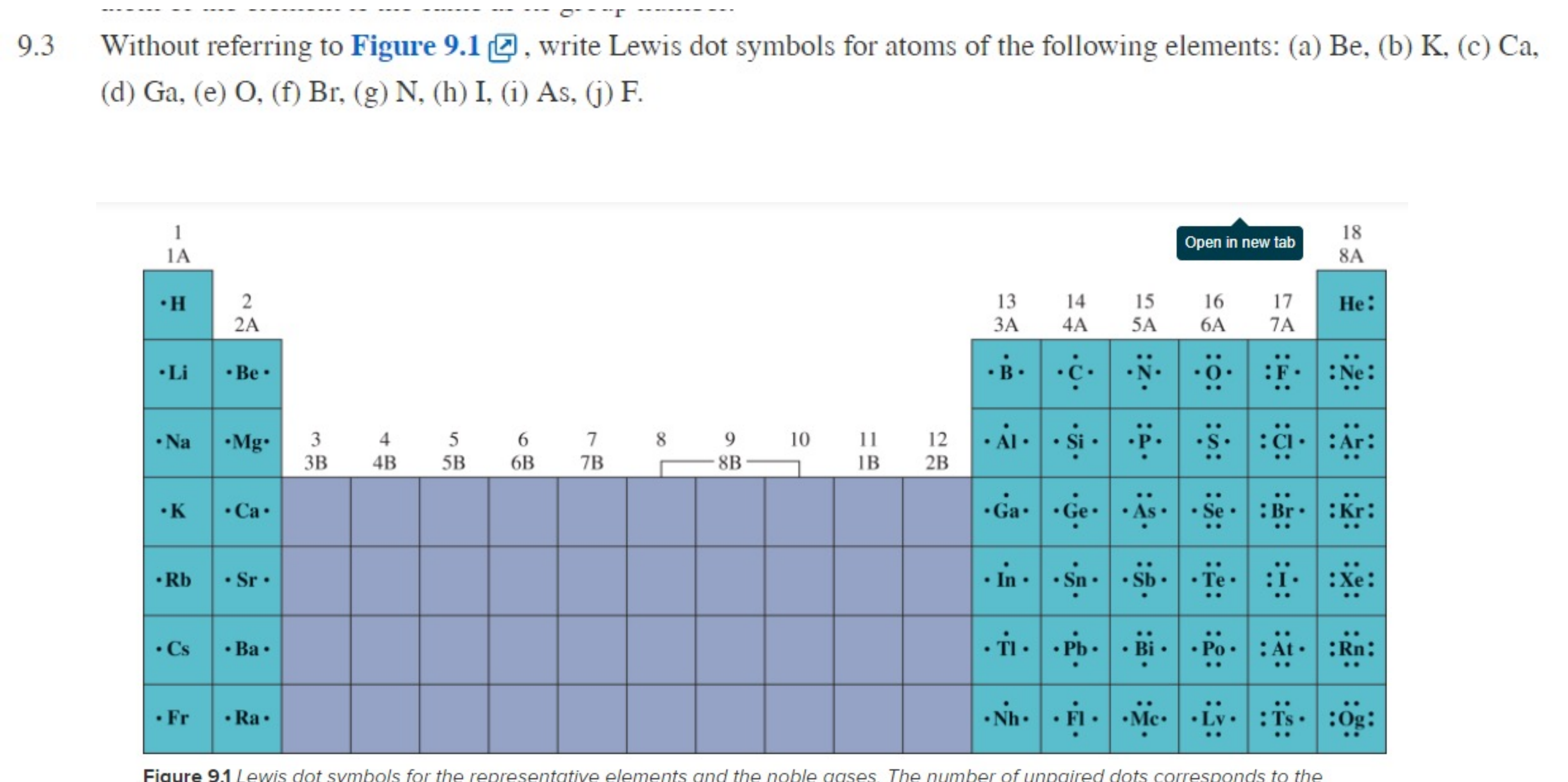
END OF TERM 3 REVISION MODULE

GRADE10-ELITE STREAM-
2021-22

تعليم

مؤسسة الإمارات للتعليم المدرسي
EMIRATES SCHOOLS ESTABLISHMENT

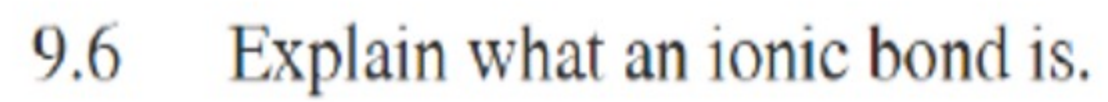
Objective: EOT Coverage:



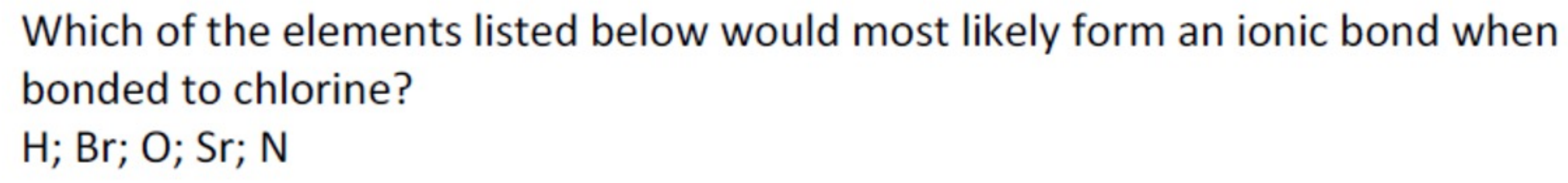
9.3 Without referring to **Figure 9.1**, write Lewis dot symbols for atoms of the following elements: (a) Be, (b) K, (c) Ca, (d) Ga, (e) O, (f) Br, (g) N, (h) I, (i) As, (j) F.

Figure 9.1 *Lewis dot symbols for the representative elements and the noble gases. The number of unpaired dots corresponds to the*

Objective: EOT Coverage:



9.6 Explain what an ionic bond is.

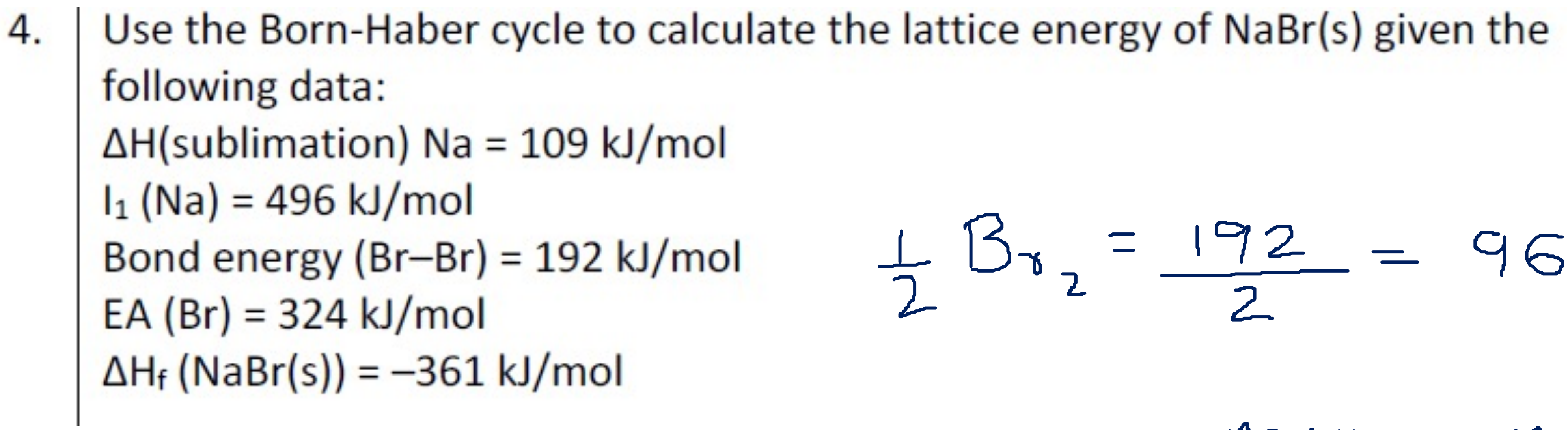


Which of the elements listed below would most likely form an ionic bond when bonded to chlorine?

H; Br; O; Sr; N

*I*onic bond, also called electrovalent bond, **type of linkage formed from the electrostatic attraction between oppositely charged ions in a chemical compound**. It is a bond between a metal and a non-metal. Sr is a metal. So it can make Ionic bond when combine with Cl

Objective: EOT Coverage:



4. Use the Born-Haber cycle to calculate the lattice energy of NaBr(s) given the following data:
ΔH(sublimation) Na = 109 kJ/mol
I_1 (Na) = 496 kJ/mol
Bond energy (Br–Br) = 192 kJ/mol
EA (Br) = 324 kJ/mol
ΔH_f (NaBr(s)) = –361 kJ/mol

