of atoms, you must first convert the mass to moles using the molar mass. Then, convert moles to the number of atoms using Avogadro's number. The given mass of the gold coin is about one-sixth the molar mass of gold (196.97 g/mol), so the number of gold atoms should be approximately one-sixth Avogadro's number.

Known	Unknown
mass = 31.1 g Au	number of atoms Au = ?
molar mass Au = 196.97 g/mol Au	

2 SOLVE FOR THE UNKNOWN

Use a conversion factor—the inverse of the molar mass—that relates moles of gold to grams of gold.

$$\text{mass Au} \times \frac{1 \text{ mol Au}}{\text{grams Au}} = \text{moles Au}$$

Apply the conversion factor.

$$31.1 \text{ g } \cancel{\text{Au}} \times \frac{1 \text{ mol Au}}{196.97 \text{ g } \cancel{\text{Au}}} = 0.158 \text{ mol Au}$$

Substitute mass Au = 31.1 g and the inverse molar mass of Au = 1 mol/196.97 g. Multiply and divide numbers and units.

To convert the calculated moles of gold to atoms, multiply by Avogadro's number.

$$\text{moles Au} \times \frac{6.02 \times 10^{23} \text{ atoms Au}}{1 \text{ mol Au}} = \text{atoms Au}$$

Apply the conversion factor.

$$0.158 \cancel{\text{mol Au}} \times \frac{6.02 \times 10^{23} \text{ atoms Au}}{1 \cancel{\text{mol Au}}} = 9.51 \times 10^{22} \text{ atoms Au}$$

Substitute moles Au = 0.153 mol, and solve.

3 EVALUATE THE ANSWER

The mass of gold has the smallest number of significant figures, three, so the answer is expressed correctly with three digits. The answer is approximately one-sixth Avogadro's number, as predicted, and the correct unit, atoms, is obtained.

EXAMPLE Problem 5

ATOMS-TO-MASS CONVERSION Helium (He) is an unreactive noble gas often found in underground deposits mixed with methane. The mixture is separated by cooling the gaseous mixture until all but the helium has liquefied. A party balloon contains 5.50×10^{22} atoms of helium gas. What is the mass, in grams, of the helium?

1 ANALYZE THE PROBLEM

You are given the number of atoms of helium and must find the mass of the gas. First, convert the number of atoms to moles, then convert moles to grams.

Known	Unknown
number of atoms He = 5.50×10^{22} atoms He	mass = ? g He
molar mass He = 4.00 g/mol He	

EXAMPLE Problem 5 (continued)

2 SOLVE FOR THE UNKNOWN

Use a conversion factor—the inverse of Avogadro's number—that relates moles to number of atoms.

$$\text{atoms He} \times \frac{1 \text{ mol He}}{6.02 \times 10^{23} \text{ atoms He}} = \text{moles He}$$

Apply the conversion factor.

$$5.50 \times 10^{22} \cancel{\text{atoms He}} \times \frac{1 \text{ mol He}}{6.02 \times 10^{23} \cancel{\text{atoms He}}} = 0.0914 \text{ mol He}$$

Substitute atoms He = 5.50×10^{22} atoms. Multiply and divide numbers and units.

Next, apply a conversion factor—the molar mass of helium—that relates mass of helium to moles of helium.

$$\text{moles He} \times \frac{\text{grams He}}{1 \text{ mol He}} = \text{grams He}$$

Apply the conversion factor.

$$0.0914 \cancel{\text{mol He}} \times \frac{4.00 \text{ g He}}{1 \cancel{\text{mol He}}} = 0.366 \text{ g He}$$

Substitute moles He = 0.0914 mol, molar mass He = 4.00 g/mol, and solve.

3 EVALUATE THE ANSWER

The answer is expressed correctly with three significant figures and is in grams, a mass unit.

PRACTICE Problems ADDITIONAL PRACTICE

19. How many atoms are in each of the following samples?

a. 55.2 g Li **b.** 0.230 g Pb **c.** 11.5 g Hg

20. What is the mass in grams of each of the following?

a. 6.02×10^{24} atoms Bi **b.** 1.00×10^{24} atoms Mn **c.** 3.40×10^{22} atoms He

d. 1.50×10^{15} atoms N **e.** 1.50×10^{15} atoms U

21. CHALLENGE Convert each given mass to number of representative particles. Identify the type of representative particle, and express the number in scientific notation.

a. 4.56×10^{3} g Si **b.** 0.120 kg Ti

You should now realize that the mole is at the center of your calculations. Mass must always be converted to moles before being converted to atoms, and atoms must be converted to moles before calculating their mass.

Figure 8, shown on the next page, shows the steps to follow as you complete these conversions. In the Example Problems, two steps were used. However, these conversions can be made in one step by combining the steps together in one calculation. If you want to find out how many atoms of oxygen (O) are in 1.00 g of oxygen, two conversions are needed—mass to moles and then moles to atoms. You could set up one equation like this.

$$1.00 \cancel{\text{g O}} \times \frac{1 \cancel{\text{mol O}}}{15.999 \text{ g O}} \times \frac{6.02 \times 10^{23} \text{ atoms O}}{1 \cancel{\text{mol O}}}$$

$$= 3.76 \times 10^{22} \text{ atoms O}$$

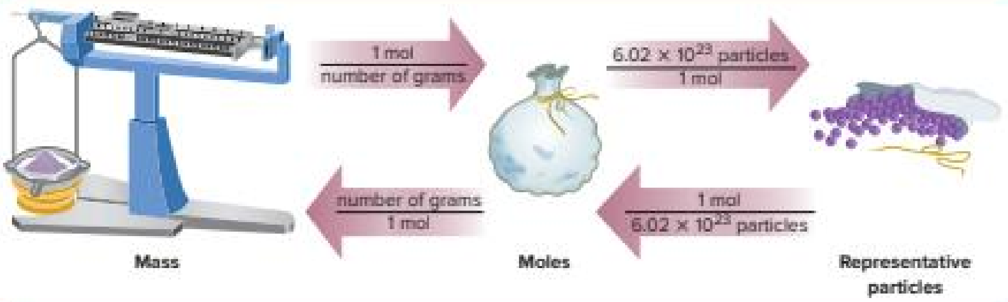

Figure 8 The mole is at the center of conversions between mass and particles (atoms, ions, or molecules). In the figure, mass is represented by a balance, moles by a bag of particles, and representative particles by the contents that are spilling out of the bag. Two steps are needed in the conversions.

The result of this calculation will be the exact same value as you would obtain if you were to do the calculation in two steps. A major advantage to this one step calculation is that you do not round off your answer until the final value is obtained. When this calculation is done in two steps, the rounding that occurs after the first step can affect the value of the final answer slightly.

Check Your Progress

Summary

- The mass in grams of one mole of any pure substance is called its molar mass.
- The molar mass of an element is numerically equal to its atomic mass.
- The molar mass of any substance is the mass in grams of Avogadro's number of representative particles of the substance.
- Molar mass is used to convert from moles to mass. The inverse of molar mass is used to convert from mass to moles.
- Understanding systems on an atomic scale leads to the ability of scientists to predict cause and effect relationships on a larger scale.

Demonstrate Understanding

22. **Summarize** why a scientist's understanding of a system at an atomic scale is important in identifying system relationships at a larger scale.
23. **State** the conversion factor needed to convert between mass and moles of the atom fluorine.
24. **Explain** how molar mass relates the mass of an atom to the mass of a mole of atoms.
25. **Describe** the steps used to convert the mass of an element to the number of atoms of the element.
26. **Arrange** these quantities from smallest to largest in terms of mass: 1.0 mol of Ar, 3.0×10^{24} atoms of Ne, and 20 g of Kr.
27. **Identify** the quantity that is calculated by dividing the molar mass of an element by Avogadro's number. Show how units are used to ensure the correct conversion factor is used to obtain the desired result.
28. **Design** a concept map that shows how conversion factors are needed to convert between mass, moles, and the number of representative particles in multi-step problems.

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LESSON 3

MOLES OF COMPOUNDS

FOCUS QUESTION

What do chemical formulas say about mole relationships?

Chemical Formulas and the Mole

Different kinds of representative particles are counted using the mole. In the last lesson, you read how to use molar mass to convert among moles, mass, and number of particles of an element. Similar conversions for compounds and ions also exist, but you need to know the molar mass of the compounds and ions involved.

Recall that a chemical formula indicates the numbers and types of atoms contained in one unit of a compound. Consider the compound dichlorodifluoromethane, CCl_2F_2. The subscripts in the formula indicate that one molecule of CCl_2F_2 consists of one carbon (C) atom, two chlorine (Cl) atoms, and two fluorine (F) atoms chemically bonded together. The C-Cl-F ratio in CCl_2F_2 is 1:2:2. A mole of CCl_2F_2 contains Avogadro's number of molecules. The C-Cl-F ratio in one mole of CCl_2F_2 would still be 1:2:2. **Figure 9** illustrates this for a dozen CCl_2F_2 molecules.

Check for yourself that a dozen CCl_2F_2 molecules contains one dozen carbon atoms, two dozen chlorine atoms, and two dozen fluorine atoms. The chemical formula CCl_2F_2 not only represents an individual molecule of CCl_2F_2, it also represents a mole of the compound.

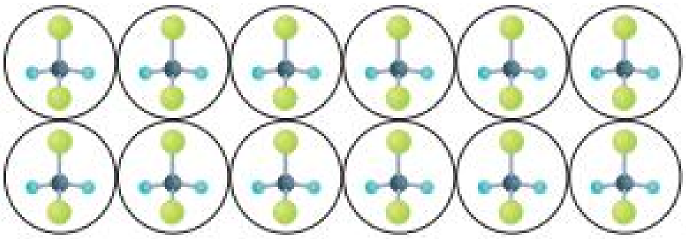


Figure 9 A dozen molecules of CCl_2F_2 contains one dozen carbon atoms, two dozen chlorine atoms, and two dozen fluorine atoms. How many of each kind of atom-carbon, chlorine, and fluorine-are contained in 1 mol of CCl_2F_2.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Virtual Investigation: Moles, Mass, and Molecules
Use mathematics and computational thinking to determine the quantity of moles in each substance.

Laboratory: Estimating the Size of a Mole
Use mathematics and computational thinking to estimate the quantity of a mole.

In some chemical calculations, you might need to convert between moles of a compound and moles of individual atoms in the compound. The following ratios, or conversion factors, can be written for use in these calculations for the molecule CCl_2F_2.

$$\frac{1 \text{ mol C atoms}}{1 \text{ mol } CCl_2F_2} \qquad \frac{2 \text{ mol Cl atoms}}{1 \text{ mol } CCl_2F_2} \qquad \frac{2 \text{ mol F atoms}}{1 \text{ mol } CCl_2F_2}$$

To find out how many moles of fluorine atoms are in 5.50 moles of freon, you multiply the moles by the conversion factor relating moles of fluorine atoms to moles.

$$\text{moles } CCl_2F_2 \times \frac{\text{moles F atoms}}{1 \text{ mol } CCl_2F_2} = \text{moles F atoms}$$

$$5.50 \cancel{\text{mol } CCl_2F_2} \times \frac{2 \text{ mol F atoms}}{1 \cancel{\text{mol } CCl_2F_2}} = 11.0 \text{ mol F atoms}$$

Conversion factors such as the one just used for fluorine can be written for any element in a compound. The number of moles of the element that goes in the numerator of the conversion factor is the subscript for that element in the chemical formula.

EXAMPLE Problem 6

MOLE RELATIONSHIPS FROM A CHEMICAL FORMULA Aluminum oxide (Al_2O_3), often called alumina, is the principal raw material for the production of aluminum (Al). Alumina occurs in the minerals corundum and bauxite. Determine the moles of aluminum ions (Al^{3+}) in 1.25 mol of Al_2O_3.

1 ANALYZE THE PROBLEM

You are given the number of moles of Al_2O_3 and must determine the number of moles of Al^{3+} ions. Use a conversion factor based on the chemical formula that relates moles of Al^{3+} ions to moles of Al_2O_3. Every mole of Al_2O_3 contains 2 mol of Al^{3+} ions. Thus, the answer should be two times the number of moles of Al_2O_3.

Known
number of moles = 1.25 mol Al_2O_3

Unknown
number of moles = ? mol Al^{3+} ions

2 SOLVE FOR THE UNKNOWN

Use the relationship that 1 mol of Al_2O_3 contains 2 mol of Al^{3+} ions to write a conversion factor.

$$\frac{2 \text{ mol } Al^{3+} \text{ ions}}{1 \text{ mol } Al_2O_3}$$

Create a conversion factor relating moles of Al^{3+} ions to moles of Al_2O_3.

To convert the known number of moles of Al_2O_3 to moles of Al^{3+} ions, multiply by the ions-to-moles conversion factor.

$$\text{moles } Al_2O_3 \times \frac{2 \text{ mol } Al^{3+} \text{ ions}}{1 \text{ mol } Al_2O_3} = \text{moles } Al^{3+} \text{ ions}$$

Apply the conversion factor.

$$1.25 \cancel{\text{mol } Al_2O_3} \times \frac{2 \text{ mol } Al^{3+} \text{ ions}}{1 \cancel{\text{mol } Al_2O_3}} = 2.50 \text{ mol } Al^{3+} \text{ ions}$$

Substitute moles Al_2O_3 = 1.25 mol Al_2O_3 and solve.

3 EVALUATE THE ANSWER

Because the conversion factor is a ratio of whole numbers, the number of significant digits is based on the moles of Al_2O_3. Therefore, the answer is expressed correctly with three significant figures. As predicted, the answer is twice the number of moles of Al_2O_3.

PRACTICE Problems ADDITIONAL PRACTICE

29. Zinc chloride ($ZnCl_2$) is used in soldering flux, an alloy used to join two metals together. Determine the moles of Cl^- ions in 2.50 mol $ZnCl_2$.
30. Plants and animals depend on glucose ($C_6H_{12}O_6$) as an energy source. Calculate the number of moles of each element in 1.25 mol $C_6H_{12}O_6$.
31. Iron(III) sulfate [$Fe_2(SO_4)_3$] is sometimes used in the water purification process. Determine the number of moles of sulfate ions present in 3.00 mol of $Fe_2(SO_4)_3$.
32. How many moles of oxygen atoms are present in 5.00 mol of diphosphorus pentoxide (P_2O_5)?
33. **CHALLENGE** Calculate the number of moles of hydrogen atoms in 1.15×10^3 mol of water. Express the answer in scientific notation.

The Molar Mass of Compounds

The mass of your backpack is the sum of the mass of the pack and the masses of the books, notebooks, pencils, lunch, and miscellaneous items you put into it. Similarly, the mass of a mole of a compound equals the sum of the masses of all the particles that make up the compound.

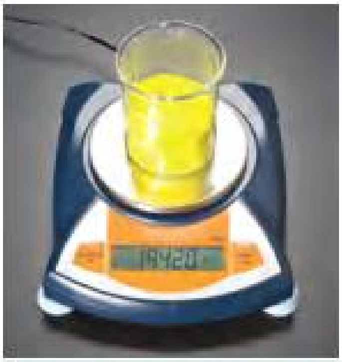
Suppose you want to determine the molar mass of the compound potassium chromate (K_2CrO_4). Start by looking up the molar mass of each element present in K_2CrO_4. Then, multiply each molar mass by the number of moles of that element in the chemical formula. Adding the masses of each element yields the molar mass of K_2CrO_4.

$$2\ \cancel{\text{mol K}} \times \frac{39.10\ \text{g K}}{1\ \cancel{\text{mol K}}} = 78.20\ \text{g}$$

$$1\ \cancel{\text{mol Cr}} \times \frac{52.00\ \text{g Cr}}{1\ \cancel{\text{mol Cr}}} = 52.00\ \text{g}$$

$$4\ \cancel{\text{mol O}} \times \frac{16.00\ \text{g O}}{1\ \cancel{\text{mol O}}} = 64.00\ \text{g}$$

$$\text{molar mass } K_2CrO_4 = 194.20\ \text{g}$$



Potassium chromate (K_2CrO_4)

Figure 10 The molar mass of each compound is the sum of the masses of elements contained in the compound.

The molar mass of a compound demonstrates the law of conservation of mass; the total mass of the reactants that reacted equals the mass of the compound formed. The fact that atoms are conserved in a reaction, together with knowledge of the chemical properties of the elements, are used to describe and predict chemical reactions. **Figure 10** shows the mass of one mole of potassium chromate.

CCC CROSSCUTTING CONCEPTS

Energy and Matter Show that the molar mass of iron(III) sulfate, and the individual ions that form this compound, demonstrate that the total amount of matter in a closed system is conserved. Cite your evidence by using the following reactants and products as an example: $2Fe^{+3} + 3SO_4^{-2} \rightarrow Fe_2(SO_4)_3$

STEM CAREER Connection

Materials Scientist

Are you interested in what things are made of? Materials scientists study substances at the atomic and molecular level. They investigate how substances interact with one another. If you go to college and graduate with a bachelor's degree, you can participate in developing new and improved products and technologies that will improve the way we live.

PRACTICE Problems ADDITIONAL PRACTICE

34. Determine the molar mass of each ionic compound.

a. NaOH **b.** $CaCl_2$ **c.** $KC_2H_3O_2$

35. Calculate the molar mass of each molecular compound.

a. C_2H_5OH **b.** HCN **c.** CCl_4

36. CHALLENGE Identify each substance as a molecular compound or an ionic compound, and then calculate its molar mass.

a. $Sr(NO_3)_2$ **b.** $(NH_4)_3PO_4$ **c.** $C_{12}H_{22}O_{11}$

Converting Moles of a Compound to Mass

Suppose you need to measure a certain number of moles of a compound for an experiment. First, you must calculate the mass in grams that corresponds to the necessary number of moles. Then, you can measure that mass on a balance. In Example Problem 2, you learned how to convert the number of moles of elements to mass using molar mass as the conversion factor. The procedure is the same for compounds, except that you must first calculate the molar mass of the compound.

EXAMPLE Problem 7

MOLE-TO-MASS CONVERSION FOR COMPOUNDS The characteristic odor of garlic is due to allyl sulfide [$(C_3H_5)_2S$]. What is the mass of 2.50 mol of $(C_3H_5)_2S$?

1 ANALYZE THE PROBLEM

You are given 2.50 mol of $(C_3H_5)_2S$ and must convert the moles to mass using the molar mass as a conversion factor. The molar mass is the sum of the molar masses of all the elements in $(C_3H_5)_2S$.

Known
number of moles = 2.50 mol $(C_3H_5)_2S$

Unknown
molar mass = ? g/mol $(C_3H_5)_2S$
mass = ? g $(C_3H_5)_2S$

2 SOLVE FOR THE UNKNOWN

Calculate the molar mass of $(C_3H_5)_2S$.

$$1\ \cancel{\text{mol S}} \times \frac{32.07\ \text{g S}}{1\ \cancel{\text{mol S}}} = 32.07\ \text{g S}$$

Multiply the moles of S in the compound by the molar mass of S.

$$6\ \cancel{\text{mol C}} \times \frac{12.01\ \text{g C}}{1\ \cancel{\text{mol C}}} = 72.06\ \text{g C}$$

Multiply the moles of C in the compound by the molar mass of C.

$$10\ \cancel{\text{mol H}} \times \frac{1.008\ \text{g H}}{1\ \cancel{\text{mol H}}} = 10.08\ \text{g H}$$

Multiply the moles of H in the compound by the molar mass of H.

molar mass = (32.07 g + 72.06 g + 10.08 g)
= 114.21 g/mol $(C_3H_5)_2S$

Total the mass values.

Use a conversion factor—the molar mass—that relates grams to moles.

$$\text{moles } (C_3H_5)_2S \times \frac{\text{grams } (C_3H_5)_2S}{1\ \text{mol } (C_3H_5)_2S} = \text{mass } (C_3H_5)_2S$$

Apply the conversion factor.

$$250\ \cancel{\text{moles } (C_3H_5)_2S} \times \frac{114.21\ \text{g } (C_3H_5)_2S}{1\ \cancel{\text{mol } (C_3H_5)_2S}} = 286\ \text{g } (C_3H_5)_2S$$

Substitute moles $(C_3H_5)_2S$ = 2.5 mol, molar mass $(C_3H_5)_2S$ = 114.21 g/mol, and solve.

PRACTICE Problems ADDITIONAL PRACTICE

37. The United States chemical industry produces more sulfuric acid (H_2SO_4), in terms of mass, than any other chemical. What is the mass of 3.25 mol of H_2SO_4?

38. What is the mass of 4.35×10^{-2} mol of zinc chloride ($ZnCl_2$)?

39. CHALLENGE Write the chemical formula for potassium permanganate, and then calculate the mass in grams of 2.55 mol of the compound.

Converting Mass of a Compound to Moles

Imagine that an experiment you are doing in the laboratory produces 5.55 g of a compound. How many moles is this? To find out, you calculate the molar mass of the compound and determine it to be 185.0 g/mol. The molar mass relates grams and moles, but this time you need the inverse of the molar mass as the conversion factor.

$$5.50\ \cancel{\text{g compound}} \times \frac{1\ \text{mol compound}}{185.0\ \cancel{\text{g compound}}} = 0.0297\ \text{mol compound}$$

EXAMPLE Problem 8

MASS-TO-MOLE-CONVERSION FOR COMPOUNDS Calcium hydroxide [$Ca(OH)_2$] is used to remove sulfur dioxide from the exhaust gases emitted by power plants and for softening water by the elimination of Ca^{2+} and Mg^{2+} ions. Calculate the number of moles of calcium hydroxide in 325 g of the compound.

1 ANALYZE THE PROBLEM

You are given 325 g of $Ca(OH)_2$ and must solve for the number of moles of $Ca(OH)_2$. You must first calculate the molar mass of $Ca(OH)_2$.

Known
mass = 325 g $Ca(OH)_2$

Unknown
molar mass = ? g/mol $Ca(OH)_2$
number of moles = ? mol $Ca(OH)_2$

2 SOLVE FOR THE UNKNOWN

Determine the molar mass of $Ca(OH)_2$.

$$1\ \cancel{\text{mol Ca}} \times \frac{40.08\ \text{g Ca}}{1\ \cancel{\text{mol Ca}}} = 40.08\ \text{g Ca}$$

Multiply the moles of Ca in the compound by the molar mass of Ca.

$$2\ \cancel{\text{mol O}} \times \frac{16.00\ \text{g O}}{1\ \cancel{\text{mol O}}} = 32.00\ \text{g O}$$

Multiply the moles of O in the compound by the molar mass of O.

$$2\ \cancel{\text{mol H}} \times \frac{1.008\ \text{g H}}{1\ \cancel{\text{mol H}}} = 2.016\ \text{g H}$$

Multiply the moles of H in the compound by the molar mass of H.

molar mass = (40.08 g + 32.00 g + 2.016 g)
= **74.10 g/mol $Ca(OH)_2$**

Total the mass values.

Use a conversion factor—the inverse of the molar mass—that relates moles to grams.

$$325\ \cancel{\text{g Ca(OH)}_2} \times \frac{1\ \text{mol Ca(OH)}_2}{74.10\ \cancel{\text{g Ca(OH)}_2}} = \mathbf{4.39\ \text{mol Ca(OH)}_2}$$

Apply the conversion factor. Substitute mass Ca = 325 g, inverse molar mass $Ca(OH)_2$ = 1 mol/74.10 g, and solve.

3 EVALUATE THE ANSWER

To check the reasonableness of the answer, round the molar mass of $Ca(OH)_2$ to 75 g/mol and the given mass of $Ca(OH)_2$ to 300 g. Seventy-five is contained in 300 four times. Thus, the answer is reasonable. The unit, moles, is correct, and there are three significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

40. Determine the number of moles present in each compound.

a. 22.6 g $AgNO_3$ b. 6.50 g $ZnSO_4$ c. 35.0 g HCl

41. CHALLENGE Identify each as an ionic or molecular compound and convert the given mass to moles. Express your answers in scientific notation.

a. 2.50 kg Fe_2O_3 b. 25.4 mg $PbCl_4$

Get It?

Explain how a chemical formula is used to determine the number of moles in a given mass of compound.

Converting Mass of a Compound to Number of Particles

Example Problem 8 illustrated how to find the number of moles of a compound contained in a given mass. Now, you will learn how to calculate the number of representative particles--molecules or formula units--contained in a given mass and, in addition, the number of atoms or ions.

No direct conversion is possible between mass and number of particles. You must first convert the given mass to moles. Then, you can convert moles to the number of representative particles. To determine the numbers of atoms or ions in a compound, you will need conversion factors that are ratios of the number of atoms or ions in the compound to 1 mol of compound. These are based on the chemical formula. Example Problem 9 provides practice in solving this type of problem.

EXAMPLE Problem 9

CONVERSION FROM MASS TO MOLES TO PARTICLES Aluminum chloride ($AlCl_3$) is used in refining petroleum and manufacturing rubber and lubricants. A sample of aluminum chloride has a mass of 35.6 g.

a. How many aluminum ions are present?

b. How many chloride ions are present?

c. What is the mass, in grams, of one formula unit of aluminum chloride?

1 ANALYZE THE PROBLEM

You are given 35.6 g of $AlCl_3$ and you must calculate the number of Al^{3+} ions, the number of Cl^- ions, and the mass in grams of one formula unit of $AlCl_3$. The ratio of Al^{3+} ions to Cl^- ions in the chemical formula is 1:3. Therefore, the calculated numbers of ions should be in that same ratio.

Known

mass = 35.6 g $AlCl_3$

Unknown

number of ions = ? Al^{3+} ions

number of ions = ? Cl^- ions

mass = ? g/formula unit $AlCl_3$

EXAMPLE Problem 9 (continued)

2 SOLVE FOR THE UNKNOWN

Determine the molar mass of $AlCl_3$.

$1\ \cancel{\text{mol Al}} \times \frac{26.98\text{ g Al}}{1\ \cancel{\text{mol Al}}} = 26.98\text{ g Al}$ — Multiply the moles of Al in the compound by the molar mass of Al.

$3\ \cancel{\text{mol Cl}} \times \frac{35.45\text{ g Cl}}{1\ \cancel{\text{mol Cl}}} = 106.35\text{ g Cl}$ — Multiply the moles of Cl in the compound by the molar mass of Cl.

molar mass $= (26.98\text{ g} + 106.35\text{ g}) = 133.33\text{ g/mol } AlCl_3$ — Total the molar mass values.

Use a conversion factor—the inverse of the molar mass—that relates moles to grams.

$\text{mass } AlCl_3 \times \frac{1\text{ mol } AlCl_3}{\text{grams } AlCl_3} = \text{moles } AlCl_3$ — Apply the conversion factor.

$35.6\ \cancel{\text{g } AlCl_3} \times \frac{1\text{ mol } AlCl_3}{133.33\ \cancel{\text{g } AlCl_3}} = 0.267\text{ mol } AlCl_3$ — Substitute mass $AlCl_3$ = 35.6 g and inverse molar mass $AlCl_3$ = 1 mol/133.33 g, and solve.

Use Avogadro's number.

$0.267\ \cancel{\text{mol } AlCl_3} \times \frac{6.02 \times 10^{23}\text{ formula units}}{1\ \cancel{\text{mol } AlCl_3}}$ — Multiply and divide numbers and units.

$= 1.61 \times 10^{23}\text{ formula units } AlCl_3$

To calculate the number of Al^{3+} and Cl^- ions, use the ratios from the chemical formula as conversion factors.

$1.61 \times 10^{23}\ \cancel{AlCl_3\text{ formula units}} \times \frac{1\ Al^{3+}\text{ ion}}{1\ \cancel{AlCl_3\text{ formula unit}}}$ — Multiply and divide numbers and units.

$= 1.61 \times 10^{23}\ Al^{3+}$ **ions**

$1.61 \times 10^{23}\ \cancel{AlCl_3\text{ formula units}} \times \frac{3\ Cl^-\text{ ions}}{1\ \cancel{AlCl_3\text{ formula unit}}}$ — Multiply and divide numbers and units.

$= 4.83 \times 10^{23}\ Cl^-$ **ions**

Calculate the mass in grams of one formula unit of $AlCl_3$. Use the inverse of Avogadro's number as a conversion factor.

$\frac{133.33\text{ g } AlCl_3}{1\ \cancel{\text{mol}}} \times \frac{1\ \cancel{\text{mol}}}{6.02 \times 10^{23}\text{ formula units}}$ — Substitute mass $AlCl_3$ = 133.33 g, and solve.

$= 2.21 \times 10^{-22}\text{ g } AlCl_3\text{/formula unit}$

3 EVALUATE THE ANSWER

A minimum of three significant figures is used in each value in the calculations. Therefore, the answers have the correct number of digits. The number of Cl^- ions is three times the number of Al^{3+} ions, as predicted. The mass of a formula unit of $AlCl_3$ can be checked by calculating it in a different way. Divide the sample mass of $AlCl_3$ (35.6 g) by the number of formula units contained in the mass (1.61×10^{23} formula units) to obtain the mass of one formula unit. The two answers are the same.

PRACTICE Problems ADDITIONAL PRACTICE

42. Ethanol (C_2H_5OH), a domestically produced fuel source, is often blended with gasoline. A sample of ethanol has a mass of 45.6 g.

- **a.** How many carbon atoms does the sample contain?
- **b.** How many hydrogen atoms are present?
- **c.** How many oxygen atoms are present?

43. A sample of sodium sulfite (Na_2SO_3) has a mass of 2.25 g.

- **a.** How many Na^+ ions are present?
- **b.** How many SO_3^{2-} ions are present?
- **c.** What is the mass in grams of one formula unit of Na_2SO_3?

44. A sample of carbon dioxide (CO_2) has a mass of 52.0 g.

- **a.** How many carbon atoms are present?
- **b.** How many oxygen atoms are present?
- **c.** What is the mass in grams of one molecule of CO_2?

45. What mass of sodium chloride (NaCl) contains 4.59×10^{24} formula units?

46. CHALLENGE A sample of silver chromate has a mass of 25.8 g.

- **a.** Write the formula for silver chromate.
- **b.** How many cations are present in the sample?
- **c.** How many anions are present in the sample?
- **d.** What is the mass in grams of one formula unit of silver chromate?

Conversions between mass, moles, and the number of particles are summarized in **Figure 11,** on the next page. Note that molar mass and the inverse of molar mass are conversion factors between mass and number of moles. Avogadro's number and its inverse are the conversion factors between moles and the number of representative particles. To convert between moles and the number of moles of atoms or ions contained in the compound, use the ratio of moles of atoms or ions to 1 mole of compound or its inverse, which are shown on the upward and downward arrows in **Figure 11.** These ratios are derived from the subscripts in the chemical formula.

Get It?

State whether it is possible to directly convert mass to the number of particles. Explain your reasoning.

SCIENCE USAGE V. COMMON USAGE

convert

Science usage: to calculate an equivalent value, such as for units of measurement

You can convert the number of moles of a substance into the number of molecules.

Common usage: to modify (something) so as to serve a different function

They are going to convert the garden shed into a workshop.

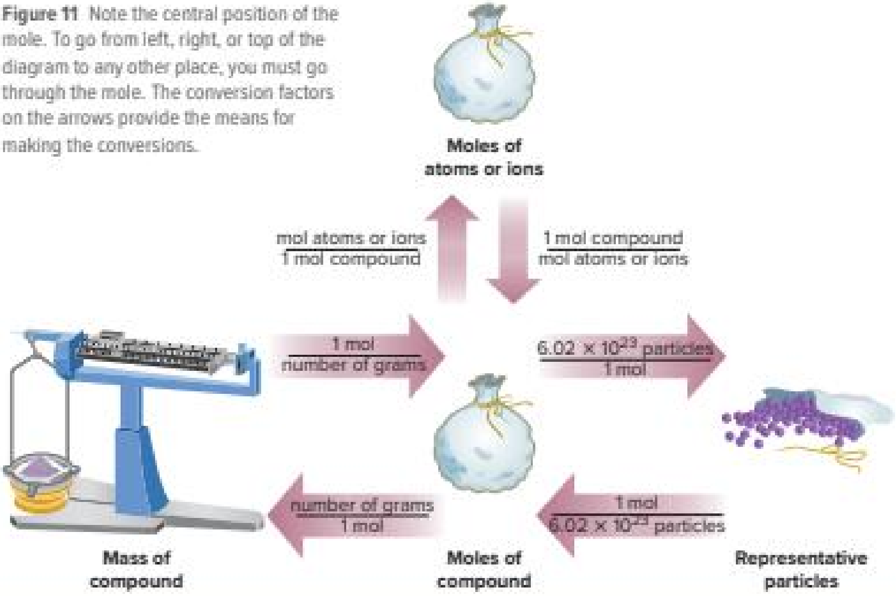


Figure 11 Note the central position of the mole. To go from left, right, or top of the diagram to any other place, you must go through the mole. The conversion factors on the arrows provide the means for making the conversions.

Check Your Progress

Summary

- Subscripts in a chemical formula indicate how many moles of each element are present in 1 mol of the compound.
- The molar mass of a compound is calculated from the molar masses of all the elements in the compound.
- Conversion factors are based on a compound's molar mass and are used to convert between moles and mass of a compound.
- The fact that atoms are conserved in a reaction is important because scientists, with knowledge of the chemical properties of the atoms, are then able to predict and describe chemical reactions.

Demonstrate Understanding

47. **Describe** how to determine the molar mass of a compound and how the molar masses of the ions that make up a compound during its formation relate to the law of conservation of mass.
48. **Use** units to create the conversion factors needed to convert between the number of moles and the mass of a compound in multi-step problems.
49. **Explain** how you can determine the number of atoms or ions in a given mass of a compound.
50. **Apply** How many moles of K, C, and O are there in 1 mol of $K_2C_2O_4$?
51. **Calculate** the molar mass of $MgBr_2$.
52. **Calculate** Calcium carbonate is the calcium source for many vitamin tablets. The recommended daily allowance is 1000 mg of Ca^{2+} ions. How many moles of Ca^{2+} does 1000 mg represent?
53. **Design** a bar graph that will show the number of moles of each element present in 500 g of a particular form of dioxin ($C_{12}H_4Cl_4O_2$), a powerful poison.

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LESSON 4

EMPIRICAL AND MOLECULAR FORMULAS

FOCUS QUESTION

Can different molecules have the same ratio of elements?

Percent Composition

Chemists, such as those shown in **Figure 12**, are often involved in developing new compounds for industrial, pharmaceutical, and home uses. After a synthetic chemist (one who makes new compounds) has produced a new compound, an analytical chemist analyzes the compound to provide experimental proof of its composition and its chemical formula. Empirical evidence, information collected by observation or experimentation, is an important part of acquiring scientifc knowledge.

It is the analytical chemist's job to identify the elements a compound contains and determine their percents by mass. Gravimetric and volumetric analyses are experimental procedures based on the measurement of mass for solids and liquids, respectively.

Figure 12 New compounds are first made on a small scale by a synthetic chemist like the one shown on the left. Then, an analytical chemist, like the one shown on the right, analyzes the compound to verify its structure and percent composition.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Laboratory: Mole Ratios

Obtain information and use mathematics to calculate the proportion and quantity of cations to anions in an ionic compound.

Quick Investigation: Analyze Chewing Gum

Analyze and interpret data to calculate the quantity of sweeteners in chewing gum.

Percent composition from experimental data

To determine the percent composition of an unknown compound, experimental analyses are first used to determine the mass of each element in the sample. For example, consider that a 100-g sample of a compound was determined to contain 55 g of Element X and 45 g of Element Y. The percent by mass of any element in a compound is then found by dividing the mass of the element by the mass of the compound and multiplying by 100.

$$\text{percent by mass (element)} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100$$

Because percent means parts per 100, the percents by mass of each elements is calculated as shown.

$$\frac{55\,\cancel{g}\text{ element X}}{100\,\cancel{g}\text{ compound}} \times 100 = 55\%\text{ element X}$$

$$\frac{45\,\cancel{g}\text{ element Y}}{100\,\cancel{g}\text{ compound}} \times 100 = 45\%\text{ element Y}$$

Thus, the compound is 55% X and 45% Y. The percent by mass of each element in a compound is the **percent composition** of a compound.

Percent composition from a chemical formula

The percent composition of a known compound can be determined from its chemical formula. To do this, assume you have exactly 1 mole of the compound and use the chemical formula to calculate the compound's molar mass. Then, determine the mass of each element in a mole of the compound by multiplying the element's molar mass by its subscript in the chemical formula. Finally, use the equation below to find the percent by mass of each element.

The percent by mass of an element in a compound is the mass of the element in 1 mol of the compound divided by the molar mass of the compound, multiplied by 100.

Percent by Mass from the Chemical Formula

$$\text{percent by mass} = \frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100$$

EXAMPLE Problem 10

CALCULATING PERCENT COMPOSITION Sodium hydrogen carbonate ($NaHCO_3$), also called baking soda, is an active ingredient in some antacids used for the relief of indigestion. Determine the percent composition of $NaHCO_3$.

1 ANALYZE THE PROBLEM

You are given only the chemical formula. Assume you have 1 mol of $NaHCO_3$. Calculate the molar mass and the mass of each element in 1 mol to determine the percent by mass of each element in the compound. The sum of all percents should be 100, although your answer might vary slightly due to rounding.

Known	Unknown
formula = $NaHCO_3$	percent Na = ?
	percent H = ?
	percent C = ?
	percent O = ?

EXAMPLE Problem 10 (continued)

2 SOLVE FOR THE UNKNOWN

Determine the molar mass of $NaHCO_3$ and each element's contribution.

$1\ \cancel{\text{mol Na}} \times \frac{22.99\ \text{g Na}}{1\ \cancel{\text{mol Na}}} = 22.99\ \text{g Na}$ Multiply the molar mass of Na by the number of Na atoms in the compound.

$1\ \cancel{\text{mol H}} \times \frac{1.008\ \text{g H}}{1\ \cancel{\text{mol H}}} = 1.008\ \text{g H}$ Multiply the molar mass of H by the number of H atoms in the compound.

$1\ \cancel{\text{mol C}} \times \frac{12.01\ \text{g C}}{1\ \cancel{\text{mol C}}} = 12.01\ \text{g C}$ Multiply the molar mass of C by the number of C atoms in the compound.

$3\ \cancel{\text{mol O}} \times \frac{16.00\ \text{g O}}{1\ \cancel{\text{mol O}}} = 48.00\ \text{g O}$ Multiply the molar mass of O by the number of O atoms in the compound.

molar mass = (22.99 g + 1.008 g + 12.01 g + 48.00 g) Total the mass values.

= 84.01 g/mol $NaHCO_3$

Use the percent by mass equation.

$\%\ \text{mass element} = \frac{\text{mass of element in 1 mole of compound}}{\text{molar mass of compound}} \times 100$ State the equation.

$\text{percent Na} = \frac{22.99\ \text{g}}{\text{mol}/84.01\ \text{g/mol}} \times 100 = 27.37\%\ \text{Na}$ Substitute mass of Na in 1 mol compound = 22.99 g/mol and molar mass $NaHCO_3$ = 84.01 g/mol. Calculate % Na.

$\text{percent H} = \frac{1.008\ \text{g}}{\text{mol}/84.01\ \text{g/mol}} \times 100 = 1.200\%\ \text{H}$ Substitute mass of H in 1 mol compound = 1.008 g/mol and molar mass $NaHCO_3$ = 84.01 g/mol. Calculate % H.

$\text{percent C} = \frac{12.01\ \text{g}}{\text{mol}/84.01\ \text{g/mol}} \times 100 = 14.30\%\ \text{C}$ Substitute mass of C in 1 mol compound = 12.01 g/mol and molar mass $NaHCO_3$ = 84.01 g/mol. Calculate % C.

$\text{percent O} = \frac{48.00\ \text{g}}{\text{mol}/84.01\ \text{g/mol}} \times 100 = 57.14\%\ \text{O}$ Substitute mass of O in 1 mol compound = 48.00 g/mol and molar mass $NaHCO_3$ = 84.01 g/mol. Calculate % O.

$NaHCO_3$ is 27.37% Na, 1.200% H, 14.30% C, and 57.14% O.

3 EVALUATE THE ANSWER

All masses and molar masses contain four significant figures. Therefore, the percents are correctly stated with four significant figures. When rounding error is accounted for, the sum of the mass percents is 100%, as required.

PRACTICE Problems ADDITIONAL PRACTICE

54. What is the percent composition of phosphoric acid (H_3PO_4)?

55. Which has the larger percent by mass of sulfur, H_2SO_3 or $H_2S_2O_8$?

56. Calcium chloride ($CaCl_2$) is sometimes used as a de-icer. Calculate the percent by mass of each element in $CaCl_2$.

57. CHALLENGE Sodium sulfate is used in the manufacture of detergents.

- **a.** Identify each of the component elements of sodium sulfate, and write the compound's chemical formula.
- **b.** Identify the compound as ionic or covalent.
- **c.** Calculate the percent by mass of each element in sodium sulfate.

Empirical Formula

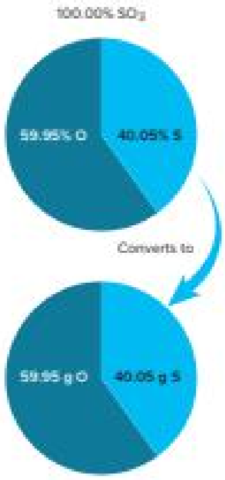


Figure 13 You can always assume that you have a 100-g sample of the compound and use the percents of the elements as masses of the elements.

When a compound's percent composition is known, its formula can be calculated. First, determine the smallest whole-number ratio of the moles of the elements in the compound. This ratio gives the subscripts in the empirical formula. The **empirical formula** for a compound is the formula with the smallest whole-number mole ratio of the elements. The empirical formula might or might not be the same as the actual molecular formula. If the two formulas are different, the molecular formula will always be a simple multiple of the empirical formula. The empirical formula for hydrogen peroxide is HO; the molecular formula is H_2O_2. In both formulas, the ratio of oxygen to hydrogen is 1:1.

Percent composition of elements in a given mass of compound can be used to determine its formula. If percent composition is given, use a total mass of the compound as 100.00 g and the percent by mass of each element as the mass of that element in grams. This is shown in **Figure 13,** where 100.00 g of the compound contains 40.05 g of S and 59.95 g of O. Each mass is then converted to moles.

$$40.05\ \cancel{g\,S} \times \frac{1\ \text{mol S}}{32.07\ \cancel{g\,S}} = 1.249\ \text{mol S}$$

$$59.95\ \cancel{g\,O} \times \frac{1\ \text{mol O}}{16.00\ \cancel{g\,O}} = 3.747\ \text{mol O}$$

Thus, the mole ratio of S atoms to O atoms in the oxide is 1.249:3.747. Since the values are not whole numbers, you convert the ratio to whole numbers by dividing by the smallest value. This does not change the ratio between the two elements because both are divided by the same number.

$$\frac{1.249\ \text{mol S}}{1.249} = 1\ \text{mol S} \qquad \frac{3.747\ \text{mol O}}{1.249} = 3\ \text{mol O}$$

Thus, the empirical formula is SO_3.

Get It?

Describe the experimental data that is needed to calculate percent composition of an unknown compound.

EXAMPLE Problem 11

EMPIRICAL FORMULA FROM PERCENT COMPOSITION Methyl acetate is a solvent commonly used in some paints, inks, and adhesives. Determine the empirical formula for methyl acetate, which has the following chemical analysis: 48.64% carbon, 8.16% hydrogen, and 43.20% oxygen.

1 ANALYZE THE PROBLEM

You are given the percent composition of methyl acetate and must find the empirical formula. Use each percent by mass as the mass, in grams, of the element in a 100.00-g sample. Then, convert from grams to moles and find the smallest whole-number ratio of moles of the elements.

Known
percent by mass C = 48.64% C
percent by mass H = 8.16% H
percent by mass O = 43.20% O

Unknown
empirical formula = ?

EXAMPLE Problem 11 (continued)

2 SOLVE FOR THE UNKNOWN

Convert each mass to moles using a conversion factor that relates moles to grams.

$48.64\ \cancel{g\ C} \times \frac{1\text{ mol C}}{12.01\ \cancel{g\ C}} = 4.050\text{ mol C}$ — Substitute mass C = 48.64 g, inverse molar mass C = 1 mol/12.01 g, and calculate moles of C.

$8.16\ \cancel{g\ H} \times \frac{1\text{ mol H}}{1.008\ \cancel{g\ H}} = 8.10\text{ mol H}$ — Substitute mass H = 8.16 g, inverse molar mass H = 1 mol/1.008 g, and calculate moles of H.

$43.20\ \cancel{g\ O} \times \frac{1\text{ mol O}}{16.00\ \cancel{g\ O}} = 2.700\text{ mol O}$ — Substitute mass O = 43.20 g, inverse molar mass O = 1 mol/16.00 g, and calculate moles of O.

Next, calculate the simplest ratio of moles by dividing the moles of each element by the smallest calculated value.

$\frac{4.050\text{ mol C}}{2.700} = 1.500\text{ mol C} = 1.5\text{ mol C}$ — Divide moles of C by 2.700.

$\frac{8.10\text{ mol H}}{2.700} = 3.00\text{ mol H} = 3\text{ mol H}$ — Divide moles of H by 2.700.

$\frac{2.700\text{ mol O}}{2.700} = 1.000\text{ mol O} = 1\text{ mol O}$ — Divide moles of O by 2.700.

The simplest mole ratio is (1.5 mol C) : (3 mol H) : (1 mol O). Multiply each number in the ratio by the smallest number—in this case 2—that yields a ratio of whole numbers.

$2 \times 1.5\text{ mol C} = 3\text{ mol C}$ — Multiply moles of C by 2 to obtain a whole number.

$2 \times 3\text{ mol H} = 6\text{ mol H}$ — Multiply moles of H by 2 to obtain a whole number.

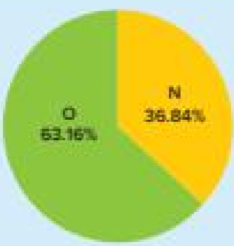
$2 \times 1\text{ mol O} = 2\text{ mol O}$ — Multiply moles of O by 2 to obtain a whole number.

The simplest whole-number ratio of atoms is (3 atoms C) : (6 atoms H) : (2 atoms O). Thus, the empirical formula of methyl acetate is $C_3H_6O_2$.

3 EVALUATE THE ANSWER

The calculations are correct, and significant figures have been observed. To check that the formula is correct, you can calculate the percent composition represented by the formula.

PRACTICE Problems — ADDITIONAL PRACTICE



58. The circle graph at the right gives the percent composition for a blue solid. What is the empirical formula for this solid?

59. Determine the empirical formula for a compound that contains 35.98% aluminum and 64.02% sulfur.

60. Propane is a hydrocarbon, a compound composed only of carbon and hydrogen. It is 81.82% carbon and 18.18% hydrogen. What is the empirical formula?

61. CHALLENGE Aspirin is the world's most-often used medication. The chemical analysis of aspirin indicates that the molecule is 60.00% carbon, 4.44% hydrogen, and 35.56% oxygen. Determine the empirical formula for aspirin.

If dividing by the smallest mole value does not yield whole numbers, each mole value must then be multiplied by the smallest factor that will make it a whole number, as is shown in Example Problem 11.

Molecular Formula

Would it surprise you to learn that substances with distinctly different properties can have the same percent composition and the same empirical formula? How is this possible? Remember that the subscripts in an empirical formula indicate the simplest whole-number ratio of moles of the elements in the compound. However, the simplest ratio does not always indicate the actual ratio in the compound.

To identify a new compound, a chemist determines the **molecular formula,** which specifies the actual number of atoms of each element in one molecule or formula unit of the substance. Consider the element in **Figure 14.** Acetylene gas has the same percent composition and the same empirical formula (CH) as benzene, which is a liquid. Yet chemically and structurally, acetylene and benzene are very different.

To determine the molecular formula for a compound, the molar mass of the compound must be determined through experimentation and compared with the mass represented by the empirical formula. For example, the molar mass of acetylene is 26.04 g/mol, and the mass of the empirical formula (CH) is 13.02 g/mol. Dividing the actual molar mass by the mass of the empirical formula indicates that the molar mass of acetylene is two times the mass of the empirical formula.

$$\frac{\text{experimentally determined molar mass of acetylene}}{\text{mass of empirical formula}} = \frac{26.04\ \cancel{\text{g/mol}}}{13.02\ \cancel{\text{g/mol}}} = 2.000$$

Because the molar mass of acetylene is two times the mass represented by the empirical formula, the molecular formula of acetylene must contain twice the number of carbon and hydrogen atoms as represented by the empirical formula.

Similarly, when the experimentally determined molar mass of benzene, 78.12 g/mol, is compared with the mass of the empirical formula, the molar mass of benzene is found to be six times the mass of the empirical formula.

$$\frac{\text{experimentally determined molar mass of benzene}}{\text{mass of the empirical formula CH}} = \frac{78.12\ \text{g mol}}{13.02\ \text{g mol}} = 6.000$$

The molar mass of benzene is six times the mass represented by the empirical formula, so the molecular formula for benzene must represent six times the number of carbon atoms and hydrogen atoms shown in the empirical formula. You can conclude that the molecular formula for acetylene is $2 \times CH$, or C_2H_2, and the molecular formula for benzene is $6 \times CH$, or C_6H_6. A molecular formula can be represented as the empirical formula multiplied by an integer n.

molecular formula = (empirical formula)n

Figure 14 Acetylene is a gas used for welding because of the high-temperature flame produced when it is burned with oxygen.

The integer n is the factor (6 in the example of benzene above) by which the subscripts in the empirical formula must be multiplied to obtain the molecular formula. The steps in determining empirical and molecular formulas from percent composition or mass data are outlined in **Figure 15.** As in other calculations, the route leads from mass through moles because formulas are based on the relative numbers of moles of elements in each mole of compound.

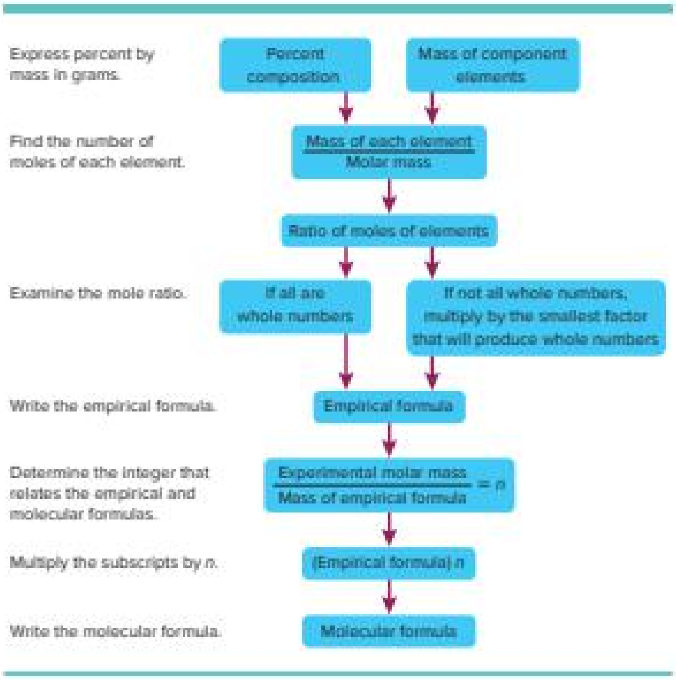


Figure 15 Use this flowchart to guide you through the steps in determining the empirical and molecular formulas for compounds.

EXAMPLE Problem 12

DETERMINING A MOLECULAR FORMULA Succinic acid is a substance produced by lichens. Chemical analysis indicates it is composed of 40.68% carbon, 5.08% hydrogen, and 54.24% oxygen and has a molar mass of 118.1 g/mol. Determine the empirical and molecular formulas for succinic acid.

1 ANALYZE THE PROBLEM

You are given the percent composition. Assume that each percent by mass represents the mass of the element in a 100.00-g sample. You can compare the given molar mass with the mass represented by the empirical formula to find n.

Known
percent by mass C = 40.68% C
percent by mass H = 5.08% H
percent by mass O = 54.24% O
molar mass = 118.1 g/mol succinic acid

Unknown
empirical formula = ?
molecular formula = ?

EXAMPLE Problem 12 (continued)

2 SOLVE FOR THE UNKNOWN

Use the percents by mass as masses in grams, and convert grams to moles by using a conversion factor—the inverse of molar mass—that relates moles to mass.

$$40.68\ \cancel{g\ C} \times \frac{1\ mol\ C}{12.01\ \cancel{g\ C}} = 3.387\ mol\ C$$

Substitute mass C = 40.68 g, inverse molar mass C = 1 mol/12.01 g, and solve for moles of C.

$$5.08\ \cancel{g\ H} \times \frac{1\ mol\ H}{1.008\ \cancel{g\ H}} = 5.04\ mol\ H$$

Substitute mass H = 5.08 g, inverse molar mass H = 1 mol/1.008 g, and solve for moles of H.

$$54.24\ \cancel{g\ O} \times \frac{1\ mol\ O}{16.00\ \cancel{g\ O}} = 3.390\ mol\ O$$

Substitute mass O = 54.24 g, inverse molar mass O = 1 mol/16.00 g, and solve for moles of O.

The mole ratio in succinic acid is (3.387 mol C) : (5.04 mol H) : (3.390 mol O).

Next, calculate the simplest ratio of moles of elements by dividing the moles of each element by the smallest value in the calculated mole ratio.

$$\frac{3.387\ mol\ C}{3.387} = 1\ mol\ C$$

Divide moles of C by 3.387.

$$\frac{5.04\ mol\ H}{3.387} = 1.49\ mol\ H \approx 1.5\ mol\ H$$

Divide moles of H by 3.387.

$$\frac{3.390\ mol\ O}{3.387} = 1.001\ mol\ O \approx 1\ mol\ O$$

Divide moles of O by 3.387.

The simplest mole ratio is 1:1.5:1. Multiply all mole values by 2 to obtain whole numbers.

$2 \times 1\ mol\ C = 2\ mol\ C$ — Multiply moles of C by 2.

$2 \times 1.5\ mol\ H = 3\ mol\ H$ — Multiply moles of H by 2.

$2 \times 1\ mol\ O = 2\ mol\ O$ — Multiply moles of O by 2.

The simplest whole-number mole ratio is 2:3:2. The empirical formula is **$C_2H_3O_2$**. Calculate the empirical formula mass using the molar mass of each element.

$$2\ \cancel{mol\ C} \times \frac{12.01\ g\ C}{1\ \cancel{mol\ C}} = 24.02\ g\ C$$

Multiply the molar mass of C by the moles of C atoms in the compound.

$$3\ \cancel{mol\ H} \times \frac{1.008\ g\ H}{1\ \cancel{mol\ H}} = 3.024\ g\ H$$

Multiply the molar mass of H by the moles of H atoms in the compound.

$$2\ \cancel{mol\ O} \times \frac{16.00\ g\ O}{1\ \cancel{mol\ O}} = 32.00\ g\ O$$

Multiply the molar mass of O by the moles of O atoms in the compound.

$$\text{molar mass } C_2H_3O_2 = (24.02\ g + 3.024\ g + 32.00\ g) = 59.04\ g/mol$$

Total the mass values.

Divide the experimentally determined molar mass of succinic acid by the mass of the empirical formula to determine *n*.

$$n = \frac{\text{molar mass of succinic acid}}{\text{molar mass of } C_2H_3O_2} = \frac{118.1\ \cancel{g/mol}}{59.04\ \cancel{g/mol}} = 2.000$$

Multiply the subscripts in the empirical formula by 2 to determine the actual subscripts in the molecular formula.

$$2 \times (C_2H_3O_2) = C_4H_6O_4$$

The molecular formula for succinic acid is **$C_4H_6O_4$**.

EXAMPLE Problem 12 (continued)

3 EVALUATE THE ANSWER

The calculation of the molar mass from the molecular formula gives the same result as the given, experimentally-determined molar mass.

EXAMPLE Problem 13

CALCULATING AN EMPIRICAL FORMULA FROM MASS DATA The mineral ilmenite is usually mined and processed for titanium, a strong, light, and flexible metal. A sample of ilmenite contains 5.41 g of iron, 4.64 g of titanium, and 4.65 g of oxygen. Determine the empirical formula for ilmenite.

1 ANALYZE THE PROBLEM

You are given the masses of the elements found in a known mass of ilmenite and must determine the empirical formula of the mineral. Convert the known masses of each element to moles, then find the smallest whole-number ratio of the moles of the elements.

Known
mass of iron = 5.41 g Fe
mass of titanium = 4.64 g Ti
mass of oxygen = 4.65 g O

Unknown
empirical formula = ?

2 SOLVE FOR THE UNKNOWN

Convert each known mass to moles by using a conversion factor—the inverse of molar mass—that relates moles to grams.

$$5.41\ \cancel{\text{g Fe}} \times \frac{1\ \text{mol Fe}}{55.85\ \cancel{\text{g Fe}}} = 0.0969\ \text{mol Fe}$$

Multiply mass Fe = 5.41 g by the inverse molar mass Fe = 1 mol/55.85 g, and calculate moles of Fe.

$$4.64\ \cancel{\text{g Ti}} \times \frac{1\ \text{mol Ti}}{47.87\ \cancel{\text{g Ti}}} = 0.0969\ \text{mol Ti}$$

Multiply mass Ti = 4.64 g by the inverse molar mass Ti = 1 mol/47.87 g, and calculate moles of Ti.

$$4.65\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.291\ \text{mol O}$$

Multiply mass O = 4.65 g by the inverse molar mass O = 1 mol/16.00 g, and calculate moles of O.

The mineral ilmenite has a mole ratio of (0.0969 mol Fe) : (0.0969 mol Ti) : (0.291 mol O).

Calculate the simplest ratio by dividing each mole value by the smallest value in the ratio.

$$\frac{0.0969\ \text{mol Fe}}{0.0969} = 1\ \text{mol Fe}$$

Divide moles of Fe by 0.0969.

$$\frac{0.0969\ \text{mol Ti}}{0.0969} = 1\ \text{mol Ti}$$

Divide moles of Ti by 0.0969.

$$\frac{0.291\ \text{mol O}}{0.0969} = 3\ \text{mol O}$$

Divide moles of O by 0.0969.

Because all the mole values are whole numbers, the simplest whole-number mole ratio is (1 mol Fe) : (1 mol Ti) : (3 mol O). The empirical formula for ilmenite is **$FeTiO_3$**.

3 EVALUATE THE ANSWER

The mass of iron is slightly greater than the mass of titanium, but the molar mass of iron is also slightly greater than that of titanium. Thus, it is reasonable that the numbers of moles of iron and titanium are equal. The mass of titanium is approximately the same as the mass of oxygen, but the molar mass of oxygen is about one-third that of titanium. Thus, a 3:1 ratio of oxygen to titanium is reasonable.

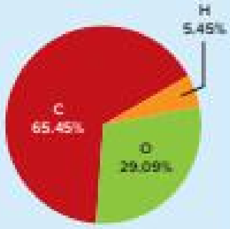
PRACTICE Problems ADDITIONAL PRACTICE

62. A compound was found to contain 49.98 g of carbon and 10.47 g of hydrogen. The molar mass of the compound is 58.12 g/mol. Determine the molecular formula.

63. A colorless liquid composed of 46.68% nitrogen and 53.32% oxygen has a molar mass of 60.01 g/mol. What is the molecular formula?

64. When an oxide of potassium is decomposed, 19.55 g of K and 4.00 g of O are obtained. What is the empirical formula for the compound?

65. CHALLENGE Analysis of a chemical used in photographic developing fluid yielded the percent composition data shown in the circle graph to the right. If the chemical's molar mass is 110.0 g/mol, what is its molecular formula?



66. CHALLENGE Analysis of the pain reliever morphine yielded the data shown in the table. Determine the empirical formula of morphine.

Element	Mass (g)
carbon	17.900
hydrogen	1.680
oxygen	4.225
nitrogen	1.228

Check Your Progress

Summary

- The percent by mass of an element in a compound gives the percentage of the compound's total mass due to that element.
- The subscripts in an empirical formula give the smallest whole-number ratio of moles of elements in the compound.
- A molecular formula gives the actual number of atoms of each element in a molecule or formula unit of a substance.
- A molecular formula is a whole number multiple of the empirical formula.

Demonstrate Understanding

67. **Assess** A classmate tells you that based on empirical evidence a compound's molecular formula is 2.5 times its empirical formula. Explain whether he is correct or not. Describe what experimental data was needed to calculate the formula.

68. **Calculate** Analysis of a compound composed of iron and oxygen yields 174.86 g of Fe, 75.14 g of O. What is the empirical formula for this compound?

69. **Calculate** An oxide of aluminum contains 0.545 g of Al and 0.845 g of O. Find the empirical formula for the oxide.

70. **Explain** how percent composition data for a compound are related to the masses of the elements in the compound.

71. **Explain** how you can find the mole ratio in a chemical compound.

72. **Apply** The molar mass of a compound is twice that of its empirical formula. How are the compound's molecular and empirical formulas related?

73. **Analyze** Hematite (Fe_2O_3) and magnetite (Fe_3O_4) are two ores used as sources of iron. Which ore provides the greater percent iron per kilogram?

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LESSON 5

FORMULAS OF HYDRATES

FOCUS QUESTION

How can you determine the amount of water in a hydrate?

Naming Hydrates

Have you ever watched crystals slowly form from a water solution? Sometimes, water molecules adhere to the ions as the solid forms. The water that becomes part of the crystal is called water of hydration. Solids in which water molecules are trapped are called hydrates. A **hydrate** is a compound that has a specific number of water molecules bound to its atoms. **Figure 16** shows the gemstone known as opal, which is hydrated silicon dioxide (SiO_2). The coloring is the result of water in the mineral.

In the formula of a hydrate, the number of water molecules associated with each formula unit of the compound is written following a dot in the molecular formula – $Na_2CO_3 \cdot 10H_2O$. This compound is called sodium carbonate decahydrate. In the word *decahydrate*, the prefix *deca-* means *ten* and the root word *hydrate* refers to *water*. A decahydrate has ten water molecules associated with one formula unit of compound.

The mass of water associated with a formula unit must be included in all molar mass calculations. The number of water molecules associated with different hydrates varies widely. Some common hydrates are listed in **Table 1**, on the next page.

Figure 16 The presence of water and various mineral impurities accounts for the variety of different-colored opals. Further color changes occur when opals are allowed to dry out.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

ChemLAB: Determine the Formula of a Hydrate

Analyze and interpret data to calculate the proportion of moles of water in a hydrated compound.

Inquiry into Chemistry: Solve It: Mystery of the Missing Mass

Use mathematics and computational thinking to calculate the quantity of percent water of Epsom salt and washing soda.

Table 1 Formulas of Hydrates

Prefix	Molecules H_2O	Formula	Name
Mono-	1	$(NH_4)_2C_2O_4 \cdot H_2O$	ammonium oxalate monohydrate
Di-	2	$CaCl_2 \cdot 2H_2O$	calcium chloride dihydrate
Tri-	3	$NaC_2H_3O_2 \cdot 3H_2O$	sodium acetate trihydrate
Tetra-	4	$FePO_4 \cdot 4H_2O$	iron(III) phosphate tetrahydrate
Penta-	5	$CuSO_4 \cdot 5H_2O$	copper(II) sulfate pentahydrate
Hexa-	6	$CoCl_2 \cdot 6H_2O$	cobalt(II) chloride hexahydrate
Hepta-	7	$MgSO_4 \cdot 7H_2O$	magnesium sulfate heptahydrate
Octa-	8	$Ba(OH)_2 \cdot 8H_2O$	barium hydroxide octahydrate
Deca-	10	$Na_2CO_3 \cdot 10H_2O$	sodium carbonate decahydrate

Figure 17 Water of hydration can be removed by heating a hydrate, producing an anhydrous compound that can look very different from its hydrated form.

Analyzing a Hydrate

When a hydrate is heated, water molecules are driven off, leaving an anhydrous compound, or one "without water." See **Figure 17.** The series of photos show that when pink cobalt(II) chloride hexahydrate is heated, blue anhydrous cobalt(II) chloride is produced.

It is possible to determine the formula of a hydrate by performing experiments and collecting empirical evidence, or data through observations, that can be used to calculate the formula. You must find the number of moles of water associated with 1 mol of the hydrate. Suppose you have a 5.00-g sample of a hydrate of barium chloride. You know that the formula is $BaCl_2 \cdot xH_2O$. You must determine x, the coefficient of H_2O in the hydrate formula that indicates the number of moles of water associated with 1 mol of $BaCl_2$. To find x, you would heat the sample of the hydrate to drive off the water of hydration. After heating, the dried substance, which is anhydrous $BaCl_2$, has a mass of 4.26 g. The mass of the water of hydration is the difference between the mass of the hydrate (5.00 g) and the mass of the anhydrous compound (4.26 g).

$$5.00 \text{ g } BaCl_2 \text{ hydrate} - 4.26 \text{ g anhydrous } BaCl_2 = 0.74 \text{ g } H_2O$$

You now know the masses of $BaCl_2$ and H_2O in the sample. You can convert these masses to moles using the molar masses. The molar mass of $BaCl_2$ is 208.23 g/mol, and the molar mass of H_2O is 18.02 g/mol.

$$4.26\ \cancel{g\ BaCl_2} \times \frac{1\ mol\ BaCl_2}{208.23\ \cancel{g\ BaCl_2}} = 0.0205\ mol\ BaCl_2$$

$$0.74\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.02\ \cancel{g\ H_2O}} = 0.041\ mol\ H_2O$$

Now that the moles of $BaCl_2$ and H_2O have been determined, you can calculate the ratio of moles of H_2O to moles of $BaCl_2$ which is x, the coefficient that precedes H_2O in the formula for the hydrate.

$$x = \frac{moles\ H_2O}{moles\ BaCl_2} = \frac{0.041\ mol\ H_2O}{0.0205\ mol\ BaCl_2} = \frac{2.0\ mol\ H_2O}{1.00\ mol\ BaCl_2} = \frac{2}{1}$$

The ratio of moles of H_2O to moles of $BaCl_2$ is 2:1, so 2 mol of water is associated with 1 mol of barium chloride. The value of the coefficient x is 2 and the formula of the hydrate is $BaCl_2 \cdot 2H_2O$. What is the name of the hydrate?

Get It?

Explain what empirical evidence is needed to determine the formula of a hydrate, and describe how the data is obtained.

Uses of Hydrates

Have you ever taken a bath in Epsom salt to sooth your sore muscles? Epsom salt is magnesium sulfate heptahydrate, with 7 water molecules ($MgSO_4 \cdot 7H_2O$). Magnesium from this salt is absorbed through skin, which can reduce inflammation. This hydrated salt appears much like table salt, as shown in **Figure 18**, but is typically coarser grained when used in bathwater or aquariums. Epsom salt is also used in gardening and agriculture as a supplement for magnesium deficient soils. The mineral name is epsomite, which precipitates from mineral solutions and is found encrusted on limestone cavern and mine wall surfaces. It was first described in 1806 in deposits near the market town Epsom, in Surrey, England.

Figure 18 Epsom salt is not actually a salt but a naturally occurring pure mineral compound of magnesium and sulfate.

EXAMPLE Problem 14

DETERMINING THE FORMULA OF A HYDRATE A mass of 2.50 g of blue, hydrated copper sulfate ($CuSO_4 \cdot xH_2O$) is placed in a crucible and heated. After heating, 1.59 g of white anhydrous copper sulfate ($CuSO_4$) remains. What is the formula for the hydrate? Name the hydrate.

1 ANALYZE THE PROBLEM

You are given a mass of hydrated copper sulfate. The mass after heating is the mass of the anhydrous compound. You know the formula for the compound, except for *x*, the number of moles of water of hydration.

Known
mass of hydrated compound = 2.50 g $CuSO_4 \cdot xH_2O$
mass of anhydrous compound = 1.59 g $CuSO_4$
molar mass H_2O = 18.02 g/mol H_2O
molar mass $CuSO_4$ = 159.6 g/mol $CuSO_4$

Unknown
formula of hydrate = ?
name of hydrate = ?

2 SOLVE FOR THE UNKNOWN

Determine the mass of water lost.

mass of hydrated copper sulfate	2.50 g
mass of anhydrous copper sulfate	−1.59 g
mass of water lost	0.91 g

Subtract the mass of anhydrous $CuSO_4$ from the mass of $CuSO_4 \cdot xH_2O$.

Convert the known masses of H_2O and anhydrous $CuSO_4$ to moles using a conversion factor—the inverse of molar mass—that relates moles and mass.

$$1.59\ \cancel{g\ CuSO_4} \times \frac{1\ mol\ CuSO_4}{159.6\ \cancel{g\ CuSO_4}} = 0.00996\ mol\ CuSO_4$$

Substitute mass $CuSO_4$ = 1.59g, inverse molar mass $CuSO_4$ = 1 mol/159.6g, and solve.

$$0.91\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.02\ \cancel{g\ H_2O}} = 0.050\ mol\ H_2O$$

Substitute mass H_2O = 0.91 g, inverse molar mass H_2O = 1 mol/18.02 g, and solve.

$$x = \frac{moles\ H_2O}{moles\ CuSO_4}$$

State the ratio of moles of H_2O to moles of $CuSO_4$.

$$x = \frac{0.050\ mol\ H_2O}{0.00996\ mol\ CuSO_4} \approx \frac{5.0\ mol\ H_2O}{1\ mol\ CuSO_4} = 5$$

Substitute moles of H_2O = 0.050 mol, moles of $CuSO_4$ = 0.00996 mol. Divide numbers, and cancel units to determine the simplest whole-number ratio.

The ratio of H_2O to $CuSO_4$ is 5:1, so the formula for the hydrate is **$CuSO_4 \cdot 5H_2O$**. The name of the hydrate is **copper(II) sulfate pentahydrate**.

3 EVALUATE THE ANSWER

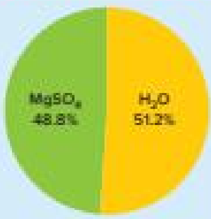
Copper (II) sulfate pentahydrate is a common hydrate listed in **Table 1.**

PRACTICE Problems

ADDITIONAL PRACTICE

74. The composition of a hydrate is given in the circle graph shown at the right. What is the formula and name of this hydrate?

75. CHALLENGE An 11.75-g sample of a common hydrate of cobalt (II) chloride is heated. After heating, 0.0712 mol of anhydrous cobalt chloride remains. What is the formula and the name of this hydrate?



Hydrates and Technologies

Calcium chloride forms three hydrates—a monohydrate, a dihydrate, and a hexahydrate. Anhydrous calcium chloride is placed in the bottom of tightly sealed containers called desiccators. The calcium chloride absorbs moisture from the air inside the desiccator, creating a dry atmosphere for other substances.

Electronic and optical equipment is often packaged with packets of desiccant. Desiccants prevent moisture from interfering with the sensitive electronic circuitry. While some types of desiccant simply absorb moisture, other types bond with moisture from the air and form hydrates.

Some hydrates such as $Na_2SO_4 \cdot 10H_2O$, for example, are used to store solar energy. When the Sun's energy heats the hydrate to a temperature greater than 32°C, the single formula unit of Na_2SO_4 in the hydrate dissolves in the 10 mol of water of hydration. In the process, energy is absorbed by the hydrate. This energy is released when the temperature decreases and the hydrate crystallizes again.

Check Your Progress

Summary

- The formula of a hydrate consists of the formula of the ionic compound and the number of water molecules associated with one formula unit.
- The name of a hydrate consists of the compound name followed by the word hydrate with a prefix indicating the number of water molecules associated with one formula unit.
- Anhydrous compounds are formed when hydrates are heated.

Demonstrate Understanding

76. **Summarize** the composition of a hydrate.
77. **Name** the compound that has the formula $SrCl_2 \cdot 6H_2O$, and state the meaning of the symbol $\cdot 6H_2O$ in the compound.
78. **Describe** the empirical evidence and procedures needed for obtaining data for calculating the proportion of water in a hydrated compound. Explain the reason for each step.
79. **Apply** A hydrate contains 0.050 mol of H_2O for every 0.00998 mol of ionic compound. Write a generalized formula of the hydrate.
80. **Calculate** the mass of water if a hydrate loses 0.025 mol of H_2O when heated.
81. **Arrange** these hydrates in order of increasing percent water content: $MgSO_4 \cdot 7H_2O$, $Ba(OH)_2 \cdot 8H_2O$, and $CoCl_2 \cdot 6H_2O$.
82. **Apply** What empirical evidence can be used to explain how the hydrate in **Figure 17** might be used as a means of roughly determining the probability of rain?

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SCIENCE & SOCIETY

Making Cents

The copper color of pennies in a handful of change makes them stand out from other coins. That color comes from copper metal, but the amount of copper in pennies has changed dramatically since the U.S. Mint first started producing the cent over 200 years ago.

Over the past 200 years, the U.S. Mint has produced pennies made of several different compositions of copper and other metals. Since 1982, the Mint has only produced the copper-coated zinc cent.

The Penny: Changes in Copper Composition

The U.S. Mint began to produce pure copper pennies in 1793. From 1837 to 1857, the Mint decided to use a copper metal alloy, bronze. An alloy is a combination of metals and often has many advantages over pure metals, including harder, more durable, and corrosion resistant properties. Bronze pennies contained 95 percent copper and 5 percent tin and zinc by mass.

For a few years the Mint switched to using brass, an alloy of copper and nickel, and then once again returned to making bronze pennies until 1982. One exception occurred in 1943, during World War II, when copper was in high demand. In response, the Mint produced a cent made of zinc-coated steel. These 1943 silver pennies are the only pennies attracted to a magnet, because of the iron in the steel. Accidentally, about 40 bronze pennies were also made in this year. Their scarcity makes the 1943 bronze pennies highly prized by coin collectors. In 1996, a collector bought one for $82,500!

A significant change occurred in 1982, when the rising cost of copper prompted the Mint to look for ways to reduce copper content. Their solution was to produce copper-plated zinc pennies. Overall, these pennies are made of 97.5 percent zinc and just 2.5 percent copper by mass, but the pennies retain their familiar copper color.

Copper-plated zinc pennies are produced to this day, but the cost of producing a penny is higher than its value: it cost 1.50 cents to produce a penny in 2016. Due to the continually decreasing value of a cent, some are calling for the abolition of the penny. The debate continues, but for now, you'll keep seeing that copper-colored coin in your handful of change.

ENGAGE IN ARGUMENT FROM EVIDENCE

Form teams to debate whether the penny should be eliminated. Research points for and against eliminating the penny to prepare for the debate.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 MEASURING MATTER

- One mole of a substance contains Avogadro's number of representative particles.
- One mole of carbon-12 atoms has a mass of exactly 12 g.
- Conversion factors written from Avogadro's relationship can be used to convert between moles and number of representative particles.

- mole
- Avogadro's number

Lesson 2 MASS AND THE MOLE

- The mass in grams of 1 mol of any pure substance is called its molar mass.
- The molar mass of any substance is the mass in grams of Avogadro's number of representative particles of the substance.
- Molar mass and its inverse are used to convert between moles and mass.

- molar mass

Lesson 3 MOLES OF COMPOUNDS

- Subscripts in a chemical formula indicate how many moles of each element are present in 1 mol of the compound.
- The molar mass of a compound is calculated from the molar masses of all of the elements in the compound.
- Conversion factors based on a compound's molar mass are used to convert between moles and mass of a compound.

Lesson 4 EMPIRICAL AND MOLECULAR FORMULAS

- The percent by mass of an element in a compound gives the percentage of the compound's total mass due to that element.
- The subscripts in an empirical formula give the smallest whole-number ratio of moles of elements in the compound.
- The molecular formula gives the actual number of atoms of each element in a molecule or formula unit of a substance.

- percent composition
- empirical formula
- molecular formula

Lesson 5 FORMULAS OF HYDRATES

- The formula of a hydrate consists of the formula of the ionic compound and the number of water molecules associated with one formula unit.
- Hydrate naming consists of the compound name and the word hydrate with a prefix indicating the number of water molecules in one formula unit.

- hydrate

THREE-DIMENSIONAL THINKING
Module Wrap-Up

REVISIT THE PHENOMENON

How is counting pennies like counting atoms?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

SEP Data Analysis Lab

How are molar mass, Avogadro's number, and the atomic nucleus related?

A nuclear model of mass can provide a simple picture of the connections among the mole, molar mass, and the number of representative particles in a mole.

Data and Observations The hydrogen-1 nucleus contains one proton with a mass of 1.007 amu. The mass of a proton, in grams, has been determined experimentally to be 1.672×10^{-24} g. The helium-4 nucleus contains two protons and two neutrons and has a mass of approximately 4 amu.

CER Analyze and Interpret Data

1. **Claim** What is the mass in grams of one helium atom? (The mass of a neutron is approximately the same as the mass of a proton.)
2. **Claim** Carbon-12 contains six protons and six neutrons. Draw the carbon-12 nucleus and calculate the mass of one atom in amu and g.
3. **Claim** How many atoms of hydrogen-1 are in a 1.007-g sample? Recall that 1.007 amu is the mass of one atom of hydrogen-1. Round your answer to two significant digits.
4. **Claim** If you had samples of helium and carbon that contained the same number of atoms as you calculated in Question 3, what would be the mass in grams of each sample?
5. **Evidence, Reasoning** What can you conclude about the relationship between the number of atoms and the mass of each sample?

STOICHIOMETRY

ENCOUNTER THE PHENOMENON

How much carbon dioxide did this field of corn need to grow?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

CER Claim, Evidence, Reasoning

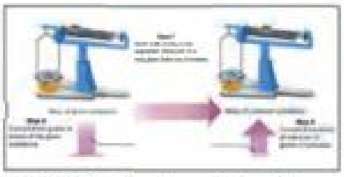
Make Your Claim Use your CER chart to make a claim about how you could figure out how much carbon dioxide this field of corn needed to grow.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.

LESSON 1: Explore & Explain: Stoichiometry and the Conservation of Mass



LESSON 2: Explore & Explain: Mass-to-Mass Conversions

LESSON 1
DEFINING STOICHIOMETRY

FOCUS QUESTION

What can you learn from balanced chemical equations?

Particle and Mole Relationships

When a colorless sodium hydrogen sulfite ($NaHSO_3$) solution is added to a purple potassium permanganate ($KMnO_4$) solution, the purple color eventually disappears. Why do you think this happens? If you concluded that the potassium permanganate had been used up and the reaction had stopped, you are right. Chemical reactions stop when one of the reactants is used up. When planning the reaction of potassium permanganate and sodium hydrogen sulfite, a chemist might ask, "How many grams of potassium permanganate are needed to react completely with a known mass of sodium hydrogen sulfite?" Or, when analyzing a photosynthesis reaction, you might ask, "How much oxygen and carbon dioxide are needed to form a known mass of sugar?"

Figure 1 The balanced chemical equation for this reaction between potassium and chlorine provides the relationships between the amounts of reactants and products.

Stoichiometry

The study of quantitative relationships between the amounts of reactants used and amounts of products formed by a chemical reaction is called **stoichiometry.** Stoichiometry is based on the law of conservation of mass. Recall that the law states that matter is neither created nor destroyed in a chemical reaction, but can change form. In any chemical reaction, the amount of matter present at the end of the reaction is the same as the amount of matter present at the beginning, given that atoms are conserved during a chemical reaction. Therefore, the mass of the reactants equals the mass of the products. Note the reaction of potassium with chlorine shown in **Figure 1.** Although potassium reacts with chlorine to form a new compound, potassium chloride, the total mass is unchanged.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Review the News

Obtain information from a current news story about chemical reactions and the conservation of matter. Evaluate your source and communicate your findings to your class.

Table 1 Relationships Derived from a Balanced Chemical Equation

$4Fe(s)$	+	$3O_2(g)$	→	$2Fe_2O_3(s)$
iron	+	oxygen	→	iron (III) oxide
4 atoms of Fe	+	3 molecules O_2	→	2 formula units of Fe_2O_3
4 mol Fe	+	3 mol O_2	→	2 mol Fe_2O_3
223.4 g Fe	+	96.00 g O_2	→	319.4 g Fe_2O_3
319.4 g reactants			→	319.4 g products

The balanced chemical equation for the chemical reaction shown in **Figure 1** is as follows.

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

You can interpret this equation in terms of representative particles by saying that four atoms of iron react with three molecules of oxygen to produce two formula units of iron(III) oxide. Remember that coefficients in an equation represent not only numbers of individual particles but also numbers of moles of particles. Therefore, you can also say that four moles of iron react with three moles of oxygen to produce two moles of iron(III) oxide.

The chemical equation does not directly tell you anything about the masses of the reactants and products. However, by converting the known mole quantities to mass, the mass relationships become obvious. Recall that moles are converted to mass by multiplying by the molar mass. The masses of the reactants are as follows.

$$4\ \cancel{\text{mol Fe}} \times \frac{55.85\ \text{g Fe}}{1\ \cancel{\text{mol Fe}}} = 223.4\ \text{g Fe}$$

$$3\ \cancel{\text{mol O}_2} \times \frac{32.00\ \text{g O}_2}{1\ \cancel{\text{mol O}_2}} = 96.00\ \text{g O}_2$$

The total mass of the reactants is: (223.4 g + 96.00 g) = 319.4 g. Similarly, the mass of the product is calculated as follows:

$$2\ \cancel{\text{mol Fe}_2\text{O}_3} \times \frac{159.7\ \text{g Fe}_2\text{O}_3}{1\ \cancel{\text{mol Fe}_2\text{O}_3}} = 319.4\ \text{g Fe}_2\text{O}_3$$

Note that the mass of the reactants equals the mass of the product.

$$\text{mass of reactants} = \text{mass of products}$$

$$319.4\ \text{g} = 319.4\ \text{g}$$

WORD ORIGIN

stoichiometry

comes from the Greek words *stoikheion*, which means *element*, and *metron*, which means *to measure*

CCC CROSSCUTTING CONCEPTS

Energy and Matter Using **Table 1** as a guide, create a similar table for the following reaction.

$$2AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2KNO_3(aq)$$

Cite evidence that the total amount of matter is conserved in this reaction.

By the law of conservation of mass, the total mass of the reactants equals the mass of the product. The relationships that can be determined from a balanced chemical equation are summarized in **Table 1.**

Get It?

List the types of relationships that can be derived from the coefficients in a balanced chemical equation.

EXAMPLE Problem 1

INTERPRETING CHEMICAL EQUATIONS The combustion of propane (C_3H_8) provides energy for heating homes, cooking food, and soldering metal parts. Interpret the equation for the combustion of propane in terms of representative particles, moles, and mass. Show that the law of conservation of mass is observed.

1 ANALYZE THE PROBLEM

The coefficients in the balanced chemical equation shown below represent both moles and representative particles, in this case molecules. Therefore, the equation can be interpreted in terms of molecules and moles. The law of conservation of mass will be verified if the masses of the reactants and products are equal.

Known

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

Unknown

Equation interpreted in terms of molecules = ?
Equation interpreted in terms of moles = ?
Equation interpreted in terms of mass = ?

2 SOLVE FOR THE UNKNOWN

The coefficients in the chemical equation indicate the number of molecules.

1 molecule C_3H_8 + 5 molecules O_2 → 3 molecules CO_2 + 4 molecules H_2O

The coefficients in the chemical equation also indicate the number of moles.

1 mol C_3H_8 + 5 mol O_2 → 3 mol CO_2 + 4 mol H_2O

To verify that mass is conserved, first convert moles of reactant and product to mass by multiplying by a conversion factor—the molar mass—that relates grams to moles.

$$\text{moles of reactant or product} \times \frac{\text{grams reactant or product}}{\text{1 mol reactant or product}} = \text{grams of reactant or product}$$

$$1\ \cancel{\text{mol } C_3H_8} \times \frac{44.09\text{ g } C_3H_8}{1\ \cancel{\text{mol } C_3H_8}} = 44.09\text{ g } C_3H_8$$
Calculate the mass of the reactant C_3H_8.

$$5\ \cancel{\text{mol } O_2} \times \frac{32.00\text{ g } O_2}{1\ \cancel{\text{mol } O_2}} = 160.0\text{ g } O_2$$
Calculate the mass of the reactant O_2.

$$3\ \cancel{\text{mol } CO_2} \times \frac{44.01\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}} = 132.0\text{ g } CO_2$$
Calculate the mass of the product CO_2.

$$4\ \cancel{\text{mol } H_2O} \times \frac{18.02\text{ g } H_2O}{1\ \cancel{\text{mol } H_2O}} = 72.08\text{ g } H_2O$$
Calculate the mass of the product H_2O.

44.09 g C_3H_8 + 160.0 g O_2 = 204.1 g reactants — Add the masses of the reactants.

132.0 g CO_2 + 72.08 g H_2O = 204.1 g products — Add the masses of the products.

204.1 g reactants = 204.1 g products — The law of conservation of mass is observed.

EXAMPLE Problem 1 (continued)

3 EVALUATE THE ANSWER

The sums of the reactants and the products are correctly stated to the first decimal place because each mass is accurate to the first decimal place. The mass of reactants equals the mass of products, as predicted by the law of conservation of mass.

PRACTICE Problems ADDITIONAL PRACTICE

1. Interpret the following balanced chemical equations in terms of particles, moles, and mass. Show that the law of conservation of mass is observed.
   a. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
   b. $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$
   c. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
2. **CHALLENGE** For each of the following, balance the chemical equation; interpret the equation in terms of particles, moles, and mass; and show that the law of conservation of mass is observed.
   a. ___$Na(s)$ + ___$H_2O(l) \rightarrow$ ___$NaOH(aq)$ + ___$H_2(g)$
   b. ___$Zn(s)$ + ___$HNO_3(aq) \rightarrow$ ___$Zn(NO_3)_2(aq)$ + ___$N_2O(g)$ + ___$H_2O(l)$

Mole ratios

You have read that the coefficients in a chemical equation indicate the relationships between moles of reactants and products. You can use the relationships between coefficients to derive conversion factors called mole ratios. A **mole ratio** is a ratio between the numbers of moles of any two of the substances in a balanced chemical equation. Consider the reaction between potassium (K) and bromine (Br_2) to form potassium bromide (KBr). The product of the reaction, the ionic salt potassium bromide, is prescribed by veterinarians, like the one in **Figure 2**, as an antiepileptic medication for dogs.

$$2K(s) + Br_2(l) \rightarrow 2KBr(s)$$

What mole ratios can be written for this reaction? Starting with the reactant potassium, you can write a mole ratio that relates the moles of potassium to each of the other two substances in the equation.

Figure 2 Potassium metal and liquid bromine react vigorously to form the ionic compound potassium bromide. Bromine is one of the two elements that are liquids at room temperature (mercury is the other). Potassium is a highly reactive metal. Potassium bromide is an ionic salt that is used to treat epilepsy in dogs.

ACADEMIC VOCABULARY

derive

to obtain from a specified source

The researcher was able to derive the meaning of the illustration from ancient texts.

Thus, one mole ratio relates the moles of potassium used to the moles of bromine used. The other mole ratio relates the moles of potassium used to the moles of potassium bromide formed.

$$\frac{2 \text{ mol K}}{1 \text{ mol } Br_2} \text{ and } \frac{2 \text{ mol K}}{2 \text{ mol KBr}}$$

Two other mole ratios show how the moles of bromine relate to the moles of the other two substances in the equation–potassium and potassium bromide.

$$\frac{1 \text{ mol } Br_2}{2 \text{ mol K}} \text{ and } \frac{1 \text{ mol } Br_2}{2 \text{ mol KBr}}$$

Similarly, two ratios relate the moles of potassium bromide to the moles of potassium and bromine.

$$\frac{2 \text{ mol KBr}}{2 \text{ mol K}} \text{ and } \frac{2 \text{ mol KBr}}{1 \text{ mol } Br_2}$$

These six ratios define all the mole relationships in this equation. Each of the three substances in the equation forms a ratio with the two other substances.

Get It?

Identify the source from which a chemical reaction's mole ratios are derived.

The decomposition of potassium chlorate ($KClO_3$) is sometimes used to obtain small amounts of oxygen in the laboratory.

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

The mole ratios that can be written for this reaction are as follows.

$$\frac{2 \text{ mol } KClO_3}{2 \text{ mol KCl}} \text{ and } \frac{2 \text{ mol } KClO_3}{3 \text{ mol } O_2}$$

$$\frac{2 \text{ mol KCl}}{2 \text{ mol } KClO_3} \text{ and } \frac{2 \text{ mol KCl}}{3 \text{ mol } O_2}$$

$$\frac{3 \text{ mol } O_2}{2 \text{ mol } KClO_3} \text{ and } \frac{3 \text{ mol } O_2}{2 \text{ mol KCl}}$$

Note that the number of mole ratios you can write for a chemical reaction involving a total of n substances is $(n)(n-1)$. In the reaction of potassium chlorate, decomposing to potassium chloride and oxygen gas, there are three substances. So, $(n)(n-1)$ becomes $(3)(2) = 6$, and you can see that all six mole ratios have been written above.

If, a reaction involves four, five, or even six substances, you can write 12, 20, and 30 mole ratios, respectively.

Four substances: $(4)(3) = 12$ mole ratios

Five substances: $(5)(4) = 20$ mole ratios

Six substances: $(6)(5) = 30$ mole ratios

Get It?

Calculate From a 5-hectare corn farm, how many tons of carbon dioxide could be removed from the air? A farm this size would also produce enough oxygen to meet the needs of how many people? Use the information from the module opener to answer these questions.

PRACTICE Problems ADDITIONAL PRACTICE

3. Determine all possible mole ratios for the following balanced chemical equations.

a. $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

b. $3Fe(s) + 4H_2O(l) \rightarrow Fe_3O_4(s) + 4H_2(g)$

c. $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

4. CHALLENGE Balance the following equations, and determine the possible mole ratios.

a. $ZnO(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$

b. butane (C_4H_{10}) + oxygen → carbon dioxide + water

It should be noted that the coefficients of the balanced equation are used in the mole ratios, but are not used to detemine how many mole ratios can be written. You count the number of reactant and product compounds that exist in the reaction to find the number of mole ratios that can be written. This means that the number of mole ratios that can be written can be determined before the equation for a chemical reaction is even balanced.

Check Your Progress

Summary

- Balanced chemical equations can be interpreted in terms of moles, mass, and representative particles (atoms, molecules, formula units).
- The law of conservation of mass applies to all chemical reactions.
- Mole ratios are derived from the coefficients of a balanced chemical equation. Each mole ratio relates the number of moles of one reactant or product to the number of moles of another reactant or product in the chemical reaction.

Demonstrate Understanding

5. **Compare** the mass of the reactants and the mass of the products in a chemical reaction, and explain how these masses are related.
6. **State** how many mole ratios can be written for a chemical reaction involving three substances.
7. **Categorize** the ways in which a balanced chemical equation can be interpreted.
8. **Apply** The general form of a chemical reaction is $xA + yB \rightarrow zAB$. In the equation, A and B are elements and x, y and z are coefficients. State the mole ratios for this reaction.
9. **Apply** Hydrogen peroxide, H_2O_2, decomposes to produce water and oxygen. Write a balanced chemical equation for this reaction, and determine the possible mole ratios.
10. **Model** Write the mole ratios for the reaction of hydrogen gas and oxygen gas, $2H_2(g) + O_2(g) \rightarrow 2H_2O$. Make a sketch of six hydrogen molecules reacting with the correct number of oxygen molecules. Show the water molecules produced.

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LESSON 2
STOICHIOMETRIC CALCULATIONS

FOCUS QUESTION

How do you determine the amounts of products and reactants involved in chemical reactions?

Using Stoichiometry

All stoichiometric calculations use mole ratios based on a balanced chemical equation. Mass-to-mole conversions are also needed.

Stoichiometric mole-to-mole conversion

The reaction between potassium and water is shown in **Figure 3.** The equation is:

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

Here, you know that two moles of potassium yield one mole of hydrogen. But how much hydrogen would be produced if only 0.0400 mol of potassium were used? To answer this question, identify the given substance and the substance that you need to determine. The given substance is 0.0400 mol of potassium. The unknown is the number of moles of hydrogen.

Figure 3 Potassium metal reacts vigorously with water, releasing so much heat that the hydrogen gas formed in the reaction catches fire.

To solve the problem, you need to know how the unknown moles of hydrogen are related to the known moles of potassium. Mole ratios are used as conversion factors to convert the known number of moles of one substance to the unknown number of moles of another substance in the same reaction. Several mole ratios can be written from the equation, but how do you choose the correct one?

As shown below, the correct mole ratio, 1 mol H_2 to 2 mol K, has moles of unknown in the numerator and moles of known in the denominator. Using this mole ratio converts the moles of potassium to the unknown number of moles of hydrogen.

$$\text{moles of known} \times \frac{\text{moles of unknown}}{\text{moles of known}} = \text{moles of unknown}$$

$$0.0400\,\cancel{\text{mol K}} \times \frac{1\text{ mol H}_2}{2\,\cancel{\text{mol K}}} = 0.0200\text{ mol H}_2$$

3D THINKING DCI CCC SEP

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

ChemLAB: Determine the Mole Ratio

Analyze and interpret data to determine the quantity of the limiting reactant and percent yield by using your knowledge of the conservation of matter.

Quick Investigation: Apply Stoichiometry

Carry out an investigation to determine the quantity of sodium carbonate that is produced when baking soda decomposes.

The following Example Problems show mole-to-mole, mole-to-mass, and mass-to-mass stoichiometry problems. The process used to solve these problems is outlined in the Problem-Solving Strategy below.

PROBLEM-SOLVING STRATEGY

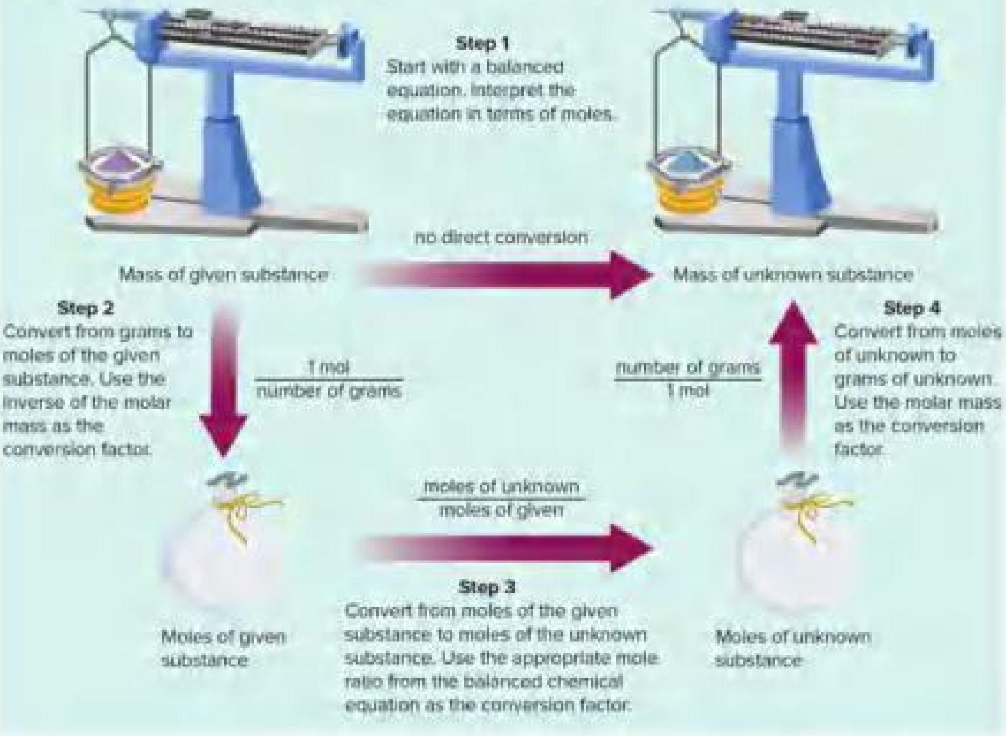
Mastering Stoichiometry

The flowchart below outlines the steps used to solve mole-to-mole, mole-to-mass, and mass-to-mass stoichiometric problems.

1. Complete Step 1 by writing the balanced chemical equation for the reaction.
2. To determine where to start your calculations, note the unit of the given substance.
   - If mass (in grams) of the given substance is the starting unit, begin your calculations with Step 2.
   - If amount (in moles) of the given substance is the starting unit, skip Step 2 and begin your calculations with Step 3.
3. The end point of the calculation depends on the desired unit of the unknown substance.
   - If the answer must be in moles, stop after completing Step 3.
   - If the answer must be in grams, stop after completing Step 4.

Apply the Strategy

Apply the Problem-Solving Strategy to Example Problems 2, 3, and 4.



EXAMPLE Problem 2

MOLE-TO-MOLE STOICHIOMETRY One disadvantage of burning propane (C_3H_8) is that carbon dioxide (CO_2) is one of the products. The released carbon dioxide increases the concentration of CO_2 in the atmosphere. How many moles of CO_2 are produced when 10.0 mol of C_3H_8 are burned in excess oxygen in a gas grill?

1 ANALYZE THE PROBLEM

You are given moles of the reactant, C_3H_8 and must find the moles of the product, CO_2. First write the balanced chemical equation, then convert from moles of C_3H_8 to moles of CO_2. The correct mole ratio has moles of unknown substance in the numerator and moles of known substance in the denominator.

Known

moles C_3H_8 = 10.0 mol C_3H_8

Unknown

moles CO_2 = ? mol CO_2

2 SOLVE FOR THE UNKNOWN

Write the balanced chemical equation for the combustion of C_3H_8. Use the correct mole ratio to convert moles of known (C_3H_8) to moles of unknown (CO_2).

10.0 mol ? mol

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

Mole ratio: $\dfrac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}$

$$10.0 \cancel{\text{mol } C_3H_8} \times \frac{3 \text{ mol } CO_2}{1 \cancel{\text{mol } C_3H_8}} = 30.0 \text{ mol } CO_2$$

Burning 10.0 moles of C_3H_8 produces 30.0 moles CO_2.

3 EVALUATE THE ANSWER

Because the given number of moles has three significant figures, the answer also has three figures. The balanced chemical equation indicates that 1 mol of C_3H_8 produces 3 mol of CO_2. Thus, 10.0 mol of C_3H_8 produces three times as many moles of CO_2, or 30.0 mol.

PRACTICE Problems

ADDITIONAL PRACTICE

11. Methane and sulfur react to produce carbon disulfide (CS_2), a liquid often used in the production of cellophane.

$$\_\_\_CH_4(g) + \_\_\_S_8(s) \rightarrow \_\_\_CS_2(l) + \_\_\_H_2S(g)$$

a. Balance the equation.

b. Calculate the moles of CS_2 produced when 1.50 mol S_8 is used.

c. How many moles of H_2S are produced?

12. CHALLENGE Sulfuric acid (H_2SO_4) is formed when sulfur dioxide (SO_2) reacts with oxygen and water.

a. Write the balanced chemical equation for the reaction.

b. How many moles of H_2SO_4 are produced from 12.5 moles of SO_2?

c. How many moles of O_2 are needed?

Stoichiometric mole-to-mass conversion

Now, suppose you know the number of moles of a reactant or product in a reaction and you want to calculate the mass of another product or reactant. This is an example of a mole-to-mass conversion.

EXAMPLE Problem 3

MOLE-TO-MASS STOICHIOMETRY Determine the mass of sodium chloride (NaCl), commonly called table salt, produced when 1.25 mol of chlorine gas (Cl_2) reacts vigorously with excess sodium.

1 ANALYZE THE PROBLEM

You are given the moles of the reactant, Cl_2, and must determine the mass of the product, NaCl. You must convert from moles of Cl_2 to moles of NaCl using the mole ratio from the equation. Then, you need to convert moles of NaCl to grams of NaCl using the molar mass as the conversion factor.

Known	Unknown
moles of chlorine = 1.25 mol Cl_2	mass of sodium chloride = ? g NaCl

2 SOLVE FOR THE UNKNOWN

$$\underset{}{2Na(s)} + \overset{1.25\text{ mol}}{Cl_2(g)} \rightarrow \overset{?\text{ g}}{2NaCl(s)}$$

Write the balanced chemical equation, and identify the known and the unknown values.

Mole ratio: $\frac{2\text{ mol NaCl}}{1\text{ mol }Cl_2}$

$$1.25\ \cancel{\text{mol }Cl_2} \times \frac{2\text{ mol NaCl}}{1\ \cancel{\text{mol }Cl_2}} = 2.50\text{ mol NaCl}$$

Multiply moles of Cl_2 by the mole ratio to get moles of NaCl.

$$2.50\ \cancel{\text{mol NaCl}} \times \frac{58.44\text{ g NaCl}}{1\ \cancel{\text{mol NaCl}}} = 146\text{ g NaCl}$$

Multiply moles of NaCl by the molar mass to get grams of NaCl.

3 EVALUATE THE ANSWER

Because the given number of moles has three significant figures, the mass of NaCl also has three. To quickly assess whether the calculated mass value for NaCl is correct, perform the calculations in reverse: divide the mass of NaCl by the molar mass of NaCl, and then divide the result by 2. You will obtain the given number of moles of Cl_2.

PRACTICE Problems

ADDITIONAL PRACTICE

13. Sodium chloride is decomposed into the elements sodium and chlorine by means of electrical energy. How much chlorine gas, in grams, is obtained from the process diagrammed at right?

14. CHALLENGE Titanium is a transition metal used in many alloys because it is extremely strong and lightweight. Titanium tetrachloride ($TiCl_4$) is extracted from titanium oxide (TiO_2) using chlorine and coke (carbon).

$$TiO_2(s) + C(s) + 2Cl_2(g) \rightarrow TiCl_4(s) + CO_2(g)$$

a. What mass of Cl_2 gas is needed to react with 1.25 mol of TiO_2?

b. What mass of C is needed to react with 1.25 mol of TiO_2?

c. What is the mass of all of the products formed by reaction with 1.25 mol of TiO_2?

Stoichiometric mass-to-mass calculations

If you were preparing to carry out a chemical reaction in the laboratory, you would need to know how much of each reactant to use in order to produce the mass of product you required. You can use the balanced chemical reaction to determine this amount. You could also use the balanced chemical reaction and mole ratios to determine how much of a product you would obtain when starting with a known mass of a reactant.

Example Problem 4 demonstrates how you can use a measured mass of the known substance, the balanced chemical equation, and mole ratios from the equation to find the mass of the unknown substance. Notice, as always, a balanced chemical equation is essential in the calculations associated with any chemical reaction. Without the ratios between reactants and products, it would not be possible to convert between moles of a reactant and the moles of a product.

EXAMPLE Problem 4

MASS-TO-MASS STOICHIOMETRY Ammonium nitrate (NH_4NO_3), an important fertilizer, produces dinitrogen monoxide (N_2O) gas and H_2O when it decomposes. Determine the mass of H_2O produced from the decomposition of 25.0 g of solid NH_4NO_3.

1 ANALYZE THE PROBLEM

Write the balanced equation and convert the known mass of the reactant to moles of the reactant. Next, use a mole ratio to relate moles of the reactant to moles of the product. Then, use the molar mass to convert from moles of the product to the mass of the product.

Known

mass of ammonium nitrate = 25.0 g NH_4NO_3

Unknown

mass of water = ? g H_2O

2 SOLVE FOR THE UNKNOWN

$$\underset{NH_4NO_3(s)}{25.0\text{ g}} \rightarrow N_2O(g) + \underset{2H_2O(g)}{?\text{ g}}$$

Write the balanced chemical equation and identify the known and unknown values.

$$25.0\text{ g }\cancel{NH_4NO_3} \times \frac{1\text{ mol }NH_4NO_3}{80.04\text{ g }\cancel{NH_4NO_3}} = 0.312\text{ mol }NH_4NO_3$$

Multiply grams of NH_4NO_3 by the inverse of molar mass to get moles of NH_4NO_3.

$$\text{Mole ratio: } \frac{2\text{ mol }H_2O}{1\text{ mol }NH_4NO_3}$$

$$0.312\text{ }\cancel{\text{mol }NH_4NO_3} \times \frac{2\text{ mol }H_2O}{1\text{ }\cancel{\text{mol }NH_4NO_3}} = 0.624\text{ mol }H_2O$$

Multiply moles of NH_4NO_3 by the mole ratio to get moles of H_2O.

$$0.624\text{ }\cancel{\text{mol }H_2O} \times \frac{18.02\text{ g }H_2O}{1\text{ }\cancel{\text{mol }H_2O}} = \mathbf{11.2\text{ g }H_2O}$$

Multiply moles of H_2O by the molar mass to get grams.

EXAMPLE Problem 4 (continued)

3 EVALUATE THE ANSWER

The number of significant figures in the answer, three, is determined by the given grams of NH_4NO_3. To verify that the mass of H_2O is correct, perform the calculations in reverse.

PRACTICE Problems ADDITIONAL PRACTICE

15. One of the reactions used to inflate automobile air bags involves sodium azide (NaN_3): $2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$. Determine the mass of N_2 produced from the decomposition of NaN_3 shown below.

100.0 g $NaN_3 \rightarrow$? g $N_2(g)$

16. CHALLENGE In the formation of acid rain, sulfur dioxide (SO_2) reacts with oxygen and water in the air to form sulfuric acid (H_2SO_4). Write the balanced chemical equation for the reaction. If 2.50 g of SO_2 reacts with excess oxygen and water, how much H_2SO_4, in grams, is produced?

Check Your Progress

Summary

- Chemists use stoichiometric calculations to predict the amounts of reactants used and products formed in specific reactions.
- The first step in solving stoichiometric problems is writing the balanced chemical equation.
- Mole ratios derived from the balanced chemical equation are used in stoichiometric calculations.
- Stoichiometric problems make use of mole ratios to convert between mass and moles.

Demonstrate Understanding

17. **Explain** why a balanced chemical equation is needed to solve a stoichiometric problem.
18. **Explain** how units can be used to guide calculations in the four steps used in solving stoichiometric problems.
19. **Describe** how a mole ratio is correctly expressed when it is used to solve a stoichiometric problem.
20. **Apply** How can you determine the mass of liquid bromine (Br_2) needed to react completely with a given mass of magnesium?
21. **Calculate** Hydrogen reacts with excess nitrogen as follows:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

If 27.0 g of H_2 reacts, what mass of NH_3 is formed?

22. **Design** a concept map for the following reaction.

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

The concept map should explain how to determine the mass of $CaCl_2$ produced from a given mass of HCl.

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LESSON 3
LIMITING REACTANTS

FOCUS QUESTION

How do you know if you have enough of a reactant?

Why do reactions stop?

Rarely in nature are the reactants present in the exact ratios specified by the balanced chemical equation. Generally, one or more reactnts are in excess and the reaction proceeds until all of one reactant is used up.

When a reaction is carried out in the laboratory, the same principle applies. Usually, one or more of the reactants are in excess, while one is limited. The amount of product depends on the reactant that is limited.

Limiting and excess reactants

Recall the reaction of sodium hydrogen sulfite with potassium permanganate from Lesson 1. After a colorless solution forms, adding more sodium hydrogen sulfite has no effect because there is no potassium permanganate left to react. Potassium permanganate is a limiting reactant. As the name implies, the **limiting reactant** limits the extent of the reaction and thereby determines the amount of product formed. A portion of all the other reactants remains after the reaction stops. Reactants leftover when a reaction stops are **excess reactants.**

To help you understand limiting and excess reactants, consider the analogy in **Figure 4.** From the available tools, four complete sets consisting of a pair of pliers, a hammer, and two screwdrivers can be assembled. The number of sets is limited by the number of available hammers. Pliers and screwdrivers remain in excess.

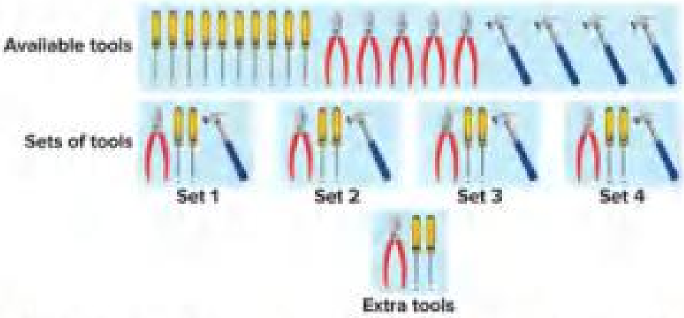

Figure 4 Each tool set must have one hammer, so only four sets can be assembled.

3D THINKING DCI CCC SEP

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Conservation of Mass

HS-PS1-7. Use mathematical representations to support the claim that atoms, and therefore mass, are conserved during a chemical reaction.

Virtual Investigation: Limiting Reactants

Use mathematics and computational thinking to determine the quantity of the limiting reactant.

Figure 5 If you check all of the atoms present before and after the reaction, you will find that some of the nitrogen molecules are unchanged. These nitrogen molecules are the excess reactant.

Determining the limiting reactant

The calculations you did in the previous section were based on having the reactants present in the ratio described by the balanced chemical equation. When this is not the case, the first thing you must do is determine which reactant is limiting.

Consider the reaction shown in **Figure 5,** in which three molecules of nitrogen (N_2) and three molecules of hydrogen (H_2) react to form ammonia (NH_3). First, all the nitrogen and hydrogen molecules are separated into individual atoms. These atoms are available for reassembling into ammonia molecules, just as the tools in **Figure 4** can be assembled into tool kits. How many molecules of ammonia can be produced from the available atoms? Two ammonia molecules can be assembled from the hydrogen atoms and nitrogen atoms because only six hydrogen atoms are available—three for each ammonia molecule. When the hydrogen is gone, two unreacted molecules of nitrogen remain. Thus, hydrogen is the limiting reactant and nitrogen is the excess reactant. It is important to know which reactant is limiting because the amount of product formed depends on this reactant.

Get It?

Extend How many more hydrogen molecules would be needed to completely react with excess nitrogen molecules shown in **Figure 5?**

Calculating the Amount of Product When a Reactant Is Limiting

Figure 6 Natural rubber, which is soft and sticky, is hardened in a chemical process called vulcanization. During vulcanization, molecules become linked together, forming a durable material that is harder, smoother, and less sticky, making it ideal for many products, such as this caster.

How can you calculate the amount of product formed when one of the reactants is limiting? Consider the formation of disulfur dichloride (S_2Cl_2), which is used to vulcanize rubber. As shown in **Figure 6,** the properties of vulcanized rubber make it useful for many products. In the production of disulfur dichloride, molten sulfur reacts with chlorine gas according to the following equation.

$$S_8(l) + 4Cl_2(g) \rightarrow 4S_2Cl_2(l)$$

If 200.0 g of sulfur reacts with 100.0 g of chlorine, what mass of disulfur dichloride is produced?

Calculating the limiting reactant

The masses of both reactants are given. First, determine which is the limiting reactant, since the reaction stops forming product when the limiting reactant is used up.

Moles of reactants Identifying the limiting reactant involves finding the number of moles of each reactant. You do this by converting the masses of chlorine and sulfur to moles. Multiply each mass by the inverse of molar mass.

$$100.0\ \cancel{g\ Cl_2} \times \frac{1\ mol\ Cl_2}{70.91\ \cancel{g\ Cl_2}} = 1.410\ mol\ Cl_2$$

$$200.0\ \cancel{g\ S_8} \times \frac{1\ mol\ S_8}{256.5\ \cancel{g\ S_8}} = 0.7797\ mol\ S_8$$

Using mole ratios Next, determine whether the two reactants are in the correct mole ratio, derived from the balanced equation. The coefficients in the equation indicate that 4 mol of chlorine is needed to react with 1 mol of sulfur. This 4:1 ratio must be compared with the calculated ratio of the moles of available reactants above. To determine the actual ratio of moles, divide the number of available moles of chlorine by the number of available moles of sulfur.

$$\frac{1.410\ mol\ Cl_2\ available}{0.7797\ mol\ S_8\ available} = \frac{1.808\ mol\ Cl_2\ available}{1\ mol\ S_8\ available}$$

Only 1.808 mol of chlorine is available for every 1 mol of sulfur, instead of the 4 mol required by the balanced chemical equation. Therefore, chlorine is the limiting reactant.

Analyzing the excess reactant

Now that you have determined the limiting reactant and the amount of product formed, what about the excess reactant, sulfur? How much of it reacted?