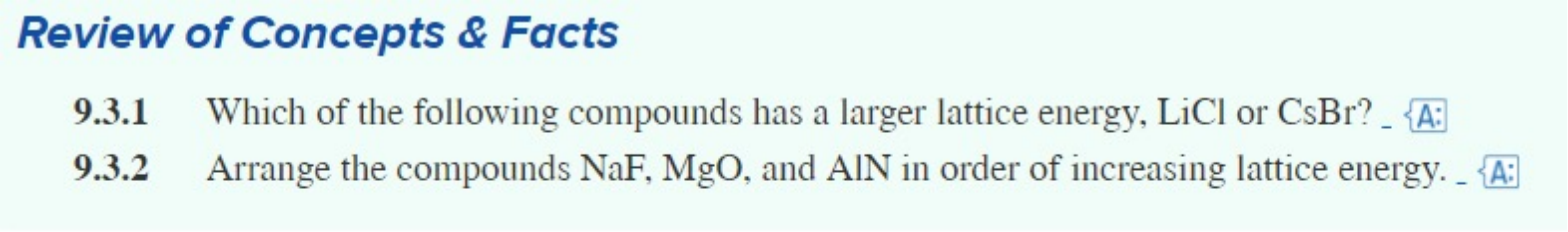
$$\frac{1}{2} Br_2 = \frac{192}{2} = 96$$

energy removed

Hf = Hsub + I1 + Bond dissociation energy + EA - Lattice Energy

-361 = 109 + 496 + 96 + (- 324) - Lattice energy
-361 = 377 – Lattice energy
Lattice Energy = 738 kj/mol.

Objective: EOT Coverage:



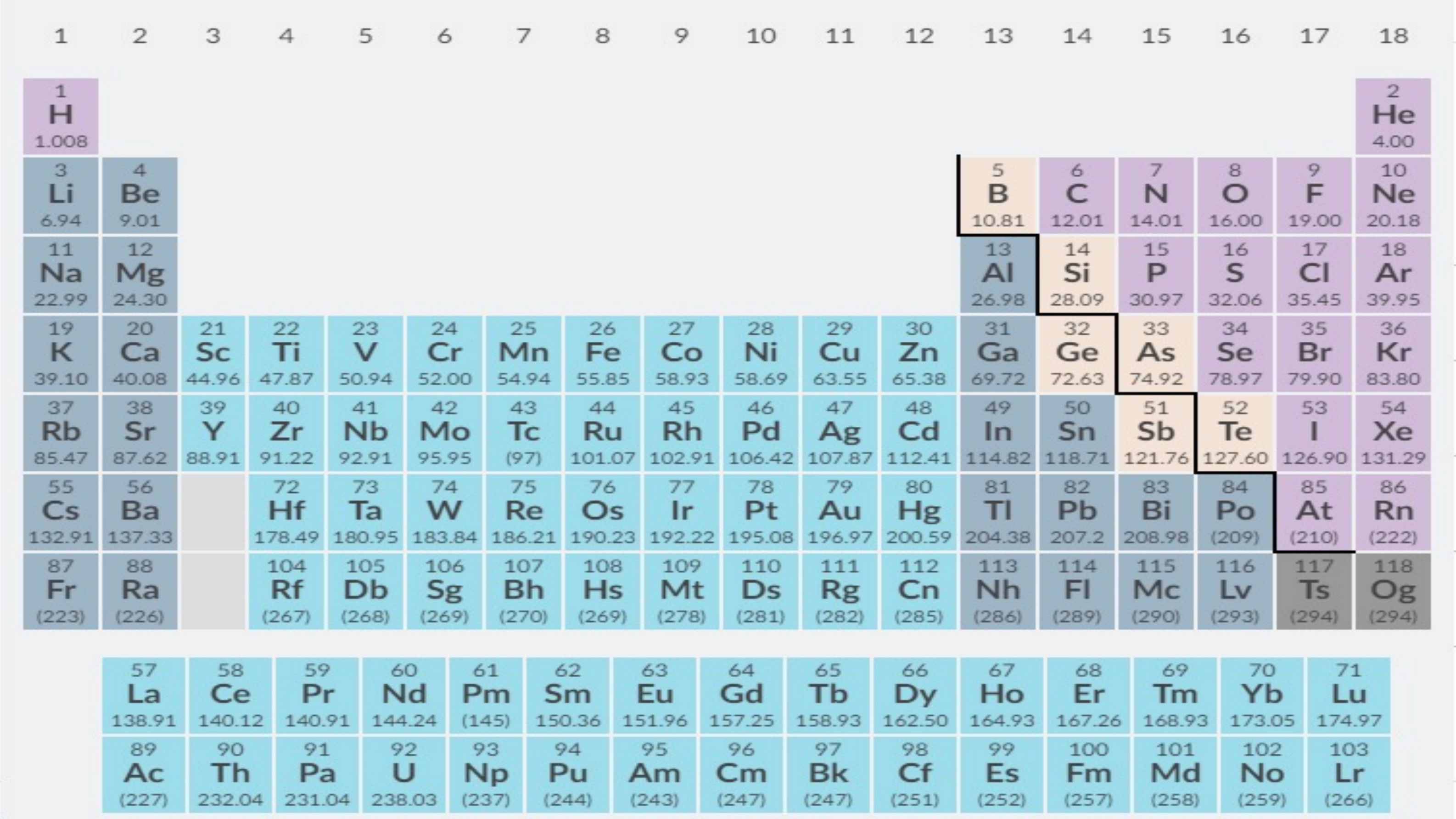
Review of Concepts & Facts

9.3.1 Which of the following compounds has a larger lattice energy, LiCl or CsBr? _

9.3.2 Arrange the compounds NaF, MgO, and AlN in order of increasing lattice energy. _

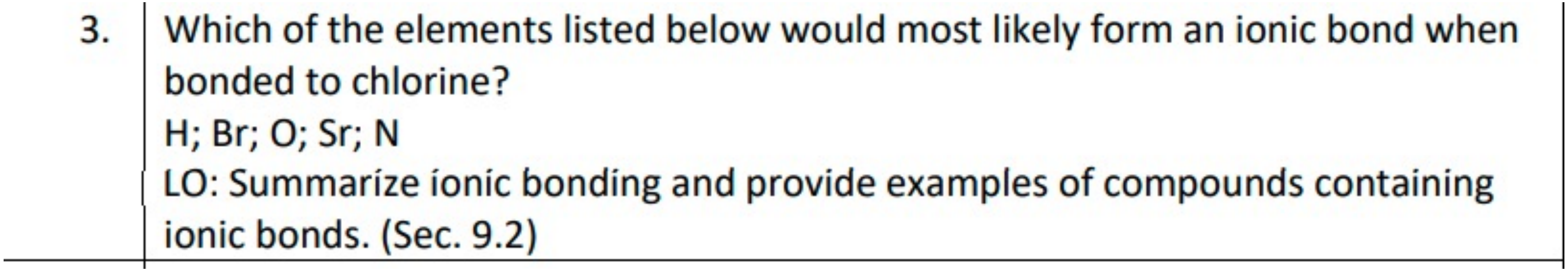
1. LiCl
2.

NaF < MgO < AlN.



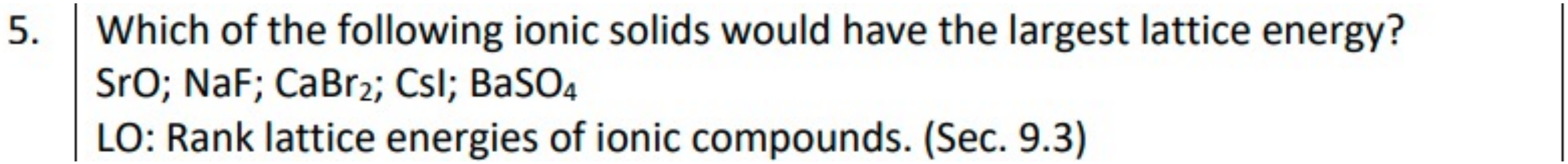
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.008																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc (97)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33		72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)		104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (282)	112 Cn (285)	113 Nh (286)	114 Fl (289)	115 Mc (290)	116 Lv (293)	117 Ts (294)	118 Og (294)

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (266)



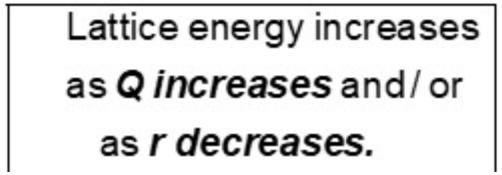
3. Which of the elements listed below would most likely form an ionic bond when bonded to chlorine?
H; Br; O; Sr; N
LO: Summarize ionic bonding and provide examples of compounds containing ionic bonds. (Sec. 9.2)

Sr Can make ionic bond with chlorine Since Sr is a metal



5. Which of the following ionic solids would have the largest lattice energy?
SrO; NaF; $CaBr_2$; CsI; $BaSO_4$
LO: Rank lattice energies of ionic compounds. (Sec. 9.3)

$$E = k\frac{Q_+Q_-}{r}$$

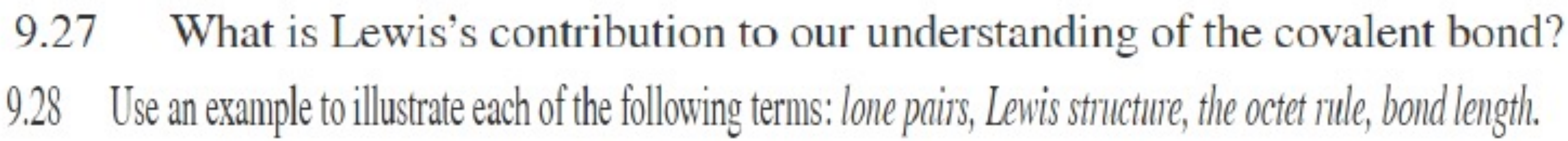


Lattice energy increases as ***Q increases*** and/ or as ***r decreases.***

Ca ,Sr and Ba are having 2+ as oxidation state, since the ionic radius of Ca is smaller than Ba and Sr the lattice enthalpy for CaBr2 is higher

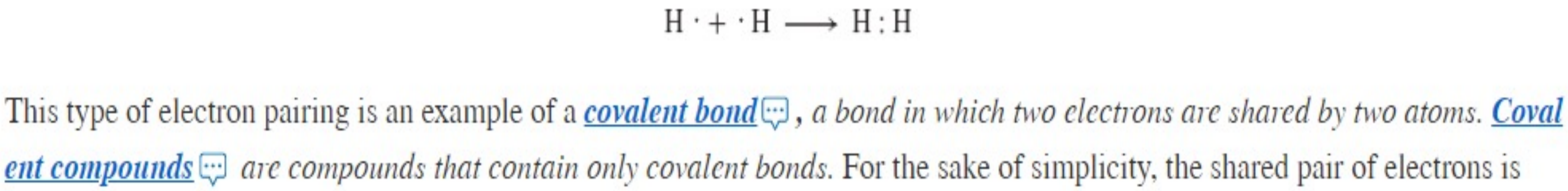
Objective: EOT Coverage:

6.



9.27 What is Lewis's contribution to our understanding of the covalent bond?

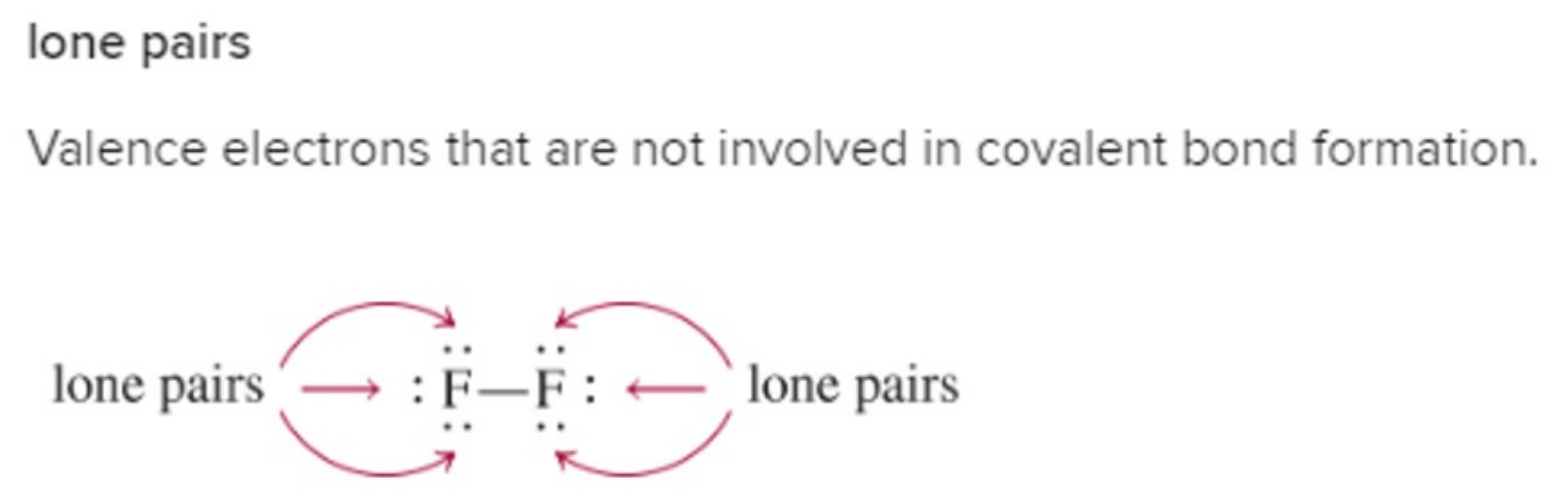
9.28 Use an example to illustrate each of the following terms: *lone pairs, Lewis structure, the octet rule, bond length.*



$$H\cdot + \cdot H \longrightarrow H:H$$

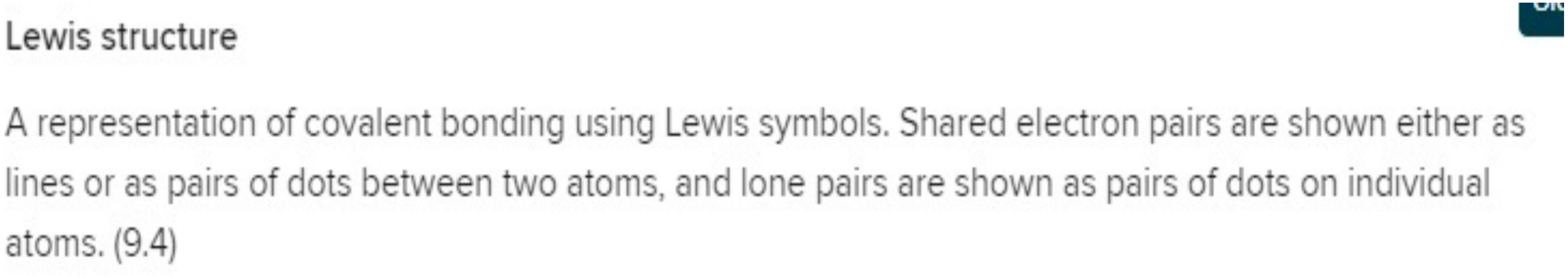
This type of electron pairing is an example of a ***covalent bond***, *a bond in which two electrons are shared by two atoms.* ***Covalent compounds*** *are compounds that contain only covalent bonds.* For the sake of simplicity, the shared pair of electrons is