Moles reacted You need to make a mole-to-mass calculation to determine the mass of sulfur needed to react completely with 1.410 mol of chlorine. First, obtain the number of moles of sulfur by multiplying the moles of chlorine by the S_8-to-Cl_2 mole ratio.

$$1.410\ \cancel{mol\ Cl_2} \times \frac{1\ mol\ S_8}{4\ \cancel{mol\ Cl_2}} = 0.3525\ mol\ S_8$$

Mass reacted Next, to obtain the mass of sulfur needed, multiply by its molar mass.

$$0.3525\ \cancel{mol\ S_8} \times \frac{265.5\ g\ S_8}{1\ \cancel{mol\ S_8}} = 90.42\ g\ S_8\ needed$$

SCIENCE USAGE V. COMMON USAGE

product

Science usage: a new substance formed during a chemical reaction *The sole reaction product was a colorless gas.*

Common usage: something produced *The cosmetics counter in the department store had hundreds of products from which to choose.*

STEM CAREER Connection

Industrial Engineer

Do you like finding ways to save time and money? An industrial engineer first observes a production process and then explores ways to make the process more efficient. Industrial engineers typically have at least a bachelor's degree in industrial engineering, but many have degrees in mechanical engineering, electrical engineering, manufacturing engineering, industrial engineering technology, or general engineering.

Excess remaining Knowing that 200.0 g of sulfur is available and that only 90.42 g of sulfur is needed, you can calculate the amount of sulfur left unreacted as follows.

$$200.0 \text{ g } S_8 \text{ available} - 90.42 \text{ g } S_8 \text{ needed} = 109.6 \text{ g } S_8 \text{ in excess}$$

EXAMPLE Problem 5

DETERMINING THE LIMITING REACTANT The reaction between solid white phosphorus (P_4) and oxygen produces solid tetraphosphorus decoxide (P_4O_{10}). This compound is often called diphosphorus pentoxide because its empirical formula is P_2O_5.

a. Determine the mass of P_4O_{10} formed if 25.0 g of P_4 and 50.0 g of oxygen are combined.

b. How much of the excess reactant remains after the reaction stops?

1 ANALYZE THE PROBLEM

You are given the masses of both reactants, so you must identify the limiting reactant and find the mass of the product. From the limiting reactant, the moles of the excess reactant used in the reaction can be determined. This can be converted to mass and subtracted from the given mass to find the amount in excess.

Known	Unknown
mass of phosphorus = 25.0 g P_4	mass of tetraphosphorus decoxide = ? g P_4O_{10}
mass of oxygen = 50.0 g O_2	mass of excess reactant = ? g excess reactant

2 SOLVE FOR THE UNKNOWN

Determine the limiting reactant.

$$\overset{25.0\text{ g}}{P_4(s)} + \overset{50.0\text{ g}}{5O_2(g)} \rightarrow \overset{?\text{ g}}{P_4O_{10}(s)}$$

Write the balanced chemical equation, and identify the known and the unknown.

Determine the number of moles of the reactants by using the inverse of molar mass.

$$25.0 \cancel{\text{ g } P_4} \times \frac{1 \text{ mol } P_4}{123.9 \cancel{\text{ g } P_4}} = 0.202 \text{ mol } P_4$$

Calculate the moles of P_4.

$$50.0 \cancel{\text{ g } O_2} \times \frac{1 \text{ mol } O_2}{32.00 \cancel{\text{ g } O_2}} = 1.56 \text{ mol } O_2$$

Calculate the moles of O_2.

Calculate the actual ratio of available moles of O_2 and available moles of P_4.

$$\frac{1.56 \text{ mol } O_2}{0.202 \text{ mol } P_4} = \frac{7.72 \text{ mol } O_2}{1 \text{ mol } P_4}$$

Calculate the ratio of moles of O_2 to moles of P_4.

Determine the mole ratio of the two reactants from the balanced chemical equation.

Mole ratio: $\frac{5 \text{ mol } O_2}{\text{mol } P_4}$

Because 7.72 mol of O_2 is available but only 5 mol is needed to react with 1 mol of P_4, O_2 is in excess and P_4 is the limiting reactant. Use the moles of P_4 to determine the moles of P_4O_{10} that will be produced. Multiply the number of moles of P_4 by the mole ratio of P_4O_{10} to P_4.

$$0.202 \cancel{\text{ mol } P_4} \times \frac{1 \text{ mol } P_4O_{10}}{1 \cancel{\text{ mol } P_4}} = 0.202 \text{ mol } P_4O_{10}$$

Calculate the moles of product (P_4O_{10}) formed.

To calculate the mass of P_4O_{10}, multiply moles of P_4O_{10} by the molar mass.

$$0.202 \cancel{\text{ mol } P_4O_{10}} \times \frac{283.9 \text{ g } P_4O_{10}}{1 \cancel{\text{ mol } P_4O_{10}}} = 57.3 \text{ g } P_4O_{10}$$

Calculate the mass of the product P_4O_{10}.

EXAMPLE Problem 5 (continued)

Because O_2 is in excess, only part of the available O_2 is consumed. Use the limiting reactant, P_4, to determine the moles and mass of O_2 used.

$$0.202\ \cancel{\text{mol } P_4} \times \frac{5 \text{ mol } O_2}{1\ \cancel{\text{mol } P_4}} = 1.01 \text{ mol } O_2$$

Multiply the moles of limiting reactant by the mole ratio to find moles of excess reactant.

Convert moles of O_2 consumed to mass of O_2 consumed.

$$1.01\ \cancel{\text{mol } O_2} \times \frac{32.00 \text{ g } O_2}{1\ \cancel{\text{mol } O_2}} = 32.3 \text{ g } O_2$$

Multiply the moles of O_2 by the molar mass.

Calculate the amount of excess O_2.

50.0 g O_2 available − 32.3 g O_2 consumed = **17.7 g O_2 in excess**

Subtract mass used from the initial mass.

3 EVALUATE THE ANSWER

All values have a minimum of three significant figures, so the mass of P_4O_{10} is correctly stated with three digits. The mass of excess O_2 (17.7 g) is found by subtracting two numbers that are accurate to the first decimal place, so the mass of excess O_2 correctly shows one decimal place. The sum of the O_2 consumed (32.3 g) and the given mass of P_4 (25.0 g) is 57.3 g, the calculated mass of the product P_4O_{10}.

PRACTICE Problems ADDITIONAL PRACTICE

23. The reaction between solid sodium and iron(III) oxide is one in a series of reactions that inflates an automobile airbag: $6Na(s) + Fe_2O_3(s) \rightarrow 3Na_2O(s) + 2Fe(s)$. If 100.0 g of Na and 100.0 g of Fe_2O_3 are used in this reaction, determine the following.
- **a.** limiting reactant
- **b.** reactant in excess
- **c.** mass of solid iron produced
- **d.** mass of excess reactant that remains after the reaction is complete

24. CHALLENGE Photosynthesis reactions in green plants use carbon dioxide and water to produce glucose ($C_6H_{12}O_6$) and oxygen. A plant has 88.0 g of carbon dioxide and 64.0 g of water available for photosynthesis. Be sure to report the correct level of accuracy based on measurements given in the question.
- **a.** Write the balanced chemical equation for the reaction.
- **b.** Determine the limiting reactant and the excess reactant.
- **c.** Determine the mass in excess.
- **d.** Determine the mass of glucose produced.

BIOLOGY Connection Your body needs vitamins, minerals, and elements in small amounts to facilitate normal metabolic reactions. A lack of these substances can lead to abnormalities in growth, development, and cell function. Phosphorus, for example, is an essential element, as phosphate groups occur regularly in strands of DNA. Potassium is needed for proper nerve function, muscle control, and blood pressure. A diet low in potassium and high in sodium might be a factor in high blood pressure. Another example is vitamin B-12. Without adequate vitamin B-12, the body is unable to synthesize DNA properly, affecting the production of red blood cells.

Why use an excess of reactant?

Many reactions stop while some of the reactants are still present. Because this is inefficient and wasteful, using an excess of one reactant causes reactions to be driven to continue until all of the limiting reactant is used up. This technique can also speed up a reaction.

Figure 7 shows an example of how you can increase efficiency. The type of Bunsen burner shown in the figure has a control that lets you adjust the amount of air that mixes with the methane gas. The burner efficiency depends on the ratio of oxygen to methane gas in the mixture. When the air is limited, the flame is yellow because of glowing bits of unburned fuel. This leaves carbon deposits on glassware. The amount of energy released is less than the amount that could have been produced if enough oxygen were available. When sufficient oxygen is present in the combustion mixture, the burner produces a hot, intense blue flame. No soot is deposited because the fuel is completely reacted.

Figure 7 With insufficient oxygen, a Bunsen burner burns with a yellow, sooty flame.

Check Your Progress

Summary

- The limiting reactant is the reactant that is completely consumed during a chemical reaction. Reactants that remain after the reaction stops are called excess reactants.
- To determine the limiting reactant, the actual mole ratio of the available reactants must be compared with the ratio of the reactants obtained from the coefficients in the balanced chemical equation.
- Stoichiometric calculations must be based on the limiting reactant.

Demonstrate Understanding

25. **Describe** the reason why a reaction between two substances comes to an end.
26. **Identify** the limiting and excess reactant in each reaction.
    a. Wood burns in a campfire.
    b. Airborne sulfur reacts with the silver plating on a teapot to produce tarnish (silver sulfide)
    c. Baking powder in batter decomposes to produce carbon dioxide.
27. **Analyze** Tetraphosphorus trisulfide (P_4S_3) is used in the match heads of some matches. It is produced in the reaction $8P_4 + 3S_8 \rightarrow 8P_4S_3$. Determine which of the following statements are incorrect, and rewrite the incorrect statements to make them correct.
    a. 4 mol of P_4 reacts with 1.5 mol of S_8 to form 4 mol P_4S_3.
    b. Sulfur is the limiting reactant when 4 mol P_4 reacts with 4 mol S_8.
    c. 6 mol P_4 reacts with 6 mol of S_8, forming 1320 g of P_4S_3.

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LESSON 4
PERCENT YIELD

FOCUS QUESTION

Do you always get the calculated amount of product out of a reaction?

How much product?

While solving stoichiometric problems, you might have concluded that chemical reactions always proceed according to the balanced equation and produce the calculated amount of product. This, however, is not the case. Most reactions never succeed in producing the predicted amount of product.

Reactions do not go to completion or yield as expected for many reasons. Reactants and products might adhere to the surfaces of their containers or evaporate. In some cases, products other than the intended ones might be formed by competing reactions, thus reducing the yield of the desired product. Or, as shown in **Figure 8,** some amount of a solid product can be left behind on filter paper or lost in the purification process. Because of this, chemists need to know how to gauge the yield of a reaction.

Figure 8 Silver chromate is formed when potassium chromate is added to silver nitrate. Note that not all of the precipitate can be removed from the filter paper. Still more of the precipitate is lost because it adheres to the sides of the beaker.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Laboratory: Stoichiometry of a Chemical Reaction
Analyze and interpret data to determine the quantity and identity of an unknown carbonate.

Identify Crosscutting Concepts
Create a table of the crosscutting concepts and fill in examples you find as you read.

Theoretical and actual yields

In many of the calculations you have performed, you have calculated the amount of product given an amount of reactant. This value is the **theoretical yield** of the reaction, and represents the maximum amount of product that can be produced from a given amount of reactant.

A chemical reaction rarely produces the theoretical yield. A chemist determines the actual yield of a reaction through an experiment in which the mass of the product is measured. The **actual yield** is the amount produced when the experiment is performed.

Percent yield

Percent yield of product is the ratio of the actual yield to the theoretical yield expressed as a percent. Chemists use this to measure the efficiency of the reaction.

Percent Yield

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

The actual yield divided by the theoretical yield multiplied by 100 is the percent yield.

EXAMPLE Problem 6

PERCENT YIELD Solid silver chromate (Ag_2CrO_4) forms when excess potassium chromate (K_2CrO_4) is added to a solution containing 0.500 g of silver nitrate ($AgNO_3$). Determine the theoretical yield of Ag_2CrO_4. Find the percent yield if the reaction yields 0.455 g of Ag_2CrO_4.

1 ANALYZE THE PROBLEM

You know the mass of a reactant and the actual yield of the product. Write the balanced chemical equation, and calculate theoretical yield by converting grams of $AgNO_3$ to moles of $AgNO_3$, moles of $AgNO_3$ to moles of Ag_2CrO_4, and moles of Ag_2CrO_4 to grams of Ag_2CrO_4. Calculate the percent yield from the actual yield and the theoretical yield.

Known
mass of silver nitrate = 0.500 g $AgNO_3$
actual yield = 0.455 g Ag_2CrO_4

Unknown
theoretical yield = ? g Ag_2CrO_4
percent yield = ? % Ag_2CrO_4

2 SOLVE FOR THE UNKNOWN

$$\overset{0.500\text{ g}}{2AgNO_3(aq)} + K_2CrO_4(aq) \rightarrow \overset{?\text{ g}}{Ag_2CrO_4(s)} + 2KNO_3(aq)$$

Write the balanced chemical equation, and identify the known and the unknown.

$$0.500\ \cancel{\text{g AgNO}_3} \times \frac{1\text{ mol AgNO}_3}{169.9\ \cancel{\text{g AgNO}_3}} = 2.94 \times 10^{-3}\text{ mol AgNO}_3$$

Convert grams of $AgNO_3$ to moles.

$$2.94 \times 10^{-3}\ \cancel{\text{mol AgNO}_3} \times \frac{1\text{ mol Ag}_2\text{CrO}_4}{2\ \cancel{\text{mol AgNO}_3}} = 1.47 \times 10^{-3}\text{ mol Ag}_2\text{CrO}_4$$

Use the mole ratio to convert moles of $AgNO_3$ to moles of Ag_2CrO_4.

$$1.47 \times 10^{-3}\ \cancel{\text{mol Ag}_2\text{CrO}_4} \times \frac{331.7\text{ g Ag}_2\text{CrO}_4}{1\ \cancel{\text{mol Ag}_2\text{CrO}_4}} = 0.488\text{ g Ag}_2\text{CrO}_4$$

Calculate the theoretical yield.

$$\frac{0.455\ \cancel{\text{g Ag}_2\text{CrO}_4}}{0.488\ \cancel{\text{g Ag}_2\text{CrO}_4}} \times 100 = 93.2\%\ \text{Ag}_2\text{CrO}_4$$

Calculate the percent yield.

EXAMPLE Problem 6 (continued)

3 EVALUATE THE ANSWER

The quantity with the fewest significant figures has three, so the percent is correctly stated with three digits. The molar mass of Ag_2CrO_4 is about twice the molar mass of $AgNO_3$, and the ratio of moles of $AgNO_3$ to moles of Ag_2CrO_4 in the equation is 2:1. Therefore, 0.500 g of $AgNO_3$ should produce about the same mass of Ag_2CrO_4. The actual yield of Ag_2CrO_4 is close to 0.500 g, so a percent yield of 93.2% is reasonable.

PRACTICE Problems ADDITIONAL PRACTICE

28. Aluminum hydroxide ($Al(OH)_3$) is often present in antacids to neutralize stomach acid (HCl). The reaction occurs as follows: $Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$. If 14.0 g of $Al(OH)_3$ is present in an antacid tablet, determine the theoretical yield of $AlCl_3$ produced when the tablet reacts with HCl.

29. Zinc reacts with iodine in a synthesis reaction: $Zn + I_2 \rightarrow ZnI_2$.

a. Determine the theoretical yield if 1.912 mol of zinc is used.

b. Determine the percent yield if 515.6 g of product is recovered.

30. CHALLENGE When copper wire is placed into a silver nitrate solution ($AgNO_3$), silver crystals and copper(II) nitrate ($Cu(NO_3)_2$) solution form.

a. Write the balanced chemical equation for the reaction.

b. If a 20.0-g sample of copper is used, determine the theoretical yield of silver.

c. If 60.0 g of silver is recovered from the reaction, determine the percent yield of the reaction.

Figure 9 Sulfur, such as these piles at Vancouver Harbor, can be extracted from petroleum products by a chemical process. Sulfur is also mined by forcing hot water into underground deposits and pumping the liquid sulfur to the surface.

Percent Yield in the Marketplace

Percent yield is important in the cost effectiveness of many industrial processes. For example, the sulfur shown in **Figure 9** is used to make sulfuric acid (H_2SO_4), an important chemical as it is a raw material used to make products such as fertilizers, detergents, pigments, and textiles. The cost of sulfuric acid affects the cost of many of the consumer items. The first two steps in the manufacturing process are shown below.

Step 1 $\quad S_8(s) + 8O_2(g) \rightarrow 8SO_2(g)$

Step 2 $\quad 2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

In the final step, SO_3 combines with water to produce H_2SO_4.

Step 3 $\quad SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

The first step, the combustion of sulfur, produces an almost 100% yield. The second step also produces a high yield if a catalyst is used at the relatively low temperature of 400°C. A catalyst is a substance that speeds a reaction but does not appear in the chemical equation. Under these conditions, the reaction is slow. Raising the temperature increases the reaction rate but decreases the yield. Often, a balance between two competing factors must be looked at in industrial applications.

To maximize yield and minimize time in the second step, engineers have devised a system in which the reactants, O_2 and SO_2, are passed over a catalyst at 400°C. Because the reaction releases a great deal of heat, the temperature gradually increases with an accompanying decrease in yield. Thus, when the temperature reaches approximately 600°C, the mixture is cooled and then passed over the catalyst again. A total of four passes over the catalyst with cooling between passes results in a yield greater than 98%.

Often, chemical engineers must look at factors such as these when dealing with industrial processes. Industry requires as high a yield as possible, in as short a period of time and at as low a cost as possible for the process to be as economical as possible. A high yield of desired products is great, but not at the expense of having to wait weeks for this to happen or to need to heat a reaction to such a high temperature that the cost to heat the reaction exceeds the revenue when selling the product. A slightly lower yield is acceptable, if cost is kept low and time needed to generate this yield is reasonable.

Chemical engineers must be in constant contact with company management to be sure that their suggestions to manage the percent yield of the chemical process meets with management expectations for the company.

Check Your Progress

Summary

- The theoretical yield of a chemical reaction is the maximum amount of product that can be produced from a given amount of reactant. Theoretical yield is calculated from the balanced chemical equation.
- The actual yield is the amount of product produced. Actual yield must be obtained through experimentation.
- Percent yield is the ratio of actual yield to theoretical yield expressed as a percent. High percent yield is important in reducing the cost of every product produced through chemical processes.

Demonstrate Understanding

31. **Identify** which type of yield – theoretical yield, actual yield, or percent yield – is a measure of the efficiency of a chemical reaction.
32. **List** several reasons why the actual yield from a chemical reaction is not usually equal to the theoretical yield.
33. **Explain** how percent yield is calculated.
34. **Apply** In an experiment, you combine 83.77 g of iron with an excess of sulfur and then heat the mixture to obtain iron (III) sulfide.

$$2Fe(s) + 3S(s) \longrightarrow Fe_2S_3(s)$$

What is the theoretical yield, in grams, of iron (III) sulfide?

35. **Calculate** the percent yield of the reaction of magnesium with excess oxygen:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Reaction Data	
Mass of empty crucible	35.67 g
Mass of crucible and Mg	38.06 g
Mass of crucible and MgO (after heated)	39.15 g

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NATURE OF SCIENCE

The Stoichiometry That Just Might Save Your Life

Millions of people around the world take aspirin every day. Most of them probably don't think much about the little tablet of acetylsalicylic acid they're swallowing—they just hope it cures their headache. Aspirin may be inexpensive and easy to find today, but this wasn't always the case. Humans and aspirin share a long history—one that tells as much about the scientific process as it does about chemistry.

Today's aspirin was influenced by the work of many people over thousands of years.

The World's Wonder Drug

References to the use of willow bark as medicine have been found in the records of almost every early civilization. For centuries, people who suffered from pain or fever consumed the bark in the form of powder, made it into tea, or even chewed it whole.

In 1838, organic chemists successfully extracted aspirin's primary active compound from willow bark: salicylic acid. However, the extraction process used resulted low yields. In 1860, scientist Hermann Kolbe synthesized salicylic acid in the lab. The percent yield of salicylic acid from Kolbe's chemical reaction was much greater than its extraction from bark. This increased the availability of salicylic acid and led to its widespread use.

Salicylic acid was an effective medicine, but it caused unpleasant side effects. Patients treated with dosages high enough to combat pain and inflammation often experienced stomach irritation, nausea, and vomiting.

One patient was the father of Felix Hoffmann, a chemist at the Bayer chemical company. Hoffmann reacted salicylic acid with excess acetic anhydride, forming acetylsalicylic acid. Acetylsalicylic acid still effectively combatted pain, but with milder side effects. Eventually, Bayer officially trademarked the name aspirin for the drug. By 1915, the tablets were available to purchase over the counter, and the world's first wonder drug was born.

USE A MODEL TO DESCRIBE

Make and use a model to describe Hoffmann's chemical reaction for aspirin synthesis. Explain how stoichiometry can be used to find the percent yield in the reaction.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 DEFINING STOICHIOMETRY

- Balanced chemical equations can be interpreted in terms of moles, mass, and representative particles (atoms, molecules, formula units).
- The law of conservation of mass applies to all chemical reactions.
- Mole ratios are derived from the coefficients of a balanced chemical equation. Each mole ratio relates the number of moles of one reactant or product to the number of moles of another reactant or product in the chemical reaction.

- stoichiometry
- mole ratio

Lesson 2 STOICHIOMETRIC CALCULATIONS

- Chemists use stoichiometric calculations to predict the amounts of reactants used and products formed in specific reactions.
- The first step in solving stoichiometric problems is writing the balanced chemical equation.
- Mole ratios derived from the balanced chemical equation are used in stoichiometric calculations.
- Stoichiometric problems make use of mole ratios to convert between mass and moles.

Lesson 3 LIMITING REACTANTS

- The limiting reactant is the reactant that is completely consumed during a chemical reaction.
- Reactants that remain after the reaction stops are called excess reactants.
- To determine the limiting reactant, the actual mole ratio of the available reactants must be compared with the ratio of the reactants obtained from the coefficients in the balanced chemical equation.
- Stoichiometric calculations must be based on the limiting reactant.

- limiting reactant
- excess reactant

Lesson 4 PERCENT YIELD

- The theoretical yield of a chemical reaction is the maximum amount of product that can be produced from a given amount of reactant. Theoretical yield is calculated from the balanced chemical equation.
- The actual yield is the amount of product produced. Actual yield must be obtained through experimentation.
- Percent yield is the ratio of actual yield to theoretical yield expressed as a percent. High percent yield is important in reducing the cost of every product produced through chemical processes.

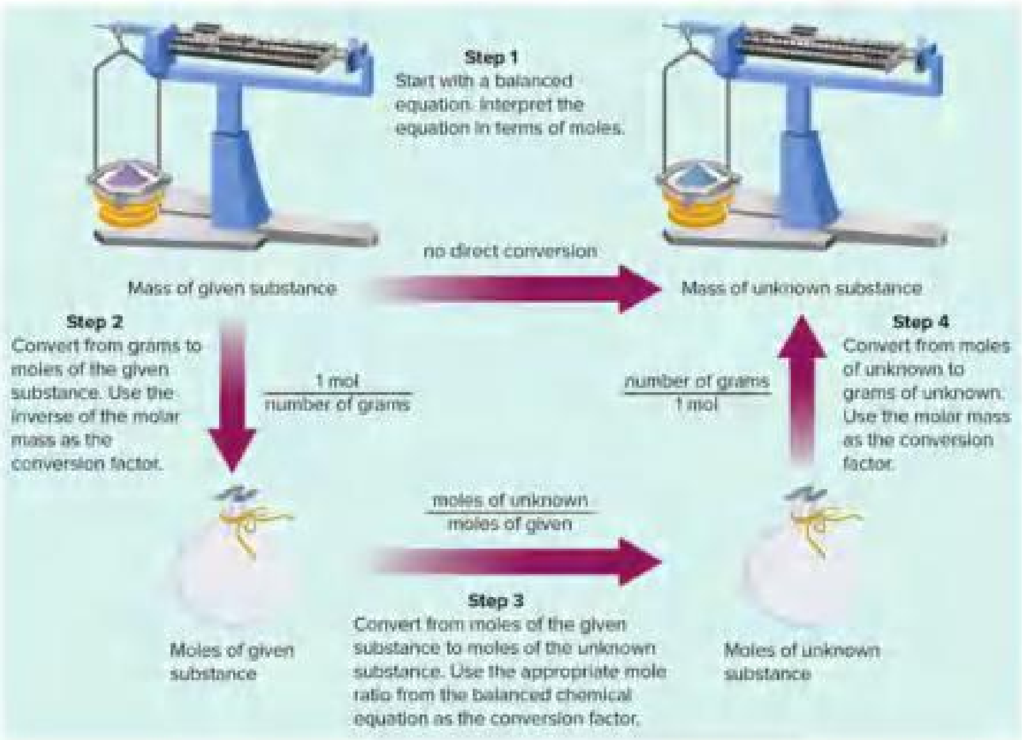
$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

- theoretical yield
- actual yield
- percent yield

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Summarizing stoichiometry Always remember that every stoichiometry problem begins with a balanced chemical equation. After that, it's simply a matter of conversion factors—molar mass or its inverse to navigate between mass and moles; the coefficients from the balanced chemical equation to navigate between moles of reactants and products. Refer to the chart below as a reminder of the basic stoichiometry problem-solving process.



THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

How much CO_2 did this field of corn need to grow?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will apply your evidence from this module and complete your project.

GO FURTHER

Based on Real Data[1,2]

SEP Data Analysis Lab

Can rocks on the Moon provide an effective oxygen source?

Scientists, looking for an oxygen source for future long-duration lunar missions, are researching ways to extract oxygen from lunar soil and rock. Analysis of samples identifies the oxides in lunar soil as well as each oxide's percent-by-weight of the soil.

CER Analyze and Interpret Data

1. **Claim, Evidence, Reasoning** Scientists want to release the oxygen from its metal oxide using a decomposition reaction: metal oxide → metal + oxygen. To assess the viability of this idea, determine the amount of oxygen per kilogram contained in each of the oxides found in lunar soil.
2. **Claim, Evidence, Reasoning** Determine the theoretical yield of oxygen from the oxides present in a 1.00-kg sample of lunar soil.
3. **Claim, Evidence, Reasoning** Using methods currently available, scientists can produce 15 kg of oxygen from 100 kg of lunar soil. What is the percent yield of the process?

Data and Observations

Moon Rock Data[1]

Oxide	% Mass of Soil
SiO_2	47.3%
Al_2O_3	17.8%
CaO	11.4%
FeO	10.5%
MgO	9.6%
TiO_2	1.6%
Na_2O	0.7%
K_2O	0.6%
Cr_2O_3	0.2%
MnO	0.1%

[1]Data obtained from: McKay, et al. 1994. JSC-1: A new lunar soil stimulant. *Engineering, Construction, and Operations in Space IV*: 857–866, American Society of Civil Engineers.

[2]Data obtained from: Berggren, et al. 2005. Carbon monoxide silicate reduction system. *Space Resources Roundtable VII*.

STATES OF MATTER

ENCOUNTER THE PHENOMENON

Why does water naturally exist as a solid, liquid, and gas on Earth?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

CER Claim, Evidence, Reasoning

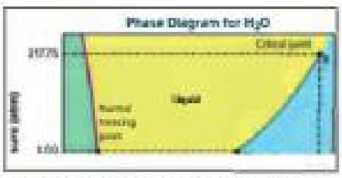
Make Your Claim Use your CER chart to make a claim about why water naturally exists as a solid, liquid, and gas on Earth.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.

LESSON 2: Explore & Explain: Polar Interactions



LESSON 4: Explore & Explain: Phase Diagrams

LESSON 1
GASES

FOCUS QUESTION

Do all gases behave the same way?

The Kinetic-Molecular Theory

You have learned that composition (the types of atoms present) and structure (their arrangement) determine the chemical properties of matter. Composition and structure also affect the physical properties of matter. Based solely on physical appearance, you can distinguish between the solids and liquids, as shown in **Figure 1.** By contrast, substances that are gases at room temperature usually display similar physical properties despite their different compositions. Why is there so little variation in behavior among gases? Why are the physical properties of gases different from those of liquids and solids?

Gold Graphite Mercury

Figure 1 You can distinguish some materials by looking at them, but this is not true for many gases.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Virtual Investigation: Kinetic Theory

Use a model to discover the pattern found in the Maxwell-Boltzmann distribution and describe properties of gases.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

By the eighteenth century, scientists knew how to collect gaseous products by displacing water. Now, they could observe and measure properties of individual gases. About 1860, chemists Ludwig Boltzmann and James Maxwell, who were working in different countries, each proposed a model to explain the properties of gases. That model is the kinetic-molecular theory. Because all of the gases known to Boltzmann and Maxwell contained molecules, the name of the model refers to molecules. The word *kinetic* comes from a Greek word meaning *to move*. Objects in motion have energy called kinetic energy. The **kinetic-molecular theory** describes the behavior of matter in terms of particles in motion. The model makes several assumptions about the size, motion, and energy of gas particles.

Particle size

Gases consist of small particles that are separated from one another by empty space. The volume of the particles is small compared with the volume of the empty space. Because gas particles are far apart, they experience no significant attractive or repulsive forces.

Particle motion

Gas particles are in constant, random motion. Particles move in a straight line until they collide with other particles or with the walls of their container, as shown in **Figure 2.** Collisions between gas particles are elastic. An **elastic collision** is one in which no kinetic energy is lost. Kinetic energy can be transferred between colliding particles, but the total kinetic energy of the two particles does not change.

Particle energy

Two factors determine the kinetic energy of a particle: mass and velocity. The kinetic energy of a particle can be represented by the following equation.

$$KE = \frac{1}{2}mv^2$$

KE is kinetic energy, m is the mass of the particle, and v is its velocity. Velocity reflects both the speed and the direction of motion. In a sample of a single gas, all particles have the same mass, but all particles do not have the same velocity. Therefore, all particles do not have the same kinetic energy. **Temperature** is a measure of the average kinetic energy of the particles in a sample of matter.

Figure 2 Kinetic energy can be transferred between gas particles during an elastic collision. Between collisions, the particles move in straight lines.

Explain *the influence that gas particles have on each other, both in terms of collisions and what happens to particles between collisions.*

Explaining the Behavior of Gases

The kinetic-molecular theory helps explain the behavior of gases. For example, the constant motion of gas particles allows a gas to expand until it fills its container, such as when you inflate a beach ball. As you blow air into the ball, the air particles spread out and fill the inside of the container—the beach ball.

Low density

Remember that density is mass per unit volume. The density of chlorine gas is 2.898×10^{-3} g/mL at 20°C; the density of solid gold is 19.3 g/mL. Gold is more than 6700 times as dense as chlorine. This large difference cannot be due only to the difference in mass between gold atoms and chlorine molecules (about 3:1). As the kinetic-molecular theory states, a great deal of space exists between gas particles. Thus, there are fewer chlorine molecules than gold atoms in the same volume.

Compression and expansion

If you squeeze a pillow made of foam, you can compress it; that is, you can reduce its volume. Air, which is a mixture of gases, is also compressible. The large amount of empty space between the particles in the air allows air to be squeezed into a smaller volume. When the volume of a container is made larger, the random motion of the particles fills the available space. **Figure 3** illustrates what happens to the density of a gas in a container as it is compressed and as it is allowed to expand.

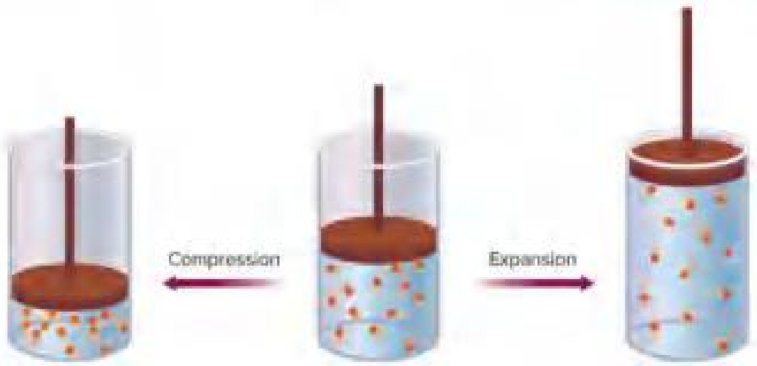


Figure 3 In a closed container, compression and expansion change the volume occupied by a constant mass of particles.

Relate *the change in volume to the density of the gas particles in each cylinder.*

CCC CROSSCUTTING CONCEPTS

Systems and System Models Review the assumptions made by the kinetic-molecular theory. Develop your own visual or physical representation to help others understand the kinetic-molecular theory and the behavior of gases. What are the limitations of your model?

WORD ORIGINS

gas

comes from the Latin word *chaos*, which means *space*

Diffusion and effusion

According to the kinetic-molecular theory, there are no significant forces of attraction between gas particles. Thus, gas particles can flow easily past each other. Often, the space into which a gas flows is already occupied by another gas. The random motion of the gas particles causes the gases to mix until they are evenly distributed. **Diffusion** is the term used to describe the movement of one material through another. The term might be new, but you are probably familiar with the process.

If food is cooking in the kitchen, you can smell it throughout the house because the gas particles diffuse. Particles diffuse from an area of high concentration (the kitchen) to one of low concentration (the other rooms in the house).

Effusion is a process related to diffusion. During effusion, a gas escapes through a tiny opening. What happens when you puncture a container, such as a balloon or a tire?

In 1846, Thomas Graham conducted experiments to measure the rates of effusion for different gases at the same temperature. Graham designed his experiments so that the gases effused into a vacuum–space containing no matter. He discovered an inverse relationship between effusion rates and molar mass. **Graham's law of effusion** states that the rate of effusion for a gas is inversely proportional to the square root of its molar mass.

Graham's Law

$$\text{Rate of effusion} \propto \frac{1}{\sqrt{\text{molar mass}}}$$

The rate of diffusion or effusion of a gas is inversely proportional to the square root of its molar mass.

The rate of diffusion depends mainly on the mass of the particles involved. Lighter particles diffuse more rapidly than heavier particles. Recall that different gases at the same temperature have the same average kinetic energy as described by the equation $KE = \frac{1}{2}mv^2$. However, the mass of gas particles varies from gas to gas. For lighter particles to have the same average kinetic energy as heavier particles, they must have, on average, a greater velocity.

Graham's law also applies to rates of diffusion, which is logical because heavier particles diffuse more slowly than lighter particles at the same temperature. Using Graham's law, you can set up a proportion to compare the diffusion rates for two gases.

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

Get It?

Explain why the rate of diffusion depends on the mass of the particles.

EXAMPLE Problem 1

GRAHAM'S LAW Ammonia has a molar mass of 17.0 g/mol; hydrogen chloride has a molar mass of 36.5 g/mol. What is the ratio of their diffusion rates?

1 ANALYZE THE PROBLEM

You are given the molar masses for ammonia and hydrogen chloride. To find the ratio of the diffusion rates for ammonia and hydrogen chloride, use the equation for Graham's law of effusion.

Known
molar mass_{HCl} = 36.5 g/mol
molar mass_{NH_3} = 17.0 g/mol

Unknown
ratio of diffusion rates = ?

2 SOLVE FOR THE UNKNOWN

$$\frac{\text{Rate}_{NH_3}}{\text{Rate}_{HCl}} = \sqrt{\frac{\text{molar mass}_{HCl}}{\text{molar mass}_{NH_3}}}$$

State the ratio derived from Graham's law.

$$= \sqrt{\frac{36.5\ \text{g/mol}}{17.0\ \text{g/mol}}} = \mathbf{1.47}$$

Substitute molar mass_{HCl} = 36.5 g/mol and molar mass_{NH_3} = 17.0 g/mol.

The ratio of diffusion rates is 1.47.

3 EVALUATE THE ANSWER

A ratio of roughly 1.5 is logical because molecules of ammonia are about half as massive as molecules of hydrogen chloride. Because the molar masses have three significant figures, the answer also does. Note that the units cancel, and the answer is stated correctly without any units.

PRACTICE Problems ADDITIONAL PRACTICE

1. Calculate the ratio of effusion rates for nitrogen (N_2) and neon (Ne).
2. Calculate the ratio of diffusion rates for carbon monoxide and carbon dioxide.
3. **CHALLENGE** What is the rate of effusion for a gas that has a molar mass twice that of a gas that effuses at a rate of 3.6 mol/min?

Gas Pressure

Have you watched someone try to walk across snow, mud, or hot asphalt in high heels? If so, you might have noticed that the heels sank into the soft surface. But that same person walking across the same surface while wearing shoes or boots does not have the same problem. Why does a person sink when wearing high heels but does not sink when wearing boots?.

Figure 4, on the next page, shows that in each case, the force pressing down on the soft surface is related to the person's mass. With boots, the force is spread out over a larger area. **Pressure** is defined as force per unit area. The area of the bottom of a boot is much larger than the area of the bottom of a high-heeled shoe. So, the pressure on the soft surface is less with a boot than it is with high heels.

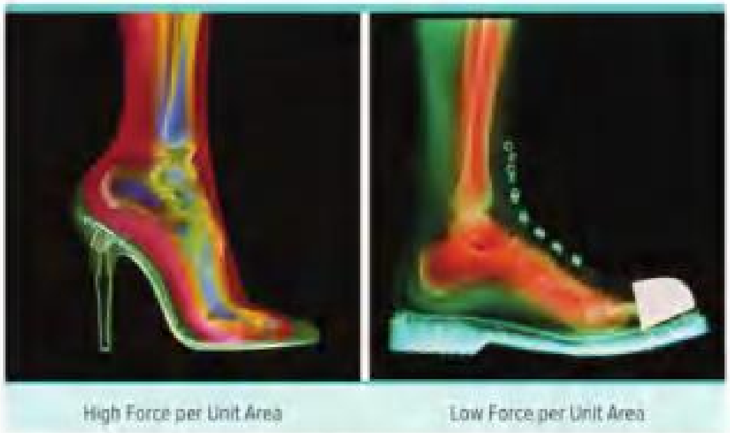


Figure 4 High-heeled shoes increase the pressure on a surface because the area touching the floor is reduced. In flatter-heeled shoes, such as boots, the force is applied over a larger area.

Infer *where the highest pressure is located between the floor and high-heel shoe.*

Gas particles also exert pressure when they collide with the walls of their container. Because an individual gas particle has little mass, it can exert little pressure. However, a liter-sized container could hold 10^{22} gas particles. With this many particles colliding, the pressure can be high.

Air pressure

Earth is surrounded by an atmosphere that extends into space for hundreds of kilometers. Because the particles in air move in every direction, they exert pressure in all directions. This pressure is called atmospheric pressure, or air pressure. Air pressure varies at different points on Earth. Because gravity is greater at the surface of Earth, there are more particles than at higher altitudes where the force of gravity is less. Fewer particles at higher elevations exert less force than the greater concentration of particles at lower altitudes. Therefore, air pressure is less at higher altitudes than it is at sea level. At sea level, atmospheric pressure is about one-kilogram per square centimeter.

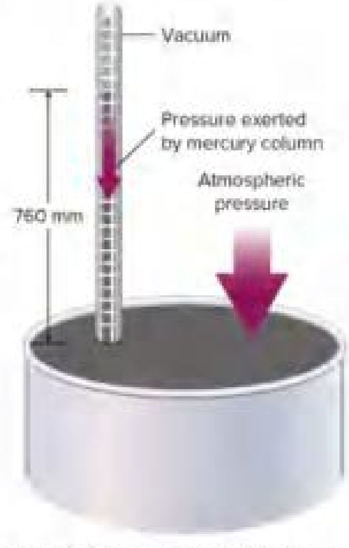
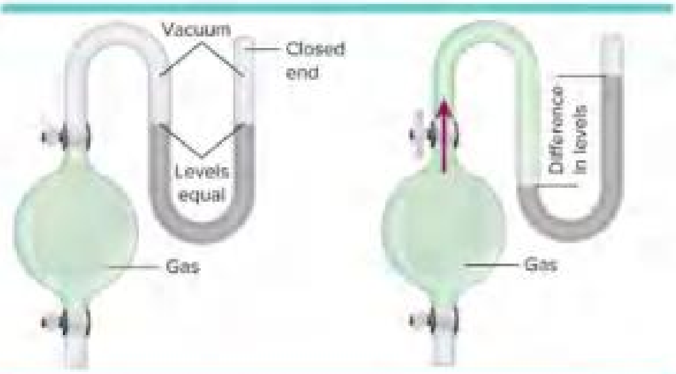


Figure 5 Torricelli was the first to design equipment to show that the atmosphere exerted pressure.

Measuring air pressure Italian physicist Evangelista Torricelli (1608–1647) was the first to demonstrate that air exerted pressure. He noticed that water pumps were unable to pump water higher than about 10 m. He hypothesized that the height of a column of liquid would vary with the density of the liquid. To test this idea, Torricelli designed the equipment shown in **Figure 5.** He filled a thin glass tube that was closed at one end with mercury. While covering the open end so that air could not enter, he inverted the tube and placed it (open end down) in a dish of mercury. The height of the mercury column fell to about one-fourteenth of a similar water column. This validated Torricelli's hypothesis because mercury is approximately fourteen times more dense than water.



Before gas is released into the U-tube, the mercury is at the same height in each arm.

After gas is released into the U-tube, the heights in the two arms are no longer equal.

Figure 6 A manometer measures the pressure of an enclosed gas.

Barometers The device that Torricelli invented is called a barometer. A **barometer** is an instrument used to measure atmospheric pressure. As Torricelli demonstrated, at sea level the height of the mercury in a barometer is usually about 760 mm.

The exact height of the mercury is determined by two forces. Gravity exerts a constant downward force on the mercury. This force is opposed by an upward force exerted by air pressing down on the surface of the mercury. Changes in air temperature or humidity cause air pressure to vary.

Manometers A manometer is an instrument used to measure gas pressure in a closed container. In a manometer, a flask is connected to a U-tube that contains mercury, as shown in **Figure 6.** When the valve between the flask and the U-tube is opened, gas particles diffuse out of the flask into the U-tube. The released gas particles push down on the mercury in the tube. The difference in the height of the mercury in the two arms is used to calculate the pressure of the gas in the flask.

Units of pressure

The SI unit of pressure is the pascal (Pa). It is named for Blaise Pascal (1623–1662), a French mathematician and philosopher. The pascal is derived from the SI unit of force, the newton (N). One **pascal** is equal to a force of one newton per square meter: 1 Pa equals 1 N/m^2.

Many fields of science still use more traditional units of pressure. For example, engineers often report pressure as pounds per square inch (psi). You might see this unit of pressure listed on the tires for bicycles and cars. The pressures measured by barometers and manometers can be reported in millimeters of mercury (mmHg). However, if you listen to your local weather report, you might hear the meteorologist report the pressure in inches, which is understood to mean inches of mercury. It is similar to the unit millimeters of mercury as it describes the height of a mercury column that could be supported by the atmospheric pressure.

Table 1 Comparison of Pressure Units

Unit	Number Equivalent to 1 atm	Number Equivalent to 1 kPa
Kilopascal (kPa)	101.3 kPa	—
Atmosphere (atm)	—	0.009869 atm
Millimeters of mercury (mmHg)	760 mmHg	7.501 mmHg
Torr	760 torr	7.501 torr
Pounds per square inch (psi or lb/in^2)	14.7 psi	0.145 psi
Bar	1.01 bar	0.01 bar

At sea level, the average air pressure is 101.3 kPa when the temperature is 0°C. Air pressure is often reported in a unit called an atmosphere (atm). One **atmosphere** is equal to 760 mmHg or 760 torr or 101.3 kilopascals (kPa). **Table 1** compares different units of pressure. Because the units 1 atm, 760 mmHg, and 760 torr are defined units, they should have as many significant figures as needed when used in calculations.

Dalton's law of partial pressures

When Dalton studied the properties of gases, he found that each gas in a mixture exerts pressure independently of the other gases present. Illustrated in **Figure 7**, **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of the pressures of all the gases in the mixture. The portion of the total pressure contributed by a single gas is called its partial pressure. The partial pressure of a gas depends on the number of moles of gas, the size of the container, and the temperature of the mixture. It does not depend on the identity of the gas. At a given temperature and pressure, the partial pressure of 1 mol of any gas is the same.

Look again at **Figure 7**. What happens when 1 mol of helium and 1 mol of nitrogen are combined in a single closed container? Because neither the volume nor the number of particles changes, the total pressure equals the sum of the two partial pressures.

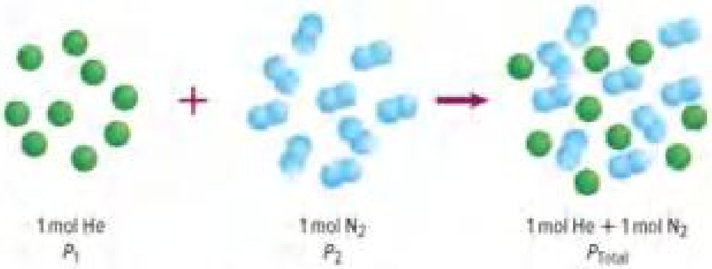


Figure 7 When gases mix, the total pressure of the mixture is equal to the sum of the partial pressures of the individual gases.

Determine *How do the partial pressures of nitrogen gas and helium gas compare when a mole of nitrogen gas and a mole of helium gas are in the same closed container?*

Dalton's Law of Partial Pressures

Dalton's law of partial pressures can be summarized by the following equation.

$$P_{total} = P_1 + P_2 + P_3 + \ldots P_n$$

P_{total} represents total pressure. P_1, P_2, and P_3 represent the partial pressures of each gas up to the final gas, P_n.

To calculate the total pressure of a mixture of gases, add the partial pressures of each of the gases in the mixture.

EXAMPLE Problem 2

THE PARTIAL PRESSURE OF A GAS A mixture of oxygen (O_2), carbon dioxide (CO_2), and nitrogen (N_2) has a total pressure of 0.97 atm. What is the partial pressure of O_2 if the partial pressure of CO_2 is 0.70 atm and the partial pressure of N_2 is 0.12 atm?

1 ANALYZE THE PROBLEM

You are given the total pressure of a mixture and the partial pressure of two gases in the mixture. To find the partial pressure of the third gas, use the equation that relates partial pressures to total pressure.

Known

$P_{N_2} = 0.12$ atm

$P_{CO_2} = 0.70$ atm

$P_{total} = 0.97$ atm

Unknown

$P_{O_2} = ?$ atm

2 SOLVE FOR THE UNKNOWN

$P_{total} = P_{N_2} + P_{CO_2} + P_{O_2}$ State Dalton's law of partial pressures.

$P_{O_2} = P_{total} - P_{CO_2} - P_{N_2}$ Solve for P_{O_2}.

$P_{O_2} = 0.97 \text{ atm} - 0.70 \text{ atm} - 0.12 \text{ atm}$ Substitute $P_{N_2} = 0.12$ atm, $P_{CO_2} = 0.70$ atm, and $P_{total} = 0.97$ atm.

$P_{O_2} = 0.15$ atm

3 EVALUATE THE ANSWER

Adding the calculated value for the partial pressure of oxygen to the known partial pressures gives the total pressure, 0.97 atm. The answer has two significant figures to match the data.

PRACTICE Problems

ADDITIONAL PRACTICE

4. What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mmHg and the partial pressure of helium is 439 mmHg?
5. Find the total pressure for a mixture that contains four gases with partial pressures of 5.00 kPa, 4.56 kPa, 3.02 kPa, and 1.20 kPa.
6. Find the partial pressure of carbon dioxide in a gas mixture with a total pressure of 30.4 kPa if the partial pressures of the other two gases in the mixture are 16.5 kPa and 3.7 kPa.
7. **CHALLENGE** Air is a mixture of gases. By percentage, it is roughly 78 percent nitrogen, 21 percent oxygen, and 1 percent argon. (There are trace amounts of many other gases in air.) If the atmospheric pressure is 760 mmHg, what are the partial pressures of nitrogen, oxygen, and argon in the atmosphere?

Figure 8 In the flask, sulfuric acid (H_2SO_4) reacts with zinc to produce hydrogen gas. The hydrogen is collected at 20°C.
Calculate *the partial pressure of hydrogen at 20°C if the total pressure of the hydrogen and water vapor mixture is 100.0 kPa.*

Using Dalton's law

Partial pressures can be used to determine the amount of gas produced by a reaction. The gas produced is bubbled into an inverted container of water, as shown in **Figure 8.** As the gas collects, it displaces the water. The gas collected in the container will be a mixture of hydrogen and water vapor. Therefore, the total pressure inside the container will be the sum of the partial pressures of hydrogen and water vapor.

The partial pressures of gases at the same temperature are related to their concentration. The partial pressure of water vapor has a fixed value at a given temperature. You can look up the value in a reference table. At 20°C, the partial pressure of water vapor is 2.3 kPa. You can calculate the partial pressure of hydrogen by subtracting the partial pressure of water vapor from the total pressure.

As you will read later, knowing the pressure, volume, and temperature of a gas allows you to calculate the number of moles of the gas. Temperature and volume can be measured during an experiment. Once the temperature is known, the partial pressure of water vapor is used to calculate the pressure of the gas. The known values for volume, temperature, and pressure are then used to find the number of moles.

Check Your Progress

Summary

- The kinetic-molecular theory explains the properties of gases in terms of the size, motion, and energy of their particles.
- Dalton's law of partial pressures is used to determine the pressures of individual gases in gas mixtures.
- Graham's law is used to compare the diffusion rates of two gases.

Demonstrate Understanding

8. **Explain** Use the kinetic theory to explain the behavior of gases.
9. **Describe** how the mass of a gas particle affects its rate of effusion and diffusion.
10. **Explain** how gas pressure is measured.
11. **Explain** why the container of water must be inverted when a gas is collected by displacement of water.
12. **Calculate** Suppose two gases in a container have a total pressure of 1.20 atm. What is the pressure of Gas B if the partial pressure of Gas A is 0.75 atm?
13. **Infer** whether or not temperature has any effect on the diffusion rate of a gas. Explain your answer.

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LESSON 2
FORCES OF ATTRACTION

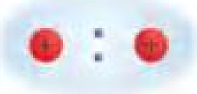
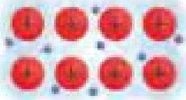
FOCUS QUESTION

What forces exist between molecules?

Intermolecular Forces

If all particles of matter at room temperature have the same average kinetic energy, why are some materials gases while others are liquids or solids? The answer lies with the attractive forces within and between particles. The structure and interactions of matter at the bulk scale are determined by electrical forces within and between atoms. The attractive forces that hold particles together in ionic, covalent, and metallic bonds are called intramolecular forces. The prefix *intra-* means *within*. For example, intramural sports are competitions among teams from within a single school or district. The term *molecular* can refer to atoms, ions, or molecules. **Table 2** summarizes what you read previously about intramolecular forces.

Table 2 Comparison of Intramolecular Forces

Force	Model	Basis of Attraction	Example
Ionic		cations and anions	NaCl
Covalent		positive nuclei and shared electrons	H_2
Metallic		metal cations and mobile electrons	Fe

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Review the News

Obtain information from a current news story about gases. Evaluate your source and communicate your findings to your class.

Revisit the Encounter the Phenomenon Question

What information from this lesson can help you answer the module question?

Intramolecular forces do not account for all attractions between particles. There are forces of attraction called intermolecular forces. The prefix *inter-* means *between* or *among*. For example, an interview is a conversation between two people. These forces can hold together identical particles, such as water molecules in a drop of water, or two different types of particles, such as carbon atoms in graphite and the cellulose particles in paper. The three intermolecular forces that will be discussed in this section are dispersion forces, dipole-dipole forces, and hydrogen bonds. Although some intermolecular forces are stronger than others, all intermolecular forces are weaker than the intramolecular forces involved in bonding.

Dispersion forces

Recall that oxygen molecules are nonpolar because electrons are evenly distributed between the equally electronegative oxygen atoms. Under the right conditions, however, oxygen molecules can be compressed into a liquid. For oxygen to condense, there must be some force of attraction between its molecules.

The force of attraction between oxygen molecules is called a dispersion force. **Dispersion forces** are weak forces that result from temporary shifts in the density of electrons in electron clouds. Dispersion forces are sometimes called London forces after the German-American physicist who first described them, Fritz London.

Get It?
Identify What is another name for dispersion forces?

Remember that the electrons in an electron cloud are in constant motion. When two molecules are in close contact, especially when they collide, the electron cloud of one molecule repels the electron cloud of the other molecule. The electron density around each nucleus is, for a moment, greater in one region of each cloud. Each molecule forms a temporary dipole. When temporary dipoles are close together, a weak dispersion force exists between oppositely charged regions of the dipoles, as shown in **Figure 9**.

Get It?
Explain why dispersion forces form.

Dispersion forces exist between all particles. Dispersion forces are weak for small particles, and these forces have an increasing effect as the number of electrons involved increases. Thus, dispersion forces tend to become stronger as the size of the particles increase.

Figure 9 When two molecules are close together, the electron clouds repel each other, creating temporary dipoles. The δ sign represents an area of partial charge on the molecule.

Explain *what the δ+ and δ− signs on a temporary dipole represent.*

For example, fluorine, chlorine, bromine, and iodine exist as diatomic molecules. Recall that the number of nonvalence electrons increases from fluorine to chlorine to bromine to iodine. Because the larger halogen molecules have more electrons, there can be a greater difference between the positive and negative regions of their temporary dipoles and, thus, stronger dispersion forces. This difference in the forces explains why fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid at room temperature.

Get It?

Infer the physical state of the element astatine at room temperature and explain your reasoning.

Dipole-dipole forces

Polar molecules contain permanent dipoles; that is, some regions of a polar molecule are always partially negative and some regions of the molecule are always partially positive. These attractions between oppositely charged regions of polar molecules are called **dipole-dipole forces.** Neighboring polar molecules orient themselves so that oppositely charged regions align.

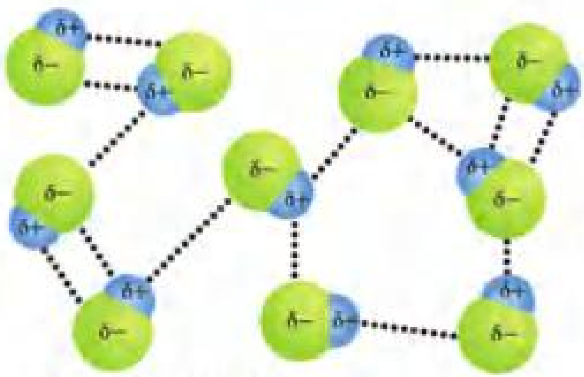


Figure 10 Neighboring polar molecules orient themselves so that oppositely charged regions align.

Identify *the types of forces that are represented in this figure.*

When hydrogen-chloride gas molecules approach, the partially positive hydrogen atom in one molecule is attracted to the partially negative chlorine atom in another molecule. **Figure 10** shows multiple attractions among hydrogen-chloride molecules. Because the dipoles are permanent, you might expect dipole-dipole forces to be stronger than dispersion forces. This prediction holds true for small polar molecules with large dipoles. However, for many polar molecules, including the HCl molecules in **Figure 10,** dispersion forces dominate dipole-dipole forces.

Get It?

Compare dipole-dipole forces and dispersion forces.

ACADEMIC VOCABULARY

orient

to arrange in a specific position; to align in the same direction

The blooms of the flowers were all oriented toward the setting Sun.

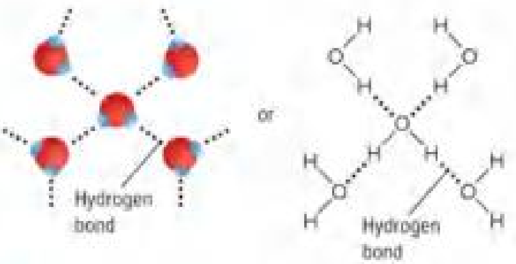


Figure 11 The hydrogen bonds between water molecules are stronger than typical dipole-dipole attractions because the bond between hydrogen and oxygen is highly polar.

Hydrogen bonds

One special type of dipole-dipole attraction is called a hydrogen bond. A **hydrogen bond** is a dipole-dipole attraction that occurs between molecules containing a hydrogen atom bonded to a small, highly electronegative atom with at least one lone electron pair. Hydrogen bonds typically dominate both dispersion forces and dipole-dipole forces. For a hydrogen bond to form, hydrogen must be bonded to either a fluorine, oxygen, or nitrogen atom. These atoms are electronegative enough to cause a large partial positive charge on the hydrogen atom, yet small enough that their lone pairs of electrons can come close to hydrogen atoms. For example, in a water molecule, the hydrogen atoms have a large partial positive charge and the oxygen atom has a large partial negative charge. When water molecules approach, a hydrogen atom on one molecule is attracted to the lone pair of electrons on the oxygen atom on the other molecule, as shown in **Figure 11.**

Get It?