Objective: EOT Coverage:



lone pairs

Valence electrons that are not involved in covalent bond formation.



Lewis structure

A representation of covalent bonding using Lewis symbols. Shared electron pairs are shown either as lines or as pairs of dots between two atoms, and lone pairs are shown as pairs of dots on individual atoms. (9.4)

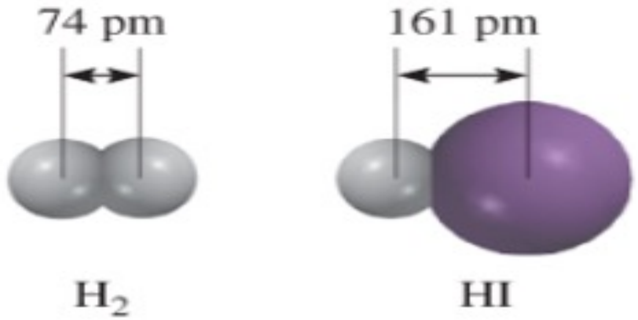
Objective: EOT Coverage:

octet rule

An atom other than hydrogen tends to form bonds until it is surrounded by eight valence electrons.

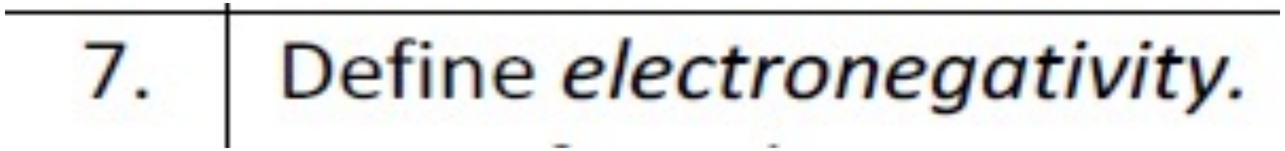
bond length

The distance between the nuclei of two bonded atoms in a molecule.



Multiple bonds are shorter than single covalent bonds.

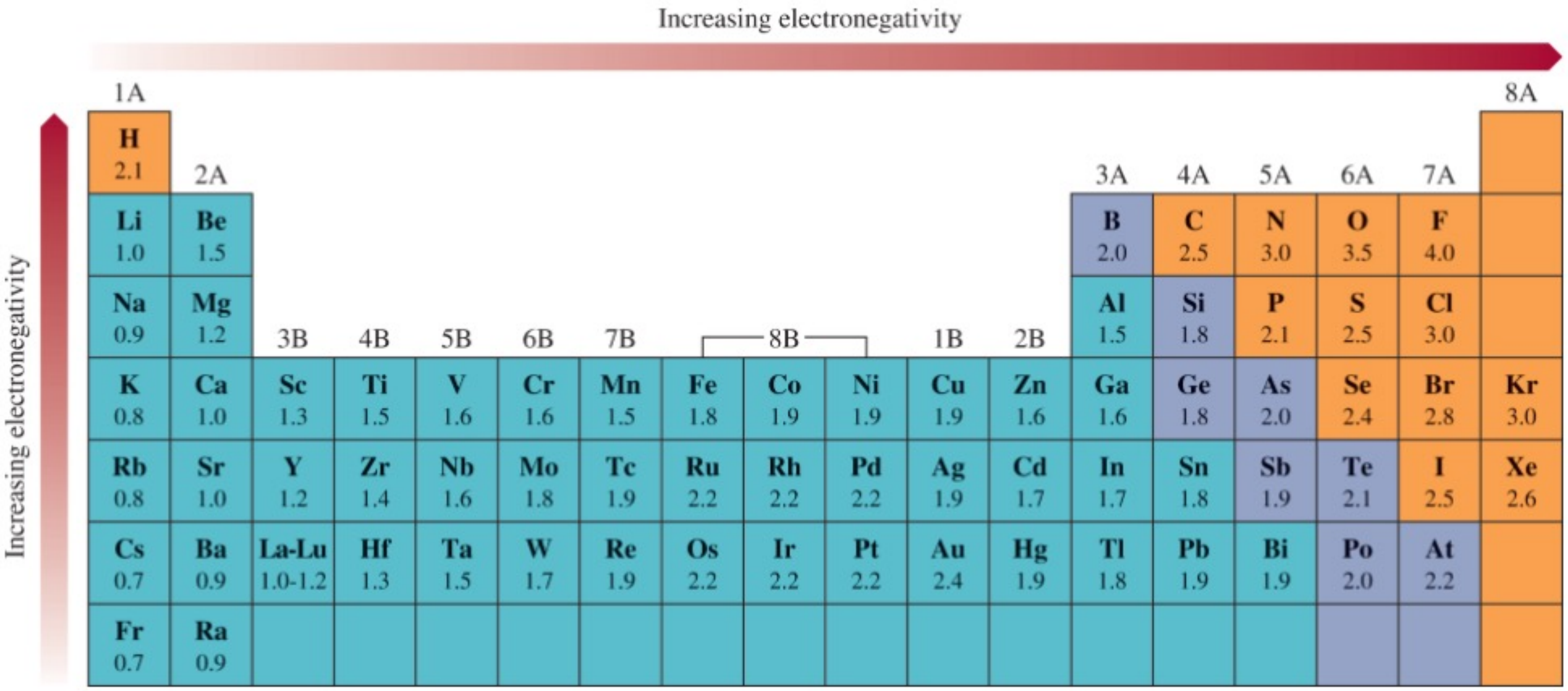
Objective: EOT Coverage:



7. Define *electronegativity.*

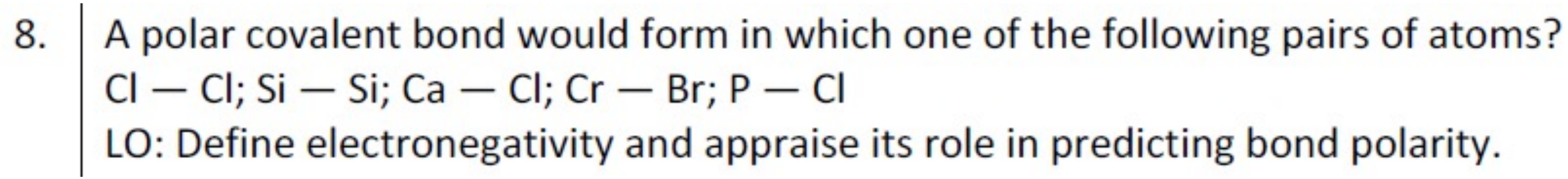
electronegativity

The ability of an atom to attract electrons toward itself in a chemical bond.



1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
H 2.1																	
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La-Lu 1.0-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	
Fr 0.7	Ra 0.9																

Objective: EOT Coverage:



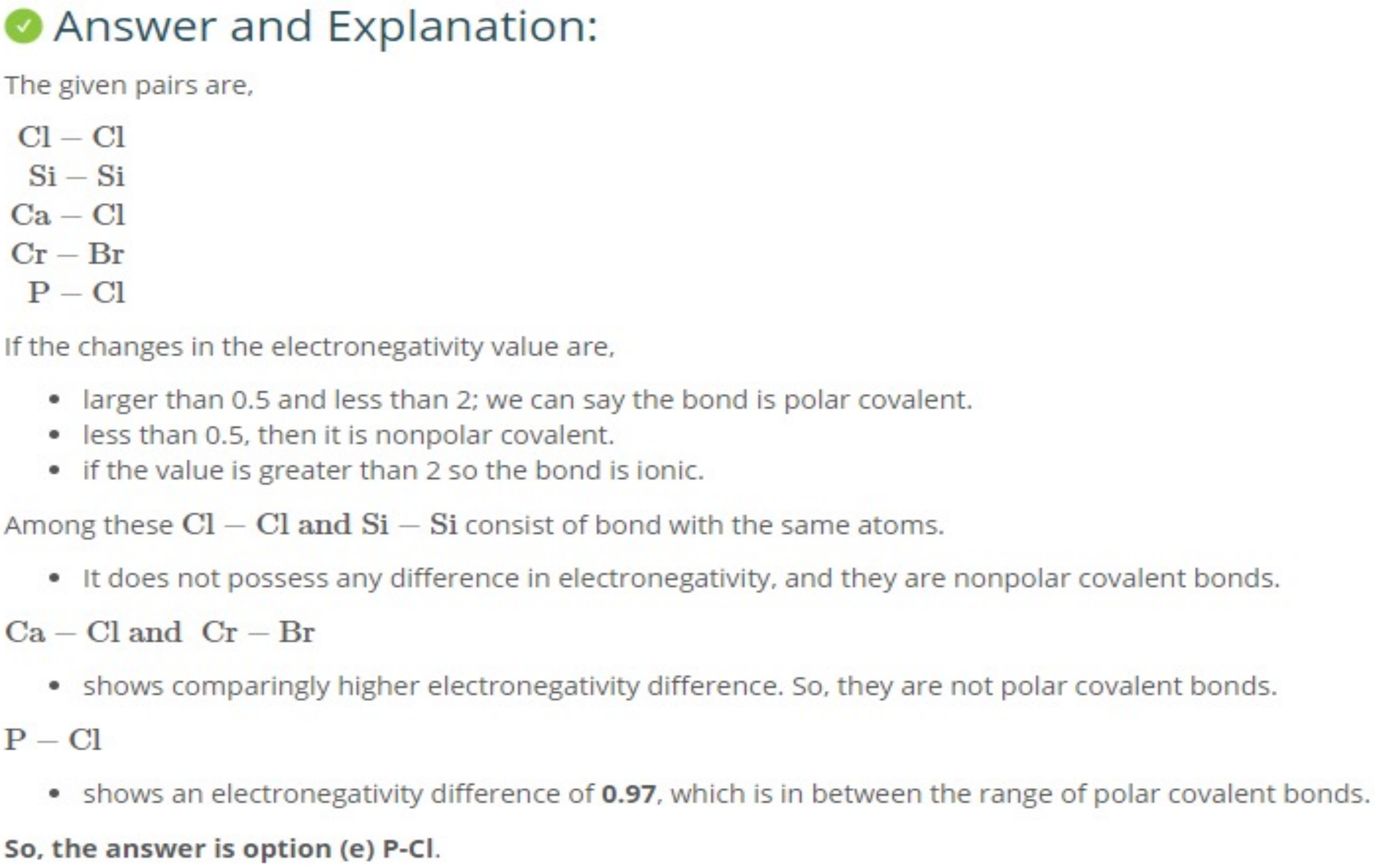
8. A polar covalent bond would form in which one of the following pairs of atoms?