Distinguish between the forces that hold atoms together in a water molecule and the attractive forces that act between water molecules.

Hydrogen bonds explain why water is a liquid at room temperature, while compounds of comparable mass are gases. Look at the data in **Table 3** on the next page. The difference between methane and water is easy to explain. Because methane molecules are nonpolar, the only forces holding the molecules together are relatively weak dispersion forces. The difference between ammonia and water is not as obvious. Molecules of both compounds can form hydrogen bonds. Yet, ammonia is a gas at room temperature, which indicates that the attractive forces between ammonia molecules are not as strong. Because oxygen atoms are more electronegative than nitrogen atoms, the O–H bonds in water are more polar than the N–H bonds in ammonia. As a result, the hydrogen bonds between water molecules are stronger than the hydrogen bonds between ammonia molecules.

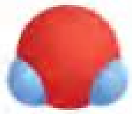
ACADEMIC VOCABULARY

approach

to come near or nearer to someone or something in distance

He made sure to approach the injured dog slowly.

Table 3 Properties of Three Molecular Compounds

Compound	Molecular Structure	Molar Mass (g)	Boiling Point (°C)
Water (H_2O)		18.0	100
Methane (CH_4)		16.0	−161.5
Ammonia (NH_3)		17.0	−33.3

Check Your Progress

Summary

- Intramolecular forces are stronger than intermolecular forces.
- Dispersion forces are intermolecular forces between temporary dipoles.
- Dipole-dipole forces occur between polar molecules.

Demonstrate Understanding

14. **Explain** what determines a substance's state at a given temperature.
15. **Compare and contrast** intermolecular forces and describe intramolecular forces.
16. **Evaluate** Which of the molecules listed below can form hydrogen bonds? For which of the molecules would dispersion forces be the only intermolecular force? Give reasons for your answers.

    a. H_2 **b.** H_2S **c.** HCl **d.** HF
17. **Intepret Data** In a methane molecule (CH_4), there are four single covalent bonds. In an octane molecule (C_8H_{18}), there are 25 single covalent bonds. How does the number of bonds affect the dispersion forces in samples of methane and octane? Which compound is a gas at room temperature? Which is a liquid?

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LESSON 3
LIQUIDS AND SOLIDS

FOCUS QUESTION

What are the properties of liquids and solids?

Liquids

Although the kinetic-molecular theory was developed to explain the behavior of gases, the model also applies to liquids and solids. When applying the kinetic-molecular theory to the solid and liquid states of matter, you must consider the forces of attraction between particles as well as their energy of motion.

Previously, you read that a liquid can take the shape of its container but its volume is fixed. In other words, the particles can flow to adjust to the shape of a container, but the liquid cannot expand to fill its container, as shown in **Figure 12.** According to the kinetic-molecular theory, individual particles do not have fixed positions in the liquid. Forces of attraction between particles in the liquid limit their range of motion so that the particles remain closely packed in a fixed volume.

Figure 12 Liquids flow and take the shape of their container, but they do not expand to fill their container like gases.

Infer *the reason that the liquid is at the same level in each of the interconnected tubes.*

3D THINKING DCI CCC SEP

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Touching the Future

HS-PS2-6 Communicate scientific and technical information about why the molecular-level structure is important in the functioning of designed materials.

Quick Investigation: Model Crystal Unit Cells

Develop and use models to illustrate the structure of crystals.

Density and compression

At 25°C and 1 atm of air pressure, liquids are much denser than gases. The density of a liquid is much greater than that of its vapor at the same conditions. For example, liquid water is about 1250 times denser than water vapor at 25°C and 1 atm of pressure. Because they are at the same temperature, both gas and liquid particles have the same average kinetic energy. Thus, the higher density of liquids is due to the intermolecular forces that hold particles together. Unlike gases, liquids are considered incompressible in many applications. The change in volume for liquids is much smaller because liquid particles are already tightly packed. An enormous amount of pressure must be applied to reduce the volume of a liquid by a very small amount.

Fluidity

Gases and liquids are classified as fluids because they can flow and diffuse. **Figure 13** shows one liquid diffusing through another liquid. Liquids usually diffuse more slowly than gases at the same temperature, because intermolecular attractions interfere with the flow. Thus, liquids are less fluid than gases. A comparison between water and natural gas can illustrate this difference. When there is a leak in a basement water pipe, the water remains in the basement unless the amount of water released exceeds the volume of the basement.

A gas will not stay in the basement. For example, natural gas, or methane, is a fuel burned in gas furnaces, hot-water heaters, and stoves. Gas that leaks from a gas pipe diffuses throughout the house. Because natural gas is odorless, companies that supply the fuel include a compound with a distinct odor. Adding odor to natural gas warns the homeowner of the leak. The customer has time to shut off the gas supply, open windows to allow the gas to diffuse, and call the gas company to report the leak.

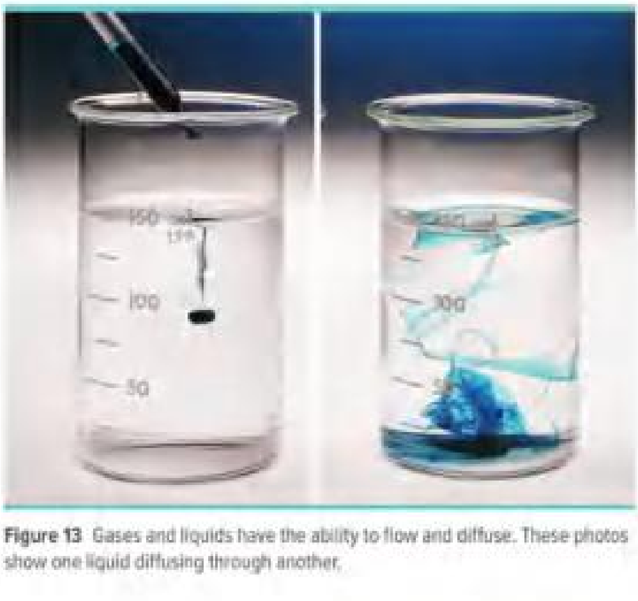


Figure 13 Gases and liquids have the ability to flow and diffuse. These photos show one liquid diffusing through another.

Viscosity

You are already familiar with viscosity if you have ever tried to get honey out of a bottle. **Viscosity** is a measure of the resistance of a liquid to flow.

The particles in a liquid are close enough for attractive forces to slow their movement as they flow past one another. The viscosity of a liquid is determined by the type of intermolecular forces in the liquid, the size and shape of the particles, and the temperature.

You should note that not all liquids have viscosity. Scientists discovered superfluids in 1937. Scientists cooled liquid helium below −270.998°C and discovered that the properties of the liquid changed. The superfluid helium lost viscosity—the resistance to flow.

Attractive forces In typical liquids, the stronger the intermolecular attractive forces, the higher the viscosity. If you have used glycerol in the laboratory to help insert a glass tube into a rubber stopper, you know that glycerol is a viscous liquid.

Figure 14 uses structural formulas to show the hydrogen bonding that makes glycerol so viscous. The hydrogen atoms attached to the oxygen atoms in each glycerol molecule are able to form hydrogen bonds with other glycerol molecules. The red dots in **Figure 14** show where the hydrogen bonds form between molecules.

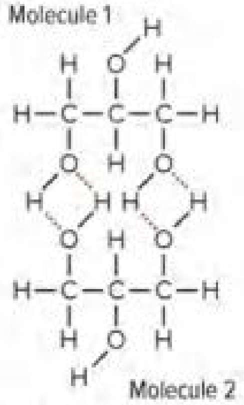


Figure 14 This diagram shows two glycerol molecules and the hydrogen bonds between them.

Determine *the possible number of hydrogen bonds a glycerol molecule can form with a second molecule.*

Particle size and shape The size and shape of particles also affect viscosity. Recall that the overall kinetic energy of a particle is determined by its mass and velocity.

Suppose the attractive forces between molecules in Liquid A and Liquid B are similar. If the molecules in Liquid A are more massive than the molecules in Liquid B, Liquid A will have a greater viscosity. Liquid A's molecules will, on average, move more slowly than the molecules in Liquid B.

Molecules with long chains, such as cooking oils and motor oil, have a higher viscosity than shorter, more-compact molecules, assuming the molecules exert the same type of attractive forces. Within the long chains, there is less distance between atoms on neighboring molecules and, thus, a greater chance for attractions between atoms.

Temperature Viscosity decreases with temperature. When you pour a small amount of cooking oil into a frying pan, the oil tends not to spread across the bottom of the pan until you heat it.

With the increase in temperature, there is an increase in the average kinetic energy of the oil molecules. The added energy makes it easier for the molecules to overcome the intermolecular forces that keep the molecules from flowing.

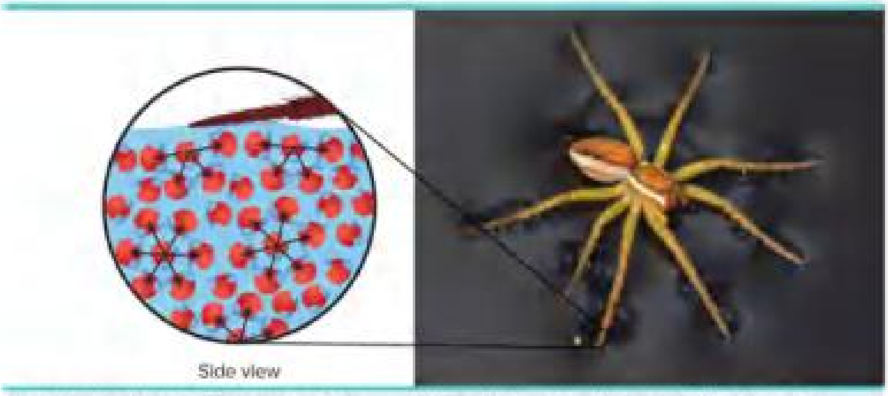
Another example of the effects of temperature on viscosity is motor oil. Motor oil keeps the moving parts of an internal combustion engine lubricated. People once used different motor-oil blends in winter and summer. The motor oil used in winter was designed to flow at low temperatures. The motor oil used in summer was more viscous so that it could maintain sufficient viscosity on extremely hot days. Today, additives in motor oil help adjust the viscosity so that the same oil blend can be used all year. Molecules in the additives are compact spheres with relatively low viscosity at cool temperatures. At high temperatures, the shape of the additive molecules changes to long strands. These strands get tangled with the oil molecules, which increases the viscosity of the oil.

Get It?

Infer why it is important for motor oil to remain viscous.

Surface tension

Intermolecular forces do not have an equal effect on all particles in a liquid, as shown in **Figure 15**. Particles in the middle of the liquid can be attracted to particles above them, below them, and to either side. For particles at the surface of the liquid, there are no attractions from above to balance those from below. Thus, there is a net attractive force pulling down on particles at the surface. The surface tends to have the smallest possible area and to act as though it is stretched tight like the head of a drum. For the surface area to increase, particles from the interior must move to the surface. It takes energy to overcome the attractions holding these particles in the interior. The energy required to increase the surface area of a liquid by a given amount is called **surface tension.** Surface tension is a measure of the inward pull by particles in the interior. In general, the stronger the attractions between particles, the greater the surface tension. Water has a high surface tension because its molecules can form multiple hydrogen bonds. Drops of water are shaped like spheres because the surface area of a sphere is smaller than the surface area of any other shape of similar volume. Water's high surface tension allows the spider in **Figure 15** to walk on the surface of the pond.



Intermolecular forces just below the surface of the water create surface tension.

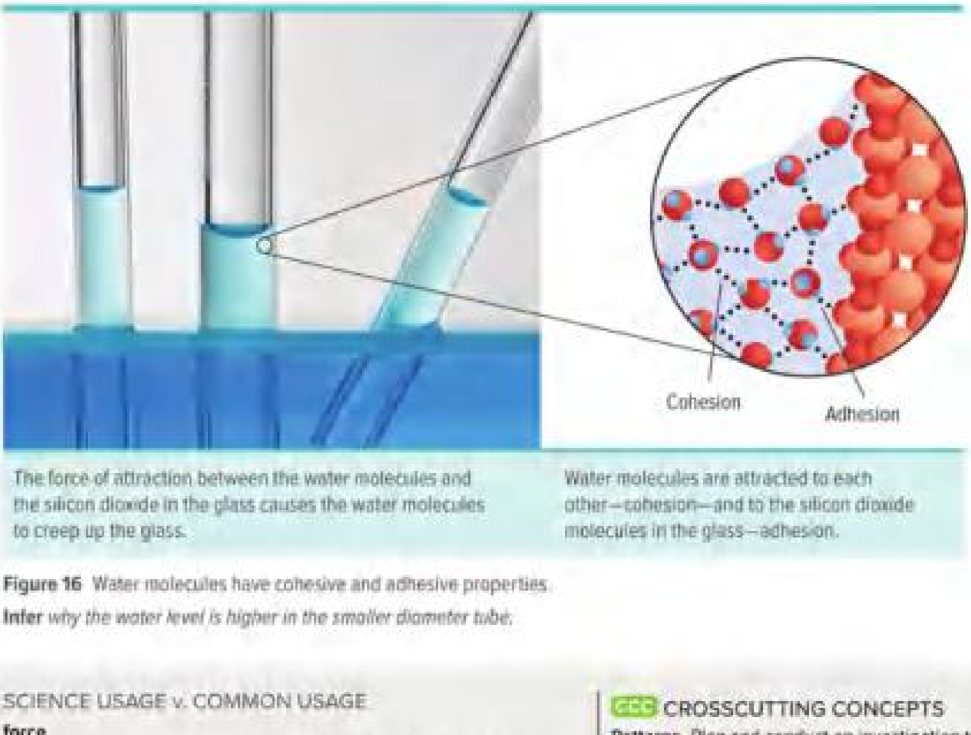
The surface tension of the water allows this spider to walk on the surface of the water.

Figure 15 At the surface of water, the particles are drawn toward the interior until attractive and repulsive forces are balanced.

The same forces that allow the spider to stay dry on the surface of a pond also makes it difficult to use water alone to remove dirt from skin and clothing. Dirt particles cannot penetrate the surface of waterdrops. Soaps and detergents decrease the surface tension of water by disrupting the hydrogen bonds between water molecules. When the bonds are broken, the water spreads out allowing the dirt to be carried away by the water. Compounds that lower the surface tension of water are called **surfactants.**

Cohesion and adhesion

When water is placed into a narrow container, such as the glass tubes in **Figure 16,** you can see that the surface of the water is not straight. The surface forms a concave meniscus; that is, the surface dips in the center. **Figure 16** models what is happening to the water at the molecular level. There are two types of forces at work: cohesion and adhesion. Cohesion describes the force of attraction between identical molecules. Adhesion describes the force of attraction between molecules that are different. Because the adhesive forces between water molecules and the silicon dioxide in glass are greater than the cohesive forces between water molecules, the water rises along the inner walls of the cylinder.



The force of attraction between the water molecules and the silicon dioxide in the glass causes the water molecules to creep up the glass.

Water molecules are attracted to each other—cohesion—and to the silicon dioxide molecules in the glass—adhesion.

Figure 16 Water molecules have cohesive and adhesive properties.
Infer *why the water level is higher in the smaller diameter tube.*

SCIENCE USAGE v. COMMON USAGE

force

Science usage: a push or a pull, having both magnitude and direction, that is exerted on an object

The gravitational force exists between any two objects with mass and is directly proportional to their masses.

Common usage: a group of people who have the power to work toward a desired outcome

The U.S. labor force increased its productivity last year.

CCC CROSSCUTTING CONCEPTS

Patterns Plan and conduct an investigation to compare the viscosity of different liquids. Decide on the evidence you will collect and the variables you will control. Use your results to explain how the kinetic-molecular theory applies to liquids.

Capillary action If the cylinder is extremely narrow, a thin film of water will be drawn upward. Narrow tubes are called capillary tubes. This movement of a liquid such as water is called capillary action, or capillarity. Capillary action helps explain how paper towels can absorb large amounts of water. The water is drawn into the narrow spaces between the cellulose fibers in paper towels by capillary action. In addition, the water molecules form hydrogen bonds with cellulose molecules.

Solids

Did you ever wonder why solids have a definite shape and volume? According to the kinetic-molecular theory, a mole of solid particles has as much kinetic energy as a mole of liquid or gas particles at the same temperature. By definition, the particles in a solid must be in constant motion. For a substance to be a solid rather than a liquid at a given temperature, there must be strong attractive forces acting between particles in the solid. These forces limit the motion of the particles to vibrations around fixed locations in the solid. Thus, there is more order in a solid than in a liquid. Because of this order, solids are not fluid. Only gases and liquids are classified as fluids.

Density of solids

In general, the particles in a solid are more closely packed than those in a liquid. Thus, most solids are more dense than most liquids. When the liquid and solid states of a substance coexist, the solid almost always sinks in the liquid. Solid cubes of benzene sink in liquid benzene because solid benzene is more dense than liquid benzene. There is about a 10% difference in density between the solid and liquid states of most substances. Because the particles in a solid are closely packed, ordinary amounts of pressure will not change the volume of a solid.

You cannot predict the relative densities of ice and liquid water based on benzene. Ice cubes and icebergs float because water is less dense as a solid than it is as a liquid. **Figure 17** shows the reason for the exception. As water freezes, each H_2O molecule can form hydrogen bonds with up to four neighboring molecules. As a result, the water molecules in ice are less-closely packed together than in liquid water.

Get It?
Describe in your own words why ice floats in water.

Figure 17 An iceberg can float because the rigid, three-dimensional structure of ice keeps water molecules farther apart than they are in liquid water. This open, symmetrical structure of ice results from hydrogen bonding.

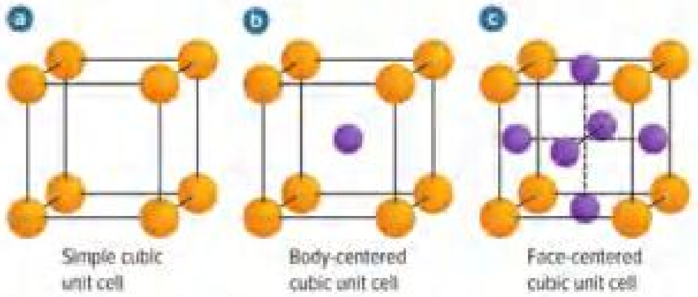


Figure 18 These drawings show three of the ways particles are arranged in crystal lattices. Each sphere represents a particle. **a.** Particles are arranged only at the corners of the cube. **b.** There is a particle in the center of the cube. **c.** There are particles in the center of each of the six cubic faces but no particle in the center of the cube itself.

Crystalline solids

Although ice is unusual in its density, ice is typical of most solids in that its molecules are packed together in a predictable way. A **crystalline solid** is a solid whose atoms, ions, or molecules are arranged in an orderly, geometric structure. The locations of particles in a crystalline solid can be represented as points on a framework called a crystal lattice. **Figure 18** shows three ways that particles in a crystal lattice can be arranged to form a cube.

A **unit cell** is the smallest arrangement of atoms in a crystal lattice that has the same symmetry as the whole crystal. Like the formula unit that you read about previously, a unit cell is a small, representative part of a larger whole. The unit cell can be thought of as a building block whose shape determines the shape of the crystal.

Get It?

Infer Imagine having a unit cell of each type of crystal lattice composed of identical atoms. How would their densities compare? Explain your reasoning.

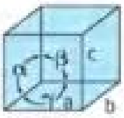
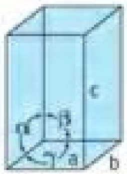
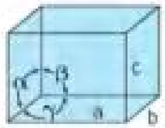
Table 4 on the next page shows seven categories of crystals based on shape. Crystal shapes differ because the surfaces, or faces, of unit cells do not always meet at right angles, and the edges of the faces vary in length. In **Table 4,** the edges are labeled *a*, *b*, and *c*; the angles at which the faces meet are labeled α, β, and γ.

STEM CAREER Connection

Geological Technician

Do you enjoy exploring remote locations? Geological technicians help scientists and engineers identify, explore, extract, and analyze natural resources, such as minerals, metals, and precious gemstones. Many technicians work for the mining industry. These technicians work on-site for days or weeks at a time, often in remote locations and under extreme weather conditions. Other technicians work primarily in laboratories.

Table 4 Unit Cells

 **Halite (rock salt)**	 **Vesuvianite**	 **Aragonite**
 a = b = c α = β = γ = 90° **Cubic**	 a = b ≠ c α = β = γ = 90° **Tetragonal**	 a ≠ b ≠ c α = β = γ = 90° **Orthorhombic**

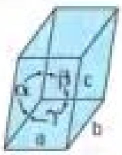
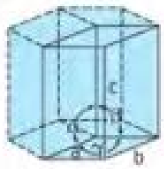
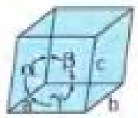
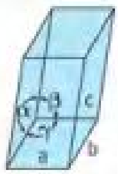
 **Microcline**	 **Beryl (aquamarine)**	 **Tourmaline**	 **Crocoite**
 a ≠ b ≠ c α ≠ β ≠ γ ≠ 90° **Triclinic**	 a = b ≠ c α = β = 90°, γ = 120° **Hexagonal**	 a = b = c α = β = γ ≠ 90° **Rhombohedral**	 a ≠ b ≠ c α = γ = 90° ≠ β **Monoclinic**

Table 5 Types of Crystalline Solids

Type	Unit Particles	Characteristics of Solid Phase	Examples
Atomic	atoms	soft to very soft; very low melting points; poor conductivity	group 18 elements
Molecular	molecules	fairly soft; low to moderately high melting points; poor conductivity	I_2, H_2O, NH_3, CO_2, $C_{12}H_{22}O_{11}$ (table sugar)
Covalent network	atoms connected by covalent bonds	very hard; very high melting points; often poor conductivity	diamond (C) and quartz (SiO_2)
Ionic	ions	hard; brittle; high melting points; poor conductivity	NaCl, KBr, $CaCO_3$
Metallic	atoms surrounded by mobile valence electrons	soft to hard; low to very high melting points; malleable and ductile; excellent conductivity	all metallic elements

Categories of crystalline solids

Crystalline solids can be classified into five categories based on the types of particles that they contain and how thoses particles are bonded together: atomic solids, molecular solids, covalent network solids, ionic solids, and metallic solids.

Table 5 summarizes the general characteristics of each category and provides examples.

Get It?

List the five categories of crystalline solids.

Atomic solids The only atomic solids are noble gases. Their properties reflect the weak dispersion forces between the atoms.

Molecular solids In molecular solids, the molecules are held together by dispersion forces, dipole-dipole forces, or hydrogen bonds. Most molecular compounds are not solids at room temperature. Even water, which can form strong hydrogen bonds, is a liquid at room temperature.

Molecular compounds such as sugar are solids at room temperature because of their large molar masses. With larger molecules, many weak attractions can combine to hold the molecules together. Because they contain no ions, molecular solids are poor conductors of heat and electricity.

Get It?

Describe the state and conductivity of molecular solids.

Covalent network solids Atoms such as carbon and silicon, which can form multiple covalent bonds, are able to form covalent network solids. The covalent network structure of quartz, which contains silicon, is shown in **Figure 19.** Carbon forms three types of covalent network solids—diamond, graphite, and buckminsterfullerene. An element, such as carbon, that exists in different forms at the same state—solid, liquid, or gas—is called an **allotrope.** For more information about carbon allotropes see the Elements Handbook.

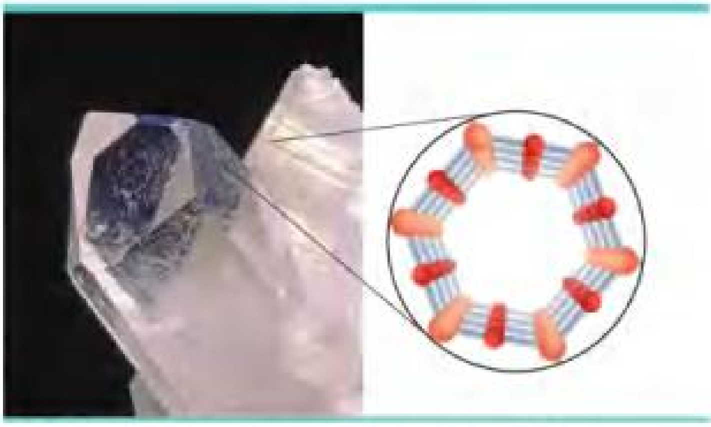


Figure 19 The most common kind of quartz has a hexagonal crystal structure.

Ionic solids Remember that each ion in an ionic solid is surrounded by ions of opposite charge. The type of ions and the ratio of ions determine the structure of the lattice and the shape of the crystal. The network of attractions that extends throughout an ionic crystal gives these compounds their high melting points and hardness. Ionic crystals are strong, but brittle. When ionic crystals are struck, the cations and anions are shifted from their fixed positions. Repulsions between ions of like charge cause the crystal to shatter.

Metallic solids Recall that metallic solids consist of positive metal ions surrounded by a sea of mobile electrons. The strength of the metallic bonds between cations and electrons varies among metals and accounts for their wide range of physical properties. For example, tin melts at 232°C, but nickel melts at 1455°C. The mobile electrons make metals malleable—easily hammered into shapes—and ductile—easily drawn into wires. When force is applied to a metal, the electrons shift and thereby keep the metal ions bonded in their new positions. Mobile electrons make metals good conductors of heat and electricity. As shown in **Figure 20,** metal wiring is used to carry electricity to businesses and homes.

Figure 20 Homes, business, and equipment of all types use metal wiring to carry electricity. The metal is usually copper, but other metals are used in special applications.

Get It?

Describe the properties of metals that make them useful for making jewelry.

Amorphous solids

An **amorphous solid** is one in which the particles are not arranged in a regular, repeating pattern. It does not contain crystals. The term *amorphous* is derived from a Greek word that means *without shape*. An amorphous solid often forms when a molten material cools too quickly to allow enough time for crystals to form.

Figure 21 shows an example of an amorphous solid. Glass, rubber, and many plastics are amorphous solids. Recent studies have shown that glass might have some structure. When X-ray diffraction is used to study glass, there appears to be no pattern to the distribution of atoms. When neutrons are used instead, an orderly pattern of silicate units can be detected in some regions. Researchers hope to use this new information to control the structure of glass for optical applications and to produce glass that can conduct electricity.

Figure 21 Native Americans used the glass-like amorphous rock obsidian to make arrowheads and knives, because it can form sharp edges when broken. Obsidian rock forms when lava cools too quickly to form crystals.

Check Your Progress

Summary

- The kinetic-molecular theory explains the behavior of solids and liquids.
- Intermolecular forces in liquids affect viscosity, surface tension, cohesion, and adhesion.
- Crystalline solids can be classified by their shape and composition.

Demonstrate Understanding

18. **Differentiate** between solids and liquids in terms of the arrangement and motion of particles.
19. **Describe** the factors that affect viscosity.
20. **Explain** why soap and water are used to clean clothing instead of water alone.
21. **Compare** a unit cell and a crystal lattice.
22. **Describe** the difference between a molecular solid and a covalent network solid.
23. **Explain** why water forms a meniscus when it is in a graduated cylinder.
24. **Infer** why the surface of mercury in a thermometer is convex; that is, the surface is higher at the center.
25. **Predict** which solid is more likely to be amorphous—one formed by allowing a molten material to cool slowly to room temperature or one formed by quickly cooling the same material in an ice bath.
26. **Design** an experiment to compare the relative abilities of water and isopropyl alcohol to support skipping stones. Include a prediction about which liquid will be better, along with a brief explanation of your prediction.

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LESSON 4
PHASE CHANGES

FOCUS QUESTION

What causes a substance to change phases?

Phase Changes That Require Energy

Most substances can exist in three states depending on the temperature and pressure. A few substances, such as water, exist in all three states under ordinary conditions. States of a substance are referred to as phases when they coexist as physically distinct parts of a mixture. Ice water is a heterogeneous mixture with two phases, solid ice and liquid water. When energy is added or removed from a system, one phase can change into another, as shown in **Figure 22.** Because you are familiar with the phases of water—ice, liquid water, and water vapor—and have observed changes between those phases, we can use water as the primary example in the discussion of phase changes. Let's start by thinking about ice cubes in a glass of water.

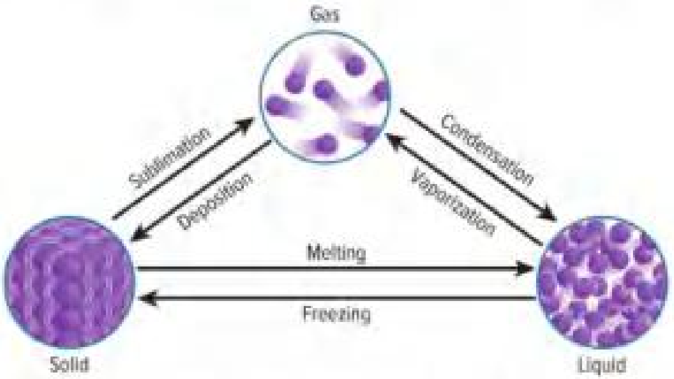


Figure 22 The diagram shows the six possible transitions between phases.
Determine *what phase changes occur between solids and liquids.*

3D THINKING DCI CCC SEP

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Investigate Interparticle Forces
HS-PS1-3 Plan and conduct an investigation to gather evidence to compare the structure of substances at the bulk scale to infer the strength of electrical forces between particles.

Revisit the Encounter the Phenomenon Question
What information from this lesson can help you answer the module question?

Melting

Ice cubes placed in warm water are at a lower temperature than the water. Heat flows from the water to the ice. Heat is the transfer of energy from an object at a higher temperature to an object at a lower temperature. At ice's melting point, the energy absorbed by the ice does not raise the ice's temperature. Instead, it disrupts the hydrogen bonds holding the water molecules together in the ice crystal. When molecules on the surface of the ice absorb enough energy to break the hydrogen bonds, they move apart and enter the liquid phase. Thus, melting is an endothermic process. As molecules are removed, the ice cube shrinks. The process continues until all of the ice melts.

The amount of energy required to melt 1 mol of a solid depends on the strength of the forces keeping the particles together in the solid. Because hydrogen bonds between water molecules are strong, a relatively large amount of energy is required. However, the energy required to melt ice is much less than the energy required to melt table salt because the ionic bonds in sodium chloride are much stronger than the hydrogen bonds in ice.

The temperature at which the liquid phase and the solid phase of a given substance can coexist is a characteristic physical property of many solids. The **melting point** of a crystalline solid is the temperature at which the forces holding its crystal lattice together are broken and it becomes a liquid. It is difficult to specify an exact melting point for an amorphous solid because they tend to melt over a temperature range.

Vaporization

Vaporization is also an endothermic process. During melting, the temperature of the ice and the water formed remains constant. Once all of the ice has melted, additional energy added to the system increases the kinetic energy of the liquid molecules. The temperature of the system rises. In liquid water, some molecules will have more kinetic energy than others. **Figure 23** shows the energy distribution among liquid molecules at 25°C. The shaded portion indicates those molecules that have the energy required to overcome the forces of attraction holding the molecules together in the liquid.

Get It?

Describe what happens to the particles in the shaded portion on the graph.

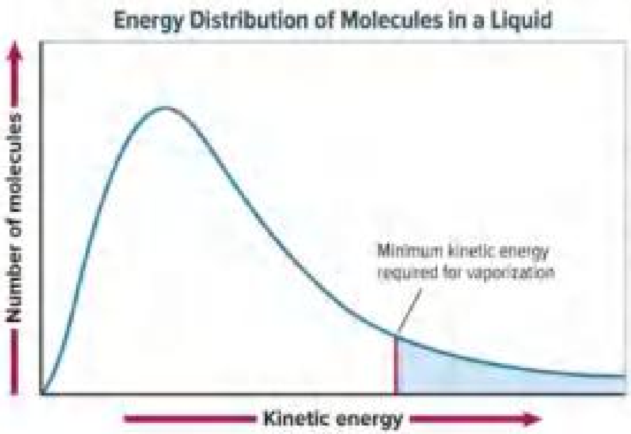


Figure 23 This graph shows a typical distribution of kinetic energy of molecules in a liquid at 25°C. The most probable amount of kinetic energy for a molecule lies at the peak of the curve.

Describe *how the curve would look for the same liquid at 30°C.*

Particles that escape from the liquid enter the gas phase. For a substance that is ordinarily a liquid at room temperature, the gas phase is called a vapor. **Vaporization** is the process by which a liquid changes to a gas or vapor. If the input of energy is gradual, the molecules tend to escape from the surface of the liquid. Remember that molecules at the surface are attracted to fewer other molecules than are molecules in the interior. When vaporization occurs only at the surface of a liquid, the process is called **evaporation.** Even at cold temperatures, some water molecules have enough energy to evaporate. As the temperature rises, more and more molecules enter the gas phase.

Get It?

Explain when the term *vapor* should be used to describe the gas phase.

Figure 24 compares evaporation in an open container with evaporation in a closed container. If water is in an open container, all the molecules will eventually evaporate. The time it takes for them to evaporate depends on the amount of water and the available energy. In a partially filled, closed container, the situation is different. Water vapor collects above the liquid and exerts pressure on the surface of the liquid. The pressure exerted by a vapor over a liquid is called **vapor pressure.**

Real-World Chemistry

Evaporation

PERSPIRATION: Evaporation, an endothermic process, is one way your body controls its temperature. When you become hot, your body releases sweat from glands in your skin. Water molecules in sweat absorb heat energy from your skin and evaporate. Blood carries excess heat from all parts of your body to your skin.

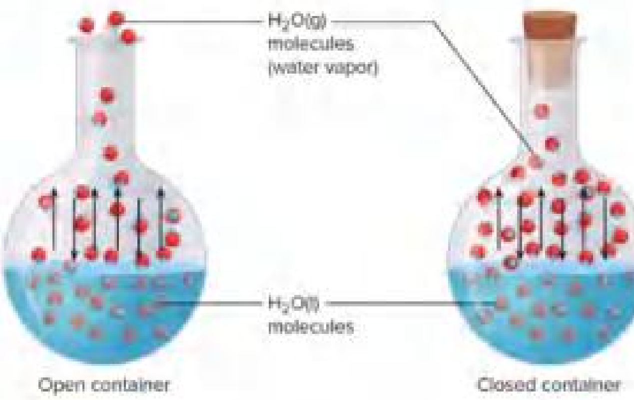


Figure 24 Evaporation occurs in both open and closed containers. In an open container, water molecules that evaporate can escape from the container. Water vapor collects above the liquid in a closed container.

ACADEMIC VOCABULARY

eventually

in the end; at some later, but unspecified, time

The company started with only one product, but eventually, it offered more than 10,000 products.

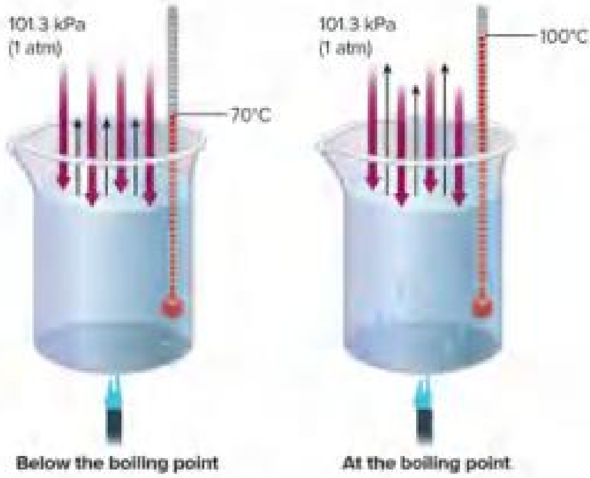


Figure 25 As temperature increases, water molecules gain kinetic energy. Vapor pressure increases (black arrows) but is less than atmospheric pressure (red arrows). A liquid has reached its boiling point when its vapor pressure is equal to atmospheric pressure. At sea level, the boiling point of water is 100°C.

Boiling The temperature at which the vapor pressure of a liquid equals the external or atmospheric pressure is called the **boiling point.** Use **Figure 25** to compare what happens to a liquid at temperatures below its boiling point with what happens to a liquid at its boiling point. At the boiling point, molecules throughout the liquid have enough energy to vaporize. Bubbles of vapor collect below the surface of the liquid and rise to the surface. Note that if large amounts of energy are added to the resulting gas, the gas may change to plasma as electrons are removed from some of the gaseous atoms.

Figure 26 Steaks, seafood, and other highly perishable foods often are shipped in a container with dry ice to keep the food cold.

Explain *why dry ice is preferred over regular ice for shipping steaks and other food products.*

Sublimation

Many substances have the ability to change directly from the solid phase to the gas phase when they absorb energy. Sublimation is an endothermic process. Recall that sublimation is the process by which a solid changes directly to a gas without first becoming a liquid. Solid iodine and solid carbon dioxide (dry ice) sublime at room temperature. Dry ice, shown in **Figure 26**, keeps objects that could be damaged by melting water cold during shipping. Mothballs, which contain the compounds naphthalene or *p*-dichlorobenzene, also sublime, as do solid air fresheners.

Phase Changes That Release Energy

Have you ever awakened on a chilly morning to see frost on your windows or the grass covered with water droplets? When you set a glass of ice water on a picnic table, do you notice beads of water on the outside of the glass? These events are examples of phase changes that release energy into the surroundings.

Freezing

Suppose you place liquid water in an ice tray into a freezer. As heat is removed from the water, the molecules lose kinetic energy and their velocity decreases. The molecules are less likely to flow past one another. When enough energy has been removed, the hydrogen bonds between water molecules keep the molecules fixed, or frozen, into set positions. The **freezing point** is the temperature at which a liquid is converted into a crystalline solid. Freezing is an exothermic process that is the reverse of melting.

Condensation

When a water vapor molecule loses energy, its velocity decreases. The water vapor molecule is more likely to form a hydrogen bond with another water molecule. The formation of a hydrogen bond releases thermal energy and indicates a change from the vapor phase to the liquid phase. The process by which a gas or a vapor becomes a liquid is called **condensation.** Condensation is an exothermic process that is the reverse of vaporization.

Different factors contribute to condensation. However, condensation always involves the transfer of thermal energy. For example, water vapor molecules can come in contact with a cold surface, such as the side of a glass of ice water. Thermal energy transfers from the water vapor molecules to the cool glass, causing condensation on the outside of the glass. A similar process can occur during the night when water vapor in the air condenses and dew forms on blades of grass.

EARTH SCIENCE Connection Precipitation, clouds, and fog all result from condensation. They form as air cools when it rises or passes over cooler land or water. Their formations require a second factor, microscopic particles suspended in the air called condensation nuclei. These can be particles, such as soot and dust, or aerosols, such as sulfur dioxide and nitrogen oxide, on which water vapor condenses. In some circumstances, warm air can settle on top of cooler air, which is called a temperature inversion. **Figure 27** shows fog trapped in a mountain valley by such an inversion.

Figure 27 Normally, air becomes cooler as elevation increases. A temperature inversion occurs when the situation is reversed and the air becomes warmer at higher elevations. Inversions can trap smog over cities and fog in mountain valleys.

Get It?

Describe the condensation of water vapor in the atmosphere.

Deposition

When water vapor comes in contact with a cold window in winter, it forms a solid deposit on the window called frost. **Deposition** is the process by which a substance changes from a gas or vapor to a solid without first becoming a liquid.

Deposition is an exothermic process that is the reverse of sublimation. Snowflakes form when water vapor high up in the atmosphere changes directly into solid ice crystals. Energy is released as the crystals form.

Phase Diagrams

Two variables combine to control the phase of a substance: temperature and pressure. These variables can have opposite effects on a substance. For example, a temperature increase causes more liquid to vaporize, but an increase in pressure causes more vapor to condense. A **phase diagram** is a graph of pressure versus temperature that shows in which phase a substance exists under different conditions of temperature and pressure.

Figure 28 shows the phase diagram for water. You can use this graph to predict what phase water will be in for any combination of temperature and pressure. Note that there are three regions representing the solid, liquid, and vapor phases of water and three curves that separate the regions from one another. At points that fall along the curves, two phases of water can coexist. The short, yellow curve shows the temperature and pressure conditions under which solid water and water vapor can coexist. The long, blue curve shows the temperature and pressure conditions under which liquid water and water vapor can coexist. The red curve shows the temperature and pressure conditions under which solid water and liquid water can coexist.

Point A on the phase diagram of water—the point where the yellow, blue, and red curves meet—is the triple point for water. The **triple point** is the point on a phase diagram that represents the temperature and pressure at which three phases of a substance can coexist. All six phase changes can occur at the triple point: freezing and melting; evaporation and condensation; sublimation and deposition. Point B is called the critical point. This point indicates the critical pressure and critical temperature above which water cannot exist as a liquid. If water vapor is at the critical temperature, an increase in pressure will not change the vapor into a liquid.

Phase diagrams can provide important information for substances. For example, the phase diagram for carbon dioxide in **Figure 28** shows why carbon dioxide sublimes at normal conditions. Find 1.0 atm on the carbon dioxide graph and follow the dashed line to the yellow line. The graph shows that carbon dioxide changes from a solid to a gas at 1 atm. If you extend the dashed line past the yellow line, the graph shows that carbon dioxide does not liquefy as temperature increases. It remains a gas.

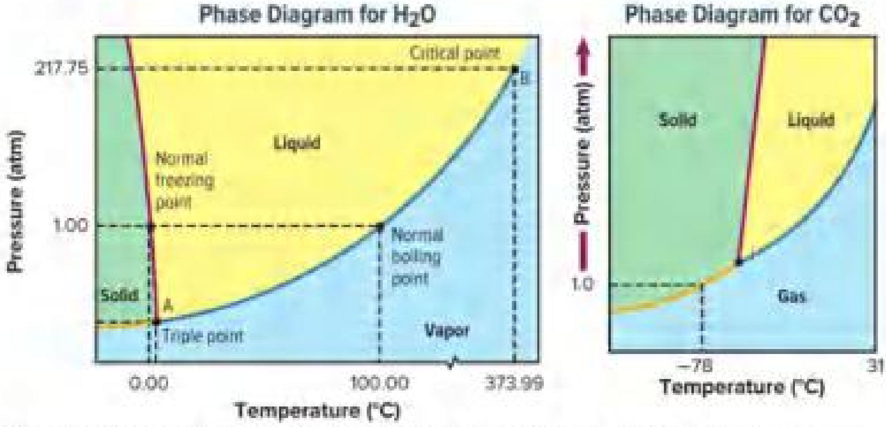


Figure 28 A phase diagram shows the phase of a substance at different temperatures and pressures.
Determine *the phase of water at 2.00 atm and 100.00°C.*

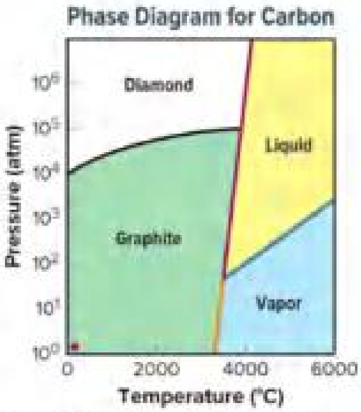


Figure 29 The phase diagram for carbon shows that two forms of solid carbon exist.

The diagram in **Figure 29** is a phase diagram for carbon. Notice that the graph contains two allotropes of carbon in the solid region. Graphite is the standard state of carbon at normal temperatures and pressures, designated by a red dot. Diamond is more stable at higher temperatures and pressures. Diamonds that exist at normal room conditions originally formed at high temperature and pressure.

Get It?

Contrast the slope of the red line in water's phase diagram with that of the red line in carbon dioxide's phase diagram. How do water and carbon dioxide differ in their reaction to increased pressure at the solid/liquid boundary?

The phase diagram for each substance is different because the normal boiling and freezing points of substances are different. However, each diagram will supply the same type of data for the phases, including a triple point. Of course, the range of temperatures chosen will vary to reflect the physical properties of the substance.

Check Your Progress

Summary

- States of a substance are referred to as phases when they coexist as physically distinct parts of a mixture.
- Energy changes occur during phase changes.
- Phase diagrams show how different temperatures and pressures affect the phase of a substance.

Demonstrate Understanding

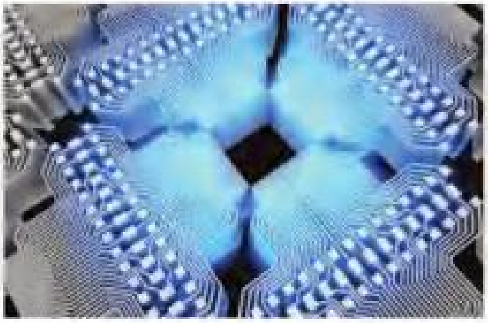
27. **Distinguish** between how endothermic and exothermic processes can result in phase changes.
28. **Explain** the difference between the processes of melting and freezing.
29. **Differentiate** the phase changes between a solid and a gas.
30. **Differentiate** solids, liquids, and gases based on the phase changes of sublimation and evaporation.
31. **Describe** the information that a phase diagram supplies.
32. **Explain** what the triple point and the critical point on a phase diagram represent.
33. **Determine** the phase of water at 75.00°C and 3.00 atm using **Figure 28.**

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

SCIENTIFIC BREAKTHROUGHS

New Matter

Most people are familiar with the three most common states of matter on Earth: solid, liquid, and gas. A fourth state—plasma—is also relatively well-known and occurs naturally on Earth. But the list of known states of matter continues to grow well beyond these most familiar states.



Scientists create new forms of matter using extreme conditions and new technology.

Solid, liquid, gas, and plasma are the states of matter that we can observe under natural conditions on Earth. Advances in technology and knowledge have allowed scientists to propose and observe many other states of matter, such as superfluids, supersolids, supercritical fluids, dropletons, and Bose-Einstein condensates. What do these states have in common? They only exist under extreme conditions. For example, Bose-Einstein condensates have been observed in laboratories under extremely cold temperatures at the point when molecular motion stops. Supercritical liquids can be observed under extremely high pressures.

Light Molecules

Scientists at MIT and Harvard University recently observed a new form of matter, called *light molecules*, or *photonic matter*. The matter is a pair of photons that interact or "stick" together. Photons usually pass through one another without interacting, but the researchers created conditions that allowed the photons to collide and interact.

They used a laser to fire one photon into a metal box, where it moved into a cloud of rubidium atoms. The rubidium gas had been cooled to an extremely low temperature so that the atoms were motionless. A second photon was then fired into the box. It, too, entered the cloud of atoms, and the photons stuck together and moved together like a molecule.

The discovery will help researchers develop quantum computers. Quantum computers use photons instead of electrons to encode and process information. Quantum computers have the potential to be much faster and more efficient than today's computers. However, much more research is required.

Scientists are continuing to study photonic matter. For now, photonic matter can be added to the growing list of new types of matter.

DEVELOP AND USE MODELS TO ILLUSTRATE

Find out more about a lesser known state of matter. Develop a model that you can use to describe the new type of matter to a classmate.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 GASES

- The kinetic-molecular theory explains the properties of gases in terms of the size, motion, and energy of their particles.
- Dalton's law of partial pressures is used to determine the pressures of individual gases in gas mixtures.
- Graham's law is used to compare the diffusion rates of two gases.

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

- kinetic-molecular theory
- elastic collision
- temperature
- diffusion
- Graham's law of effusion
- pressure
- barometer
- pascal
- atmosphere
- Dalton's law of partial pressures

Lesson 2 FORCES OF ATTRACTION

- Intramolecular forces are stronger than intermolecular forces.
- Dispersion forces are intermolecular forces between temporary dipoles.
- Dipole-dipole forces occur between polar molecules.

- dispersion force
- dipole-dipole force
- hydrogen bond

Lesson 3 LIQUIDS AND SOLIDS

- The kinetic-molecular theory explains the behavior of solids and liquids.
- Intermolecular forces in liquids affect viscosity, surface tension, cohesion, and adhesion.
- Crystalline solids can be classified by their shape and composition.

- viscosity
- surface tension
- surfactant
- crystalline solid
- unit cell
- allotrope
- amorphous solid

Lesson 4 PHASE CHANGES

- States of a substance are referred to as phases when they coexist as physically distinct parts of a mixture.
- Energy changes occur during phase changes.
- Phase diagrams show how different temperatures and pressures affect the phase of a substance.

- melting point
- vaporization
- evaporation
- vapor pressure
- boiling point
- freezing point
- condensation
- deposition
- phase diagram
- triple point

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

Why does water naturally exist as a solid, liquid, and gas on Earth?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

Based on Real Data*

SEP Data Analysis Lab

How are the depth of an underwater dive and altitude related?

Most divers dive at locations at or near sea level. In some locations, divers dive at higher altitudes. The table shows the pressure gauge correction factor for high altitude diving.

CER Analyze and Interpret Data

1. **Compare** Use the data to make a graph of atmospheric pressure versus altitude.
2. **Claim** What is your actual diving depth if your depth gauge reads 18 m, but you are at an altitude of 1800 m, and your gauge does not compensate for altitude?
3. **Claim, Evidence, Reasoning** Dive tables are used to determine how long it is safe to stay under water at a specific depth. Why is it important to know the correct depth of the dive?

Data and Observations

Altitude Diving Correction Factors

Altitude (m)	Atmospheric Pressure (atm)	Pressure Gauge Correction Factor (m)
0	1.000	0.0
600	0.930	0.7
1200	0.864	1.4
1800	0.801	2.0
2400	0.743	2.7
3000	0.688	3.2

*Data obtained from: Sawatzky, D. 2000. Diving at Altitude Part I. *Diver Magazine*. June 2000.

GASES

ENCOUNTER THE PHENOMENON

How do hot air balloons fly?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about how hot air balloons fly.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.

LESSON 1: Explore & Explain: The Combined Gas Law

LESSON 2: Explore & Explain: The Ideal Gas Law, Molar Mass, and Density

LESSON 1

THE GAS LAWS

FOCUS QUESTION

How are a gas's temperature, pressure, and volume related?

Boyle's Law

Robert Boyle (1627–1691), an Irish chemist, described the relationship between the pressure and the volume of a gas, such as the air in a balloon.

How are pressure and volume related?

Boyle designed experiments like the one shown in **Figure 1.** He showed that if the temperature and the amount of gas are constant, doubling the pressure decreases the volume by one-half. On the other hand, reducing the pressure by one-half doubles the volume. A relationship in which one variable increases proportionally as the other variable decreases is known as an inversely proportional relationship.

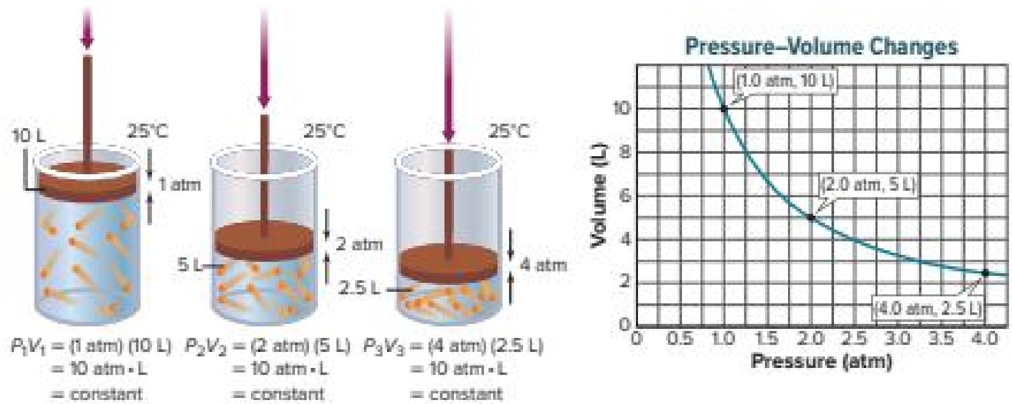


Figure 1 As the external pressure on the cylinder's piston increases, the volume inside the cylinder decreases. The graph shows the inverse relationship between pressure and volume.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Virtual Investigation: Gas Laws
Use a model to discover the patterns of the variables related to the properties of gases.

Laboratory: Charles's Law
Analyze and interpret data to measure the change in volume of a gas when the temperature is changed.

Boyle's law states that the volume of a fixed amount of gas held at a constant temperature varies inversely with the pressure. Look at the graph in **Figure 1**, on the previous page, in which volume versus pressure is plotted for a gas. The plot of an inversely proportional relationship results in a downward curve.

Note that the product of the pressure and the volume for each point in **Figure 1** is 10 atm·L. Boyle's law can be expressed mathematically as follows.

Boyle's Law

$$P_1V_1 = P_2V_2$$ *P* represents pressure. *V* represents volume.

For a given amount of gas held at constant temperature, the product of pressure and volume is constant.

P_1 and V_1 represent the initial conditions, and P_2 and V_2 represent new conditions. If you know any three of these values, you can solve for the fourth by rearranging the equation.

EXAMPLE Problem 1

BOYLE'S LAW A diver blows a 0.75-L air bubble 10 m under water. As it rises to the surface, the pressure goes from 2.25 atm to 1.03 atm. What will be the volume of air in the bubble at the surface?

1 ANALYZE THE PROBLEM

According to Boyle's law, the decrease in pressure on the bubble will result in an increase in volume, so the initial volume should be multiplied by a pressure ratio greater than 1.

Known	Unknown
$V_1 = 0.75$ L	$V_2 = ?$ L
$P_1 = 2.25$ atm	
$P_2 = 1.03$ atm	

2 SOLVE FOR THE UNKNOWN

Use Boyle's law. Solve for V_2, and calculate the new volume.

$P_1V_1 = P_2V_2$ State Boyle's law.

$V_2 = V_1\left(\frac{P_1}{P_2}\right)$ Solve for V_2.

$V_2 = 0.75\text{ L}\left(\frac{2.25\text{ atm}}{1.03\text{ atm}}\right)$ Substitute $V_1 = 0.75$ L, $P_1 = 2.25$ atm, and $P_2 = 1.03$ atm.

$V_2 = 0.75\text{ L}\left(\frac{2.25\cancel{\text{ atm}}}{1.03\cancel{\text{ atm}}}\right) = 1.6\text{ L}$ Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The pressure decreases by roughly half, so the volume should roughly double. The answer is expressed in liters, a unit of volume, and correctly contains two significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

Assume that the temperature and the amount of gas are constant in the following problems.

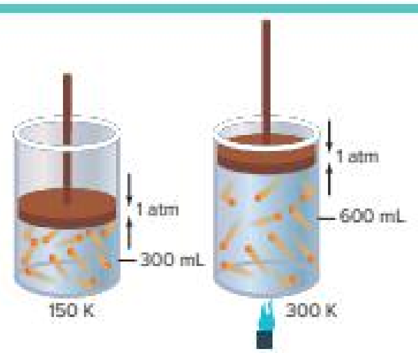
1. The volume of a gas at 99.0 kPa is 300.0 mL. If the pressure is increased to 188 kPa, what will be the new volume?
2. The pressure of a sample of helium in a 1.00-L container is 0.988 atm. What is the new pressure if the sample is placed in a 2.00-L container?
3. **CHALLENGE** Air trapped in a cylinder fitted with a piston occupies 145.7 mL at 1.08 atm pressure. What is the new volume when the piston is depressed, increasing the pressure by 25%?

Charles's Law

After a cool evening, a rubber pool raft can appear partially inflated. During a sunny afternoon, the same raft can appear fully inflated. Why does the appearance of the raft change? This can be answered by applying a second gas law—Charles's law.

How are temperature and volume related?

Jacques Charles (1746–1823), a French physicist, studied the relationship between volume and temperature. He observed that as temperature increases, so does the volume of a gas sample when the amount of gas and the pressure remain constant. The cylinders in **Figure 2** show how the volume of a fixed amount of gas changes as the gas is heated. Unlike **Figure 1,** where pressure in addition to that of the atmosphere was applied to the piston, the piston in **Figure 2** is free to float. The piston is supported by the gas in the cylinder at a level where the pressure of the gas exactly matches that of the atmosphere.



$$\frac{V_1}{T_1} = \frac{300\ \text{mL}}{150\ \text{K}} = 2\ \text{mL/K} = \text{constant}$$

$$\frac{V_2}{T_2} = \frac{600\ \text{mL}}{300\ \text{K}} = 2\ \text{mL/K} = \text{constant}$$

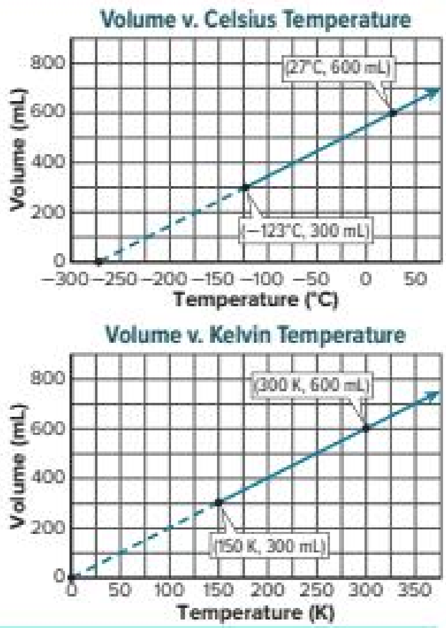


Figure 2 When the cylinder is heated, the kinetic energy of the gas particles increases, causing them to push the piston outward. The graphs show the relationship of volume to Celsius and Kelvin temperatures.

As you can see in **Figure 2,** the volume occupied by a gas at 1 atm increases as the temperature in the cylinder increases. The distance the piston moves is a measure of the increase in volume of the gas as it is heated. This behavior is explained by the kinetic-molecular theory: as temperature increases, gas particles move faster, striking the walls of their container more frequently and with greater force. Because pressure depends on the frequency and force with which gas particles strike the walls of their container, this would increase the pressure. For the pressure to stay constant, volume must increase so that the particles have farther to travel before striking the walls. Having to travel farther decreases the frequency with which the particles strike the walls of the container.

Graphing the relationship of temperature and volume

Figure 2 also shows graphs of the relationship between the temperature and the volume of a fixed amount of gas at constant pressure. The plot of volume versus temperature is a straight line. Note that you can predict the temperature at which the volume will reach 0 L by extrapolating the line to temperatures below the values that were measured.

In the first graph, the temperature that corresponds to 0 L is −273.15°C. This relationship is linear, but it is not a direct proportion. For example, you can see that the graph of the line does not pass through the origin and that doubling the temperature from 25°C to 50°C does not double the volume.

The second graph in **Figure 2**, which plots the Kelvin (K) temperature against volume, does show a direct proportion. A temperature of 0 K corresponds to 0 mL, and doubling the temperature doubles the volume. Zero on the Kelvin scale is also known as **absolute zero.** Absolute zero represents the lowest possible theoretical temperature. At absolute zero, the atoms are all in the lowest possible energy state.

Get It?

Explain why the second graph in **Figure 2** shows a direct proportion, but the first graph does not.

Using Charles's law

Charles's law states that the volume of a given amount of gas is directly proportional to its Kelvin temperature at constant pressure. Charles's law can be expressed as follows.

Charles's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

V represents volume.
T represents Temperature.

For a given amount of gas at constant pressure, the quotient of the volume and Kelvin temperature is a constant.

V_1 and T_1 represent initial conditions, while V_2 and T_2 are new conditions. As with Boyle's law, if you know three of the values, you can calculate the fourth. The temperature must be expressed in kelvin when using the equation for Charles's law. To convert a temperature from Celsius degrees to kelvin, add 273 to the Celsius temperature:

$$T_K = 273 + T_C$$

EXAMPLE Problem 2

CHARLES'S LAW A helium balloon in a closed car occupies a volume of 2.32 L at 40.0°C. If the car is parked on a hot day and the temperature inside rises to 75.0°C, what is the new volume of the balloon, assuming the pressure remains constant?

1 ANALYZE THE PROBLEM

Charles's law states that as the temperature of a fixed amount of gas increases, so does its volume, assuming constant pressure. Therefore, the volume of the balloon will increase. The initial volume should be multiplied by a temperature ratio greater than 1.

Known	Unknown
$T_1 = 40.0°C$	$V_2 = ?\ L$
$V_1 = 2.32\ L$	
$T_2 = 75.0°C$	

2 SOLVE FOR THE UNKNOWN

Convert degrees Celsius to kelvin.

$T_K = 273 + T_C$ — Apply the conversion factor.

$T_1 = 273 + 40.0°C = 313.0\ K$ — Substitute $T_1 = 40.0°C$.

$T_2 = 273 + 75.0°C = 348.0\ K$ — Substitute $T_2 = 75.0°C$.

Use Charles's law. Solve for V_2, and substitute the known values into the rearranged equation.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ — State Charles's law.

$V_2 = V_1\left(\frac{T_2}{T_1}\right)$ — Solve for V_2.

$V_2 = 2.32\ L\left(\frac{348.0\ K}{313.0\ K}\right)$ — Substitute $V_1 = 2.32\ L$, $T_1 = 313.0\ K$, and $T_2 = 348.0\ K$.

$V_2 = 2.32\ L\left(\frac{348.0\ \cancel{K}}{313.0\ \cancel{K}}\right) = 2.58\ L$ — Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The increase in temperature is relatively small, so the volume should show a small increase. The unit of the answer is liters, a volume unit, and there are three significant figures.

PRACTICE Problems

ADDITIONAL PRACTICE

Assume that the pressure and the amount of gas remain constant in the following problems.

4. What volume will the gas in the balloon at right occupy at 250 K?
5. A gas at 89°C occupies a volume of 0.67 L. At what Celsius temperature will the volume increase to 1.12 L?
6. The Celsius temperature of a 3.00-L sample of gas is lowered from 80.0°C to 30.0°C. What will be the resulting volume of this gas?
7. **CHALLENGE** A gas occupies 0.67 L at 350 K. What temperature is required to reduce the volume by 45%?

Gay-Lussac's Law

If volume is constant, is there a relationship between temperature and pressure? The answer to that question is found in Gay-Lussac's law.

How are temperature and pressure of a gas related?

Pressure is a direct result of collisions between gas particles and the walls of their container. An increase in temperature increases collision frequency and energy, so raising the temperature should also raise the pressure if the volume is not changed. Joseph Gay-Lussac (1778–1850) found that a direct proportion exists between Kelvin temperature and pressure, as illustrated in **Figure 3.** **Gay-Lussac's law** states that the pressure of a fixed amount of gas varies directly with the Kelvin temperature when the volume remains constant. It can be expressed mathematically as follows.

Gay-Lussac's Law

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

P represents pressure.
T represents temperature.

For a given amount of gas held at constant volume, the quotient of the pressure and the Kelvin temperature is a constant.

As with Boyle's and Charles's laws, if you know any three of the four variables, you can calculate the fourth using this equation. Remember that temperature must be in kelvin whenever it is used in a gas law equation.

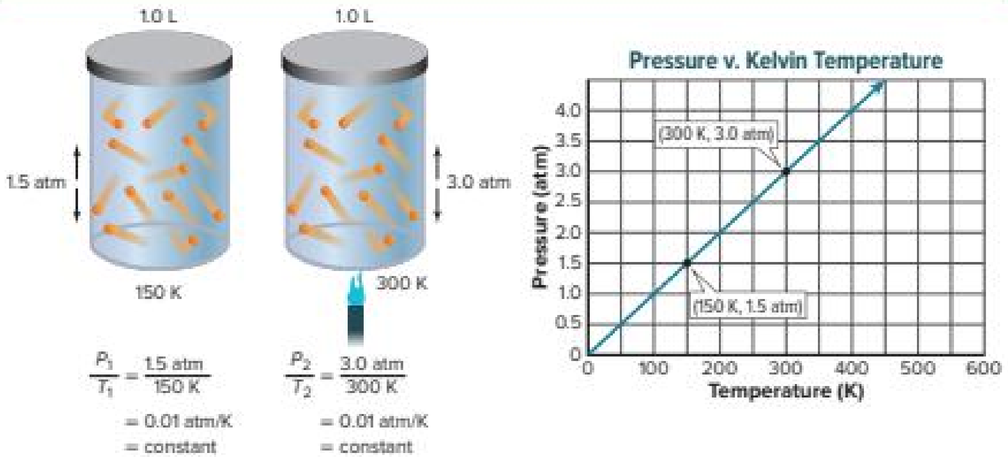


Figure 3 When the cylinder is heated, the kinetic energy of the particles increases, increasing both the frequency and energy of the collisions with the container wall. The volume of the cylinder is fixed, so the pressure exerted by the gas increases.

Compare and contrast *the graphs in* ***Figure 2****, earlier in the lesson, and* ***Figure 3****.*

EXAMPLE Problem 3

GAY-LUSSAC'S LAW The pressure of the oxygen gas inside a canister is 5.00 atm at 25.0°C. The canister is located at a camp high on Mount Everest. If the temperature there falls to −10.0°C, what is the new pressure inside the canister?

1 ANALYZE THE PROBLEM

Gay-Lussac's law states that if the temperature of a gas decreases, so does its pressure when volume is constant. Therefore, the pressure in the oxygen canister will decrease. The initial pressure should be multiplied by a temperature ratio less than 1.

Known
$P_1 = 5.00$ atm
$T_1 = 25.0°C$
$T_2 = -10.0°C$

Unknown
$P_2 = ?$ atm

2 SOLVE FOR THE UNKNOWN

Convert degrees Celsius to kelvin.

$T_K = 273 + T_C$ — Apply the conversion factor.

$T_1 = 273 + 25.0°C = 298.0$ K — Substitute $T_1 = 25.0°C$.

$T_2 = 273 + (-10.0°C) = 263.0$ K — Substitute $T_2 = -10.0°C$.

Use Gay-Lussac's law. Solve for P_2, and substitute the known values into the rearranged equation.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$ State Gay-Lussac's law.

$$P_2 = P_1\left(\frac{T_2}{T_1}\right)$$ Solve for P_2.

$$P_2 = 5.00 \text{ atm}\left(\frac{263.0 \text{ K}}{298.0 \text{ K}}\right)$$ Substitute $P_1 = 5.00$ atm, $T_1 = 298.0$ K, and $T_2 = 263.0$ K.

$$P_2 = 5.00 \text{ atm}\left(\frac{263.0 \cancel{\text{K}}}{298.0 \cancel{\text{K}}}\right) = \mathbf{4.41\ atm}$$ Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

Kelvin temperature decreases, so the pressure should decrease. The unit is atm, a pressure unit, and there are three significant figures.

Real-World Chemistry
Gay-Lussac's Law

PRESSURE COOKERS A pressure cooker is a pot with a lid that locks into place. This seals the container, which keeps its volume constant. Heating the pot increases the pressure in the cooker. As pressure increases, the temperature continues to increase and foods cook faster.

PRACTICE Problems

ADDITIONAL PRACTICE

Assume that the volume and the amount of gas are constant in the following problems.

8. The pressure in an automobile tire is 1.88 atm at 25.0°C. What will be the pressure if the temperature increases to 37.0°C?

9. Helium gas in a 2.00-L cylinder is under 1.12 atm pressure. At 36.5°C, that same gas sample has a pressure of 2.56 atm. What was the initial temperature in degrees Celsius of the gas in the cylinder?

10. CHALLENGE If a gas sample has a pressure of 30.7 kPa at 0.00°C, by how many degrees Celsius does the temperature have to increase to cause the pressure to double?

The Combined Gas Law

In a number of applications involving gases, such as the weather balloon in **Figure 4**, pressure, temperature, and volume might all change. Boyle's, Charles's, and Gay-Lussac's laws can be combined into a single law. This **combined gas law** states the relationships between pressure, temperature, and volume of a fixed amount of gas. All three variables have the same relationship to each other as they have in the other gas laws: pressure is inversely proportional to volume and directly proportional to temperature, and volume is directly proportional to temperature. The combined gas law can be expressed mathematically as follows.

Figure 4 Weather balloons carry instruments that send data, such as air temperature, pressure, and humidity, to receivers on the ground. As the balloon rises, its volume responds to changes in temperature and pressure, expanding until the sides burst. A small parachute returns the instruments to Earth.

The Combined Gas Law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P represents pressure. *V* represents volume. *T* represents temperature.

For a given amount of gas, the product of pressure and volume, divided by the Kelvin temperature, is a constant.

Using the combined gas law

The combined gas law enables you to solve problems involving changes in more than one variable. It also provides a way for you to remember the other three laws without memorizing each equation. If you can write out the combined gas law equation, equations for the other laws can be derived from it by remembering which variable is held constant in each case.

For example, if temperature remains constant as pressure and volume vary, then $T_1 = T_2$. After simplifying the combined gas law under these conditions, you are left with $P_1V_1 = P_2V_2$, which you should recognize as the equation for Boyle's law.

Get It?

Derive Charles's and Gay-Lussac's laws from the combined gas law.

STEM CAREER Connection

Occupational Health and Safety Technician

Would you like to travel and work in the field? A career in occupational health and safety might be for you. These technicians work in a variety of settings, such as offices, factories, and mines. They collect data on the health and safety conditions of the workplace in an effort to protect the workers, property, environment, and the public from harm.

EXAMPLE Problem 4

THE COMBINED GAS LAW A gas at 110 kPa and 30.0°C fills a flexible container with an initial volume of 2.00 L. If the temperature is raised to 80.0°C and the pressure increases to 440 kPa, what is the new volume?

1 ANALYZE THE PROBLEM

Both pressure and temperature change, so you will need to use the combined gas law. The pressure quadruples, but the temperature does not increase by such a large factor. Therefore, the new volume will be smaller than the starting volume.

Known

$P_1 = 110$ kPa $P_2 = 440$ kPa

$T_1 = 30.0°C$ $T_2 = 80.0°C$

$V_1 = 2.00$ L

Unknown

$V_2 = ?$ L

2 SOLVE FOR THE UNKNOWN

Convert degrees Celsius to kelvin.

$T_K = 273 + T_C$ — Apply the conversion factor.

$T_1 = 273 + 30.0°C = 303.0$ K — Substitute $T_1 = 30.0°C$.

$T_2 = 273 + 80.0°C = 353.0$ K — Substitute $T_2 = 80.0°C$.

Use the combined gas law. Solve for V_2, and substitute the known values into the rearranged equation.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$ State the combined gas law.

$$V_2 = V_1\left(\frac{P_1}{P_2}\right)\left(\frac{T_2}{T_1}\right)$$ Solve for V_2.

$$V_2 = 2.00\ \text{L}\left(\frac{110\ \text{kPa}}{440\ \text{kPa}}\right)\left(\frac{353.0\ \text{K}}{303.0\ \text{K}}\right)$$ Substitute $V_1 = 2.00$ L, $P_1 = 110.0$ kPa, $P_2 = 440$ kPa, $T_2 = 353.0$ K, and $T_1 = 303.0$ K.

$$V_2 = 2.00\ \text{L}\left(\frac{110\ \text{k}\cancel{\text{Pa}}}{440\ \text{k}\cancel{\text{Pa}}}\right)\left(\frac{353.0\ \cancel{\text{K}}}{303.0\ \cancel{\text{K}}}\right) = 0.58\ \text{L}$$ Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

Because the pressure change is much greater than the temperature change, the volume undergoes a net decrease. The unit is liters, a volume unit, and there are two significant figures.

PRACTICE Problems

ADDITIONAL PRACTICE

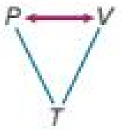
Assume that the amount of gas is constant in the following problems.

11. A sample of air in a syringe exerts a pressure of 1.02 atm at 22.0°C. The syringe is placed in a boiling-water bath at 100.0°C. The pressure is increased to 1.23 atm by pushing the plunger in, which reduces the volume to 0.224 mL. What was the initial volume?

12. A balloon contains 146.0 mL of gas confined at a pressure of 1.30 atm and a temperature of 5.0°C. If the pressure doubles and the temperature decreases to 2.0°C, what will be the volume of gas in the balloon?

13. CHALLENGE If the temperature in the gas cylinder at right increases to 30.0°C and the pressure increases to 1.20 atm, will the cylinder's piston move up or down?

Table 1 The Gas Laws

Law	Boyle's	Charles's	Gay-Lussac's	Combined
Formula	$P_1V_1 = P_2V_2$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
What is constant?	amount of gas, temperature	amount of gas, pressure	amount of gas, volume	amount of gas
Graphic organizer				

You have now seen how pressure, temperature, and volume affect a gas sample. You can use the gas laws, summarized in **Table 1**, as long as the amount of gas remains constant. But what happens if the amount of gas changes? In the next lesson, you will add the fourth variable, amount of gas present, to the gas laws.

Check Your Progress

Summary

- Boyle's law states that the volume of a fixed amount of gas is inversely proportional to its pressure at constant temperature.
- Charles's law states that the volume of a fixed amount of gas is directly proportional to its Kelvin temperature at constant pressure.
- Gay-Lussac's law states that the pressure of a fixed amount of gas is directly proportional to its Kelvin temperature at constant volume.
- The combined gas law relates pressure, temperature, and volume in a single statement.

Demonstrate Understanding

14. **State** the relationships between pressure, temperature, and volume of a fixed amount of gas.
15. **Explain** Which of the three variables that apply to equal amounts of gases are directly proportional? Which are inversely proportional?
16. **Analyze** A weather balloon is released into the atmosphere. You know the initial volume, temperature, and air pressure. What information will you need to predict its volume when it reaches its final altitude?
17. **Infer** Why must compressed gases be shielded from high temperatures? What must happen to compressed oxygen before it can be inhaled?
18. **Calculate** A rigid plastic container holds 1.00 L of methane gas at 660 torr pressure when the temperature is 22.0°C. How much pressure will the gas exert if the temperature is raised to 44.6°C?
19. **Design** a concept map that shows the relationships between pressure, volume, and temperature in Boyle's, Charles's, and Gay-Lussac's laws.

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LESSON 2
THE IDEAL GAS LAW

FOCUS QUESTION

What happens when you change the amount of gas present?

Avogadro's Principle

The particles that make up different gases can vary greatly in size. However, kinetic-molecular theory assumes that the particles in a gas sample are far enough apart that size has very little influence on the volume occupied by a gas. For example, 1000 relatively large krypton gas particles occupy the same volume as 1000 smaller helium gas particles at the same temperature and pressure. It was Avogadro who first proposed this idea in 1811. **Avogadro's principle** states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. **Figure 5** shows equal volumes of carbon dioxide, helium, and oxygen.

Figure 5 Gas tanks of equal volume that are at the same pressure and temperature contain equal numbers of gas particles, regardless of which gas they contain.

Infer *Why doesn't Avogadro's principle apply to liquids and solids?*

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

ChemLAB: Determine Pressure in Popcorn Kernels
Analyze and interpret data to determine the quantity of pressure needed to bust a kernel of popcorn.

Quick Investigation: Model a Fire Extinguisher
Analyze and interpret data to understand the function of carbon dioxide in fire extinguishers.

Volume and moles

Recall that one mole of a substance contains 6.02×10^{23} particles. The **molar volume** of a gas is the volume that 1 mol occupies at 0.00°C and 1.00 atm pressure. The conditions of 0.00°C and 1.00 atm are known as **standard temperature and pressure (STP).** Avogadro showed experimentally that 1 mol of any gas occupies a volume of 22.4 L at STP.

Because the volume of 1 mol of a gas at STP is 22.4 L, you can use 22.4 L/mol as a conversion factor whenever a gas is at STP. For example, suppose you want to find the number of moles in a sample of gas that has a volume of 3.72 L at STP. Use the molar volume to convert from volume to moles.

$$3.72 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.166 \text{ mol}$$

EXAMPLE Problem 5

MOLAR VOLUME The main component of natural gas used for home heating and cooking is methane (CH_4). Calculate the volume that 2.00 kg of methane gas will occupy at STP.

1 ANALYZE THE PROBLEM

The number of moles can be calculated by dividing the mass of the sample, *m*, by its molar mass, M. The gas is at STP (0.00°C and 1.00 atm pressure), so you can use the molar volume to convert from the number of moles to the volume.

Known	Unknown
$m = 2.00$ kg	$V = ?$ L
$T = 0.00$°C	
$P = 1.00$ atm	

2 SOLVE FOR THE UNKNOWN

Determine the molar mass for methane.

$$\text{M} = 1\text{ C } \cancel{\text{atom}}\left(\frac{12.01 \text{ amu}}{1 \text{ C } \cancel{\text{atom}}}\right) + 4 \text{ H } \cancel{\text{atoms}}\left(\frac{1.01 \text{ amu}}{1 \text{ H } \cancel{\text{atom}}}\right)$$

Determine the molecular mass.

$$= 12.01 \text{ amu} + 4.04 \text{ amu} = 16.05 \text{ amu}$$

$$= 16.05 \text{ g/mol}$$

Express the molecular mass as g/mol to arrive at the molar mass.

Determine the number of moles of methane.

$$2.00 \cancel{\text{kg}}\left(\frac{1000 \text{ g}}{1 \cancel{\text{kg}}}\right) = 2.00 \times 10^3 \text{ g}$$

Convert the mass from kg to g.

$$\frac{m}{\text{M}} = \frac{2.00 \times 10^3 \cancel{\text{g}}}{16.05 \cancel{\text{g}}/\text{mol}} = 125 \text{ mol}$$

Divide mass by molar mass to determine the number of moles.

Use the molar volume to determine the volume of methane at STP.

$$V = 125 \cancel{\text{mol}} \times \frac{22.4 \text{ L}}{1 \cancel{\text{mol}}} = 2.80 \times 10^3 \text{ L}$$

Use the molar volume, 22.4 L/mol, to convert from moles to the volume.

3 EVALUATE THE ANSWER

The amount of methane present is much more than 1 mol, so you should expect a large volume, which is in agreement with the answer. The unit is liters, a volume unit, and there are three significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

20. What size container do you need to hold 0.0459 mol of N_2 gas at STP?
21. How much carbon dioxide gas, in grams, is in a 1.0-L balloon at STP?
22. What volume in milliliters will 0.00922 g of H_2 gas occupy at STP?
23. What volume will 0.416 g of krypton gas occupy at STP?
24. Calculate the volume that 4.5 kg of ethylene gas (C_2H_4) will occupy at STP.
25. **CHALLENGE** A flexible plastic container contains 0.860 g of helium gas in a volume of 19.2 L. If 0.205 g of helium is removed at constant pressure and temperature, what will be the new volume?

The Ideal Gas Law

Avogadro's principle and the laws of Boyle, Charles, and Gay-Lussac can be combined into a single mathematical statement that describes the relationships between pressure, volume, temperature, and number of moles of a gas. This formula works best for gases that obey the assumptions of the kinetic-molecular theory. Known as ideal gases, their particles occupy a negligible volume and are far enough apart that they exert minimal attractive or repulsive forces on one another.

Figure 6 The volume and temperature of this tire stay the same as air is added. However, the pressure in the tire increases as the amount of air present increases.

From the combined gas law to the ideal gas law

The combined gas law relates the variables of pressure, volume, and temperature for a given amount of gas.

$$\frac{P_1V_1}{T_1} = \frac{V_2V_2}{T_2}$$

For a specific sample of gas, this relationship of pressure, volume, and temperature is always the same. You could rewrite the relationship represented in the combined gas law as follows.

$$\frac{PV}{T} = \text{constant}$$

As **Figure 6** illustrates, increasing the amount of gas present in a sample will raise the pressure if temperature and volume are constant. Likewise, if pressure and temperature remain constant, the volume will increase as more particles of a gas are added. In fact, we know that both volume and pressure are directly proportional to the number of moles, n, so n can be incorporated into the combined gas law as follows.

$$\frac{PV}{nT} = \text{constant}$$

Table 2 Values of R

Value of R	Units of R
0.0821	$\frac{\text{L·atm}}{\text{mol·K}}$
8.314	$\frac{\text{L·kPa}}{\text{mol·K}}$
62.4	$\frac{\text{L·mm Hg}}{\text{mol·K}}$

Experiments using known values of P, T, V, and n have determined the value of this constant. It is called the **ideal gas constant,** and it is represented by the symbol R. If pressure is in atmospheres, the value of R is 0.0821 L·atm/mol·K. Note that the units for R are simply the combined units for each of the four variables. **Table 2** shows the numerical values for R in different units of pressure.

Substituting R for the constant in the equation on the previous page and rearranging the variables gives the most familiar form of the ideal gas law. The **ideal gas law** describes the physical behavior of an ideal gas in terms of the pressure, volume, temperature, and number of moles of gas present.

The Ideal Gas Law

$$PV = nRT$$

P represents pressure. *V* represents volume.
n represents number of moles. R is the ideal gas constant.
T represents temperature.

For a given amount of gas held at constant temperature, the product of pressure and volume is a constant.

If you know any three of the four variables, you can rearrange the equation to solve for the unknown.

EXAMPLE Problem 6

THE IDEAL GAS LAW Calculate the number of moles of ammonia gas (NH_3) contained in a 3.0-L vessel at 3.00×10^2 K with a pressure of 1.50 atm.

1 ANALYZE THE PROBLEM

You are given the volume, temperature, and pressure of a gas sample. Use the ideal gas law, and select the value of R that contains the pressure units given in the problem. Because the pressure and temperature are close to STP, but the volume is much smaller than 22.4 L, it would make sense if the calculated answer were much smaller than 1 mol.

Known
$V = 3.0$ L
$T = 3.00 \times 10^2$ K
$P = 1.50$ atm
$R = 0.0821 \frac{\text{L·atm}}{\text{mol·K}}$

Unknown
$n = ?$ mol

2 SOLVE FOR THE UNKNOWN

Use the ideal gas law. Solve for *n*, and substitute the known values.

$PV = nRT$ — State the ideal gas law.

$n = \frac{PV}{RT}$ — Solve for *n*.

$$n = \frac{(1.50 \text{ atm})(3.0 \text{ L})}{\left(0.0821 \frac{\text{L·atm}}{\text{mol·K}}\right)(3.00 \times 10^2 \text{ K})}$$

Substitute $V = 3.0$ L, $T = 3.00 \times 10^2$ K, $P = 1.50$ atm, and $R = 0.0821$ L·atm/mol·K.

$$n = \frac{(1.50 \cancel{\text{atm}})(3.0 \cancel{\text{L}})}{\left(0.0821 \frac{\cancel{\text{L·atm}}}{\text{mol·}\cancel{\text{K}}}\right)(3.00 \times 10^2 \cancel{\text{K}})} = 0.18 \text{ mol}$$

Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The answer agrees with the prediction that the number of moles present will be significantly less than 1 mol. The unit of the answer is the mole, and there are two significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

26. Determine the Celsius temperature of 2.49 mol of a gas contained in a 1.00-L vessel at a pressure of 143 kPa.
27. Calculate the volume of a 0.323-mol sample of a gas at 265 K and 0.900 atm.
28. What is the pressure, in atmospheres, of a 0.108-mol sample of helium gas at a temperature of 20.0°C if its volume is 0.505 L?
29. If the pressure exerted by a gas at 25°C in a volume of 0.044 L is 3.81 atm, how many moles of gas are present?
30. **CHALLENGE** An ideal gas has a volume of 3.0 L. If the number of moles of gas and the temperature are doubled, while the pressure remains constant, what is the new volume?

The Ideal Gas Law—Molar Mass and Density

The ideal gas law can be used to solve for the value of any one of the four variables P, V, T, or n if the values of the other three are known. However, you can also rearrange the $PV = nRT$ equation to calculate the molar mass and density of a gas sample.

Molar mass and the ideal gas law

To find the molar mass of a gas sample, the mass, temperature, pressure, and volume of the gas must be known. Recall that the number of moles of a gas (n) is equal to the mass (m) divided by the molar mass (M). Therefore, the n in the equation can be replaced by m/M.

$$PV = nRT \quad \xrightarrow{\text{substitute } n = \frac{m}{\mathrm{M}}} \quad PV = \frac{mRT}{\mathrm{M}}$$

You can rearrange the new equation to solve for the molar mass.

$$\mathrm{M} = \frac{mRT}{PV}$$

Density and the ideal gas law

Recall that the density (D) of a substance is defined as mass (m) per unit volume (V). After rearranging the ideal gas equation to solve for molar mass, you can substitute D for m/V.

$$\mathrm{M} = \frac{mRT}{PV} \quad \xrightarrow{\text{substitute } \frac{m}{V} = D} \quad \mathrm{M} = \frac{DRT}{P}$$

You can rearrange the new equation to solve for density.

$$D = \frac{\mathrm{M}P}{RT}$$

Figure 7 To extinguish a fire, you need to take away fuel, oxygen, or heat. The fire extinguishers in the photo contain carbon dioxide, which displaces oxygen but does not burn. It also has a cooling effect due to the rapid expansion of the carbon dioxide as it is released from the nozzle.

Explain *Why does carbon dioxide displace oxygen?*

Why might you need to know the density of a gas? Consider fighting a fire. One way to put out a fire is to remove its oxygen source by covering it with another gas that will neither burn nor support combustion, as shown in **Figure 7.** This gas must have a greater density than oxygen so that it will displace the oxygen at the source of the fire.

PROBLEM-SOLVING STRATEGY

Deriving Gas Laws

If you master the following strategy, you will need to remember only one gas law—the ideal gas law. Consider the example of a fixed amount of gas held at constant pressure. You need Charles's law to solve problems involving volume and temperature.

1. Use the ideal gas law to write two equations that describe the gas sample at two different volumes and temperatures. (Quantities that do not change are shown in red.)
2. Isolate volume and temperature—the two conditions that vary—on the same side of each equation.
3. Because *n*, R, and *P* are constant under these conditions, you can set the volume and temperature conditions equal, deriving Charles's law.

$$PV_1 = nRT_1 \qquad PV_2 = nRT_2$$

$$\frac{V_1}{T_1} = \frac{nR}{P} \qquad \frac{V_2}{T_2} = \frac{nR}{P}$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Apply the Strategy

Derive Boyle's law, Gay-Lussac's law, and the combined gas law based on the example above.

CCC CROSSCUTTING CONCEPTS

Scale, Proportion, and Quantity Algebraic thinking is used to examine data and predict the effect of a change in one variable on another. Write a paragraph citing examples of how scientists used algebraic thinking in their study of gases.

WORD ORIGIN

mole
comes from the German word *Mol*, which is short for *Molekulargewicht*, meaning *molecular weight*

Real Versus Ideal Gases

What does the term ***ideal gas*** mean? Ideal gases follow the assumptions of the kinetic-molecular theory. According to this theory, an ideal gas is one whose particles take up no space. Ideal gases experience no intermolecular attractive (electrical) forces, nor are they attracted or repelled by the walls of their containers. The particles of an ideal gas are in constant, random motion, moving in straight lines until they collide with each other or with the walls of the container. Additionally, these collisions are perfectly elastic, which means that the kinetic energy of the system does not change. An ideal gas follows the gas laws under all conditions of temperature and pressure.

In reality, no gas is truly ideal. All gas particles have some volume, however small, and are subject to intermolecular interactions. Also, the collisions that particles make with each other and with the container are not perfectly elastic. Despite that, most gases behave like ideal gases at a wide range of temperatures and pressures. Calculations made using the ideal gas law often closely approximate experimental measurements.

Get It?

Explain how the interactions of real gases are determined by attractive (electrical) forces between particles.

Extreme pressure and temperature

When is the ideal gas law not likely to work for a real gas? Real gases deviate most from ideal gas behavior at high pressures and low temperatures. The nitrogen gas shown in **Figure 8 (left)** behaves as a real gas. Lowering the temperature of nitrogen gas results in less kinetic energy of the gas particles, which means their intermolecular attractive (electric) forces are strong enough to affect their behavior. When the temperature is low enough, this real gas condenses to form a liquid.

The propane gas in the tanks shown in **Figure 8 (right)** also behaves as a real gas. Increasing the pressure on a gas forces the gas particles closer together until the volume occupied by the gas particles themselves is no longer negligible. Real gases such as propane will liquefy if enough pressure is applied.

Nitrogen gas turns to liquid at −196°C. At this temperature, scientists can preserve biological specimens, such as body tissues, for future research or medical procedures.

About 270 times more propane can be stored as a liquid than as a gas in the same amount of space. Your family might use small tanks of liquid propane as fuel for your barbecue grill or larger tanks for heating and cooking.

Figure 8 Real gases do not follow the ideal gas law at all pressures and temperatures.

Polarity and size of particles

The nature of the particles making up a gas also affects how ideally the gas behaves. For example, polar gas molecules, such as water vapor, generally have larger attractive forces between their particles than nonpolar gases, such as helium. The oppositely charged ends of polar molecules are pulled together through electrostatic forces, as shown in **Figure 9**. Therefore, polar gases do not behave as ideal gases. Also, the particles of gases composed of larger nonpolar molecules, such as butane (C_4H_{10}), occupy more actual volume than an equal number of smaller gas particles in gases such as helium (He). Therefore, larger gas particles tend to exhibit a greater departure from ideal behavior than do smaller gas particles.

Figure 9 Polar gases, such as the water vapor shown, experience forces of attraction between particles. In a nonpolar gas, there is minimal attraction between particles.

Check Your Progress

Summary

- Avogadro's principle states that equal volumes of gases at the same pressure and temperature contain equal numbers of particles.
- The ideal gas law relates the amount of a gas present to its pressure, temperature, and volume.
- The ideal gas law can be used to find molar mass, if the mass of the gas is known, or the density of the gas, if its molar mass is known.
- At very high pressures and very low temperatures, real gases behave differently than ideal gases.

Demonstrate Understanding

31. **Explain** why Avogadro's principle holds true for ideal gases that have small particles and for ideal gases that have large particles.
32. **State** the equation for the ideal gas law.
33. **Interpret** the behavior of ideal gases in terms of kinetic molecular theory.
34. **Predict** the conditions under which a real gas might deviate from ideal behavior.
35. **List** common units for each variable in the ideal gas law.
36. **Calculate** A 2.00-L flask is filled with propane gas (C_3H_8) at a pressure of 1.00 atm and a temperature of −15.0°C. What is the mass of the propane in the flask?
37. **Make and Use Graphs** For every 6°C drop in temperature, the air pressure in a car's tires goes down by about 1 psi (14.7 psi = 1.00 atm). Make a graph illustrating the change in tire pressure from 20°C to −20°C (assume 30.0 psi at 20°C).

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LESSON 3
GAS STOICHIOMETRY

FOCUS QUESTION

How are the amounts of gaseous reactants and products in a chemical reaction calculated?

Stoichiometry of Reactions Involving Gases

The gas laws can be applied to calculate the stoichiometry of reactions in which gases are reactants or products. Recall that the coefficients in chemical equations represent molar amounts of substances taking part in the reaction. For example, hydrogen gas can react with oxygen gas to produce water vapor.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

From the balanced chemical equation, you know that 2 mol of hydrogen gas reacts with 1 mol of oxygen gas, producing 2 mol of water vapor. This tells you the molar ratios of substances in this reaction. Avogadro's principle states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. Thus, for gases, the coefficients in a balanced chemical equation represent not only molar amounts but also relative volumes. Therefore, 2 L of hydrogen gas would react with 1 L of oxygen gas to produce 2 L of water vapor.

Stoichiometry and Volume–Volume Problems

To find the volume of a gaseous reactant or product in a reaction, you must know the balanced equation for the reaction and the volume of at least one other gas involved in the reaction. Examine the reaction in **Figure 10,** which shows the combustion of methane.

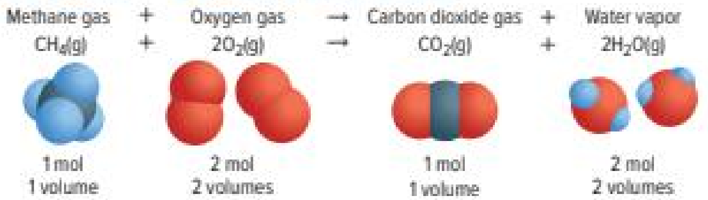

Figure 10 The coefficients in a balanced equation show the relationships between numbers of moles of all reactants and products, and the relationships between volumes of any gaseous reactants or products. From these coefficients, volume ratios can be set up for any pair of gases in the reaction.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Review the News

Obtain information from a current news story about gas stoichiometry. Evaluate your source and communicate your findings to your class.

Because the coefficients represent volume ratios for gases taking part in the reaction, you can determine that it takes 2 L of oxygen to react completely with 1 L of methane. The complete combustion of 1 L of methane will produce 1 L of carbon dioxide and 2 L of water vapor. Note that no conditions of temperature and pressure are listed. They are not needed as part of the calculation because after mixing, both gases are at the same temperature and pressure. The temperature of the entire reaction might change during the reaction, but a change in temperature would affect all gases in the reaction the same way.

EXAMPLE Problem 7

VOLUME–VOLUME PROBLEMS What volume of oxygen gas is needed for the complete combustion of 4.00 L of propane gas (C_3H_8)? Assume that pressure and temperature remain constant.

1 ANALYZE THE PROBLEM

You are given the volume of a gaseous reactant in a chemical reaction. Remember that the coefficients in a balanced chemical equation provide the volume relationships of gaseous reactants and products.

Known	Unknown
$V_{C_3H_8} = 4.00\ L$	$V_{O_2} = ?\ L$

2 SOLVE FOR THE UNKNOWN

Use the balanced equation for the combustion of C_3H_8. Find the volume ratio for O_2 and C_3H_8, then solve for V_{O_2}.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

Write the balanced equation.

$$\frac{5\ \text{volumes}\ O_2}{1\ \text{volume}\ C_3H_8}$$

Find the volume ratio for O_2 and C_3H_8.

$$V_{O_2} = (4.00\ L\ \cancel{C_3H_8}) \times \frac{5\ \cancel{\text{volumes}}\ O_2}{1\ \cancel{\text{volume}\ C_3H_8}}$$

Multiply the known volume of C_3H_8 by the volume ratio to find the volume of O_2.

$$= 20.0\ L\ O_2$$

3 EVALUATE THE ANSWER

The coefficients in the combustion equation show that a much larger volume of O_2 than C_3H_8 is used up in the reaction, which is in agreement with the calculated answer. The unit of the answer is liters, a unit of volume, and there are three significant figures.

Real-World Chemistry

Using Stoichiometry

KILNS Correct proportions of gases are needed for many chemical reactions. Although many pottery kilns are fueled by methane, a precise mixture of propane and air can be used to fuel a kiln if methane is unavailable.

PRACTICE Problems

ADDITIONAL PRACTICE

38. How many liters of propane gas (C_3H_8) will undergo complete combustion with 34.0 L of oxygen gas?

39. Determine the volume of hydrogen gas needed to react completely with 5.00 L of oxygen gas to form water.

40. What volume of oxygen is needed to completely combust 2.36 L of methane gas (CH_4)?

41. CHALLENGE Nitrogen and oxygen gases react to form dinitrogen monoxide gas (N_2O). What volume of O_2 is needed to produce 34 L of N_2O?

Figure 11 Ammonia is essential in the production of fertilizers containing nitrogen. Proper levels of nitrogen in the soil lead to increased crop yields.

Stoichiometry and Volume–Mass Problems

BIOLOGY Connection What you have learned about stoichiometry can be applied to the production of ammonia (NH_3) from nitrogen gas (N_2). Fertilizer manufacturers use ammonia to make nitrogen-based fertilizers. Nitrogen is an essential element for plant growth. Natural sources of nitrogen in soil, such as nitrogen fixation by plants, the decomposition of organic matter, and animal wastes, do not always supply enough nitrogen for optimum crop yields. **Figure 11** shows a farmer applying fertilizer rich in nitrogen to the soil. This enables the farmer to produce a crop with a higher yield.

Example Problem 8 shows how to use a volume of nitrogen gas to produce a certain amount of ammonia. In doing this type of problem, remember that the balanced chemical equation allows you to find ratios for only moles and gas volumes, not for masses. All masses given must be converted to moles or volumes before being used as part of a ratio. Also, remember that the temperature units used must be kelvin.

EXAMPLE Problem 8

VOLUME–MASS PROBLEMS Ammonia is synthesized from hydrogen and nitrogen.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

If 5.00 L of nitrogen reacts completely with hydrogen at a pressure of 3.00 atm and a temperature of 298 K, how much ammonia, in grams, is produced?

1 ANALYZE THE PROBLEM

You are given the volume, pressure, and temperature of a gas sample. The mole and volume ratios of gaseous reactants and products are given by the coefficients in the balanced chemical equation. Volume can be converted to moles and thus related to mass by using molar mass and the ideal gas law.

Known	Unknown
$V_{N_2} = 5.00$ L	$m_{NH_3} = ?$ g
$P = 3.00$ atm	
$T = 298$ K	

EXAMPLE Problem 8 (continued)

2 SOLVE FOR THE UNKNOWN

Determine how many liters of gaseous ammonia will be made from 5.00 L of nitrogen gas.

$$\frac{1 \text{ volume } N_2}{2 \text{ volumes } NH_3}$$ Find the volume ratio for N_2 and NH_3 using the balanced equation.

$$5.00 \text{ L } \cancel{N_2}\left(\frac{2 \cancel{\text{ volumes}} NH_3}{1 \cancel{\text{ volume } N_2}}\right) = 10.0 \text{ L } NH_3$$ Multiply the known volume of N_2 by the volume ratio to find the volume of NH_3.

Use the ideal gas law. Solve for n, and calculate the number of moles of NH_3.

$$PV = nRT$$ State the ideal gas law.

$$n = \frac{PV}{RT}$$ Solve for n.

$$n = \frac{(3.00 \text{ atm})(10.0 \text{ L})}{\left(0.0821 \frac{\text{L·atm}}{\text{mol·K}}\right)(298 \text{ K})}$$ Substitute $P = 3.00$ atm, $V_{NH_3} = 10.0$ L, and $T = 298$ K.

$$n = \frac{(3.00 \cancel{\text{atm}})(10.0 \cancel{\text{L}})}{\left(0.0821 \frac{\cancel{\text{L·atm}}}{\text{mol·}\cancel{\text{K}}}\right)(298 \cancel{\text{K}})} = 1.23 \text{ mol } NH_3$$ Multiply and divide numbers and units.

$$M = \left(\frac{1 \cancel{\text{N atoms}} \times 14.01 \text{ amu}}{1 \cancel{\text{N atom}}}\right) + \left(\frac{3 \cancel{\text{H atoms}} \times 1.01 \text{ amu}}{1 \cancel{\text{H atom}}}\right)$$ Find the molecular mass of NH_3.

$$= 17.04 \text{ amu}$$

$$M = 17.04 \text{ g/mol}$$ Express molar mass in units of g/mol.

Convert moles of ammonia to grams of ammonia.

$$1.23 \cancel{\text{mol } NH_3} \times \frac{17.04 \text{ g } NH_3}{1 \cancel{\text{mol } NH_3}} = 21.0 \text{ g } NH_3$$ Use the molar mass of ammonia as a conversion factor.

3 SOLVE FOR THE UNKNOWN

To check your answer, calculate the volume of reactant nitrogen at STP. Then, use molar volume and the mole ratio between N_2 and NH_3 to determine how many moles of NH_3 were produced. The unit of the answer is grams, a unit of mass. There are three significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

42. Ammonium nitrate is a common ingredient in chemical fertilizers. Use the reaction shown to calculate the mass of solid ammonium nitrate that must be used to obtain 0.100 L of dinitrogen monoxide gas at STP.

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(g)$$

43. When solid calcium carbonate ($CaCO_3$) is heated, it decomposes to form solid calcium oxide (CaO) and carbon dioxide gas (CO_2). How many liters of carbon dioxide will be produced at STP if 2.38 kg of calcium carbonate reacts completely?

44. When iron rusts, it undergoes a reaction with oxygen to form iron(III) oxide.

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

Calculate the volume of oxygen gas at STP that is required to completely react with 52.0 g of iron.

45. CHALLENGE An excess of acetic acid is added to 28 g of sodium bicarbonate at 25°C and 1 atm pressure. During the reaction, the gas cools to 20°C. What volume of carbon dioxide will be produced? The balanced equation for the reaction is shown below.

$$NaHCO_3(aq) + CH_3COOH(aq) \rightarrow NaCH_3COO(aq) + CO_2(g) + H_2O(l)$$

Stoichiometric problems, such as the ones in this lesson, are considered in industrial processes. For example, ethene gas (C_2H_4), also called ethylene, is the raw material for making polyethylene polymers. Polyethylene is produced when numerous ethene molecules join together in chains of repeating $-CH_2-CH_2-$ units. These polymers are used to make many everyday items, such as the ones shown in **Figure 12**. The general formula for this polymerization reaction is shown below. In this formula, n is the number of units used.

$$n(C_2H_4)(g) \rightarrow -(CH_2-CH_2)_n-(s)$$

If you were a process engineer for a polyethylene manufacturing plant, you would need to know about the properties of ethene gas and the polymerization reaction. Knowledge of the gas laws would help you calculate both the mass and volume of raw material needed under different temperature and pressure conditions to make different types of polyethylene.

Figure 12 To effectively manufacture a product, such as many of these plastics, it is essential to answer the following questions. How much of a reactant should be purchased? How much of a product will be produced?

Check Your Progress

Summary

- The coefficients in a balanced chemical equation specify volume ratios for gaseous reactants and products.
- The gas laws can be used along with balanced chemical equations to calculate the amount of a gaseous reactant or product in a reaction.

Demonstrate Understanding

46. **Explain** When fluorine gas combines with water vapor, the following reaction occurs.

$$2F_2(g) + 2H_2O(g) \rightarrow O_2(g) + 4HF(g)$$

If the reaction starts with 2 L of fluorine gas, how many liters of water vapor react with the fluorine, and how many liters of oxygen and hydrogen fluoride are produced?

47. **Analyze** Is the volume of a gas directly or inversely proportional to the number of moles of a gas at constant temperature and pressure? Explain.

48. **Calculate** One mole of a gas occupies a volume of 22.4 L at STP. Calculate the temperature and pressure conditions needed to fit 2 mol of a gas into a volume of 22.4 L.

49. **Interpret Data** Ethene gas (C_2H_4) reacts with oxygen to form carbon dioxide and water. Write a balanced equation for this reaction, then find the mole ratios of substances on each side of the equation.

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SCIENCE & SOCIETY

What Goes Up Doesn't Always Come Down

Party balloons and squeaky voices are what most people associate with helium gas. But helium has more serious applications too—some of them lifesaving. Should people try harder to conserve this important element?

Should people be allowed to use helium for party balloons (left) when it is also essential for applications such as MRI imaging (right)?

A nonrenewable resource

Helium is one of the most abundant elements in the universe, but it is relatively scarce on Earth. Produced over millennia by the radioactive decay of elements such as uranium and thorium, helium is extracted for human use from natural underground pockets that have existed for millions or billions of years. Helium also can be removed by distillation from natural gas.

Helium's low density and chemical inertness, make it perfect for making buoyant and safe party balloons. But when helium is released into the atmosphere, its low density also means that the gas continues to rise, and the gas eventually escapes into space.

Most helium applications are based on another property of helium: its very low melting point (−272°C; almost 0 K). Helium can be cooled to very low temperatures and still remain a liquid. This property makes helium a good coolant, which is an important application of helium. MRI scanners use helium as a coolant because they use superconducting magnets that must be cooled to function properly. MRI images help physicians diagnose and monitor diseases, including cancer.

Helium also is used as a coolant in nuclear reactors and the large hadron collider, which is used for research in particle physics.

Supplies of helium fluctuate as some sources are depleted and new sources are discovered. Some scientists have called for an end to "frivolous" uses of helium and some industries have developed ways to recapture and recycle the helium they use. Still, with current sources continuing to produce and new ones being discovered, a true helium shortage is not likely any time soon.

OBTAIN, EVALUATE, AND COMMUNICATE INFORMATION

What is the United States National Helium Reserve? Conduct research to discover why it was created, how it is being used now, and how it relates to helium supplies in the United States. Make a digital slideshow to share your findings.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 THE GAS LAWS

- Boyle's law states that the volume of a fixed amount of gas is inversely proportional to its pressure at constant temperature.

$$P_1V_1 = P_2V_2$$

- Charles's law states that the volume of a fixed amount of gas is directly proportional to its Kelvin temperature at constant pressure.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

- Gay-Lussac's law states that the pressure of a fixed amount of gas is directly proportional to its Kelvin temperature at constant volume.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

- The combined gas law relates pressure, temperature, and volume in a single statement.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

- Boyle's law
- absolute zero
- Charles's law
- Gay-Lussac's law
- combined gas law

Lesson 2 THE IDEAL GAS LAW

- Avogadro's principle states that equal volumes of gases at the same pressure and temperature contain equal numbers of particles.
- The ideal gas law relates the amount of a gas present to its pressure, temperature, and volume.

$$PV = nRT$$

- The ideal gas law can be used to find molar mass, if the mass of the gas is known, or the density of the gas, if its molar mass is known.

$$\mathrm{M} = \frac{mRT}{PV} \qquad D = \frac{\mathrm{M}P}{RT}$$

- At very high pressures and very low temperatures, real gases behave differently than ideal gases.

- Avogadro's principle
- molar volume
- standard temperature and pressure (STP)
- ideal gas constant (R)
- ideal gas law

Lesson 3 GAS STOICHIOMETRY

- The coefficients in a balanced chemical equation specify volume ratios for gaseous reactants and products.
- The gas laws can be used along with balanced chemical equations to calculate the amount of a gaseous reactant or product in a reaction.

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

How do hot air balloons fly?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

SEP Data Analysis Lab

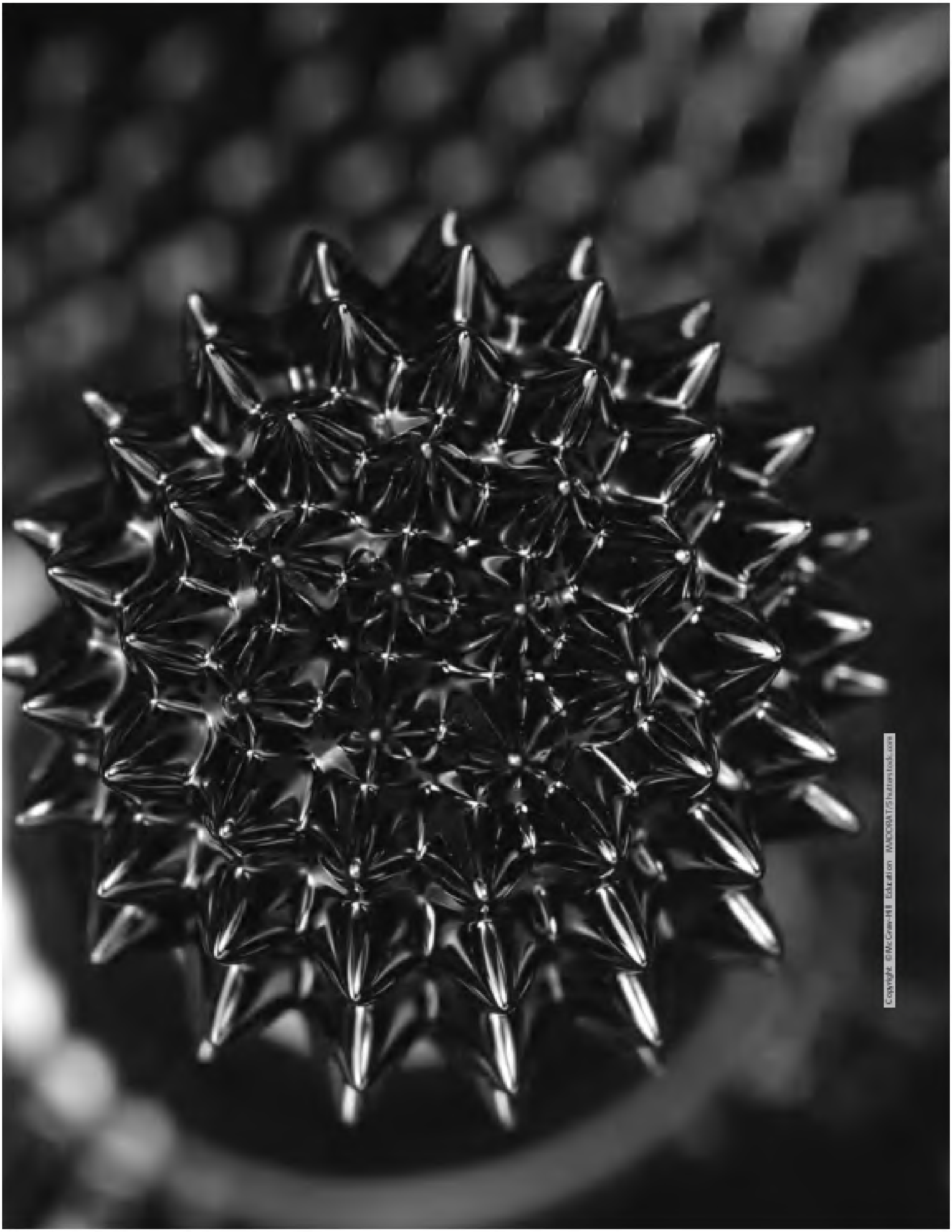
What does Boyle's law have to do with breathing?

You take a breath about 20 times per minute. How do pressure and volume change in your lungs as you breathe?

Data and Observations The spongy, elastic tissue that makes up your lungs allows them to expand and contract in response to movement of the diaphragm, a strong muscle beneath the lungs. As your diaphragm moves downward, increasing lung volume, you inhale. As your diaphragm moves upward, decreasing lung volume, you exhale.

CER Analyze and Interpret Data

1. **Claim, Evidence, Reasoning** Apply Boyle's law to explain why air enters your lungs when you inhale and leaves when you exhale.
2. **Claim, Evidence, Reasoning** Explain what happens inside the lungs when a blow to the abdomen knocks the wind out of a person. Use Boyle's law to determine your answer.
3. **Claim, Evidence, Reasoning** Parts of the lungs lose elasticity and become enlarged when a person has emphysema. From what you know about Boyle's law, why does this condition affect breathing?
4. **Claim, Evidence, Reasoning** Explain why beginning scuba divers are taught never to hold their breath while ascending from deep water.



MIXTURES AND SOLUTIONS

ENCOUNTER THE PHENOMENON

How it is possible for a liquid to hold this shape?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about how a liquid could hold this shape.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.

LESSON 1: Explore & Explain: Heterogeneous Mixtures

LESSON 1: Explore & Explain: Homogeneous Mixtures

LESSON 1

TYPES OF MIXTURES

FOCUS QUESTION

Do all mixtures have a uniform composition?

Heterogeneous Mixtures

Recall that a mixture is a combination of two or more pure substances in which each pure substance retains its individual chemical properties. Homogeneous mixtures are called solutions, where the particles are evenly distributed or blended. Particles in a solution are very small, and occur on an atomic-scale. Heterogeneous mixtures, however, do not blend smoothly throughout, and the individual substances remain distinct. Two types of heterogeneous mixtures are suspensions and colloids.

Suspensions

A **suspension** is a mixture containing particles that are large enough that, if left undisturbed, will settle out over time due to gravity. The muddy water shown in **Figure 1** is a suspension of clay and silt particles. Pouring a liquid suspension through a filter will also separate out the suspended particles.

Figure 1 A suspension can be separated by allowing it to sit for a period of time. A liquid suspension can also be separated by pouring it through a filter.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Inquiry into Chemistry: Make the Water Clean

Carry out an investigation at a small-scale to understand the interactions of water molecules.

Review the News

Obtain information from a current news story about types of mixtures. Evaluate your source and communicate your findings to your class.

Some suspensions separate into a solidlike mixture on the bottom of a container with water on the top. When solidlike mixtures are stirred or agitated, they flow like liquids. Substances that behave in this way are called thixotropic suspensions. For example, toothpaste is thixotropic—it acts as a liquid when it is squeezed from the tube but as a solid when it sits on your brush. Some paints are thixotropic—you can stir them but they don't flow down the stirring stick or brush when you hold them up. Builders in earthquake zones must be aware that some clays are thixotropic. Thixotropic clays form liquids in response to the agitation of an earthquake, which in turn causes structures built on them to collapse.

Colloids

A heterogeneous mixture of intermediate-sized particles, between atomic scale solution particles and larger suspension particles, is called a **colloid.** Colloid particles are between 1 nm and 1000 nm in diameter and do not settle out of the mixture. Milk is a colloid. The components of homogenized milk cannot be separated by settling or by filtration.

The most abundant substance in a colloid is the dispersion medium. Colloids are categorized according to the phases of their dispersed particles and dispersing mediums. Milk is a colloidal emulsion because liquid particles are dispersed in a liquid medium. Other types of colloids are described in **Table 1**.

Brownian motion Why do particles in a colloid stay suspended instead of settling out? One reason is that the dispersed particles of liquid colloids make jerky, random movements. This erratic movement of colloid particles is called **Brownian motion** and is used to describe the motion of particles when observed under microscopic conditions. It was first observed by, and later named for, the Scottish botanist Robert Brown (1773–1858), who noticed the random movements of pollen grains dispersed in water while observing them under a microscope. Brownian motion results from collisions of particles of the dispersion medium with the dispersed particles. These collisions prevent the colloid particles from settling out of the mixture and, over time, the particles will spread evenly throughout the medium.

Table 1 Types of Colloids

Category	Example	Dispersed Particles	Dispersing Medium
Solid sol	colored gems	solid	solid
Sol	blood, gelatin	solid	liquid
Solid emulsion	butter, cheese	liquid	solid
Emulsion	milk, mayonnaise	liquid	liquid
Solid foam	marshmallow, soaps that float	gas	solid
Foam	whipped cream, beaten egg white	gas	liquid
Solid aerosol	smoke, dust in air	solid	gas
Liquid aerosol	spray deodorant, fog, clouds	liquid	gas

Electrostatic layering The dispersed particles in a colloid often have polar or charged atomic groups on their surfaces. These groups are another factor that prevents particles from settling out of a colloid. The charged or polar areas on the surfaces of the particles attract the positively or negatively charged areas of the dispersing-medium particles. This results in the formation of electrostatic layers around the particles. The layers repel each other when the dispersed particles collide; thus, the particles remain suspended in the colloid.

Figure 2 This milk has been heated and acid has been added to it. As a result, the colloid is destroyed and particles in the milk clump together, forming curds.

If you interfere with the electrostatic layering, colloid particles will settle out of the mixture. For example, if you stir an electrolyte into a colloid, the dispersed particles clump together, destroying the colloid. Heating also destroys a colloid because it gives colliding particles enough kinetic energy to overcome the electrostatic forces and settle out. **Figure 2** shows how heating and adding acid, an electrolyte, to milk destroys the colloid. Recipes for homemade cheese involve heating milk and adding lemon juice. Once the colloid separates, the mixture is passed through a cloth to remove the liquid, called the whey. The solids, the curds, can then be pressed to make cheese. This is a good example of how the structure and interactions of matter at a bulk scale are determined by electrical forces between atoms.

Get It?

Identify a property of the ferrofluid in the photograph at the beginning of the module that would distinguish it as a colloid.

Tyndall effect Concentrated colloids are often cloudy or opaque. Dilute colloids sometimes appear as clear as solutions. Dilute colloids appear to be homogeneous solutions because their dispersed particles are so small. However, dispersed colloid particles scatter light, a phenomenon known as the **Tyndall effect.** In **Figure 3**, a beam of light is shone through two unknown mixtures. You can observe that dispersed colloid particles scatter the light, unlike particles in the solution. Suspensions also exhibit the Tyndall effect, but solutions never exhibit the Tyndall effect. You have observed the Tyndall effect if you have observed rays of sunlight passing through smoke-filled air, or viewed lights through fog. The Tyndall effect can be used to determine the amount of colloid particles in suspension.

Figure 3 Particles in a colloid scatter light, unlike particles in a solution. Called the Tyndall effect, the beam of light is visible in the colloid because of light scattering.

Determine *which mixture is a colloid.*

Table 2 Types and Examples of Solutions

Type of Solution	Example	Solvent	Solute
Gas	air	nitrogen (gas)	oxygen (gas)
Liquid	carbonated water	water (liquid)	carbon dioxide (gas)
	ocean water	water (liquid)	oxygen gas (gas)
	antifreeze	water (liquid)	ethylene glycol (liquid)
	vinegar	water (liquid)	acetic acid (liquid)
	ocean water	water (liquid)	sodium chloride (solid)
Solid	dental amalgam	silver (solid)	mercury (liquid)
	steel	iron (solid)	carbon (solid)

Homogeneous Mixtures

Cell solutions, ocean water, and steel might appear dissimilar, but they share certain characteristics. You learned earlier that solutions are homogeneous mixtures that contain two or more substances called the solute and the solvent. The solute is the substance that dissolves. The solvent is the dissolving medium. When you look at a solution, it is not possible to distinguish the solute from the solvent.

Types of solutions

A solution might exist as a gas, a liquid, or a solid, depending on the state of its solvent, as shown in **Table 2.** Air is a gaseous solution, and its solvent is nitrogen gas. Braces that you wear on your teeth might be made of nitinol, which is a solid solution of titanium in nickel. Another application of nitinol is shown in **Figure 4.** Most solutions, however, are liquids. You read previously that reactions can take place in aqueous solutions, solutions in which the solvent is water. Water is the most common solvent among liquid solutions.

Just as solutions can exist in different forms, the solutes in the solutions can be gases, liquids, or solids, also shown in **Table 2.** Many solutions, such as ocean water, can contain more than one solute.

Figure 4 This structure is a stent, designed to be inserted into an artery or vein to provide support and allow blood to flow freely. It is made of nitinol, a solid solution of titanium in nickel that has high elasticity.

Forming solutions

Some combinations of substances readily form solutions, and others do not. A substance that dissolves in a solvent is said to be **soluble** in that solvent. For example, sugar is soluble in water—a fact you might have learned by dissolving sugar in flavored water to make a sweetened beverage, such as tea or lemonade. Two liquids that are soluble in each other in any proportion are said to be **miscible.** For example, the antifreeze liquid that is shown in **Figure 5** is made up primarily of ethylene glycol and water. These two liquids are said to be miscible in one another.

A substance that does not dissolve in a solvent is said to be **insoluble** in that solvent. Sand is insoluble in water. The liquids in a bottle of oil and vinegar separate shortly after they are mixed. Oil is insoluble in vinegar. Two liquids that can be mixed together but separate shortly after are said to be **immiscible.**

Figure 5 This antifreeze liquid is a solution of water and ethylene glycol. Green dye makes it easy to identify.

Check Your Progress

Summary

- The individual substances in a heterogeneous mixture remain distinct.
- Two types of heterogeneous mixtures are suspensions and colloids.
- Brownian motion is the erratic movement of colloid particles on a microscopic level.
- Colloid particles exhibit the Tyndall affect.
- A solution can exist as a gas, a liquid, or a solid, depending on the solvent.
- Solutes in a solution can be gases, liquids, or solids.

Demonstrate Understanding

1. **Explain** Use the properties of seawater to describe the characteristics of mixtures.
2. **Distinguish** between suspensions and colloids.
3. **Identify** the various types of solutions. Describe the characteristics of each type of solution.
4. **Explain** Use the Tyndall effect to explain why it is more difficult to drive through fog using high beams than using low beams.
5. **Describe** the different types of colloids.
6. **Explain** how the electrical forces between atomic groups affect the dispersion of colloid particles.
7. **Summarize** What causes Brownian motion?
8. **Compare and Contrast** Make a table that compares the properties of suspensions, colloids, and solutions.

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LESSON 2

SOLUTION CONCENTRATION

FOCUS QUESTION

How can you describe the concentration of a solution?

Expressing Concentration

The **concentration** of a solution is a measure of how much solute is dissolved in a specific amount of solvent or solution. Visual observations can be used to describe solutions qualitatively by using the words concentrated or dilute. Notice the containers of tea in **Figure 6.** One of the tea solutions is more concentrated than the other. In general, a concentrated solution contains a large amount of solute. The darker tea has more tea particles than the lighter tea. Conversely, a dilute solution contains a small amount of solute. The lighter tea in **Figure 6** is dilute and contains fewer tea particles than the darker tea.

Although qualitative descriptions of concentration can be useful, solutions are more often described quantitatively. Some commonly used quantitative descriptions are percent by mass, percent by volume, molarity, and molality. These descriptions express concentration as a ratio of measured amounts of solute and solvent or solution.

Figure 6 The strength of the tea corresponds to its concentration. The lighter tea is less concentrated than the darker tea.

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COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Review the News

Obtain information from a current news story about the importance of solution concentration. Evaluate your source and communicate your findings to your class.

Table 3 Concentration Ratios

Concentration Description	Ratio
Percent by mass	$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
Percent by volume	$\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$
Molarity	$\frac{\text{moles of solute}}{\text{liter of solution}}$
Molality	$\frac{\text{moles of solute}}{\text{kilogram of solvent}}$
Mole fraction	$\frac{\text{mole of solute}}{\text{moles of solute + moles of solvent}}$

Table 3 shows ratios for common quantitative descriptions. Which quantitative description should be used? The description used depends on the type of solution analyzed and the reason for describing it. For example, a chemist working with a reaction in an aqueous solution most likely refers to the molarity of the solution because he or she needs to know the number of particles involved in the reaction.

Percent by mass

The percent by mass is the ratio of the solute's mass to the solution's mass expressed as a percent. The mass of the solution is the sum of the masses of the solute and solvent.

Percent by Mass

$$\text{percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

Percent by mass equals the mass of the solute divided by the mass of the whole solution, multiplied by 100.

EXAMPLE Problem 1

CALCULATE PERCENT BY MASS In order to maintain a sodium chloride (NaCl) concentration similar to ocean water, an aquarium must contain 3.6 g NaCl per 100.0 g of water. What is the percent by mass of NaCl in the solution?

1 ANALYZE THE PROBLEM

You are given the amount of sodium chloride dissolved in 100.0 g of water. The percent by mass of a solute is the ratio of the solute's mass to the solution's mass, which is the sum of the masses of the solute and the solvent.

Known

mass of solute = 3.6 g NaCl

mass of solvent = 100.0 g H_2O

Unknown

percent by mass = ?

EXAMPLE Problem 1 (continued)

2 SOLVE FOR THE UNKNOWN

Find the mass of the solution.

mass of solution = grams of solute + grams of solvent

mass of solution = 3.6 g + 100.0 g = 103.6 g — Substitute mass of solute = 3.6 g, and mass of solvent = 100.0 g.

Calculate the percent by mass.

$$\text{percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

State the equation for percent by mass.

$$\text{percent by mass} = \frac{3.6\ \text{g}}{103.6\ \text{g}} \times 100 = 3.5\%$$

Substitute mass of solute = 3.6 g, and mass of solution = 103.6 g.

3 EVALUATE THE ANSWER

Because only a small mass of sodium chloride is dissolved per 100.0 g of water, the percent by mass should be a small value, which it is. The mass of sodium chloride was given with two significant figures; therefore, the answer is also expressed with two significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

9. What is the percent by mass of $NaHCO_3$ in a solution containing 20.0 g of $NaHCO_3$ dissolved in 600.0 mL of H_2O?
10. You have 1500.0 g of a bleach solution. The percent by mass of the solute sodium hypochlorite (NaOCl) is 3.62%. How many grams of NaOCl are in the solution?
11. In Question 10, how many grams of solvent are in the solution?
12. **CHALLENGE** The percent by mass of calcium chloride in a solution is found to be 2.65%. If 50.0 g of calcium chloride is used, what is the mass of the solution?

Percent by volume

Percent by volume usually describes solutions in which both solute and solvent are liquids. The percent by volume is the ratio of the volume of the solute to the volume of the solution, expressed as a percent. The volume of the solution is the sum of the volumes of the solute and the solvent. Calculations of percent by volume are similar to those involving percent by mass.

Percent by Volume

$$\text{percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

Percent by volume equals the volume of solute divided by the volume of the solution, multiplied by 100.

Biodiesel, shown in **Figure 7**, is an alternative fuel that is produced from renewable resources. Biodiesel can be used in diesel engines with little or no modifications. Biodiesel is simple to use, biodegradable, nontoxic, and it does not contain some of the pollutants found in regular gasoline.

Figure 7 B20 is 20% by volume biodiesel and 80% by volume petroleum diesel. Biodiesel is a clean-burning alternative fuel that can be produced from renewable resources, such as vegetable oil.

PRACTICE Problems ADDITIONAL PRACTICE

13. What is the percent by volume of ethanol in a solution that contains 35 mL of ethanol dissolved in 155 mL of water?
14. What is the percent by volume of isopropyl alcohol in a solution that contains 24 mL of isopropyl alcohol in 1.1 L of water?
15. **CHALLENGE** If 18 mL of methanol is used to make an aqueous solution that is 15% methanol by volume, how many milliliters of solution is produced?

Molarity

Percent by volume and percent by mass are only two of the commonly used ways to quantitatively describe the concentrations of liquid solutions. One of the most common units of solution concentration is molarity. **Molarity** (*M*) is the number of moles of solute dissolved per liter of solution. Molarity is also known as molar concentration, and the unit *M* is read as molar. A liter of solution containing 1 mol of solute is a 1*M* solution, which is read as a one-molar solution. A liter of solution containing 0.1 mol of solute is a 0.1*M* solution. To calculate a solution's molarity, you must know the volume of the solution in liters and the amount of dissolved solute in moles.

Molarity

$$\text{molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

The molarity of a solution equals the moles of solute divided by the liters of solution.

Molarity is the measure of concentration that you will likely use most in your studies of chemistry. The molarity of a solution tells you the amount of solute present in a solution, and this is the information that is important to know when you are planning and analyzing reactions in solution. Knowing the amounts of different solutes in solution allows you to predict how they will react based on a balanced chemical equation.

When performing calculations involving concentrations, you can think of molarity as a conversion factor that you can use to convert between volume of solution and amount of solute in moles. (Note that sometimes you will see molarity expressed using units of mol/L instead of *M*.)

Get It?

Determine What is the molar concentration of a liter solution with 0.5 mol of solute?

STEM CAREER Connection

Chemical Technician

If you like conducting experiments and working with instruments, you can get an associate's degree at a community college and work in a laboratory or production facility where new products, such as medicines, are manufactured. You would participate in researching, producing, and testing chemical products.

When calculating the molarity of a solution, keep in mind that the unit *M* is equivalent to moles per liter (mol/L). Remembering this will help you remember to convert volumes given in other units, such as milliliters, to liters. Write the units mol/L as *M* as a final step.

EXAMPLE Problem 2

CALCULATING MOLARITY A 100.5-mL intravenous (IV) solution contains 5.10 g of glucose ($C_6H_{12}O_6$). What is the molarity of this solution? The molar mass of glucose is 180.16 g/mol.

1 ANALYZE THE PROBLEM

You are given the mass of glucose dissolved in a volume of water. The molarity of the solution is the ratio of moles of solute per liter of solution.

Known

mass of solute = 5.10 g $C_6H_{12}O_6$

molar mass of $C_6H_{12}O_6$ = 180.16 g/mol

volume of solution = 100.5 mL

Unknown

solution concentration = ? *M*

2 SOLVE FOR THE UNKNOWN

Calculate the number of moles of $C_6H_{12}O_6$.

$$(5.10\ \cancel{g\ C_6H_{12}O_6})\left(\frac{1\ mol\ C_6H_{12}O_6}{180.16\ \cancel{g\ C_6H_{12}O_6}}\right)$$

Multiply grams of $C_6H_{12}O_6$ by the molar mass of $C_6H_{12}O_6$.

$$= 0.0283\ mol\ C_6H_{12}O_6$$

Convert the volume of H_2O to liters.

$$(100.5\ \cancel{mL}\ solution)\left(\frac{1L}{1000\ \cancel{mL}}\right) = 0.1005\ L\ solution$$

Use the conversion factor 1 L/1000 mL.

Solve for the molarity.

$$M = \frac{\text{moles of solute}}{\text{liters of solutions}}$$

State the molarity equation.

$$M = \left(\frac{0.0283\ mol\ C_6H_{12}O_6}{0.1005\ L\ solution}\right)$$

Substitute moles of $C_6H_{12}O_6$ = 0.0283 and volume of solution = liters of solution = 0.1005 L.

$$M = \left(\frac{0.0282\ mol\ C_6H_{12}O_6}{1\ L\ solution}\right) = 0.282M$$

Divide numbers and units.

3 EVALUATE THE ANSWER

The molarity value will be small because only a small mass of glucose was dissolved in the solution. The mass of glucose used in the problem has three significant figures; therefore, the value of the molarity also has three significant figures.

PRACTICE Problems

ADDITIONAL PRACTICE

16. What is the molarity of an aqueous solution containing 40.0 g of glucose ($C_6H_{12}O_6$) in 1.5 L of solution?
17. Calculate the molarity of 1.60 L of a solution containing 1.55 g of dissolved KBr.
18. What is the molarity of a bleach solution containing 9.5 g of NaOCl per liter of bleach?
19. **CHALLENGE** How much calcium hydroxide ($Ca(OH)_2$), in grams, is needed to produce 1.5 L of a 0.25*M* solution?

Step 1: The mass of the solute is measured.

Step 2: The solute is placed in a volumetric flask of the correct volume.

Step 3: Distilled water is added to the flask to bring the solution level up to the calibration mark.

Figure 8 Accurately preparing a solution of copper(II) sulfate involves several steps.
Explain *why you cannot add 375 g of copper(II) sulfate directly to 1 L of water to make a 1.5M solution.*

Preparing molar solutions Now that you know how to calculate the molarity of a solution, how do you think you would prepare 1 L of a 1.50*M* aqueous solution of copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$)? A 1.50*M* aqueous solution of $CuSO_4 \cdot 5H_2O$ contains 1.50 mol of $CuSO_4 \cdot 5H_2O$ dissolved in 1 L of solution. The molar mass of $CuSO_4 \cdot 5H_2O$ is about 249.70 g. Thus, 1.50 mol of $CuSO_4 \cdot 5H_2O$ has a mass of 375 g, an amount that you can measure on a balance. You cannot simply add 375 g of $CuSO_4 \cdot 5H_2O$ to 1 L of water to make the 1.50*M* solution. Like all substances, $CuSO_4 \cdot 5H_2O$ takes up space and will add volume to the solution. Therefore, you must use slightly less than 1 L of water to make 1 L of solution, as shown in **Figure 8**.

Suppose you needed only 100 mL of a 1.50*M* $CuSO_4 \cdot 5H_2O$ solution for an experiment. As calculated above, 1 L of this solution contains 375 g of $CuSO_4 \cdot 5H_2O$. This relationship can be used to calculate how much solute to use to make 100 mL of the solution.

$$100\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} \times \frac{375\ \text{g}\ CuSO_4 \cdot 5H_2O}{1\ \text{L solution}} = 37.5\ \text{g}\ CuSO_4 \cdot 5H_2O$$

You would need 37.5 g of $CuSO_4 \cdot 5H_2O$ to make 100 mL of a 1.50*M* solution.

PRACTICE Problems

ADDITIONAL PRACTICE

20. How many grams of $CaCl_2$ would be dissolved in 1.0 L of a 0.10*M* solution of $CaCl_2$?
21. How many grams of $CaCl_2$ should be dissolved in 500.0 mL of water to make a 0.20*M* solution of $CaCl_2$?
22. What mass of NaOH is in 250 mL of a 3.0*M* NaOH solution?
23. **CHALLENGE** What volume of ethanol (C_2H_5OH) is in 100.0 mL of 0.15*M* solution? The density of ethanol is 0.7893 g/mL.

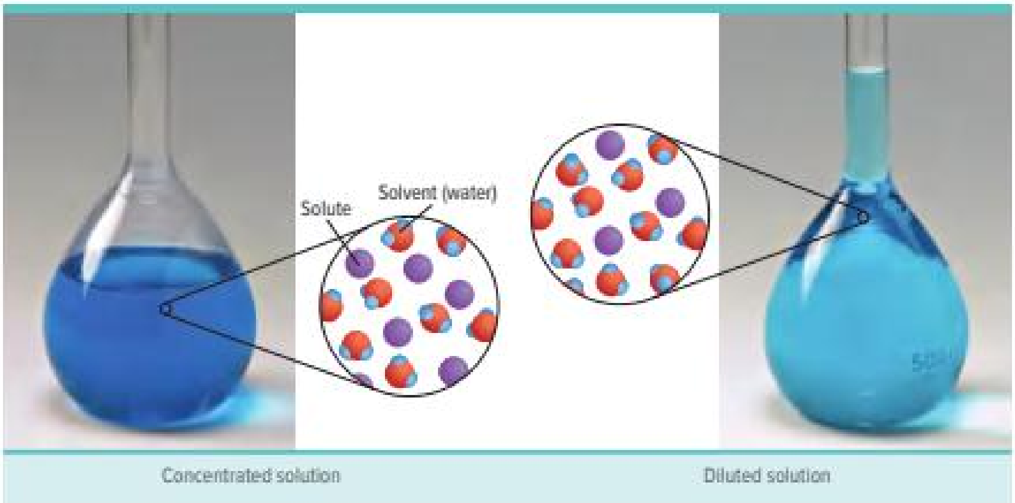


Figure 9 A concentrated solution can be diluted by adding solvent. The number of moles of solute does not change when a concentrated solution is diluted.

Diluting molar solutions In the laboratory, you might use concentrated solutions of standard molarities, called stock solutions. For example, concentrated hydrochloric acid (HCl) is 12*M*. Recall that a concentrated solution has a large amount of solute. You can prepare a less-concentrated solution by diluting the stock solution with additional solvent. When you add solvent, you increase the number of solvent particles among which the solute particles move, as shown in **Figure 9,** thereby decreasing the solution's concentration.

How do you determine the volume of stock solution you must dilute? You can rearrange the expression of molarity to solve for moles of solute.

$$\text{molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{moles of solute} = \text{molarity} \times \text{liters of solution}$$

Because the total number of moles of solute does not change during dilution,

$$\text{moles of solute in the stock solution} = \text{moles of solute after dilution.}$$

Substituting moles of solute with molarity times liters of solution, the relationship can be expressed in the dilution equation on the following page.

ACADEMIC VOCABULARY

diluted

solutions containing relatively low amounts of solute in a given volume of solvent

Adding too much water diluted the lemonade concentration.

Dilution Equation

$$M_1V_1 = M_2V_2$$

M represents molarity.
V represents volume.

For a given amount of solute, the product of the molarity and volume of the stock solution equals the product of the molarity and the volume of the dilute solution.

In the dilution equation, M_1 and V_1 represent the molarity and volume of the stock solution, and M_2 and V_2 represent the molarity and volume of the dilute solution. Before dilution, a concentrated solution contains a fairly high ratio of solute particles to solvent particles. After adding more solvent, however, the ratio of solute particles to solvent particles has decreased. However, because the total number of solute particles would not change during dilution, the total number of moles of solute would remain the same.

EXAMPLE Problem 3

DILUTING STOCK SOLUTIONS If you know the concentration and volume of the solution you want to prepare, you can calculate the volume of stock solution you will need. What volume, in milliliters, of 2.00*M* calcium chloride ($CaCl_2$) stock solution would you use to make 0.50 L of 0.300*M* calcium chloride solution?

1 ANALYZE THE PROBLEM

You are given the molarity of a stock solution of $CaCl_2$ and the volume and molarity of a dilute solution of $CaCl_2$. Use the relationship between molarities and volumes to find the volume, in liters, of the stock solution required. Then, convert the volume to milliliters.

Known	Unknown
$M_1 = 2.00M\ CaCl_2$	$V_1 = ?\ mL\ 2.00M\ CaCl_2$
$M_2 = 0.300M$	
$V_2 = 0.50\ L$	

2 SOLVE FOR THE UNKNOWN

Solve the molarity-volume relationship for the volume of the stock solution V_1.

$M_1V_1 = M_2V_2$ — State the dilution equation.

$V_1 = V_2\left(\frac{M_2}{M_1}\right)$ — Solve for V_1.

$V_1 = (0.50\ L)\left(\frac{0.300M}{2.00M}\right)$ — Substitute $M_1 = 2.00M$, $M_2 = 0.300M$, and $V_2 = 0.50\ L$.

$V_1 = (0.50\ L)\left(\frac{0.300\cancel{M}}{2.00\cancel{M}}\right) = 0.075\ L$ — Multiply and divide numbers and units.

$V_1 = (0.075\ \cancel{L})\left(\frac{1000\ mL}{1\ \cancel{L}}\right) = 75\ mL$ — Convert to milliliters using the conversion factor 1000 mL/1 L.

To make the dilution, measure out 75 mL of the stock solution and dilute it with enough water to make the final volume 0.50 L.

3 EVALUATE THE ANSWER

The volume V_1 was calculated, and then its value was converted to milliliters. This volume should be less than the final volume of the dilute solution, and it is. Of the given information, V_2 had the fewest number of significant figures, with two. Thus, the volume V_1 should also have two significant figures, and it does.

PRACTICE Problems ADDITIONAL PRACTICE

24. What volume of a 3.00*M* KI stock solution would you use to make 0.300 L of a 1.25*M* KI solution?
25. How many milliliters of a 5.0*M* H_2SO_4 stock solution would you need to prepare 100.0 mL of 0.25*M* H_2SO_4?
26. **CHALLENGE** If 0.50 L of 5.00*M* stock solution of HCl is diluted to make 2.0 L of solution, how much HCl, in grams, is in the solution?

Molality

The volume of a solution changes with temperature as it expands or contracts. This change in volume alters the molarity of the solution. Masses, however, do not change with temperature. It is sometimes more useful to describe solutions in terms of how many moles of solute are dissolved in a specific mass of solvent. Such a description is called **molality**—the ratio of the number of moles of solute dissolved in 1 kg of solvent. The unit *m* is read as molal. A solution containing 1 mol of solute per kilogram of solvent is a one-molal solution.

Molality

$$\text{molality } (m) = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

The molality of a solution equals the moles of solute divided by kilograms of solvent.

EXAMPLE Problem 4

CALCULATING MOLALITY In the lab, a student adds 4.5 g of sodium chloride (NaCl) to 100.0 g of water. Calculate the molality of the solution.

1 ANALYZE THE PROBLEM

You are given the mass of solute and solvent. Determine the number of moles of solute. Then, you can calculate the molality.

Known
mass of water (H_2O) = 100.0 g
mass of sodium chloride (NaCl) = 4.5 g

Unknown
m = ? mol/kg

2 SOLVE FOR THE UNKNOWN

$$4.5\ \cancel{\text{g NaCl}} \times \frac{1\ \text{mol NaCl}}{58.44\ \cancel{\text{g NaCl}}} = 0.077\ \text{mol NaCl}$$

Calculate the number of moles of solute.

$$100.0\ \cancel{\text{g } H_2O} \times \frac{1\ \text{kg } H_2O}{1000\ \cancel{\text{g } H_2O}} = 0.1000\ \text{kg } H_2O$$

Convert the mass of H_2O from grams to kilograms using the factor 1 kg/1000 g.

Substitute the known values into the expression for molality, and solve.

$$m = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

Write the equation for molality.

$$m = \frac{0.077\ \text{mol NaCl}}{0.1000\ \text{kg } H_2O} = 0.77\ \text{mol/kg}$$

Substitute moles of solute = 0.077 mol NaCl, kilograms of solvent = 0.1000 kg H_2O.

3 EVALUATE THE ANSWER

Because there was less than one-tenth of a mole of solute present in one-tenth of a kilogram of water, the molality should be less than one, and it is. The mass of sodium chloride was given with two significant figures; therefore, the molality is also expressed with two significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

27. What is the molality of a solution containing 10.0 g of Na_2SO_4 in 1000.0 g of water?

28. CHALLENGE How much $Ba(OH)_2$, in grams, is needed to make a 1.00*m* aqueous solution?

Mole fraction

If you know the number of moles of solute and solvent, you can also express the concentration of a solution as a **mole fraction**—the ratio of the number of moles of solute or solvent in solution to the total number of moles of solute and solvent, as shown in **Figure 10.** The symbol *X* is commonly used for mole fraction, with a subscript to indicate the solvent or solute.

Mole Fraction

$$X_A = \frac{n_A}{n_A + n_B} \qquad X_B = \frac{n_B}{n_A + n_B}$$

X_A and X_B represent the mole fractions of each substance.
n_A and n_B represent the number of moles of each substance.

A mole fraction equals the number of moles of solute or solvent in a solution divided by the total number of moles of solute and solvent.

For example, suppose a hydrochloric acid solution contains 36 g of HCl and 64 g of H_2O. To convert these masses to moles, you would use the molar masses as conversion factors.

$$n_{HCl} = 36\ \cancel{\text{g HCl}} \times \frac{1\ \text{mol HCl}}{36.5\ \cancel{\text{g HCl}}} = 0.99\ \text{mol HCl}$$

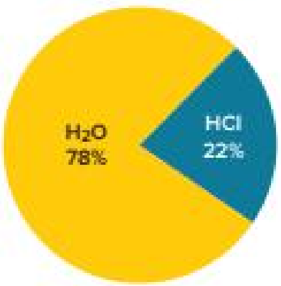
$$n_{H_2O} = 64\ \cancel{\text{g } H_2O} \times \frac{1\ \text{mol } H_2O}{18.0\ \cancel{\text{g } H_2O}} = 3.6\ \text{mol } H_2O$$

The mole fractions of HCl and water can be expressed as follows.

$$X_{HCl} = \frac{n_{HCl}}{n_{HCl} + n_{H_2O}} = \frac{0.99\ \text{mol HCl}}{0.99\ \text{mol HCl} + 3.6\ \text{mol } H_2O} = 0.22$$

$$X_{H_2O} = \frac{n_{H_2O}}{n_{HCl} + n_{H_2O}} = \frac{3.6\ \text{mol } H_2O}{0.99\ \text{mol HCl} + 3.6\ \text{mol } H_2O} = 0.78$$

Hydrochloric Acid in Aqueous Solution



$X_{HCl} + X_{H_2O} = 1.00$

$0.22 + 0.78 = 1.00$

Figure 10 The mole fraction expresses the number of moles of solute and solvent relative to the total number of moles of solution. Each mole fraction can be thought of as a percent. For example, the mole fraction of water (X_{H_2O}) is 0.78, which is equivalent to saying the solution contains 78% water (on a mole basis).

PRACTICE Problems ADDITIONAL PRACTICE

29. What is the mole fraction of NaOH in an aqueous solution that contains 22.8% NaOH by mass?

30. CHALLENGE If the mole fraction of sulfuric acid (H_2SO_4) in an aqueous solution is 0.325, what is the percent by mass of H_2SO_4?

BIOLOGY Connection People working in the field of medicine must have a good understanding of how to express, interpret, and calculate the concentration of solutions. When administering solutions containing anesthetics or other medications, a medical professional must know the concentration of that solution in order to determine the correct dose for a patient. Intravenous solutions, which are delivered to a patient directly into their bloodstream, must be have just the right concentration to be effective. Mistakes can result in ineffective or harmful treatments.

Interpreting concentrations of solutions listed on labels can sometimes be confusing, especially when percent concentrations are provided. The concentration of saline solutions for intravenous use is often given as a mass–volume percent, which tells you the number of grams of a solute per 100 mL of solution. For example, a 0.9% saline solution has 0.9 g of sodium chloride in 100 mL of solution. This measure is not a true percent because the units are g/mL and do not cancel out. Because the density of dilute aqueous solutions is very close to 1 g/mL, however, the mass–volume percent is often effectively the same as the percent by mass. You are not likely to work with the mass–volume percent very often in your study of chemistry, but you may see it when you look at labels that express concentration, such as labels on medications at a pharmacy. The potential confusion between percent by mass, percent by volume, and mass–volume percent shows why it is important to express clearly the method of measurement you are using.

Check Your Progress

Summary

- Concentrations can be measured qualitatively and quantitatively.
- Molarity is the number of moles of solute dissolved per liter of solution.
- Molality is the ratio of the number of moles of solute dissolved in 1 kg of solvent.
- The number of moles of solute does not change during a dilution.

Demonstrate Understanding

31. **Compare and contrast** five quantitative ways to describe the composition of solutions.
32. **Explain** the similarities and differences between a 1*M* solution of NaOH and a 1*m* solution of NaOH.
33. **Calculate** A can of chicken broth contains 450 mg of sodium chloride in 240.0 g of broth. What is the percent by mass of sodium chloride in the broth?
34. **Solve** How much ammonium chloride (NH_4Cl), in grams, is needed to produce 2.5 L of a 0.5*M* aqueous solution?
35. **Outline** the laboratory procedure for preparing a specific volume of a dilute solution from a concentrated stock solution.

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LESSON 3

FACTORS AFFECTING SOLVATION

FOCUS QUESTION

Why do some substances dissolve in water while others don't?

The Solvation Process

Why are some substances soluble, while others are not? To form a solution, solute particles must separate from one another and mix with solvent particles. Attractive forces between electric charges play an important role in this process. When a solid solute is placed in a solvent, the solvent particles completely surround the surface of the solute. If the attractive forces between the solvent and solute particles are greater than the attractive forces holding the solute particles together, the solvent particles pull the solute particles apart and surround them. These surrounded solute particles then move away from the solid solute and out into the solution. The process of surrounding solute particles with solvent particles to form a solution is called **solvation**, as shown in **Figure 11**. As ions dissolve in a solvent they spread out and become surrounded by solvent molecules. "Like dissolves like" is the general rule used to determine whether solvation will occur in a specific solvent. To determine whether a solvent and solute are alike, you must examine the bonding and polarity of the particles and the intermolecular forces among particles.

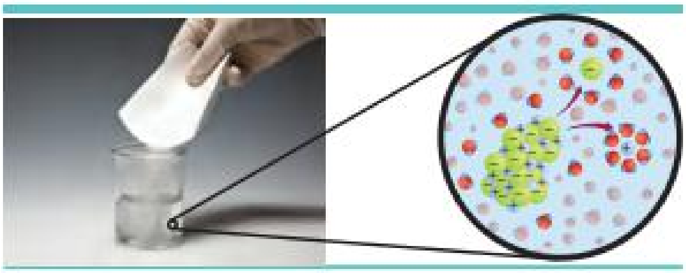


Figure 11 Salt begins to separate when it is dropped into water. The solute particles are pulled from the solid and surrounded by solvent particles.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Virtual Investigation: Salts and Solubility

Use a model to determine the cause and effect of adding a solute to a solvent in a chemical reaction.

ChemLAB: Investigate Factors Affecting Solubility

Carry out an investigation to determine what factors affect the rate of solution formation.

Solvation Process of NaCl

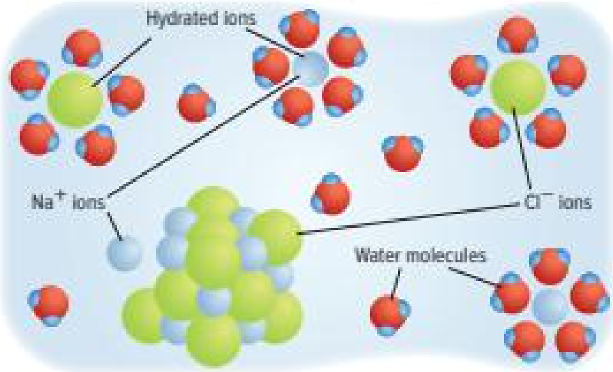


Figure 12 Sodium chloride dissolves in water as the water molecules surround the sodium and chloride ions. Note how the polar water molecules orient themselves differently around the positive and negative ion.

Aqueous solutions of ionic compounds

Recall that water molecules are polar molecules and are in constant motion, as described by the kinetic-molecular theory. When a crystal of an ionic compound, such as sodium chloride (NaCl), is placed in water, the water molecules collide with the surface of the crystal. The charged ends of the water molecules attract the positive sodium ions and negative chloride ions. This attraction between the dipoles and the ions is greater than the attraction among the ions in the crystal, so the ions break away from the surface. The water molecules surround the ions—in other words, the ions become solvated. The solvated ions move into the solution, as shown in **Figure 12**, exposing more ions on the surface of the crystal. Notice that the positively charged ends of the water molecules point inward when they surround the negatively charged chloride ions. The negatively charged ends of the water molecules point outward when surrounding the positively charged sodium ions. Solvation continues in this way until the entire crystal has dissolved. The attraction between ions and the polar molecules of water is the reason why many ionic compounds are soluble in water.

Not all ionic substances, however, are solvated by water molecules. Gypsum is insoluble in water because the attractive forces between the ions in gypsum are so strong that they cannot be overcome by the attractive forces of the water molecules. Calcium carbonate is another example of an ionic substance that is insoluble in water. This substance is found in nature as limestone, in eggshells, and in the shells of animals such as snails and clams. Although calcium carbonate does not dissolve in water, it does dissolve in acids. This is why rain with high acidity damages buildings and statues made of limestone.

Get It?

Describe how attraction between electric charges explains what happens when a soluble solid ionic compound is exposed to water molecules.

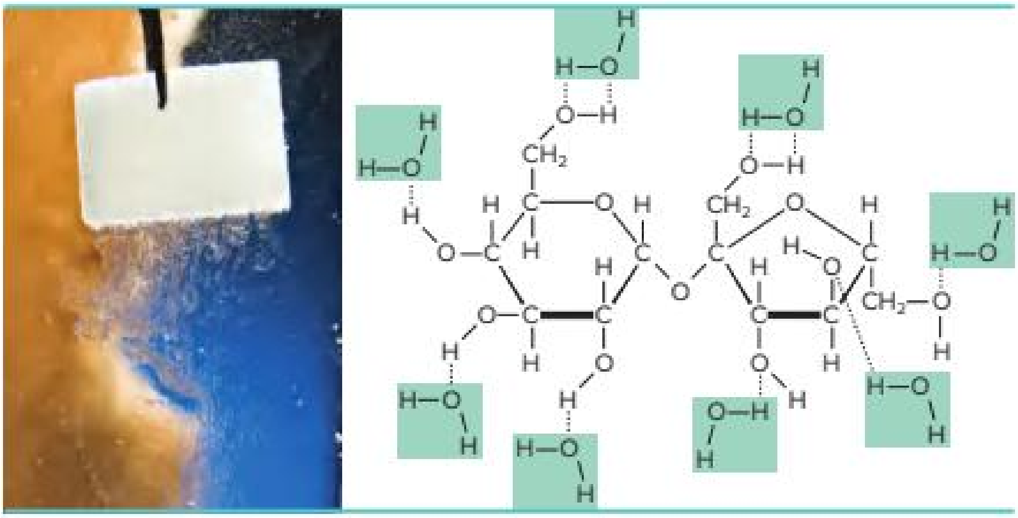


Figure 13 Sucrose molecules contain eight O—H bonds and are polar. Polar water molecules form hydrogen bonds with the O—H bonds, which pulls the sucrose into solution.

Aqueous solutions of molecular compounds

Water is also a good solvent for many molecular compounds. Table sugar is the molecular compound sucrose. As shown in **Figure 13**, sucrose molecules are polar and have several O–H bonds. As soon as the sugar crystals contact the water, water molecules collide with the outer surface of the crystal. Each O–H bond becomes a site for hydrogen bonding with water. The attractive forces among sucrose molecules are overcome by the attractive forces between polar water molecules and polar sucrose molecules. Sucrose molecules leave the crystal and become solvated by water molecules.

Oil is a substance made up primarily of carbon and hydrogen. It does not form a solution with water. There is little attraction between the polar water molecules and the nonpolar oil molecules. However, oil spills can be cleaned up with a nonpolar solvent because nonpolar solutes are more readily dissolved in nonpolar solvents.

Heat of solution

During the process of solvation, the solute must separate into particles. Solvent particles must also move apart in order to allow solute particles to come between them. Energy is required to overcome the attractive forces within the solute and within the solvent, so both steps are endothermic. When solute and solvent particles collide, the particles attract each other and energy is released. This step in the solvation process is exothermic. The overall energy change that occurs during the solution formation process is called the **heat of solution.**

Some solutions release energy as they form, whereas others absorb energy during formation. For example, after ammonium nitrate dissolves in water, its container feels cool. In contrast, after calcium chloride dissolves in water, its container feels warm.

Factors That Affect Solvation

Solvation occurs only when the solute and solvent particles collide with each other. There are three common ways to increase the collisions between solute and solvent particles and thus increase the rate at which the solute dissolves. These three common ways, shown in **Figure 14,** are agitation, increasing the surface area of the solute, and increasing the temperature of the solvent.

Agitation

Stirring or shaking—agitation of the mixture—moves dissolved solute particles away from the contact surfaces more quickly and thereby allows new collisions between solute and solvent particles to occur. Without agitation, the solvated solute particles move away from the contact areas slowly.

Surface area

Breaking the solute into small pieces increases its surface area. A greater surface area allows more collisions to occur between the solute particles and the solvent particles. This is why a teaspoon of granulated sugar dissolves more quickly than an equal amount of sugar in cube form.

Temperature

The rate of solvation is affected by temperature. For example, sugar dissolves more quickly in hot tea, shown in **Figure 14,** than it does in iced tea. Additionally, hotter solvents generally can dissolve more solid solute. Hot tea can hold more dissolved sugar than the iced tea.

Most solids act in the same way as sugar—as temperature increases, the rate of solvation also increases. Solvation of other substances, such as gases, decreases at higher temperatures. For example, a carbonated soft drink will lose its fizz (carbon dioxide) much faster when it is at room temperature than when it is cold.

Get It?
Explain how the rate of solvation can be increased.

A sugar cube in iced tea will dissolve slowly, but stirring will make the sugar cube dissolve more quickly.

Granulated sugar dissolves faster in iced tea than a sugar cube. Stirring speeds the process.

Granulated sugar dissolves very quickly in hot tea.

Figure 14 Agitation, surface area, and temperature affect the rate of solvation.

Solubility

Solubility is a property of a solute that describes its ability to dissolve in a solvent. Just as solvation can be understood at the particle level, so can solubility. When a solute comes in contact with water molecules and solvation is initiated, at first, the solute particles are carried away from the solute. As the number of dissolved particles increases, collisions between dissolved solute particles and the remaining solute also increases.

However, solvation is a reversible process, and some colliding solute particles rejoin the solid solute. As long as the solvation rate is greater than this crystallization rate, solvation continues. Eventually, the rate of solvation and crystallization may become equal, forming a state of dynamic equilibrium, as shown in **Figure 15,** as long as the temperature remains constant. During dynamic equilibrium, solute particles return to the surface of the solid solute at the same rate as they are leaving.

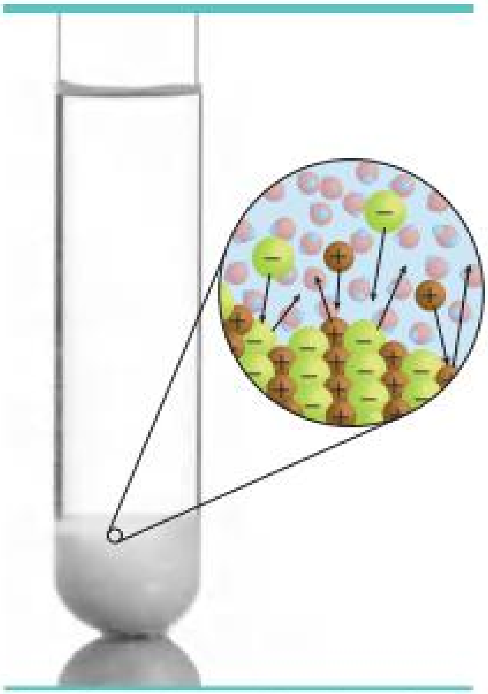


Figure 15 During dynamic equilibrium, the rate of solvation equals the rate of crystalization and the amount of dissolved solute remains unchanged.

Saturated solutions

Saturated solutions occur when no more solute can be dissolved. The solution contains the maximum amount of dissolved solute for a given amount of solvent at a specific temperature and pressure.

Unsaturated solutions

Unsaturated solutions contain less dissolved solute for a given temperature and pressure than a saturated solution. This means that more solute can be dissolved in an unsaturated solution.

CCC CROSSCUTTING CONCEPTS

Energy and Matter Create a flow chart that shows how overall energy can change during the solution formation process.

Temperature and supersaturated solutions

Solubility is affected by raising the temperature of the solvent because the kinetic energy of its particles is increased, resulting in more-frequent collisions and collisions with greater energy than those that occur at lower temperatures. The fact that many substances are more soluble at high temperatures is demonstrated in **Figure 16.** For example, calcium chloride ($CaCl_2$) has a solubility of about 64 g $CaCl_2$ per 100 g H_2O at 10°C. Increasing the temperature to approximately 27°C increases the solubility by almost 50%, to 100 g $CaCl_2$ per 100 g H_2O. For other substances, such as cerium sulfate, $Ce_2(SO_4)_3$, solubility initially decreases rapidly as temperature increases, but then levels off and remains constant.

The effect of temperature on solubility is also illustrated by the data in **Table 4.** Notice in **Table 4** that at 20°C, 203.9 g of sucrose ($C_{12}H_{22}O_{11}$) dissolves in 100 g of water. At 100°C, 487.2 g of sucrose dissolves in 100 g of water, a nearly 140% increase in solubility.

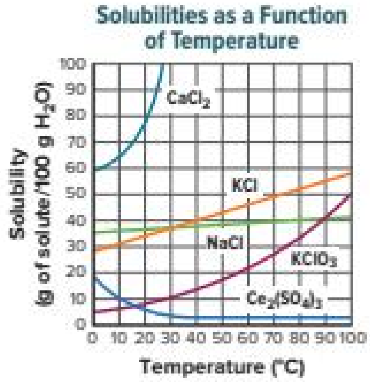


Figure 16 The solubilities of several substances as a function of temperature are shown in this graph.

Get It?

Interpret According to the graph in **Figure 16,** what is the solubility of NaCl at 80°C?

Table 4 **Solubilities of Solutes in Water at Various Temperatures**

Substance	Formula	Solubility (g/100 g H_2O)*			
		0°C	20°C	60°C	100°C
Aluminum sulfate	$Al_2(SO4)_3$	31.2	36.4	59.2	89.0
Barium hydroxide	$Ba(OH)_2$	1.67	3.89	20.94	—
Calcium hydroxide	$Ca(OH)_2$	0.189	0.173	0.121	0.076
Lithium sulfate	Li_2SO_4	36.1	34.8	32.6	—
Potassium chloride	KCl	28.0	34.2	45.8	56.3
Sodium chloride	NaCl	35.7	35.9	37.1	39.2
Silver nitrate	$AgNO_3$	122	216	440	733
Sucrose	$C_{12}H_{22}O_{11}$	179.2	203.9	287.3	487.2
Ammonia*	NH_3	1130	680	200	—
Carbon dioxide*	CO_2	1.713	0.878	0.359	—
Oxygen*	O_2	0.048	0.031	0.019	—

*L/1 L H_2O of gas at standard pressure (101 kPa)

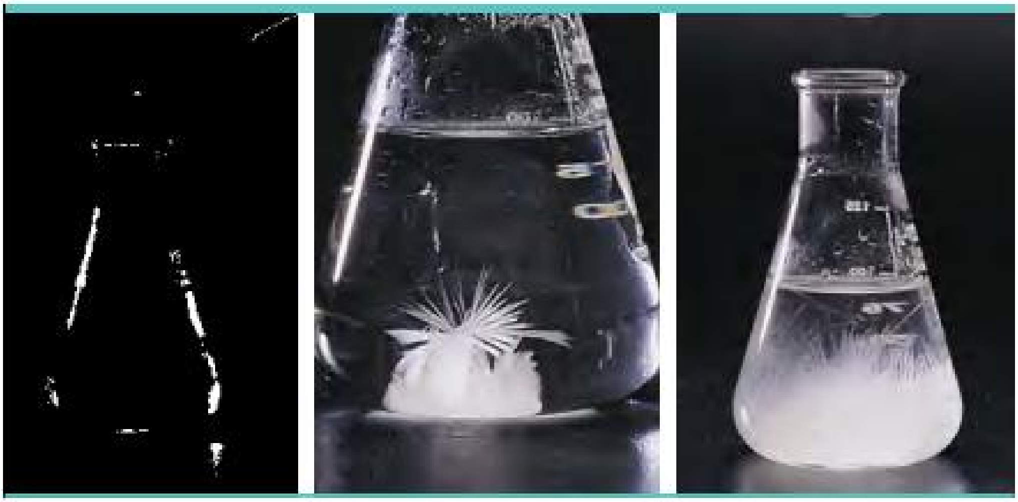


Figure 17 When a seed crystal is added to a supersaturated solution, the excess solute crystallizes out of the solution.

The fact that solubility changes with temperature and that some substances become more soluble with increasing temperature is the key to the formation of **supersaturated solutions.** A supersaturated solution contains more dissolved solute than a saturated solution at the same temperature. To make a supersaturated solution, a saturated solution is formed at a high temperature and then cooled slowly. The slow cooling allows the excess solute to remain dissolved in solution at the lower temperature.

Supersaturated solutions are unstable. If a tiny amount of solute, called a seed crystal, is added to a supersaturated solution, the excess solute crystalizes quickly, as illustrated in **Figure 17.** Crystallization can also occur if the inside of the container is scratched or the supersaturated solution undergoes a physical shock, such as stirring or tapping the container. Rock candy and mineral deposits at the edges of mineral springs, such as those shown in **Figure 18,** are both formed from supersaturated solutions.

Figure 18 Hot spring mineral deposits are an example of crystals that formed from supersaturated solutions.

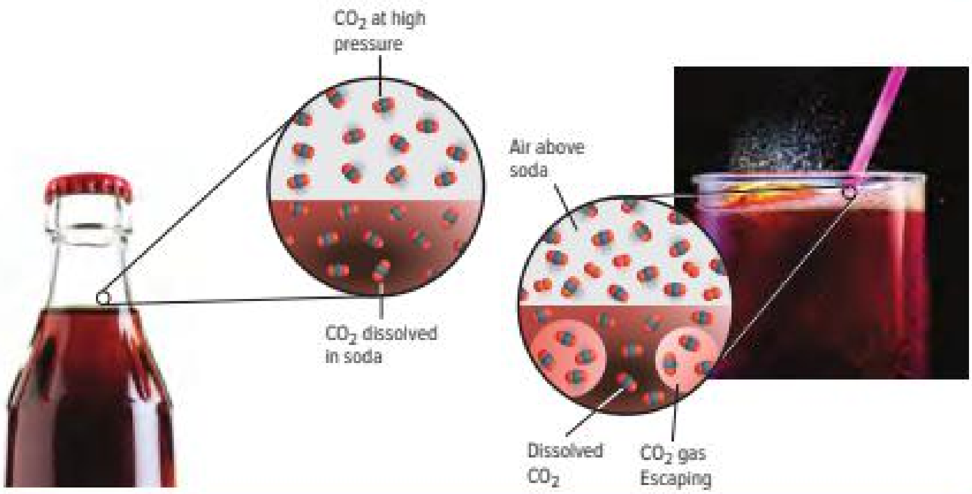
Solubility of gases

In **Table 4** you can see that the gases oxygen and carbon dioxide are less soluble at higher temperatures than at lower temperatures. This is a predictable trend for all gaseous solutes in liquid solvents. Can you explain why? Recall that the kinetic energy of gas particles allows them to escape from a solution more readily at higher temperatures. Thus, as a solution's temperature increases, the solubility of a gaseous solute decreases.

Pressure and Henry's Law

Pressure affects the solubility of gaseous solutes in solutions. The solubility of a gas in any solvent increases as its external pressure (the pressure above the solution) increases. Carbonated beverages depend on this fact. Carbonated beverages contain carbon dioxide gas dissolved in an aqueous solution. In bottling or canning the beverage, carbon dioxide is dissolved in the solution at a pressure higher than atmospheric pressure. When the beverage container is opened, the pressure of the carbon dioxide gas in the space above the liquid decreases. As a result, bubbles of carbon dioxide gas form in the solution, rise to the top, and escape. Unless the container is sealed, the process will continue until the solution loses almost all of its carbon dioxide gas and goes flat. The decreased solubility of the carbon dioxide contained in the beverage after it is opened can be described by Henry's law.

Henry's law states that at a given temperature, the solubility (*S*) of a gas in a liquid is directly proportional to the pressure (*P*) of the gas above the liquid. When the bottle of soda is closed, as illustrated in **Figure 19**, the pressure above the solution keeps carbon dioxide from escaping the solution.



The pressure above the solution of a closed soda bottle keeps excess carbon dioxide from escaping the solution.

The pressure above the solution decreases when the cap is removed, which decreases the solubility of the carbon dioxide.

Figure 19 Carbon dioxide (CO_2) is dissolved in soda. Some CO_2 also is found in the gas above the liquid.

Explain *Why does the carbon dioxide escape from the solution when the cap is removed?*

You can express this relationship in the following way.

Henry's Law

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

S represents solubility.
P represents pressure.

At a given temperature, the quotient of solubility of a gas and its pressure is constant.

Using Henry's law You will often use Henry's law to determine the solubility S_2 at a new pressure P_2, where P_2 is known. The basic rules of algebra can be used to solve Henry's law for any one specific variable. To solve for S_2, begin with the standard form of Henry's law.

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

Cross multiplying yields the following expression.

$$S_1P_2 = P_1S_2$$

Dividing both sides of the equation by P_1 yields the desired result—the equation solved for S_2.

$$\frac{S_1P_2}{P_1} = \frac{\cancel{P_1}S_2}{\cancel{P_1}}$$

$$S_2 = \frac{S_1P_2}{P_1}$$

EXAMPLE Problem 5

HENRY'S LAW If 0.85 g of a gas at 4.0 atm of pressure dissolves in 1.0 L of water at 25°C, how much will dissolve in 1.0 L of water at 1.0 atm of pressure and the same temperature?

1 ANALYZE THE PROBLEM

You are given the solubility of a gas at an initial pressure. The temperature of the gas remains constant as the pressure changes. Because decreasing pressure reduces a gas's solubility, less gas should dissolve at the lower pressure.

Known	Unknown
S_1 = 0.85 g/L	S_2 = ? g/L
P_1 = 4.0 atm	
P_2 = 1.0 atm	

SCIENCE USAGE v. COMMON USAGE

pressure

Science usage: the force exerted over an area

As carbon dioxide escapes the solution, the pressure in the closed bottle increases.

Common usage: the burden of physical or mental stress

There is a lot of pressure to do well on exams.

EXAMPLE Problem 5 (continued)

2 SOLVE FOR THE UNKNOWN

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$ State Henry's law.

$$S_2 = S_1\left(\frac{P_2}{P_1}\right)$$ Solve Henry's law for S_2.

$$S_2 = \left(\frac{0.85\ g}{1.0\ L}\right)\left(\frac{1.0\ \cancel{atm}}{4.0\ \cancel{atm}}\right) = 0.21\ g/L$$ Substitute S_1 = 0.85 g/L, P_1 = 4.0 atm, and P_2 = 1.0 atm. Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The solubility decreased as expected. The pressure on the solution was reduced from 4.0 atm to 1.0 atm, so the solubility should be reduced to one-fourth its original value, which it is. The unit g/L is a solubility unit, and there are two significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

36. If 0.55 g of a gas dissolves in 1.0 L of water at 20.0 kPa of pressure, how much will dissolve at 110.0 kPa of pressure?

37. A gas has a solubility of 0.66 g/L at 10.0 atm of pressure. What is the pressure on a 1.0-L sample that contains 1.5 g of gas?

38. CHALLENGE The solubility of a gas at 7.0 atm of pressure is 0.52 g/L. How many grams of the gas would be dissolved per 1.0 L if the pressure increased 40.0 percent?

Check Your Progress

Summary

- The process of solvation involves solute particles surrounded by solvent particles.
- Solutions can be saturated, unsaturated, or supersaturated.
- Henry's law states that at a given temperature, the solubility (*S*) of a gas in a liquid is directly proportional to the pressure (*P*) of the gas above the liquid.

Demonstrate Understanding

39. **Compare and contrast** saturated and unsaturated solutions.
40. **Define** *solubility*.
41. **Describe** the role electrical charges play in solvation.
42. **Explain** why a saturated solution containing solid crystals is an example of a dynamic equilibrium.
43. **Summarize** If a seed crystal is added to a supersaturated solution, how would you characterize the resulting solution?
44. **Make and Use Graphs** Use the information in **Table 4** to graph the solubilities of aluminum sulfate, lithium sulfate, and potassium chloride at 0°C, 20°C, 60°C, and 100°C. Which substance's solubility is most affected by increasing temperature?

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

LESSON 4

COLLIGATIVE PROPERTIES OF SOLUTIONS

FOCUS QUESTION

Why do we salt the roads when it's cold outside?

Electrolytes and Colligative Properties

Solutes affect some of the physical properties of their solvents. For example, if you add salt to water it results in a phenomenon called boiling point elevation, where the boiling point of water is slightly increased. Early researchers were puzzled to discover that the effects of a solute on a solvent in some cases, depended only on how many solute particles were in the solution, not on the specific solute dissolved. Physical properties of solutions that are affected by the number of particles but not by the identity of dissolved solute particles are called **colligative properties.** The word colligative means depending on the collection. Colligative properties include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.

Electrolytes in aqueous solution

Recall that the structure and interaction of matter at a bulk scale are determined by electrical forces within and between atoms. Ionic compounds are called electrolytes because they dissociate in water to form a solution that conducts electric current, as shown in **Figure 20.** Some molecular compounds ionize in water and are also electrolytes. By studying the small scale interactions of ions and molecules and understanding their electrical properties, we can predict which compounds form electrolytes in water.

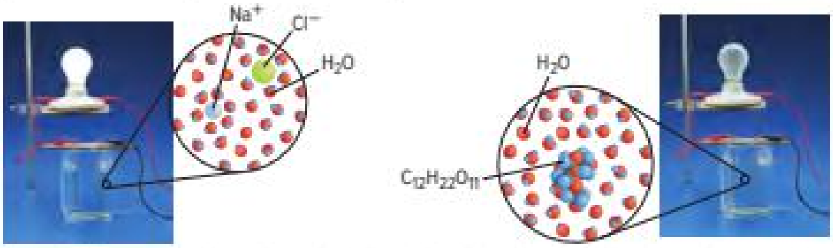


Figure 20 Sodium chloride conducts electricity well in solution because it is an electrolyte containing charged ions. Sucrose does not conduct electricity in solution because it is not an electrolyte.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Vapor Pressure Lowering

HS-PS1-3. Plan and conduct an investigation to gather evidence to compare the structure of substances at the bulk scale to infer the strength of electrical forces between particles.

Review the News

Obtain information from a current news story about applications of colligative properties. Evaluate your source and communicate your findings to your class.

Strong and weak electrolytes in aqueous solution

Electrolytes that produce many ions in a solution are called strong electrolytes; those that produce only a few ions in a solution are called weak electrolytes. Sodium chloride is a strong electrolyte. It dissociates in solution, producing Na^+ and Cl^- ions.

$$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$$

Dissolving 1 mol of NaCl in 1 kg of water would not yield a 1*m* solution of ions. Rather, there would be 2 mol of solute particles in solution—1 mol each of Na^+ and Cl^- ions.

Nonelectrolytes in aqueous solution

Many molecular compounds dissolve in solvents but do not ionize. Such solutions do not conduct an electric current, as shown in **Figure 20,** and the solutes are called nonelectrolytes. Sucrose is an example of a nonelectrolyte. A 1*m* sucrose solution contains only 1 mol of sucrose particles.

Get It?
Infer Predict which has the greater effect on colligative properties, sodium chloride or sucrose Explain your reasoning.

Vapor Pressure Lowering

You learned previously that vapor pressure is the pressure exerted in a closed container by the particles that have escaped the liquid's surface and entered the gaseous state. In a closed container at constant temperature and pressure, the solvent particles of a pure solvent reach a state of dynamic equilibrium, escaping and reentering the liquid state at the same rate.

Experiments show that adding a nonvolatile solute (one that has little tendency to become a gas) to a solvent lowers the solvent's vapor pressure. The particles that produce vapor pressure escape the liquid phase at its surface. When a solvent is pure, as shown in **Figure 21,** its particles occupy the entire surface area. When the solvent contains solute, as also shown in **Figure 21,** a mix of solute and solvent particles occupies the surface area. With fewer solvent particles at the surface, fewer particles enter the gaseous state, and the vapor pressure is lowered. The greater the number of solute particles in a solvent, the lower the resulting vapor pressure. Thus, **vapor pressure lowering** is due to the number of solute particles in solution and is a colligative property of solutions.

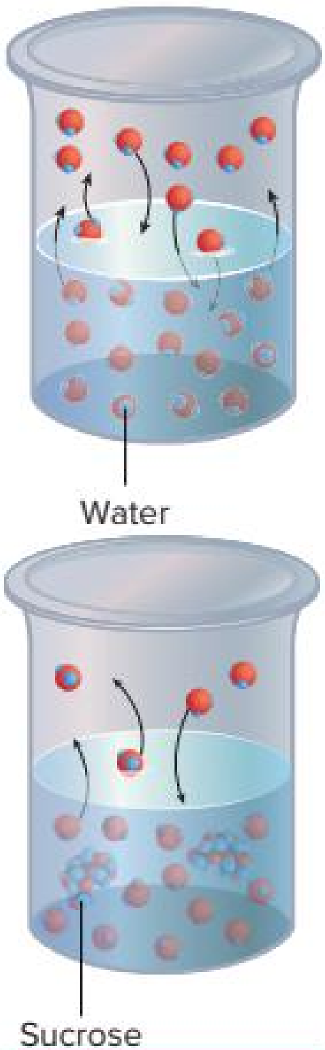


Figure 21 The vapor pressure of a pure solvent is greater than the vapor pressure of a nonvolatile solution.

You can predict the relative effect of a solute on vapor pressure based on whether the solute is an electrolyte or a nonelectrolyte. Recall that an electrolyte is a substance that dissociates into ions in solution and has the capacity to conduct electricity. For example, 1 mol each of dissociated nonelectrolyte molecules, glucose, sucrose, and ethanol, have the same relative effect on the vapor pressure. However, 1 mol each of the dissociated electrolyte molecules, sodium chloride (NaCl), and sodium sulfate (Na_2SO_4), have a greater effect on vapor pressure because they produce more ions in solution.

Boiling Point Elevation

Because a nonvolatile solute lowers a solvent's vapor pressure, it also affects the boiling point of the solvent. Recall that liquid in a pot on a stove boils when its vapor pressure equals the atmospheric pressure. When the temperature of a solution containing a nonvolatile solute is raised to the boiling point of the pure solvent, the resulting vapor pressure is still less than the atmospheric pressure and the solution will not boil. Thus, the solution must be heated to a higher temperature to supply the additional kinetic energy needed to raise the vapor pressure to atmospheric pressure. The temperature difference between a solution's boiling point and a pure solvent's boiling point is called the **boiling point elevation.**

For nonelectrolytes, the value of the boiling point elevation, which is symbolized ΔT_b, is directly proportional to the solution's molality.

Boiling Point Elevation

$$\Delta T_b = K_b m$$

ΔT_b represents the difference in boiling point elevations.
K_b represents the molal boiling elevation constant.
m represents molality.

The temperature difference is equal to the molal boiling point elevation constant multiplied by the solution's molality.

The molal boiling point elevation constant, K_b, is the difference in boiling points between a $1m$ nonvolatile, nonelectrolyte solution and a pure solvent. Boiling point elevation is expressed in units of °C/m and varies for different solvents. Values of K_b for several common solvents are found in **Table 5.** Note that water's K_b value is 0.512°C/m. This means that a $1m$ aqueous solution containing a nonvolatile, nonelectrolyte solute boils at 100.512°C—a temperature just 0.512°C higher than pure water's boiling point of 100.0°C.

Table 5 Molal Boiling Point Elevation Constants (K_b)

Solvent	Boiling Point (°C)	K_b (°C/m)
Water	100.0	0.512
Benzene	80.1	2.53
Carbon Tetrachloride	76.7	5.03
Ethanol	78.5	1.22
Chloroform	61.7	3.63

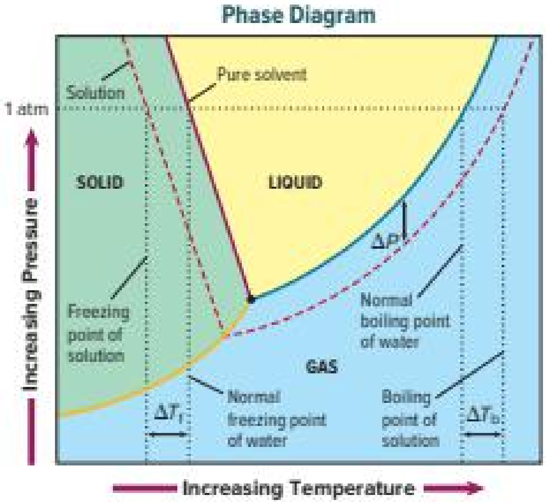


Figure 22 Temperature and pressure affect solid, liquid, and gas phases of a pure solvent (solid lines) and a solution (dashed line).

Like vapor pressure lowering, boiling point elevation is a colligative property. The value of the boiling point elevation can be predicted because it is directly proportional to the solution's solute molality; that is, the greater the number of solute particles in the solution, the greater the boiling point elevation. Because it is related to mole fraction, which involves the number of solute particles, molality is used as the concentration. Molality also uses mass of solvent rather than volume, and therefore is not affected by temperature changes. Examine **Figure 22** and notice that the curve for a solution lies below the curve for the pure solvent at any temperature.

Freezing Point Depression

At a solvent's freezing point temperature, the particles no longer have sufficient kinetic energy to overcome the interparticle attractive forces; the particles form into a more organized structure in the solid state. In a solution, the solute particles interfere with the attractive forces among the solvent particles. The solvent is prevented from entering the solid state at its normal freezing point.

The freezing point of a solution is always lower than that of a pure solvent. **Figure 22** shows the differences in boiling and melting points of pure water and an aqueous solution. Two common applications of freezing point depression are shown in **Figure 23.** Both of these applications use salt to lower the freezing point of an aqueous solution.

Figure 23 By adding salts to the ice on a road, the freezing point of the ice is lowered, which results in the ice melting. Adding salt to ice when making ice cream lowers the freezing point of the ice. More energy is absorbed from the surroundings to melt the ice, which freezes the ice cream.

A solution's **freezing point depression,** ΔT_f, is the difference in temperature between its freezing point and the freezing point of its pure solvent. Molal freezing point depression constants (K_f) for several solvents are shown in **Table 6.** For nonelectrolytes, the value of the freezing point depression is directly proportional to the solution's molality.

Table 6 Molal Freezing Point Depression Constants (K_f)

Solvent	Freezing Point (°C)	K_f (°C/m)
Water	0.0	1.86
Benzene	5.5	5.12
Carbon tetrachloride	−23.0	29.8
Ethanol	−114.1	1.99
Chloroform	−63.5	4.68

Freezing Point Depression

$$\Delta T_f = K_f m$$

ΔT_f represents temperature.
K_f is the freezing point depression constant.
m represents molality.

The temperature difference is equal to the freezing point depression constant multiplied by the solution's molality.

As with K_b values, K_f values are specific to their solvents. With water's K_f value of 1.86°C/m, a 1m aqueous solution containing a nonvolatile, nonelectrolyte solute freezes at −1.86°C rather than at pure water's freezing point of 0.0°C. Glycerol is a nonelectrolyte solute produced by many fish and insects to keep their blood from freezing during cold winters. Antifreeze and many de-icer solutions contain the nonelectrolyte solute ethylene glycol. Notice that the equations for boiling point elevation and freezing point depression specify the molality of a nonelectrolyte. For electrolytes, you must make sure to use the effective molality of the solution. Example Problem 6 illustrates this point.

EXAMPLE Problem 6

CHANGES IN BOILING AND FREEZING POINTS Sodium chloride (NaCl) is often used to prevent icy roads and to freeze ice cream. What are the boiling and freezing points of a 0.029m aqueous solution of sodium chloride?

1 ANALYZE THE PROBLEM

You are given the molality of an aqueous sodium chloride solution. First, calculate ΔT_b and ΔT_f based on the number of particles in the solution. Then, to determine the elevated boiling point and the depressed freezing point, add ΔT_b to the normal boiling point and subtract ΔT_f from the normal freezing point of water.

Known	Unknown
molality of solution = 0.029m	boiling point = ?°C
K_b = 0.512°C/m	freezing point = ?°C
K_f = 1.86°C/m	

EXAMPLE Problem 6 (continued)

2 SOLVE FOR THE UNKNOWN

Determine the molality of the particles.
particle molality = 2 × 0.029m = 0.058m
Determine ΔT_b and ΔT_f

$\Delta T_b = K_b m$
$\Delta T_f = K_f m$

State the boiling point elevation and freezing point depression formulas.

$\Delta T_b = (0.512°C/\cancel{m})(0.058\cancel{m}) = 0.030°C$
$\Delta T_f = (1.86°C/\cancel{m})(0.058\cancel{m}) = 0.11°C$

Substitute K_b = 0.512°C/m, K_f = 1.86°C/m, and m = 0.058m.

Determine the elevated boiling point and depressed freezing point of the solution.

boiling point = 100.000°C + 0.030°C
= **100.030°C**

freezing point = 0.00°C − 0.11°C
= **−0.11°C**

Add ΔT_b to the normal boiling point and subtract ΔT_f from the normal freezing point.

3 EVALUATE THE ANSWER

The boiling point is higher and the freezing point is lower, as expected. Because the molality of the solution has two significant figures, both ΔT_b and ΔT_f have two significant figures. Because the normal boiling point and freezing point are exact values, they do not affect the number of significant figures in the final answer.

Real-World Chemistry

Freezing Point Depression

SALTWATER FISH Maintaining the proper saline (salt) concentration is important to the health of saltwater fish in aquariums. In the ocean, the presence of salt in arctic areas keeps the water from freezing, allowing aquatic life to be sustained.

PRACTICE Problems ADDITIONAL PRACTICE

45. What are the boiling point and freezing point of a 0.625m aqueous solution of any nonvolatile, nonelectrolyte solute?

46. What are the boiling point and freezing point of a 0.40m solution of sucrose in ethanol?

47. CHALLENGE A 0.045m solution (consisting of a nonvolatile, nonelectrolyte solute) is experimentally found to have a freezing point depression of 0.080°C. What is the freezing point depression constant (K_f)? Which is most likely to be the solvent: water, ethanol, or chloroform?

Osmotic Pressure

BIOLOGY Connection Recall that diffusion is the mixing of gases or liquids resulting from their random motions. **Osmosis** is the diffusion of a solvent through a semipermeable membrane. Semipermeable membranes are barriers that allow some, but not all, particles to cross. Investigating how osmotic processes function in biological systems is possible by examining the different structures and properties found in living cells and plants. The membranes surrounding all living cells are semipermeable membranes. Osmosis plays an important role in many biological systems, such as the intake of water and distribution of nutrients by trees and other plants. Through the process of osmosis, water enters the roots of trees, creating pressure within the roots. This pressure helps push the tree sap up to the top of the tree. The movement of the water is driven by the difference in concentration between water that the roots absorb from the ground and the sap in the tree. The water is significantly less concentrated than the sap.

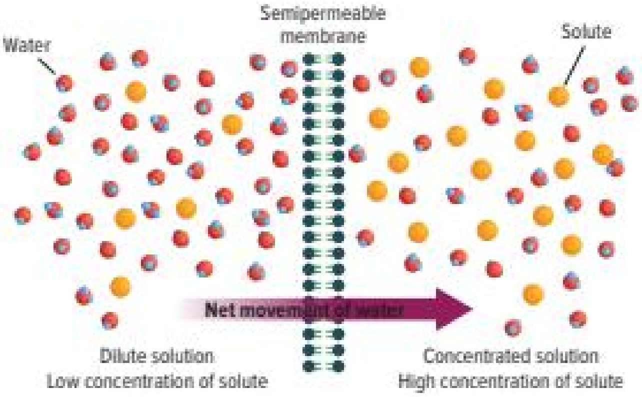


Figure 24 Due to osmosis, solvents diffuse from a lower solute concentration to a higher solute concentration through semipermeable membranes.

Examine a system in which a dilute solution is separated from a concentrated solution by a semipermeable membrane, illustrated in **Figure 24.** During osmosis, water molecules move in both directions across the membrane, but the solute molecules cannot cross it. Water molecules diffuse across the membrane from the dilute solution to the concentrated solution. The amount of additional pressure caused by the water molecules that moved into the concentrated solution is called the **osmotic pressure.** Osmotic pressure depends on the number of solute particles in a given volume of solution and is a colligative property of solutions.

Check Your Progress

Summary

- Nonvolatile solutes lower the vapor pressure of a solution.
- Boiling point elevation is directly related to the solution's molality.
- A solution's freezing point depression is always lower than that of the pure solvent.
- Osmotic pressure depends on the number of solute particles in a given volume.

Demonstrate Understanding

48. **Explain** the nature of colligative properties.
49. **Describe** four colligative properties of solutions.
50. **Describe** a cause and effect relationship on a particle scale that can be used to understand and predict the difference between boiling points for a pure solvent and a nonvolatile solution.
51. **Solve** An aqueous solution of calcium chloride ($CaCl_2$) boils at 101.3°C. How many kilograms of calcium chloride were dissolved in 1000.0 g of the solvent?
52. **Calculate** the boiling point elevation and freezing point depression of a solution containing 50.0 g of glucose ($C_6H_{12}O_6$) dissolved in 500.0 g of water.
53. **Investigation** A lab technician determines the boiling point elevation of an aqueous solution of a nonvolatile, nonelectrolyte to be 1.12°C. What is the solution's molality?

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SCIENTIFIC BREAKTHROUGHS

Blood Falls: A Salty Secret Under the Ice

Blood Falls is a unique feature in Antarctica. The reddish-brown falls flow from Taylor Glacier, the coldest glacier on Earth where water flows persistently. Scientists have recently uncovered the chemical properties of the unusual falls.

Blood Falls flow from Taylor Glacier in East Antarctica.

When explorers first discovered Blood Falls more than 100 years ago, they thought its color was caused by red algae. In fact, the liquid flowing from the falls is a very highly concentrated salt solution called brine. The color of the falls is caused by the oxidation of iron in the brine when it is exposed to oxygen at the surface of the glacier. The iron comes from the scraping of underlying bedrock beneath the glacier.

Using radio-echo sounding (RES), researchers recently found evidence that links the brine to an ancient saltwater lake trapped beneath the glacier. RES is sometimes called radioglaciology. It is a technique that uses radar (radio waves) to study ice sheets and glaciers.

Learning More About Brine

The brine in Blood Falls moves under high pressure through cracks and channels to reach the surface. Researchers mapped these cracks and channels to learn more about where the brine came from and how it reached the surface.

The brine in Blood Falls contains so much salt in solution that it does not freeze at the normal freezing point of water. The temperature of the water in the falls is –17 degrees Celsius (1.4 degrees Fahrenheit). The freezing point of the brine is much lower than the freezing point of water. In addition, the release of heat as the saltwater freezes, melts the surrounding ice.

Life in Brine

Scientists have also discovered that certain microbes can survive in brine trapped under ice. Learning more about brine and these microbes may help scientists understand more about the beginning of life on Earth and how life might exist on other planets under extreme environments.

COMMUNICATE SCIENTIFIC AND TECHNICAL INFORMATION

Research other features on Earth where brine occurs naturally. Create a poster about one of these features describing its characteristics and origin.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 TYPES OF MIXTURES

- The individual substances in a heterogeneous mixture remain distinct.
- Two types of heterogeneous mixtures are suspensions and colloids.
- Brownian motion is the erratic movement of colloid particles on a microscopic level.
- Colloids exhibit the Tyndall effect.
- A solution can exist as a gas, a liquid, or a solid, depending on the solvent.
- Solutes in a solution can be gases, liquids, or solids.

- suspension
- colloid
- Brownian motion
- Tyndall effect
- soluble
- miscible
- insoluble
- immiscible

Lesson 2 SOLUTION CONCENTRATION

- Concentrations can be measured qualitatively and quantitatively.
- Molarity is the number of moles of solute dissolved per liter of solution.

$$\text{molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

- Molality is the ratio of the number of moles of solute dissolved in 1 kg of solvent.

$$\text{molality } (m) = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

- The number of moles of solute does not change during a dilution.

$$M_1V_1 = M_2V_2$$

- concentration
- molarity
- molality
- mole fraction

Lesson 3 FACTORS AFFECTING SOLVATION

- The process of solvation involves solute particles surrounded by solvent particles.
- Solutions can be saturated, unsaturated, or supersaturated.
- Henry's law states that at a given temperature, the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid.

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

- solvation
- heat of solution
- unsaturated solution
- saturated solution
- supersaturated solution
- Henry's law

Lesson 4 COLLIGATIVE PROPERTIES OF SOLUTIONS

- Nonvolatile solutes lower the vapor pressure of a solution.
- Boiling point elevation is directly related to the solution's molality.

$$\Delta T_b = K_b m$$

- A solution's freezing point depression is always lower than that of the pure solvent.

$$\Delta T_f = K_f m$$

- Osmotic pressure depends on the number of solute particles in a given volume.

- colligative property
- vapor pressure lowering
- boiling point elevation
- freezing point depression
- osmosis
- osmotic pressure

THREE-DIMENSIONAL THINKING
Module Wrap-Up

REVISIT THE PHENOMENON

How it is possible for a liquid to hold this shape?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT
Now that you've completed the module, revisit your **STEM** unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

SEP Data Analysis Lab

How can you measure turbidity?

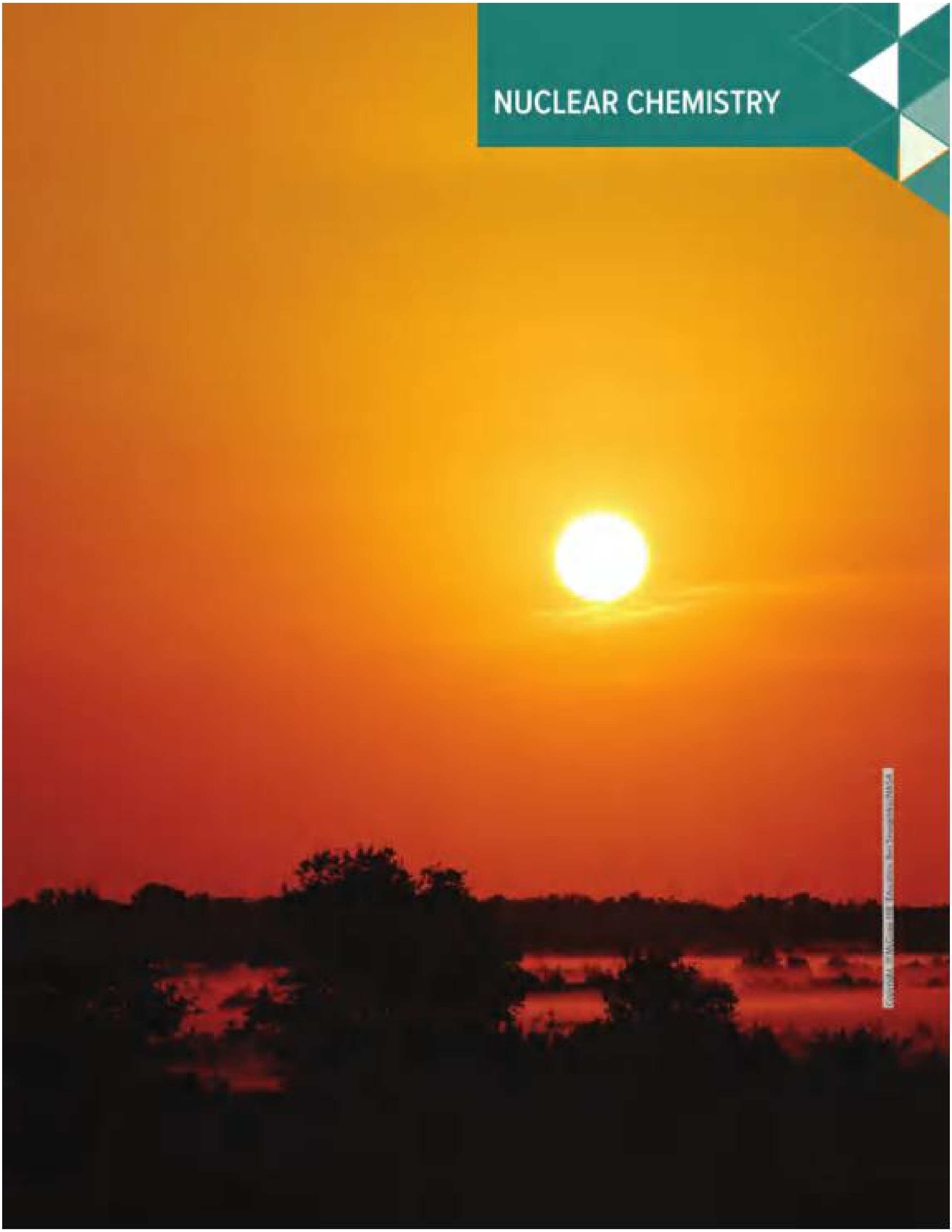
The National Primary Drinking Water Regulations set the standards for public water systems. Turbidity—a measure of the cloudiness of water that results from the suspension of solids in the water—is often associated with contamination from viruses, parasites, and bacteria. Most of these colloid particles come from erosion, industrial and human waste, algae blooms from fertilizers, and decaying organic matter.

Data and Observations The Tyndall effect can be used to measure the turbidity of water. Your goal is to plan a procedure and develop a scale to interpret data.

CER Analyze and Interpret Data

1. **Claim** Identify the variables that can be used to relate the ability of light to pass through the liquid and the number of the colloid particles present. What will you use as a control?
2. **Claim, Evidence** Relate the variables used in the experiment to the actual number of colloid particles that are present.
3. **Analyze** What safety precautions must be considered?
4. **Evidence, Reasoning** Determine the materials you need to measure the Tyndall effect. Select technology to collect and interpret data.

NUCLEAR CHEMISTRY



NUCLEAR CHEMISTRY

ENCOUNTER THE PHENOMENON

Where does the Sun get all its energy?

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

CER Claim, Evidence, Reasoning

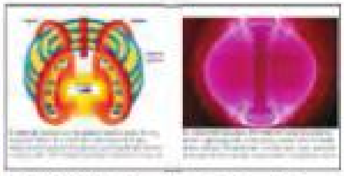
Make Your Claim Use your CER chart to make a claim about where the sun gets all its energy.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.

LESSON 1: Explore & Explain: Defining Radioactivity



LESSON 3: Explore & Explain: Nuclear Fusion

LESSON 1

NUCLEAR RADIATION

FOCUS QUESTION

How was radioactivity discovered?

The Discovery of Radioactivity

You have studied various forms of chemical reactions. In any chemical reaction, atoms can gain, lose, or share valence electrons, but the identity of the atoms does not change. Nuclear reactions, which you will study in this chapter, are different. Nuclear chemistry is concerned with the structure of atomic nuclei and the changes they undergo. Whereas chemical reactions involve only small energy changes, nuclear reactions involve much larger energy changes. **Table 1** offers a comparison of chemical reactions and nuclear reactions.

Table 1 Comparison of Chemical and Nuclear Reactions

Chemical Reactions		Nuclear Reactions	
	• Occur when bonds are broken and formed • Involve only valence electrons • Associated with small energy changes • Atoms keep the same identity although they might gain, lose, or share electrons, and form new substances • Temperature, pressure, concentration, and catalysts affect reaction rates		• Occur when nuclei combine, split, and emit radiation • Can involve protons, neutrons, and electrons • Associated with large energy changes • Atoms of one element are often converted into atoms of another element • Temperature, pressure, and catalysts do not normally affect reaction rates

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Review the News

Obtain information from a current news story about nuclear processes, including fusion and fission of unstable nuclei. Evaluate your source and communicate your findings to the class.

In 1895, German physicist Wilhelm Roentgen (1845–1923) found that invisible rays were emitted when electrons bombarded the surface of certain materials. These invisible rays caused photographic plates to darken, and Roentgen named these high-energy emissions X-rays. At that time, French physicist Henri Becquerel (1852–1908) was studying minerals that emit light after being exposed to sunlight, a phenomenon called phosphorescence. Building on Roentgen's work, Becquerel wanted to determine whether phosphorescent minerals also emitted X-rays.

Becquerel discovered by chance that phosphorescent uranium salts produced spontaneous emissions that darkened photographic plates. He observed this phenomenon even when the uranium salts were not exposed to light. Chemist Marie Curie (1867–1934) and her husband Pierre Curie (1859–1906) took Becquerel's mineral sample, called pitchblende, and isolated the components emitting the rays. They concluded that the darkening of the photographic plates was due to rays emitted from the uranium atoms present in the mineral sample. Marie Curie named the process by which materials give off such rays radioactivity; the rays and particles emitted by a radioactive source are called radiation. **Figure 1** shows the darkening of photographic film that is exposed to radiation emitted by radium salts.

The work of Marie and Pierre Curie was extremely important in establishing the origin of radioactivity and developing the field of nuclear chemistry. In 1898, the Curies identified two new elements, polonium and radium, on the basis of their radioactivity. Henri Becquerel and the Curies shared the 1903 Nobel Prize in Physics for their work. Marie Curie also received the 1911 Nobel Prize in Chemistry for her work with polonium and radium. The 1911 prize honored her achievement in discovering those elements, as well as isolating radium and contributing substantially to chemical knowledge through her study of its properties and compounds.

Get It?

Explain what Marie and Pierre Curie concluded about the darkening of the photographic plates.

Figure 1 Radium salts are placed on a special emulsion on a photographic plate. After the plate is developed, the emulsion shows the dark tracks left by radiation emitted by the radium salts.

Table 2 Properties of Alpha, Beta, and Gamma Radiation

Property	Alpha Radiation	Beta Radiation	Gamma Radiation
Symbol	α	β	γ
Composition	alpha particles	beta particles	high-energy electromagnetic radiation
Description of radiation	helium nuclei, $^{4}_{2}He$	electrons	photons
Charge	2+	1–	0
Mass	6.64×10^{-27} kg	9.11×10^{-31} kg	0
Approximate energy	5 MeV	0.05 to 1 MeV	1 MeV
Relative penetrating power	blocked by paper	blocked by metal foil	not completely blocked by lead or concrete

Types of Radiation

Recall that isotopes are atoms of the same element that have different numbers of neutrons. Isotopes of atoms with unstable nuclei are called **radioisotopes.** These unstable nuclei emit radiation to attain more stable atomic configurations in a process called radioactive decay. During radioactive decay, unstable nuclei release energy by emitting radiation. The three most common types of radiation are alpha (α), beta (β), and gamma (γ). **Table 2** summarizes some of their important properties.

Ernest Rutherford (1871–1937), who performed the famous gold foil experiment that helped define modern atomic structure, identified alpha, beta, and gamma radiation when studying the effects of an electric field on the emissions from a radioactive source. The effects of an electric field on gamma rays, alpha particles, and beta particles are shown in **Figure 2.**

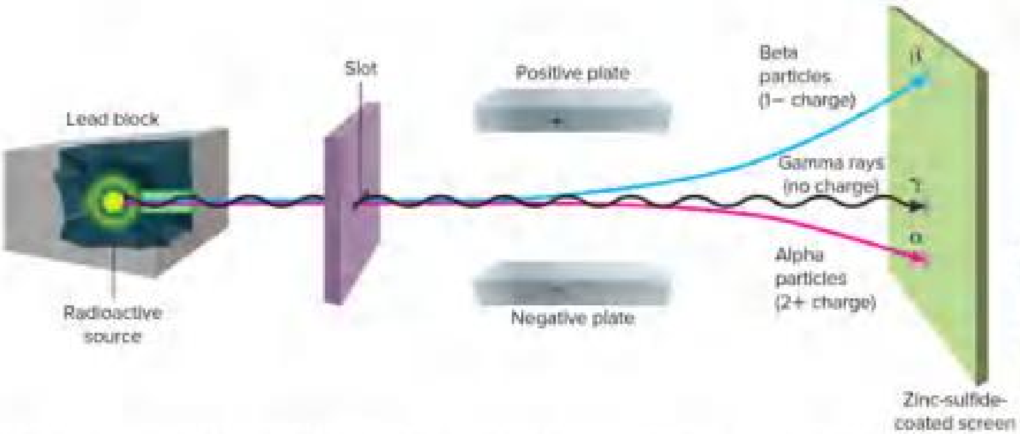


Figure 2 The effect of an electric field depends on the charge and mass of the radiation. Positively charged alpha particles deflect toward the negative plate. Negatively charged beta particles deflect toward the positive plate. The lighter beta particles undergo the larger deflection. Gamma rays have no charge and are not affected by an electric field.

Figure 3 A radium-226 nucleus undergoes alpha decay to form radon-222 and an alpha particle.
Compare *the number of protons and neutrons in radium-226 and radon-222.*

Alpha particles

An alpha particle (α) has the same composition as a helium nucleus—two protons and two neutrons—and is therefore given the symbol $^{4}_{2}He$. The charge of an alpha particle is 2+ due to the presence of the two protons. Alpha radiation consists of a stream of alpha particles. Because of their mass and charge, alpha particles are relatively slow-moving compared with other types of radiation. Thus, alpha particles are not very penetrating—a single sheet of paper stops alpha particles. In **Figure 3,** radium-226, an atom whose nucleus contains 88 protons and 138 neutrons, undergoes alpha decay by emitting an alpha particle. Note that the reaction is balanced. That is, the sum of the mass numbers (superscripts) and the sum of the atomic numbers (subscripts) on each side of the arrow are equal. The total number of neutrons plus protons does not change in the nuclear process. Also note that when a radioactive nucleus emits an alpha particle, the product nucleus has an atomic number that is lower by 2 and a mass number that is lower by 4.

Get It?

Describe what happens to the total number of neutrons plus protons in alpha decay.

Beta particles

A beta particle is a very fast-moving electron that is emitted when a neutron in an unstable nucleus converts into a proton. Beta particles are represented by the symbol β or e−. They have a 1− charge. Their mass is so small compared with the mass of nuclei involved in nuclear reactions that it can be approximated to zero.

Beta radiation consists of a stream of fast-moving electrons. An example of the beta decay process is the decay of iodine-131 into xenon-131 by beta-particle emission, as shown in **Figure 4.** Note that the mass number of the product nucleus is the same as that of the original nucleus (they are both 131), but its atomic number has increased by 1 (54 instead of 53). This change in atomic number occurs because a neutron is converted into a proton, as shown by the following equation.

$$n \rightarrow p + \beta$$

Figure 4 An iodine-131 nucleus undergoes beta decay to form xenon-131 and a beta particle.
Explain *How does beta decay affect the mass number of the decaying nucleus?*

As you might recall, the number of protons in an atom determines its identity. Thus, the formation of an additional proton results in the transformation from iodine-131 to xenon-131. Also, note that the electric charge in the equation above is conserved. The neutron is neutral. The proton has a 1+ charge and the beta particle has a 1− charge. Because beta particles are both lightweight and fast-moving, they have greater penetrating power than alpha particles. A thin sheet of metal foil is required to stop beta particles.

Gamma rays

Gamma rays are photons, which are high-energy (short wavelength) electromagnetic radiation. They are denoted by the symbol γ. Because photons have no mass and no charge, the emission of gamma rays does not change the atomic number or mass number of a nucleus. Gamma rays almost always accompany alpha and beta radiation, as they account for most of the energy loss that occurs as a nucleus decays. For example, gamma rays accompany the alpha-decay reaction of uranium-238.

$$^{238}_{92}\mathrm{U} \rightarrow {}^{234}_{90}\mathrm{Th} + {}^{4}_{2}\mathrm{He} + 2\gamma$$

The 2 in front of the γ symbol indicates that two gamma rays of different frequencies are emitted. Because gamma rays have no effect on mass number or atomic number, it is customary to omit them from nuclear equations.

As you have learned, the discovery of X-rays helped set the stage for the discovery of radioactivity. **X-rays,** like gamma rays, are a form of high-energy electromagnetic radiation. However, X-rays are not produced by radioactive sources and their energy is lower than that of gamma rays. They are emitted when inner electrons are knocked out and electrons from higher energy levels drop down to fill the vacancy. **Figure 5** shows an X-ray image taken in space. It allows astronomers to observe objects not visible in optical images. The presence of X-rays indicates phenomena such as exploding stars or black holes. Hospitals and dentists have machines that produce X-rays when a beam of electrons strikes a metal target. The familiar X-ray images are produced as the beam of X-rays passes easily through soft tissue but is partly blocked by hard tissue, such as bone.

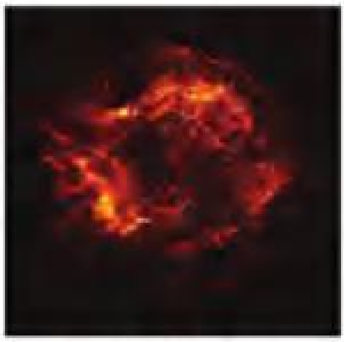


Figure 5 The *Chandra Observatory*, launched in July 1999, photographed X-rays emitted from a cool gas cloud surrounding the black hole at the center of a neighboring galaxy.

Get It?

Compare and contrast X-rays and gamma rays.

WORD ORIGIN

radiation

from the Latin word *radiare*, meaning *to radiate*

CCC CROSSCUTTING CONCEPTS

Energy and Matter In nuclear processes, atoms are not conserved, but the total number of protons plus neutrons is conserved. Create a table that shows that the above statement is true. What evidence did you use to make your table?

Penetrating power

The ability of radiation to pass through matter is called **penetrating power.** Alpha particles have a low penetrating power because they move slowly due to their large mass, and their 2+ charge causes them to lose energy quickly through interactions with other particles in matter they encounter. As a result, alpha particles can be stopped by very little shielding. Even a piece of paper, light clothing, or the outer layers of skin will stop them.

The penetrating power of beta particles is higher than alpha particles because beta particles are smaller and faster than alpha particles. However, they still interact with other particles and will be stopped fairly quickly by shielding. For example, metal foil or thick specialized clothing will stop beta particles.

Gamma rays are highly penetrating. Because they have no charge and no mass, gamma rays do not interact much with matter. Therefore, the probability of matter stopping them is low. A thick layer of concrete or lead is relatively effective at shielding against gamma rays, although some gamma radiation may still penetrate.

Penetrating power may be quantified as the depth of water that stops 50 percent of the incoming radiation. Water is used for this purpose because living tissue contains a high proportion of water. Measured in this way, the penetrating power is about 0.03 mm for alpha radiation, about 2 mm for beta radiation, and about 10 cm for gamma radiation. In the case of gamma radiation, for example, this means that a 10 cm layer of water will stop 50 percent of incoming gamma radiation.

Get It?

Explain why the penetrating power of gamma rays is greater than the penetrating power of alpha or beta particles.

Check Your Progress

Summary

- Wilhelm Roentgen discovered X-rays in 1895.
- Henri Becquerel, Marie Curie, and Pierre Curie pioneered the fields of radioactivity and nuclear chemistry.
- Radioisotopes emit radiation to attain more stable atomic configurations.

Demonstrate Understanding

1. **List** the different types of radiation and their charges.
2. **Compare** the subatomic particles involved in nuclear and chemical reactions.
3. **Explain** how you know whether the reaction is chemical or nuclear when an atom undergoes a reaction and attains a more stable form.
4. **Calculate** The approximate energy values in **Table 2** are given in units of MeV. Convert each value into Joules using the following conversion factor: $1\ MeV = 1.6 \times 10^{-13}\ J$.
5. **Summarize** Make a time line that summarizes the major events that led to the understanding of alpha, beta, and gamma radiation.

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LESSON 2
RADIOACTIVE DECAY

FOCUS QUESTION

Why are some nuclei radioactive?

Nuclear Stability

Except for the emission of gamma radiation, radioactive decay involves the conversion of an element into another element. Such a reaction, in which an atom's atomic number is altered, is called **transmutation.** Whether an atom spontaneously decays and what type of radiation it emits depends on its neutron-to-proton ratio.

The protons and neutrons in an atom's nucleus are referred to as **nucleons.** Despite the strong electrostatic repulsion forces among protons, all nucleons remain bound in the dense nucleus because of the **strong nuclear force.** As shown in **Figure 6,** the strong nuclear force acts on subatomic particles that are extremely close together and overcomes the electrostatic repulsion among protons. Nuclear stability is related to the balance between electrostatic and strong nuclear forces.

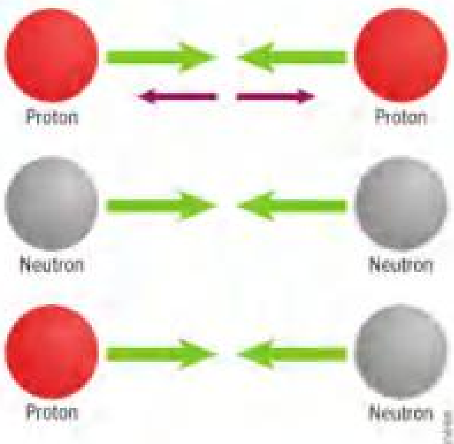


Figure 6 A repulsive electrostatic force, represented by the purple arrows, acts between the two positively charged protons. The strong nuclear force, represented by the green arrows, acts between any two or more nucleons and is always attractive. Because neutrons do not repel one another or protons, experiencing only the attractive strong force, their presence adds to the overall attraction among nucleons.

Infer *What is the effect of the electrostatic force between two neutrons? Between a proton and a neutron?*

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Laboratory: Modeling Isotopes

Construct an explanation to determine the stability and change of an atom experiencing spontaneous radioactive decay.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Neutron-to-proton ratio

To a certain degree, the stability of a nucleus can be correlated to its neutron-to-proton (n/p) ratio. For atoms with low atomic numbers (<20), the most stable nuclei are those with neutron-to-proton ratios of 1:1. For example, helium ($^{4}_{2}He$) has two neutrons and two protons, and a neutron-to-proton ratio of 1:1. As atomic number increases, more and more neutrons are needed to produce a strong nuclear force that is sufficient to balance the electrostatic repulsion force between protons. Therefore, the neutron-to-proton ratio for stable atoms gradually increases, reaching a maximum of approximately 1.5:1 for the largest atoms. An example of this is lead ($^{206}_{82}Pb$). With 124 neutrons and 82 protons, lead has a neutron-to-proton ratio of 1.51:1.

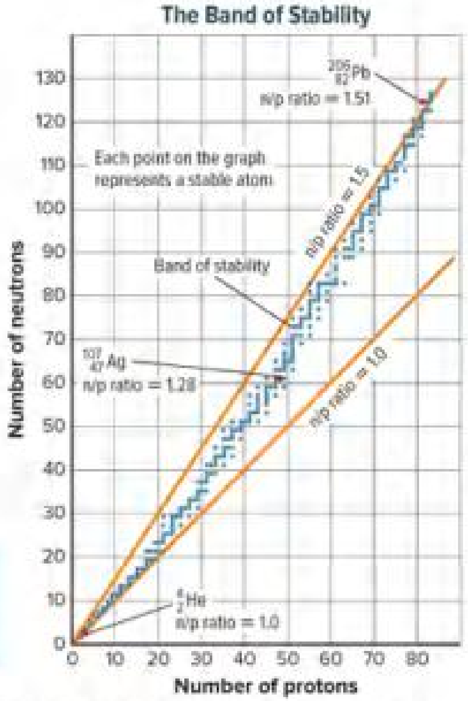


Figure 7 The band of stability is the region where all stable nuclei fall when plotting the number of neutrons versus the number of protons. As the atomic number increases, the neutron-to-proton ratio (n/p) increases from 1:1 to 1.5:1.

Get It?

Explain why the neutron-to-proton ratio of stable nuclei increases as the atomic number increases.

The band of stability

Examine the plot of the number of neutrons versus the number of protons for all known stable nuclei shown in **Figure 7.** Notice that the slope of the plot indicates that the number of neutrons required for a nucleus to be stable increases as the number of protons increases. This correlates with the increase in the neutron-to-proton ratio of stable nuclei with increasing atomic number. The area on the graph within which all stable nuclei are found is known as the **band of stability.**

As shown in **Figure 7,** $^{4}_{2}He$ and $^{206}_{82}Pb$ are both positioned within the band of stability although they have a different neutron-to-proton ratio. All nuclei outside the band of stability—either above or below—are radioactive and undergo decay in order to gain stability. After decay, the new atom is positioned more closely to, if not within, the band of stability. The band of stability ends at lead-208; all elements with atomic numbers greater than 82 are radioactive.

Analysis of the types of decay that various unstable nuclei undergo reveals a pattern. Unstable nuclei found above the band of stability undergo beta decay, which decreases the neutron-to-proton ratio. Nuclei found under the band of stability undergo other types of decay that result in an increase of their neutron-to-proton ratio. Heavy nuclei beyond the band undergo alpha decay.

Get It?

Define the band of stability and relate it to the value of the neutron-to-proton ratio.

Types of Radioactive Decay

The type of radioactive decay a particular radioisotope undergoes depends to a large degree on the underlying causes for its instability. Atoms lying above the band of stability generally have too many neutrons to be stable, whereas atoms lying below the band of stability tend to have too many protons to be stable. Depending on the relative number of neutrons and protons, atoms can undergo different types of decay—beta decay, alpha decay, positron emission, or electron capture—to gain stability.

Beta decay

A radioisotope that lies above the band of stability is unstable because it has too many neutrons relative to its number of protons. For example, unstable $^{14}_{6}C$ has a neutron-to-proton ratio of 1.33:1, whereas stable elements of similar mass, such as $^{12}_{6}C$ and $^{14}_{7}N$, have neutron-to-proton ratios of approximately 1:1. It is not surprising, then, that $^{14}_{6}C$ undergoes beta decay, as this type of decay decreases the number of neutrons in the nucleus.

$$^{14}_{6}C \rightarrow ^{14}_{7}N + \beta$$

Figure 8a shows the beta decay of carbon-14 into nitrogen-14. Note that the atomic number of the product nucleus, $^{14}_{7}N$, has increased by one. The nitrogen-14 atom now has a stable neutron-to-proton ratio of 1:1. Thus, beta emission has the effect of increasing the stability of a neutron-rich atom by increasing its atomic number, that is, by lowering its neutron-to-proton ratio. The resulting atom is closer to, if not within, the band of stability.

Get It?
Explain why radioisotopes above the band of stability are unstable.

Figure 8 Depending on where nuclei lie on the band of stability, they can emit a beta particle or an alpha particle.

Compare and contrast *beta decay and alpha decay in terms of the atomic number of the nuclei involved in the reaction.*

SCIENCE USAGE v. COMMON USAGE

unstable

Science usage: spontaneously radioactive

Unstable atoms decay to reach a more stable state.

Common usage: not firm or fixed in one place

The chair is unstable because one of its legs is shorter than the others.

Alpha decay

All nuclei with more than 82 protons are radioactive and decay spontaneously. Both the number of neutrons and the number of protons must be reduced in order to make these radioisotopes stable. These very heavy nuclei often decay by emitting alpha particles. For example, polonium-210 spontaneously decays into lead-206 by emitting an alpha particle.

$$^{210}_{84}\mathrm{Po} \rightarrow ^{206}_{82}\mathrm{Pb} + ^{4}_{2}\mathrm{He}$$

Figure 8b on the previous page shows the alpha decay of polonium-210 into lead-206. The atomic number of $^{210}_{84}\mathrm{Po}$ decreases by 2 and the mass number decreases by 4 as the nucleus decays into $^{206}_{82}\mathrm{Pb}$.

Get It?

Calculate how the neutron-to-proton ratio changes when polonium-210 decays into lead-206.

Positron emission and electron capture

For nuclei with low neutron-to-proton ratios, two common radioactive decay processes occur: positron emission and electron capture. These two processes tend to increase the neutron-to-proton ratio of the neutron-poor atom, bringing the atom closer to, if not within, the band of stability.

Positron emission is a radioactive decay process that involves the emission of a positron from a nucleus. A **positron** is a particle with the same mass as an electron but opposite charge; thus, it is represented by the symbol β^+ or e^+. During positron emission, a proton in the nucleus is converted into a neutron and a positron, and then the positron is emitted.

$$p \rightarrow n + e^+$$

Figure 9a shows the positron emission of a carbon-11 nucleus. Carbon-11 lies below the band of stability and has a low neutron-to-proton ratio of approximately 0.8:1. Carbon-11 undergoes positron emission to form boron-11. Positron emission decreases the number of protons from six to five, and increases the number of neutrons from five to six. The resulting atom, $^{11}_{5}\mathrm{B}$, has a neutron-to-proton ratio of 1.2:1, which is within the band of stability.

Get It?

Explain how positron emission increases the stability of nuclei with low neutron-to-proton ratios.

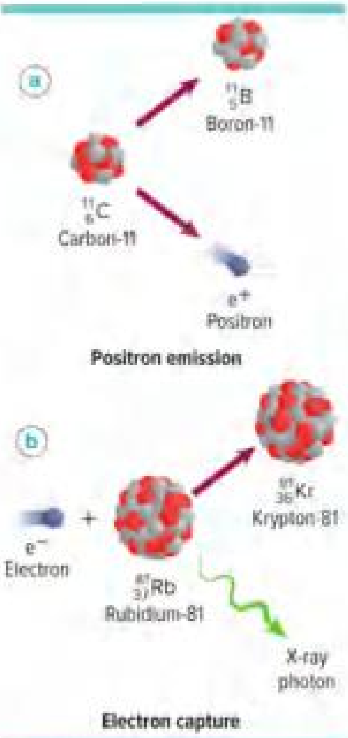


Figure 9 When a nucleus undergoes positron emission or captures an electron, the number of protons decreases by one.

Compare and contrast *how the number of protons and neutrons change during positron emission and electron capture.*

Table 3 **Summary of Radioactive Decay Processes**

Type of Radioactive Decay	Particle Emitted	Change in Mass Number	Change in Atomic Number
Alpha decay	${}^{4}_{2}He$	decreases by 4	decreases by 2
Beta decay	β or e^-	no change	increases by 1
Positron emission	β^+ or e^+	no change	decreases by 1
Electron capture	X-ray photon	no change	decreases by 1
Gamma emission	γ	no change	no change

Electron capture is the other common radioactive-decay process that decreases the number of protons in unstable nuclei lying below the band of stability. **Electron capture** occurs when the nucleus of an atom draws in a surrounding electron, usually one from the lowest energy level. This captured electron combines with a proton to form a neutron.

$$p + e^- \rightarrow n$$

The atomic number of the nucleus decreases by 1 as a consequence of electron capture. The formation of the neutron also results in an X-ray photon being emitted. These two characteristics of electron capture are shown in the electron capture of rubidium-81 in **Figure 9b.** The balanced nuclear equation for the reaction is shown below.

$$e^- + {}^{81}_{37}Rb \rightarrow {}^{81}_{36}Kr + \text{X-ray photon}$$

The five types of radioactive decay you have read about in this chapter are summarized in **Table 3.**

Get It?

List the decay processes that result in an increased neutron-to-proton ratio, and those that result in a decreased neutron-to-proton ratio.

Writing and Balancing Nuclear Equations

The radioactive decay processes you have just read about are all examples of nuclear reactions. Nuclear reactions are expressed by balanced nuclear equations just as chemical reactions are expressed by balanced chemical equations. However, in balanced chemical equations, numbers and types of atoms are conserved; in balanced nuclear equations, mass numbers and charges are conserved.

In balancing nuclear equations, you will compare the mass numbers of the product and reactant particles involved in the nuclear reaction to ensure that they are conserved. You will use atomic numbers and the charges of electrons and other charged particles to ensure that charges are conserved. In many cases, your challenge in analyzing nuclear equations will be to determine the identity of an unknown product or reactant.

EXAMPLE Problem 1

BALANCING A NUCLEAR EQUATION NASA uses the alpha decay of plutonium-238 ($^{238}_{94}Pu$) as a heat source on spacecraft. Write a balanced equation for this decay.

1 ANALYZE THE PROBLEM

You are given that a plutonium atom undergoes alpha decay and forms an unknown product. Plutonium-238 is the initial reactant, while the alpha particle is one of the products of the reaction. The reaction is summarized below.

$$^{238}_{94}Pu \rightarrow {}^{A}_{Z}X + {}^{4}_{2}He$$

You must determine the unknown product of the reaction, X.

Known	Unknown
reactant: plutonium-238 ($^{238}_{94}Pu$)	mass number of the product A = ?
decay type: alpha particle emission ($^{4}_{2}He$)	atomic number of the product Z = ?
	reaction product X = ?

2 SOLVE FOR THE UNKNOWN

$238 = A + 4$ — Apply the conservation of mass number.

$A = 238 - 4 = 234$ — Solve for A.

Thus, the mass number of X is 234.

$94 = Z + 2$ — Apply the conservation of charges.

$Z = 94 - 2 = 92$ — Solve for Z.

Thus, the atomic number of X is 92.

The periodic table identifies the element as uranium (U).

$^{238}_{94}Pu \rightarrow {}^{234}_{92}U + {}^{4}_{2}He$ — Write the balanced nuclear equation.

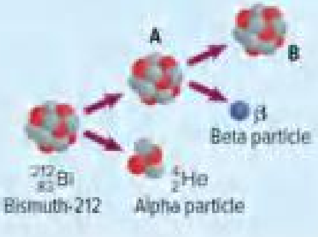
3 EVALUATE THE ANSWER

The correct formula for an alpha particle is used. The sums of the superscripts and subscripts on each side of the equation are equal. Therefore, the charge and the mass number are conserved. The nuclear equation is balanced.

PRACTICE Problems

ADDITIONAL PRACTICE

6. Write a balanced nuclear equation for the reaction in which oxygen-15 undergoes positron emission.
7. Thorium-229 is used to increase the lifetime of fluorescent bulbs. What type of decay occurs when thorium-229 decays to form radium-225?
8. **CHALLENGE** The figure at right shows one way that bismuth-212 can decay, producing isotopes A and B.
   a. Write balanced nuclear equations for this decay.
   b. Identify the isotopes A and B that are produced.



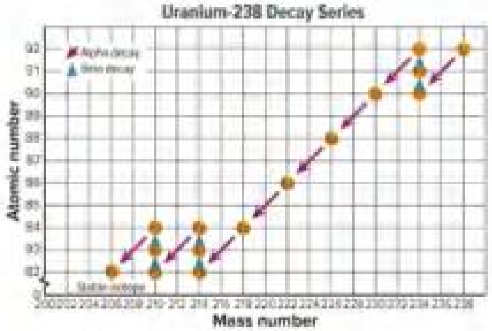


Figure 10 Uranium-238 undergoes 14 different radioactive decay steps before forming stable lead-206.

Radioactive Series

A **radioactive decay series** is a series of nuclear reactions that begins with an unstable nucleus and results in the formation of a stable nucleus. As **Figure 10** shows, uranium-238 first decays to thorium-234, which in turn decays to protactinium-234. Decay reactions continue until a stable nucleus, lead-206, is formed.

Get It?
List each step in the decay of uranium-238. Include the type of decay and the resulting product.

Radioactive Decay Rates

You might wonder how there could be any naturally occurring radioisotopes found on Earth. After all, if radioisotopes undergo continuous radioactive decay, won't they eventually disappear? Furthermore, radioisotopes have been decaying for about 4.6 billion years—the span of Earth's existence. Yet, naturally occurring radioisotopes are not uncommon on Earth. Some radioisotopes, such as carbon-14, are continuously formed in the upper atmosphere of Earth. Others are formed in the universe, during stellar nucleosynthesis for instance. Radioisotopes can also be synthesized in laboratories. The differing decay rates of isotopes also contribute to their presence on Earth.

Radioactive decay rates are measured in half-lives. A **half-life** is the time required for one-half of a radioisotope's nuclei to decay into its products. For example, the half-life of the radioisotope strontium-90 is 29 years. If you had 10.0 g of strontium-90 today, 29 years from now you would have 5.0 g left. **Table 4** on the next page shows how this decay continues through four half-lives of strontium-90. **Figure 11** on the next page presents the data from the table in terms of the percent of strontium-90 remaining after each half-life. The decay continues until a negligible amount of strontium-90 remains.

Get It?
Define the term *half-life*.

Table 4 The Decay of Strontium-90

Number of Half-Lives	Elapsed Time	Amount of Strontium-90 Present
0	0 y	10.0g
1	29 y	$10.0\ g \times \left(\frac{1}{2}\right) = 5.00\ g$
2	58 y	$10.0\ g \times \left(\frac{1}{2}\right)\left(\frac{1}{2}\right) = 2.50\ g$
3	87 y	$10.0\ g \times \left(\frac{1}{2}\right)\left(\frac{1}{2}\right)\left(\frac{1}{2}\right) = 1.25\ g$
4	116 y	$10.0\ g \times \left(\frac{1}{2}\right)\left(\frac{1}{2}\right)\left(\frac{1}{2}\right)\left(\frac{1}{2}\right) = 0.625\ g$

The data in **Table 4** can be summarized in a simple equation representing the decay of any radioactive element.

Remaining Amount of Radioactive Element

$$N = N_0\left(\frac{1}{2}\right)^n$$

N is the remaining amount.
N_0 is the initial amount.
n is the number of half-lives that have passed.

The amount remaining is equal to the initial amount times one-half raised the number of half-lives that have passed.

The exponent n can also be replaced with the equivalent quantity t/T, where t is the elapsed time and T is the duration of the half-life. Note that t and T must have the same units of time.

$$N = N_0\left(\frac{1}{2}\right)^{t/T}$$

This type of expression is known as an exponential decay function. **Figure 11** shows the graph of a typical exponential decay function—in this case, the decay curve for strontium-90.

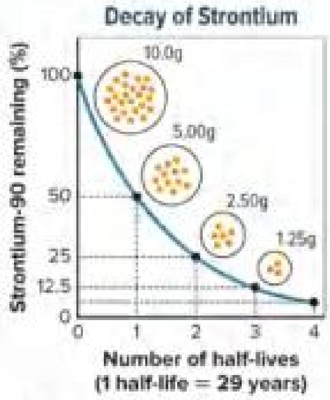


Figure 11 The graph shows how the amount of strontium in a sample changes as a function of the number of half-lives.

Get It?

Infer how much strontium remains after 1.5 half-lives.

Characteristic half-lives

Each radioisotope has its own characteristic half-life. For example, the half-life of polonium-214 is 163.7 μs, the half-life of radon-222 is 3.8 days, and the half-life of uranium-238 is 4.46×10^9 years. Notice the large range of values for half-lives, from millionths of a second to billions of years!

EXAMPLE Problem 2

CALCULATING THE AMOUNT OF REMAINING ISOTOPE Krypton-85 is used in indicator lights of appliances. The half-life of krypton-85 is 11 y. How much of a 2.000-mg sample remains after 33 y?

1 ANALYZE THE PROBLEM

You are given a known mass of a radioisotope with a known half-life. You must first determine the number of half-lives that passed during the 33-year period. Then, use the exponential decay equation to calculate the amount of the sample remaining.

Known

Initial amount = 2.000 mg

Elapsed time (t) = 33 y

Half-life (T) = 11 y

Unknown

Amount remaining = ? mg

2 SOLVE FOR THE UNKNOWN

Number of half-lives (n) = $\frac{\text{elapsed time}(t)}{\text{half-life}(T)}$ — Determine the number of half-lives.

$n = \frac{33\ \text{y}}{11\ \text{y}} = 3.0$ half-lives — Substitute t = 33 y and T = 11 y.

Amount remaining = (initial amount) $\left(\frac{1}{2}\right)^n$ — Write the exponential decay equation.

Amount remaining = (2.000 mg) $\left(\frac{1}{2}\right)^{3.0}$ — Substitute initial amount = 2.000 mg and n = 3.0.

Amount remaining = (2.000 mg) $\left(\frac{1}{8}\right)$ = **0.2500 mg**

3 EVALUATE THE ANSWER

Three half-lives are equivalent to $\left(\frac{1}{2}\right)\left(\frac{1}{2}\right)\left(\frac{1}{2}\right)$, or $\left(\frac{1}{8}\right)$. The answer (0.25 mg) is equal to $\left(\frac{1}{8}\right)$ of the initial amount. The answer has two significant figures because the number of years has two significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

9. Bandages can be sterilized by exposure to gamma radiation from cobalt-60, which has a half-life of 5.27 y. How much of a 10.0-mg sample of cobalt-60 is left after one half-life? Two half-lives? Three half-lives?

10. If the passing of five half-lives leaves 25.0 mg of a strontium-90 sample, how much was present in the beginning?

11. **CHALLENGE** The table shows the amounts of radioisotopes in three different samples. To the nearest gram, how much will be in Sample B and Sample C when Sample A has 16.2 g remaining?

Sample	Radioisotope	Half-life	Amount (g)
A	cobalt-60	5.27 y	64.8
B	tritium	12.32 y	58.4
C	strontium-90	28.79 y	37.6

Radiochemical dating

Chemical reaction rates are greatly affected by changes in temperature, pressure, and concentration, and by the presence of a catalyst. In contrast, nuclear reaction rates remain constant regardless of such changes. In fact, the half-life of any particular radioisotope is constant. Because of this, radioisotopes can be used to determine the age of an object. The process of determining the age of an object by measuring the amount of a certain radioisotope remaining in that object is called **radiochemical dating.**

LIFE SCIENCE Connection A type of radiochemical dating known as carbon dating is used to measure the age of artifacts that were once part of a living organism. Carbon dating makes use of the radioactive decay of carbon-14, which is formed by cosmic rays in the upper atmosphere at a fairly constant rate. These carbon-14 atoms become evenly spread throughout Earth's biosphere, where they mix with stable carbon-12 and carbon-13 atoms. Plants use carbon dioxide from the environment, which contains all carbon isotopes, to build more complex molecules through the process of photosynthesis. When animals eat plants, the carbon-14 atoms that were part of the plant become part of the animal. Because organisms are constantly taking in carbon compounds, they contain the same ratio of carbon- 14 to carbon-12 and carbon-13 found in the atmosphere. However, after they die, organisms no longer ingest new carbon compounds, and the carbon-14 they already contain continues to decay. The carbon-14 undergoes beta decay to form nitrogen-14.

$$^{14}_{6}C \rightarrow {}^{14}_{7}N + \beta$$

Carbon-14 has a half-life of 5730 years. Because the amount of stable carbon in the dead organism remains constant while the carbon-14 continues to decay, the ratio of unstable carbon-14 to stable carbon-12 and carbon-13 decreases. By measuring this ratio and comparing it to the nearly constant ratio present in the atmosphere, the age of an object can be estimated. For example, if an object's C-14 to (C-12 + C-13) ratio is one-quarter of the ratio measured in the atmosphere, the object is approximately two half-lives, or 11,460 years old. Carbon-14 dating is limited to accurately dating objects up to approximately 45,000 years of age. This method was used to date the Great Pyramid of Giza, shown in **Figure 12.**

Figure 12 Using the radiocarbon dating method on organic materials, such as ash and charcoal found at the Great Pyramid of Giza, scientists estimate the pyramid to be more than 4000 years old.

EARTH SCIENCE Connection The decay process of a different radioisotope, uranium-238 to lead-206, is commonly used to date objects such as rocks. Because the half-life of uranium-238 is 4.5×10^9 years, it can be used to estimate the age of objects that are too old to be dated using carbon-14. By radiochemical dating of meteorites, the age of the solar system has been estimated at 4.6×10^9 years. Certain minerals in rocks are more suitable than others for radiochemical dating by this method. The most commonly analyzed mineral is zircon, containing zirconium, silicon, and oxygen. Zircon is suitable for analysis using the decay of uranium-238 to lead-206 because uranium is naturally incorporated into its crystal structure, whereas lead is not. Therefore, any lead found in zircon can be assumed to be there as a result of the decay of uranium-238.

Check Your Progress

Summary

- The conversion of an atom of one element to an atom of another by radioactive decay processes is called *transmutation*.
- Atomic number and mass number are conserved in nuclear reactions.
- A half-life is the time required for half of the atoms in a radioactive sample to decay.
- Radiochemical dating is a technique for determining the age of an object by measuring the amount of certain radioisotopes remaining in the object.

Demonstrate Understanding

12. **Describe** what happens to unstable nuclei.
13. **Explain** how you can predict whether or not an isotope is likely to be stable if you know its number of neutrons and protons.
14. **Describe** the forces acting on the particles within a nucleus and explain why neutrons are the glue holding the nucleus together.
15. **Predict** the nuclear equation for the alpha decay of radium-226 used on the tips of older lightning rods.
16. **Calculate** how much of a 10.0-g sample of americium-241 remains after four half-lives. Americium-241 is a radioisotope commonly used in smoke detectors and has a half-life of 430 y.
17. **Calculate** After 2.00 y, 1.986 g of a radioisotope remains from a sample that had an original mass of 2.000 g.
    a. Calculate the half-life.
    b. How much of the radioisotope remains after 10.00 y?
18. **Graph** A sample of polonium-214 originally has a mass of 1.0 g. Express the mass remaining as a percent of the original sample after a period of one, two, and three half-lives. Graph the percent remaining versus the number of half-lives. Approximately how much time has elapsed when 20% of the original sample remains? The half-life of polonium-214 is 163.7 μs.

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LESSON 3

NUCLEAR REACTIONS

FOCUS QUESTION

What is the relationship between mass and energy and why is it important?

Induced Transmutation

All nuclear reactions, or transmutations, that have been described thus far are examples of radioactive decay, where one element is converted into another element by the spontaneous emission of radiation. However, transmutations can also be forced, or induced, by bombarding a stable nucleus with a neutron or with high-energy alpha, beta, or gamma radiation. In 1919, Ernest Rutherford performed the first laboratory conversion of one element into another element. By bombarding nitrogen-14 with high-speed alpha particles, oxygen-17 and hydrogen-1 were formed. This transmutation reaction is illustrated in **Figure 13** and the reaction is shown below.

$$^{14}_{7}N + ^{4}_{2}He \rightarrow ^{17}_{8}O + ^{1}_{1}H$$

As Rutherford demonstrated, nuclear reactions can be induced, in other words, produced artificially. The process, which involves striking nuclei with high-velocity particles, is called **induced transmutation.** In the case of charged particles, such as the alpha particles used by Rutherford, the incident particles must be moving at extremely high speeds to overcome the electrostatic repulsion between themselves and the target nucleus. Because of this, scientists have developed methods to accelerate charged particles to extreme speeds by using very strong electrostatic fields and magnetic fields. Particle accelerators are machines built to produce the high-speed particles needed to induce transmutation. Since Rutherford's first experiments involving induced transmutation, scientists have used the technique to synthesize hundreds of new isotopes in the laboratory.

Figure 13 When an alpha particle bombards a nitrogen-14 atom, an atom of oxygen-17 and an atom of hydrogen-1 are produced.

3D THINKING DCI CCC SEP

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practice: Modeling Fission, Fusion, and Radioactive Decay
HS-PS1-8. Develop models to illustrate the changes in the composition of the nucleus of the atom and the energy released during the process of fission, fusion, and radioactive decay.

Revisit the Encounter the Phenomenon Question
What information from this lesson can help you answer the module question?

Transuranium elements

The elements immediately following uranium in the periodic table—elements with atomic numbers 93 and greater—are known as the **transuranium elements.** All transuranium elements have been produced in the laboratory by induced transmutation and are radioactive. Many transuranium elements have been named in honor of their discoverers or the laboratories at which they were created. For example, element 117, tennessine, was so named because some of the key work leading to its discovery was carried out in Tennessee. Scientists continue their ongoing efforts to synthesize new transuranium elements and study their properties.

EXAMPLE Problem 3

INDUCED TRANSMUTATION REACTION EQUATIONS Write a balanced nuclear equation for the induced transmutation of oxygen-16 into nitrogen-13 by proton bombardment. An alpha particle is emitted in the reaction.

1 ANALYZE THE PROBLEM

You are given all of the particles involved in an induced transmutation reaction. Because the proton bombards the oxygen atom, they are reactants and must appear on the reactant side of the reaction arrow.

Known	Unknown
reactants: oxygen-16 and a proton	nuclear equation for the reaction = ?
products: nitrogen-13 and an α-particle	

2 SOLVE FOR THE UNKNOWN

Nuclear formula for oxygen-16: $^{16}_{8}O$ — Use the periodic table to obtain the atomic number of oxygen.

Nuclear formula for nitrogen-13: $^{13}_{7}N$ — Use the periodic table to obtain the atomic number of nitrogen.

Nuclear formula for proton: p

Nuclear formula for alpha particle: $^{4}_{2}He$

$^{16}_{8}O + p \rightarrow ^{13}_{7}N + ^{4}_{2}He$ — Write the balanced nuclear equation.

3 EVALUATE THE ANSWER

A proton has a charge of 1+ and a mass number of 1. Therefore, both charge and mass number are conserved. The formula for each participant in the reaction is also correct. The nuclear equation is written correctly.

PRACTICE Problems

ADDITIONAL PRACTICE

19. Write the balanced nuclear equation for the induced transmutation of aluminum-27 into sodium-24 by neutron bombardment. An alpha particle is released in the reaction.
20. Write the balanced nuclear equation for the alpha-particle bombardment of $^{239}_{94}Pu$. One of the reaction products is a neutron.
21. **CHALLENGE** Archeologists sometimes use a procedure called neutron activation analysis to identify elements in artifacts. The figure at right shows one type of reaction that can occur when an artifact is bombarded with neutrons. If the product of the process is cadmium-110, what was the target and unstable isotope? Write balanced nuclear equations for the process to support your answer.

Nuclear Reactions and Energy

In your study of chemical reactions, you read that mass is conserved. For most practical purposes this is true—but, it is not accurate.

Einstein's equation

Albert Einstein's equation relates mass and energy. It states that any reaction produces or consumes energy due to a loss or gain in mass. Energy and mass are equivalent. Note that because c^2 is large, a small change in mass results in a large change in energy.

Energy Equivalent of Mass

$$\Delta E = \Delta mc^2$$

ΔE is the change in energy, in Joules. Δm is the change in mass, in kg. c is the speed of light.

The change in energy is equal to the change in mass times the square of the speed of light.

Mass defect and binding energy

Scientists have determined that the mass of the nucleus is always less than the sum of the masses of the individual protons and neutrons that make up the nucleus. This observed difference in mass between a nucleus and its component nucleons is called the **mass defect.**

When nucleons combine together to form an atom, the energy corresponding to the mass defect is released. Conversely, energy is needed to break apart a nucleus into its component nucleons. The nuclear binding energy can be defined as the amount of energy needed to break one mole of nuclei into its individual nucleons. The larger the binding energy per nucleon, the more strongly the nucleons are held together, and the more stable the nucleus is. Less-stable atoms have lower binding energies per nucleon. In other words, it is harder to break apart a nucleus with a high binding energy than it is to break apart a nucleus with a low binding energy.

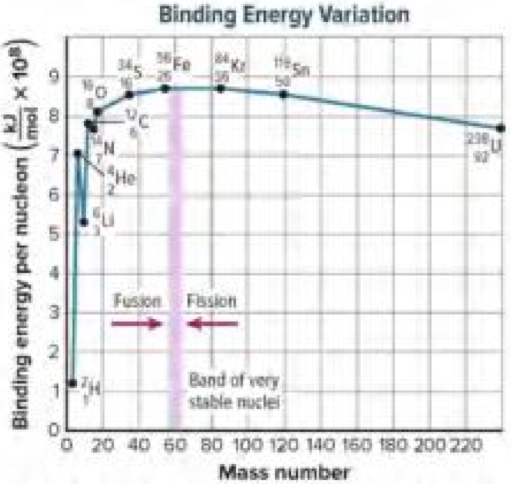


Figure 14 The binding energy per nucleon is a function of the mass number. Light nuclei gain stability by undergoing nuclear fusion. Heavy nuclei gain stability by undergoing nuclear fission.

Figure 14 shows the average binding energy per nucleon versus the mass number. Note that the binding energy per nucleon reaches a maximum around a mass number of 60. Elements with a mass number near 60 are the most stable.

Get It?

Describe how the binding energy varies as a function of the mass number.

PROBLEM-SOLVING STRATEGY

Calculating Mass Defect

You can calculate the mass defect of an isotope if you know the mass of the isotope and the number and masses of its components. Applying the equation $\Delta E = \Delta mc^2$, you can then derive the equivalent binding energy.

Mass defect = $m_{nucleus} - [N_p m_p + N_n m_n]$

where $m_{nucleus}$ is the mass of the nucleus, m_p is the mass of a proton, m_n is the mass of a neutron, N_p is the number of protons, and N_n is the number of neutrons.

If you start with the mass of the atom, you have to take into account the mass of the electrons. To do so, the mass of a hydrogen atom, which is composed of a proton and an electron, is used instead of the mass of a proton. The equation is then:

Mass defect = $m_{isotope} - [N_p m_H - N_n m_n]$

Use the following values for the calculations: $m_H = 1.007825$ amu and $m_n = 1.008665$ amu. The accepted value for c is 3.00×10^8 m/s.

To calculate the energy in Joules, you can convert the masses into kilograms using 1 amu = 1.660540×10^{-27} kg.

Apply the Strategy

Calculate the mass defect and binding energy of lithium-7. The mass of lithium-7 is 7.016003 amu.

In typical chemical reactions, the energy produced or consumed is so small that the accompanying changes in mass are negligible. In contrast, the mass changes and associated energy changes in nuclear reactions are significant. For example, the energy released from the nuclear reaction of 1 kg of uranium is equivalent to the energy released during the chemical combustion of about four billion kilograms of coal.

Nuclear Fission

The binding energies shown in **Figure 14** on the previous page indicate that heavy nuclei tend to be unstable. To gain stability, they can fragment into several smaller nuclei. Because atoms with mass numbers around 60 are the most stable, heavy atoms (those with mass numbers greater than 60) tend to fragment into smaller atoms in order to increase their stability. The splitting of a nucleus into fragments is known as **nuclear fission.** The fission of a nucleus is accompanied by a very large release of energy. Nuclear power plants use the large release of energy associated with nuclear fission to generate power. The generation of nuclear power is an example of a series of energy transformations in which nuclear energy is transformed into thermal energy and then into electrical energy.

ACADEMIC VOCABULARY

generate
to bring into existence, to originate by a physical or chemical process
Fire generates a lot of heat.

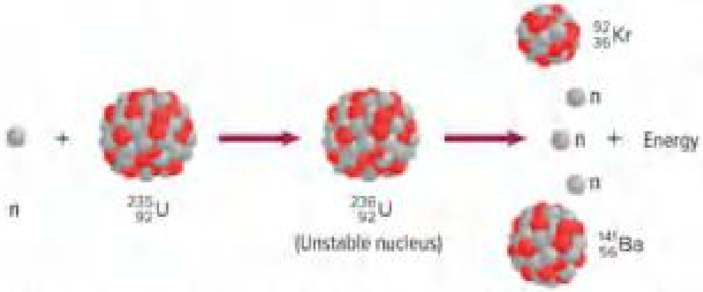


Figure 15 When bombarded with a neutron, uranium-235 forms unstable uranium-236, which then splits into two lighter nuclei and additional neutrons. The fission of uranium-235 is accompanied by a large release of energy.

The first nuclear fission reaction discovered involved uranium-235. As you can see in **Figure 15**, when a neutron strikes a uranium-235 nucleus, it undergoes fission. Barium-141 and krypton-92 are just two of the many possible products of this fission reaction. In fact, scientists have identified more than 200 different product isotopes from the fission of a uranium-235 nucleus.

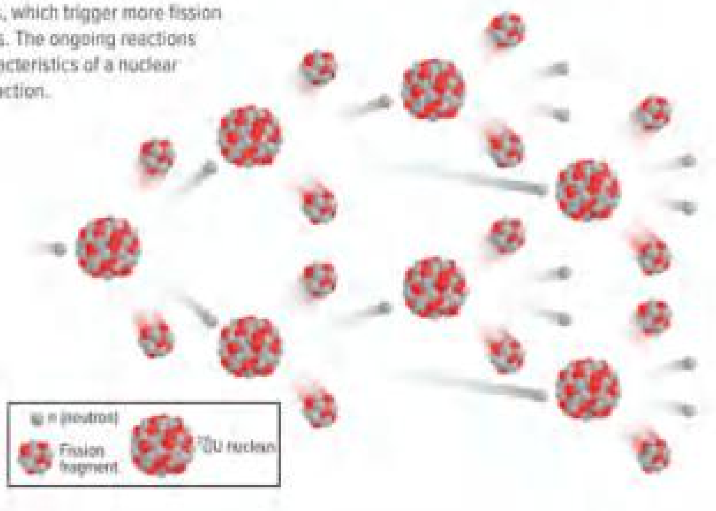
Get It?

Explain why heavy atoms undergo nuclear fission.

Chain reactions

Each fission of uranium-235 releases additional neutrons, as shown in **Figure 15.** If one fission reaction produces two neutrons, these two neutrons can cause two additional fissions. If those two fissions release four neutrons, those four neutrons could then produce four more fissions, and so on, as shown in **Figure 16.** This self-sustaining process in which one reaction initiates the next is called a chain reaction. As you might imagine, the number of fissions and the amount of energy released can increase rapidly. The explosion from an atomic bomb is an example of an uncontrolled chain reaction.

Figure 16 When uranium nuclei undergo fission, they release neutrons, which trigger more fission reactions. The ongoing reactions are characteristics of a nuclear chain reaction.



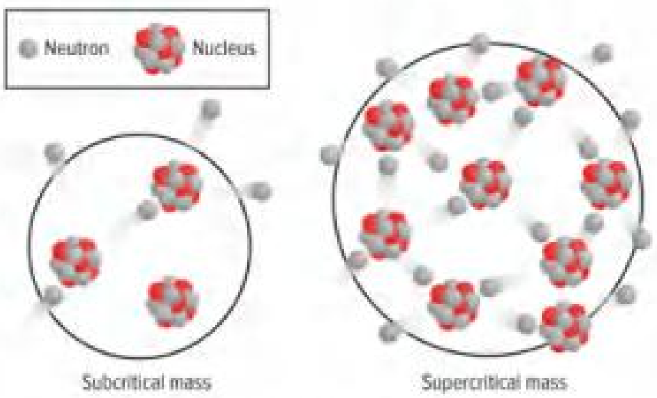


Figure 17 Whether a nuclear reaction can be sustained depends on the amount of matter present. In a subcritical mass, the chain reaction does not start because neutrons escape before causing enough fission to sustain the chain reaction. In a supercritical mass, neutrons cause more and more fissions, and the chain reaction accelerates.

A sample of fissionable material must have sufficient mass in order for a chain reaction to occur. If it does not, neutrons escape from the sample before they can start the chain reaction by striking other nuclei. A sample that is not massive enough to sustain a chain reaction is said to have subcritical mass. A sample that is massive enough to sustain a chain reaction has **critical mass.** When a critical mass is present, the neutrons released in one fission cause other fissions to occur. If much more mass than the critical mass is present, the chain reaction rapidly escalates. This can lead to a violent nuclear explosion. A sample of fissionable material with a mass greater than the critical mass is said to have supercritical mass. **Figure 17** shows the effect of mass on the initiation and progression of a fission reaction.

Nuclear Reactors

Nuclear fission produces the energy generated by nuclear reactors. This energy is primarily used to generate electricity at nuclear power plants, such as the one shown in **Figure 18.** A common fuel is fissionable uranium (IV) oxide (UO_2) encased in corrosion-resistant rods. U-238 is the most abundant isotope (99%) of uranium. U-235, which makes up 0.7% of the natural uranium, has the rare property of being able to undergo induced fission; U-235 atoms undergo fission when hit by a neutron. The fuel used in nuclear power plants is enriched to contain 3% uranium-235, the amount required to sustain a chain reaction, and is called enriched uranium. Additional rods, often made of cadmium or boron, control the fission process inside the reactor by absorbing neutrons released during the reaction.

Keeping the chain reaction going while preventing it from racing out of control requires precise monitoring and continual adjusting of the control rods. Much of the concern about nuclear power plants focuses on the risk of losing control of the nuclear reactor, possibly resulting in the accidental release of harmful levels of radiation. The Three Mile Island accident in the United States in 1979 and the Chernobyl accident in Ukraine in 1986 provide examples of why controlling the reactor is critical. **Figure 19** on the next page shows the city of Pripyat, located 3 km from Chernobyl. The city was completely abandoned after the accident.

Figure 18 The main parts of a nuclear power plant are the reactor under the dome and the cooling tower.

How a nuclear reactor works

The fission within a nuclear reactor is started by a neutron-emitting source and is stopped by positioning the control rods to absorb all of the neutrons produced in the reaction. The reactor core contains a reflector that acts to reflect neutrons back into the core, where they will react with the fuel elements, also called fuel rods. A coolant, usually water, circulates through the reactor core, to carry off the heat generated by the nuclear fission reactions. The hot coolant heats water so that it boils, producing steam that is in turn used to power turbines. The movement of the turbines is used to generate an electric current.

Figure 19 The city of Pripyat was deserted after the accident at the Chernobyl power plant.

Nuclear power plants and fossil-fuel burning power plants are similar. In both types of power plant, heat from a reaction—nuclear fission or chemical combustion of a fossil fuel such as coal—is used to generate steam. The steam then drives turbines that produce electricity, as shown in the nuclear power plant illustrated in **Figure 20.** The other major components of a nuclear power plant are also illustrated in **Figure 20.**

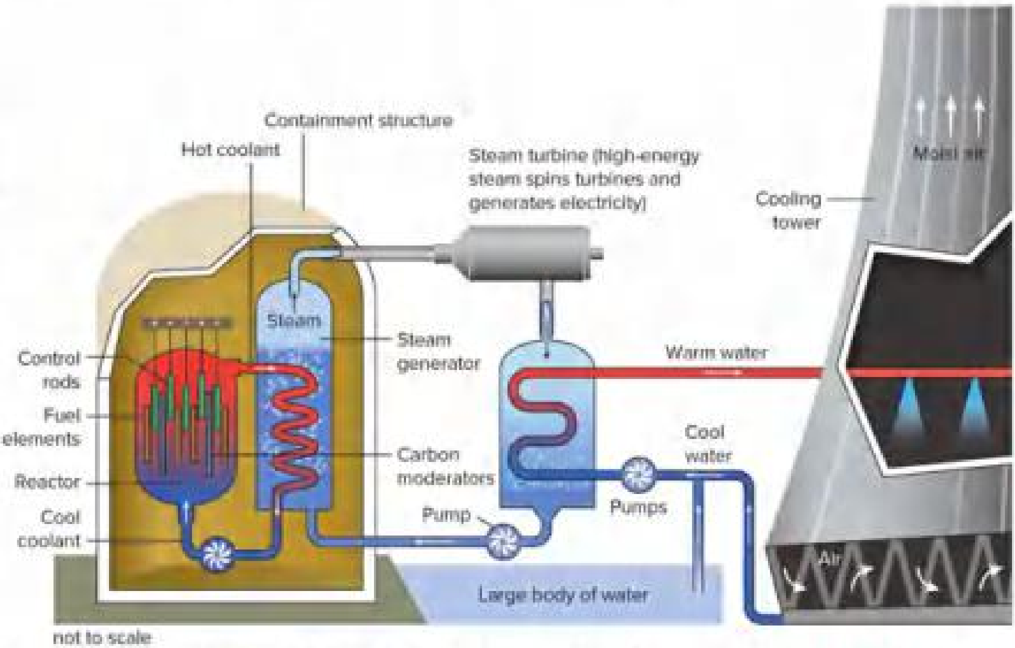


Figure 20 A nuclear reactor produces heat that drives the formation of steam. The energy from the steam spins a turbine, which produces electricity. The steam is eventually cooled and recycled. The water used to cool the steam enters the cooling tower, where steam is released to the atmosphere.

Identify *two examples of energy transformations that take place in a nuclear reactor.*

Figure 21 The interior of a reactor is filled with water. A crane is used to extract and replace fuel rods.

Because of the hazardous radioactive fuels and fission products present at nuclear power plants, a dense concrete structure is usually built to enclose the reactor. The main purpose of the containment structure is to shield personnel and nearby residents from harmful radiation.

As the reactor operates, the fuel rods are gradually depleted and products from the fission reactions accumulate. Because of this, the reactor must be serviced periodically. Spent fuel rods are extracted from the reactor, as shown in **Figure 21,** and can be reprocessed and repackaged to make new fuel rods. Some fission products, however, are extremely radioactive and cannot be used again. These products must be stored as nuclear waste.

Issues associated with nuclear power plants

Risks of accidents, such as the ones mentioned in **Figure 22** on the next page, have to be taken into account when operating nuclear power plants. However, the storage of highly radioactive nuclear waste is still one of the major issues surrounding the debate over the use of nuclear power. Approximately 20 half-lives are required for the radioactivity of nuclear waste materials to reach levels acceptable for biological exposure. For some types of nuclear fuels, the wastes remain substantially radioactive for thousands of years. A considerable amount of scientific research is devoted to the disposal of radioactive wastes. Highly radioactive materials from the reactor core are first treated with advanced technologies that ensure the materials will not deteriorate over a very long period of time. Treated wastes are then stored in sealed containers that are buried deep underground.

Another issue is the limited supply of the uranium-235 used in the fuel rods. One option is to build reactors that produce new quantities of fissionable fuels. Reactors able to produce more fuel than they use are called **breeder reactors.** Although the design of breeder reactors poses many difficult technical problems, they are currently in operation in several countries.

Get It?

Infer how the storage of nuclear wastes affects the environment.

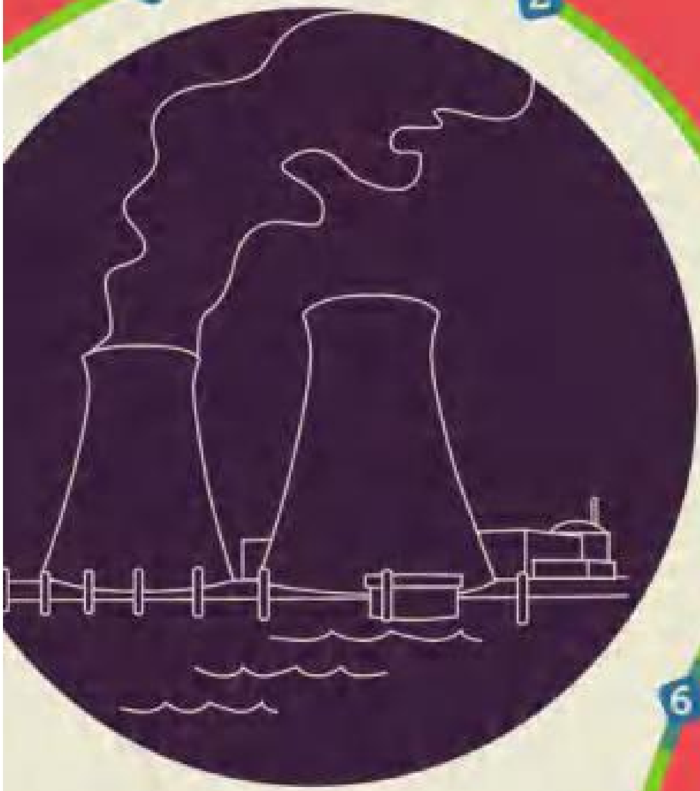


Figure 22

The Nuclear Age

The discovery of X-rays in 1895 initiated a series of breakthroughs in understanding atomic nuclei. Today, nuclear chemistry applications involving medicine, weaponry, and energy affect the lives of people worldwide.

1 **1898** Marie and Pierre Curie discover the radioactive elements polonium and radium. Their work establishes the early framework for the study of nuclear chemistry.

2 **1919** The first artificially induced nuclear reaction transmutes nitrogen into an isotope of oxygen by bombarding nitrogen gas with alpha particles.

3 **1934** Enrico Fermi bombards heavy elements with neutrons. Ida Noddack reviews these experiments and suggests that smaller nuclei might have been produced.

4 **1941–1945** Manhattan Project scientists develop uranium and plutonium bombs, which were dropped on Hiroshima and Nagasaki, Japan, in 1945 and ended World War II.

5 **1949** Radiocarbon dating allows scientists to determine the age of artifacts made from plant-based materials as old as 45,000 years.

6 **1960s** Scientists research using high-energy radiation to treat cancer. Clinical trials bring dramatic improvement in the treatment and cure of malignant tumors.

7 **1979 | 1986** Nuclear power plant accidents at Three Mile Island, Pennsylvania, and Chernobyl, Ukraine, focus world attention on the dangers associated with nuclear power.

8 **2010** Scientists using NASA's Fermi Gamma-ray Space Telescope discover two massive bubbles, one above and one below the core of the Milky Way. They are thought to be the result of activity from either a black hole or star formation.

Nuclear Fusion

Recall from the binding energy diagram in **Figure 14** that a mass number of about 60 has the most stable atomic configuration. Thus, it is possible to bind together two or more light (mass number less than 60) and less-stable nuclei to form a single more-stable nucleus. The combining of atomic nuclei is called **nuclear fusion.** Nuclear fusion reactions, which are responsible for producing the heaviest elements, can release very large amounts of energy. You already have everyday knowledge of this fact—the Sun is powered by a series of fusion reactions as hydrogen atoms fuse to form helium atoms.

$$4\,{}^{1}_{1}\mathrm{H} \rightarrow {}^{4}_{2}\mathrm{He} + 2\mathrm{e}^{+} + \text{energy}$$

Scientists have spent several decades researching nuclear fusion. It is a promising source of energy and has several advantages compared to nuclear fission. Lightweight isotopes used to fuel the reactions, such as hydrogen, are abundant. Fusion reaction products are not generally radioactive. Nuclear fusion produces large amounts of energy. Fusion reactions produce more energy per unit of mass of fuel than fission reactions. This could solve the problem of many societies' increasing needs for electricity.

Real-World CHEMISTRY

Nuclear Fusion

SOLAR FUSION Nuclear fusion reactions are responsible for the glow and heat from stars such as the Sun. The temperature of the Sun's core is about 15,000,000 K. It is so hot and dense that hydrogen nuclei fuse to produce helium. After billions of years, the Sun's hydrogen will be mostly depleted. Its temperature will rise to about 100,000,000 K, and the fusion process will then change helium into carbon.

Using nuclear fusion for energy

Unfortunately, there are major problems that must be overcome before fusion can be used to produce energy on a commercially viable scale. One such problem is that fusion requires extremely high energies to initiate and sustain a reaction. The required energy, which is achieved only at extremely high temperatures, is needed to overcome the electrostatic repulsion between the nuclei in the reaction. Because of the energy requirements, fusion reactions are also known as **thermonuclear reactions.** A temperature of 5,000,000 K is required to fuse hydrogen atoms. This temperature—and even higher temperatures—have been achieved using an atomic explosion to initiate the fusion process, but this approach is not practical for controlled electric power generation.

Another significant problem is confinement of the reaction. There are currently no materials capable of withstanding the tremendous temperatures that are required by a fusion reaction. Much of the current research centers around an apparatus called a tokamak reactor. The name *tokamak* comes from Russian and means *toroidal chamber with an axial magnetic field.* A tokamak reactor, shown in **Figure 23** on the next page, is a donut-shaped device that uses strong magnetic fields to contain the fusion reaction. While significant progress has been made in the field of fusion, temperatures high enough for continuous fusion have not yet been sustained for long periods of time.

Get It?

Explain the link between the Sun's temperature and the release of such an enormous amount of energy.

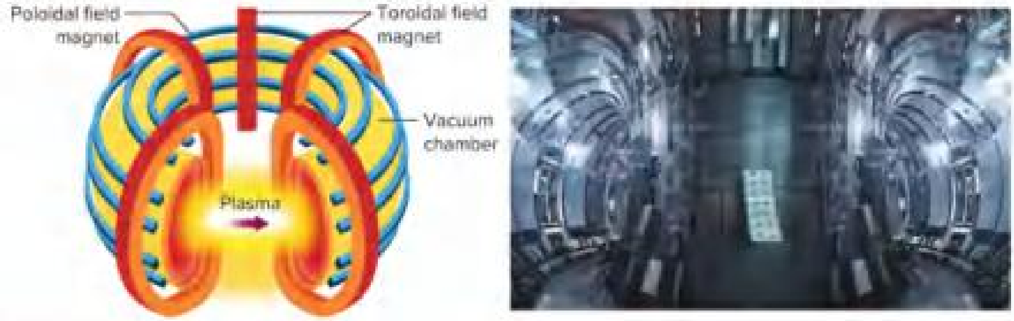


Figure 23 A tokamak reactor, a ring-shaped reactor, uses strong magnetic fields to contain the intensely hot fusion reaction and keep it from direct contact with the reactor interior walls. The poloidal magnets follow the shape of the reactor, and the toroidal magnets wrap around the reactor.

Check Your Progress

Summary

- Induced transmutation is the bombardment of nuclei with particles in order to create new elements.
- In a chain reaction, one reaction induces others to occur. A sufficient mass of fissionable material is necessary to initiate the chain reaction.
- Fission and fusion reactions release large amounts of energy.

Demonstrate Understanding

22. **Explain** why energy is released when nucleons combine to form an atom but is needed to break a nucleus apart.
23. **Compare and contrast** nuclear fission and fusion.
24. **Describe** the process that occurs during a nuclear chain reaction and explain how to monitor a chain reaction in a nuclear reactor.
25. **Explain** how nuclear fission can be used to generate electric power.
26. **Formulate an argument** supporting or opposing nuclear power as your state's primary power source. Assume the primary source of power currently is the burning of fossil fuels.
27. **Calculate** What is the energy change (ΔE) associated with a change in mass (Δm) of 1.00 mg?
28. **Interpret Graphs** Use the graph in **Figure 14** to answer the following questions.
    a. Why is the isotope $^{56}_{26}Fe$ highest on the curve?
    b. Are more stable isotopes located higher or lower on the curve?
    c. Compare the stability of Li-6 and He-4.

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LESSON 4

APPLICATIONS AND EFFECTS OF NUCLEAR REACTIONS

FOCUS QUESTION

What are some applications of nuclear reactions?

Detecting Radioactivity

Radiation energetic enough to ionize matter it contacts is called **ionizing radiation.** The Geiger counter is an ionizing radiation detection device. As shown in **Figure 24,** a Geiger counter consists of a metal tube filled with a gas. In the center of the tube is a wire connected to a power supply. When ionizing radiation penetrates the end of the tube, the gas absorbs the radiation and forms ions and free electrons. The free electrons are attracted to the wire, causing an electric current. A meter measures the current flow through the ionized gas, which indicates the amount of ionizing radiation present.

Figure 24 A Geiger counter is used to detect and measure radiation levels. Ionizing radiation produces an electric current in the counter. The current is displayed on a scaled meter, and a speaker produces audible sounds that vary according to the current flow caused by the ionizing radiation.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science and Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practice: Human Health and Radiation Frequency

HS-PS4-4. Evaluate the validity and reliability of claims in published materials of the effects that different frequencies of electromagnetic radiation have when absorbed by matter.

CCC Identify Crosscutting Concepts

Create a table of the crosscutting concepts and fill in examples you find as you read.

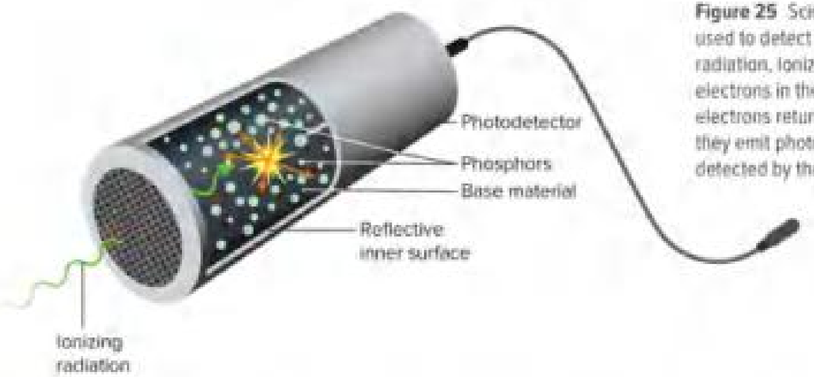

Figure 25 Scintillation counters are used to detect the presence of ionizing radiation. Ionizing radiation excites the electrons in the phosphors. As the electrons return to their ground states, they emit photons, which are then detected by the photodetector.

Monitoring the radiation dose received by people who work near radioactive sources is important to ensure their safety. People who work near radioactive sources might be required to wear a thermoluminescent dosimeter (TLD) badge, which contains a tiny crystal. Radiation excites electrons within the crystal. To determine the radiation dose, the crystal is heated, and the electrons return to their ground states, emitting light. Radioactivity readers detect this light as a measure of the radiation dose to which a worker has been exposed.

Another detection device is a scintillation counter. Scintillations are brief flashes of light produced when ionizing radiation excites the electrons in certain types of atoms or molecules called phosphors. A scintillation counter contains a base material–often a plastic, a crystal, or a liquid–containing phosphors, as shown in **Figure 25.** Ionizing radiation that strikes the scintillation counter can transfer energy either directly to the phosphors or to the base material, which then transfers the energy to the phosphors. This energy excites electrons in the phosphors. As these electrons return to their ground states, they emit light. This light is transmitted through the base material to a photodetector that converts the light to an electrical signal. The number and brightness of the scintillations indicate the amount of ionizing radiation.

Get It?

Summarize how a scintillation counter works.

Uses of Radiation

With proper safety procedures, radiation can be useful in many scientific experiments and industrial applications. For instance, neutron activation analysis is used to detect trace amounts of elements present in a sample. Computer-chip manufacturers use this technique to analyze the composition of highly purified silicon wafers. In the process, the sample is bombarded with a beam of neutrons from a radioactive source, causing some of the atoms in the sample to become radioactive. The type and amount of radiation emitted by the sample is used to determine the types and quantities of elements present. Neutron activation analysis is a highly sensitive measurement technique capable of detecting quantities of less than 1×10^{-9} atoms in a sample. Another application of radiation is the use of beta emission to measure paper thickness.

Using radioisotopes

Radioisotopes can also be used to follow the course of an element through a chemical reaction. For example, CO_2 gas containing radioactive carbon-14 isotopes has been used to study glucose formation in photosynthesis.

$$6CO_2 + 6H_2O \xrightarrow{\text{sunlight}} C_6H_{12}O_6 + 6O_2$$

Because the CO_2 containing carbon-14 is used to trace the progress of carbon through the reaction, it is referred to as a **radiotracer.** A radiotracer is a radioisotope that emits non-ionizing radiation and is used to signal the presence of an element or specific substance. The fact that all of an element's isotopes have the same chemical properties makes the use of radioisotopes possible. Thus, replacing a stable atom of an element in a reaction with one of its isotopes does not alter the reaction. Radiotracers are important in a number of areas of chemical research, particularly in analyzing the reaction mechanisms of complex, multistep reactions.

Radiotracers also have important uses in medicine. Iodine-131, for example, is commonly used to detect diseases associated with the thyroid gland. If a problem is suspected, the patient will drink a solution containing a small amount of iodine-131. After the iodine is absorbed, the amount of iodine taken up by the thyroid is measured and used to monitor the functioning of the thyroid gland.

Get It?

Define *radiotracer.*

Treating cancer

Radiation can pose serious health problems for humans because it can damage or destroy healthy cells. However, radiation can also destroy unhealthy cells, such as cancer cells. All cancers are characterized by the rapid growth of abnormal cells. This growth can produce masses of abnormal tissue, called malignant tumors. Radiation therapy is used to treat cancer by destroying the cancer cells. In fact, cancer cells are more susceptible to destruction by radiation than healthy ones. **Figure 26** shows an MRI of a malignant tumor in a patient's brain. If all goes well, the tumor will be destroyed. Unfortunately, in the process of destroying unhealthy cells, radiation also destroys some healthy cells. Despite this major drawback, radiation therapy has become one of the most effective treatment options in the fight against cancer.

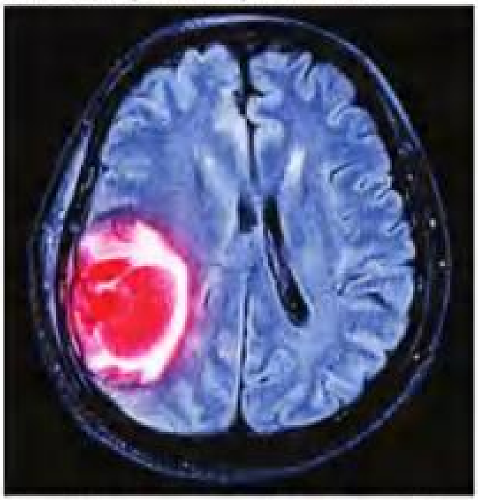

Figure 26 Radiation can be used to treat cancer. This image shows a cancerous tumor in a patient's brain. The tumor is shown in pink.

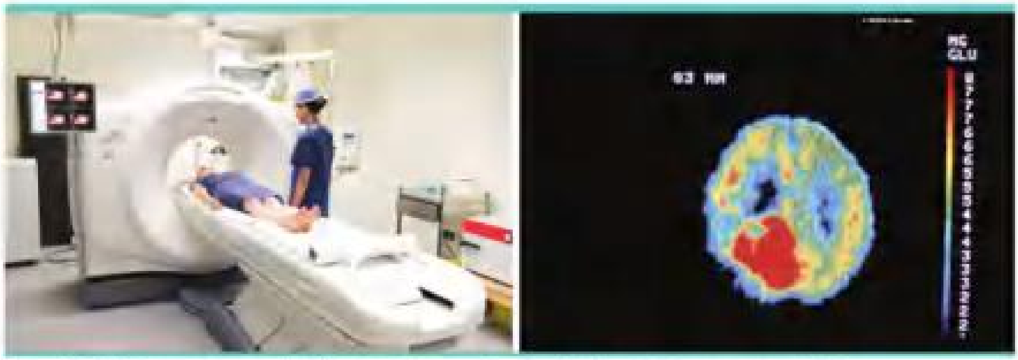

Figure 27 In PET, gamma rays emitted by the radiotracers absorbed by the patients are measured with a detector such as the one shown on the left. The PET scan on the right shows different areas of the brain emitting gamma rays. These images might help doctors locate a tumor or observe a brain function.

Using positron emission

Another radiation-based medical diagnostic tool is called positron emission transaxial tomography (PET). In this procedure, a radiotracer that decays by positron emission is injected into the patient's bloodstream. Positrons that are emitted by the radiotracer cause gamma-ray emissions. These emissions are then detected by an array of sensors surrounding the patient, as shown in **Figure 27**. PET scans can be used to diagnose diseases or study the parts of the brain that are activated under given circumstances, also shown in **Figure 27**.

Biological Effects of Radiation

Although radiation has a number of medical and scientific applications, it can be very harmful. The damage produced from ionizing radiation absorbed by the body depends on several factors, such as the type of radiation, its energy, the type of tissue absorbing the radiation, the penetrating power, and the distance from the source of the radiation. The skin lesion shown in **Figure 28** is an example of such damage.

Figure 28 Radiation can disrupt cell processes and damage skin.

Infer *Is the lesion pictured here somatic or genetic?*

STEM CAREER Connection

Biomedical Engineer

Biomedical engineers apply engineering skills to medical problems. One of their jobs is to design equipment and software, such as those used in positron emission transaxial tomography, that help treat or diagnose medical conditions. If you'd like to combine a love of technology with a desire to help people, this may be the career for you.

LIFE SCIENCE Connection High-energy ionizing radiation is dangerous because it can fragment and ionize molecules within biological tissue. A free radical is an atom or molecule that contains one or more unpaired electrons and is one example of the highly reactive products of ionizing radiation. In a biological system, free radicals can affect a large number of other molecules and ultimately disrupt the operation of normal cells. Ionizing radiation damage to living systems can be classified as either somatic or genetic. Somatic damage affects only nonreproductive body tissue. It includes burns and cancer caused by damage to the cell's growth mechanism. Genetic damage can affect offspring by damaging reproductive tissue. Such damage is difficult to study because it might not become apparent for several generations.

Dose of radiation

A dose of radiation refers to the amount of radiation a body absorbs from a radioactive source. Two units, the rad and the rem, are commonly used to measure doses. The rad, which stands for radiation-absorbed dose, is a measure of the amount of radiation that results in the absorption of 0.01 J of energy per kilogram of tissue. The dose in rads, however, does not account for the energy of the radiation, the type of living tissue absorbing the radiation, or the time of the exposure. To account for these factors, the dose in rads is multiplied by a numerical factor that is related to the radiation's effect on the tissue involved. The result of this multiplication is a unit called the rem. The rem, which stands for roentgen equivalent for man, is named after Wilhelm Roentgen, who discovered X-rays in 1895. **Table 5** summarizes the short-term effects of radiation on humans, depending on the dose.

Intensity and distance

The intensity of radiation depends on the distance from the source as shown by the equation below. The farther away the source, the lower the intensity. The intensity of radiation is measured in amount of radiation per unit of time and/or surface, such as mrem/s•m^2.

Radiation Intensity and Distance

$$I_1 d_1^2 = I_2 d_2^2$$

d_1 and d_2 are two distances from the source.
I_1 is the intensity at d_1 and I_2 is the intensity at d_2.

The intensity of radiation at a distance d_1 from the source multiplied by the square of the distance equals the intensity of the radiation at a distance d_2 multiplied by the square of the distance.

Table 5 **Effects of Short-term Radiation Exposure**

Dose (rem)	Effects on Humans
0–25	no detectable effects
25–50	temporary decrease in white-blood-cell population
100–200	nausea, substantial decrease in white-blood-cell population
500	50% chance of death within 30 days of exposure

Sources of radiation

A variety of sources constantly bombard your body with radiation. Your exposure to these sources results in an average annual radiation exposure of 100–300 millirems of high-energy radiation or 0.1–0.3 rems. **Table 6** shows your annual exposure to common radiation sources.

Table 6 Average Annual Radiation Exposure

Source	Average Exposure (mrem/y)
Cosmic radiation	20–50
Radiation from the ground	25–175
Radiation from buildings	10–160
Radiation from air	20–260
Human body (internal)	~20
Medical and dental X-rays	50–75
Nuclear weapon testing	<1
Air travel	5
Total average	100–300

Check Your Progress

Summary

- Different types of counters are used to detect and measure radiation.
- Radiotracers are used to diagnose disease and to analyze chemical reactions.
- Short-term and long-term radiation exposure can cause damage to living cells.

Demonstrate Understanding

29. **Explain** one way in which nuclear chemistry is used to diagnose or treat disease.
30. **Describe** several methods used to detect and measure radiation.
31. **Compare and contrast** somatic and genetic biological damage.
32. **Explain** why it is safe to use radioisotopes to diagnose medical problems.
33. **Calculate** A lab worker receives an average radiation dose of 21 mrem each month. Her allowed dose is 5,000 mrem/y. On average, what fraction of her yearly dose does she receive?
34. **Interpret Data** Look at the data in **Table 6.** Suppose someone is exposed to the maximum values listed for average annual radiation from the ground, from buildings, and from the air. What fraction would the person receive of the minimum short-term dose (25 rem) that causes a temporary decrease in white blood cell population?

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STEM AT WORK

Disease Detectives

Nuclear medicine technologists are medical professionals who help doctors diagnose and track patients' diseases and injuries by taking images of the body. Their work employs cutting-edge technology, including gamma cameras and positron emission tomography (PET) scanners. They work in hospitals, doctors' offices, imaging facilities, and diagnostic laboratories.

Nuclear medicine technologists prepare and administer radioactive drugs to patients before diagnostic tests to detect abnormalities in the body.

Radioactive Drugs

Nuclear medicine technologists prepare and administer drugs that contain radioisotopes. Depending on the test, patients can receive an injection, or they can inhale or swallow the drug. These drugs, which belong to a class called *radiopharmaceuticals*, emit radiation and cause parts of the body to show up differently in images taken by gamma cameras and other equipment. Areas with atypical concentrations of radioactivity signal problems. The drugs enable doctors to see parts of the body that are diseased, damaged, or not functioning properly. Tumors are detected in this way.

Nuclear medicine is also used to treat different diseases. Nuclear medicine technologists administer various radiotherapy procedures, such as radiation therapy for breast, prostate, and bone cancers.

Testing Equipment

Nuclear medicine technologists operate specialized equipment and use computers to process the data produced by the tests. In conjunction with images produced by gamma cameras and PET scanners, they may run tests using X-ray radiography, magnetic resonance imaging (MRI) scanners, computed tomography (CT), and computerized axial tomography (CAT) scanners.

Nuclear medicine technologists also work closely with patients, gathering medical information, preparing them for tests, and explaining how the tests work. Thanks to technology and dedicated professionals in nuclear medicine, diseases are detected earlier and lives are saved every day.

ASK QUESTIONS TO CLARIFY

Write several questions that you have about nuclear medicine and nuclear medicine technologists. Use print or online sources to find answers. Share the questions and answers with your classmates.

STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 NUCLEAR RADIATION

- Wilhelm Roentgen discovered X-rays in 1895.
- Henri Becquerel, Marie Curie, and Pierre Curie pioneered the fields of radioactivity and nuclear chemistry.
- Radioisotopes emit radiation to attain more stable atomic configurations.

- radioisotope
- X-ray
- penetrating power

Lesson 2 RADIOACTIVE DECAY

- The conversion of an atom of one element to an atom of another by radioactive decay processes is called *transmutation*.
- Atomic number and mass number are conserved in nuclear reactions.
- A half-life is the time required for half of the atoms in a radioactive sample to decay.

$$N = N_0\left(\frac{1}{2}\right)^n \text{ or } N = N_0\left(\frac{1}{2}\right)^{t/T}$$

- Radiochemical dating is a technique for determining the age of an object by measuring the amount of certain radioisotopes remaining in the object.

- transmutation
- nucleon
- strong nuclear force
- band of stability
- positron emission
- positron
- electron capture
- radioactive decay series
- half-life
- radiochemical dating

Lesson 3 NUCLEAR REACTIONS

- Induced transmutation is the bombardment of nuclei with particles in order to create new elements.
- In a chain reaction, one reaction induces others to occur. A sufficient mass of fissionable material is necessary to initiate the chain reaction.
- Fission and fusion reactions release large amounts of energy.

$$E = mc^2$$

- induced transmutation
- transuranium element
- mass defect
- nuclear fission
- critical mass
- breeder reactor
- nuclear fusion
- thermonuclear reaction

Lesson 4 APPLICATIONS AND EFFECTS OF NUCLEAR REACTIONS

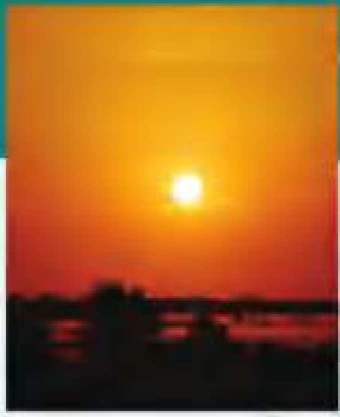
- Different types of counters are used to detect and measure radiation.
- Radiotracers are used to diagnose disease and to analyze chemical reactions.
- Short-term and long-term radiation exposure can cause damage to living cells.

$$I_1d_1^2 = I_2d_2^2$$

- ionizing radiation
- radiotracer

THREE-DIMENSIONAL THINKING

Module Wrap-Up



REVISIT THE PHENOMENON

Where does the Sun get all its energy?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will apply your evidence from this module and complete your project.

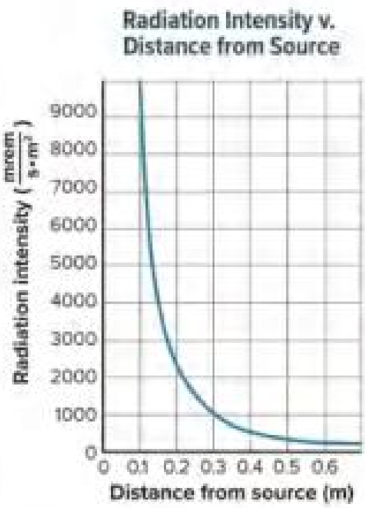
GO FURTHER

SEP Data Analysis Lab

How does distance affect radiation exposure?

When one of the reactors at the Chernobyl nuclear power plant exploded, the radiation spread over thousands of kilometers. The intensity of the radiation decreased with the distance from the reactor.

Data and Observations The graph shows the intensity of a radioactive source versus the distance from the source. The unit of radiation intensity (millirems per second per square meter) indicates the amount of radiation striking a square meter each second.



CER Analyze and Interpret Data

1. **Claim** How does the radiation exposure change as the distance doubles from 0.1 m to 0.2 m? How does it change as the distance quadruples from 0.1 m to 0.4 m?
2. **Evidence, Reasoning** Determine the distance from the source at which the radiation decreased to 0.69 mrem/s•m². This intensity is the maximum radiation exposure intensity considered safe. *(Hint: Use the equation* $\frac{I_1}{I_2} = \frac{d_2^{\,2}}{d_1^{\,2}}$.*)*

Credits