Cl — Cl; Si — Si; Ca — Cl; Cr — Br; P — Cl

LO: Define electronegativity and appraise its role in predicting bond polarity.

A polar covalent bond is **a bond formed when a shared pair of electrons are not shared equally**. This is due to one of the elements having a higher electronegativity than the other. The shared pair of electrons between an atom of Phosphorous and an atom of bromine are not shared equally

Objective: EOT Coverage:



✓ Answer and Explanation:

The given pairs are,

$$\begin{array}{r} \mathrm{Cl-Cl} \\ \mathrm{Si-Si} \\ \mathrm{Ca-Cl} \\ \mathrm{Cr-Br} \\ \mathrm{P-Cl} \end{array}$$

If the changes in the electronegativity value are,

- larger than 0.5 and less than 2; we can say the bond is polar covalent.
- less than 0.5, then it is nonpolar covalent.
- if the value is greater than 2 so the bond is ionic.

Among these $\mathrm{Cl-Cl\ and\ Si-Si}$ consist of bond with the same atoms.

- It does not possess any difference in electronegativity, and they are nonpolar covalent bonds.

$\mathrm{Ca-Cl\ and\ \ Cr-Br}$

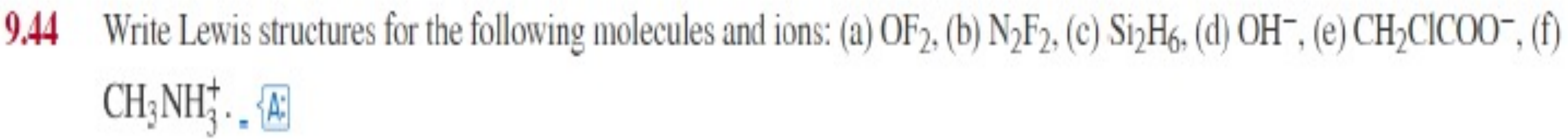
- shows comparingly higher electronegativity difference. So, they are not polar covalent bonds.

$\mathrm{P-Cl}$

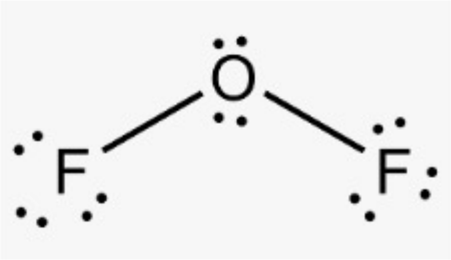
- shows an electronegativity difference of **0.97**, which is in between the range of polar covalent bonds.

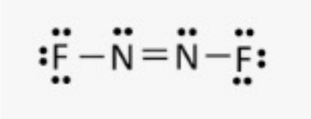
So, the answer is option (e) P-Cl.

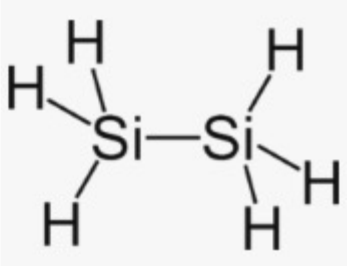
Objective: EOT Coverage:

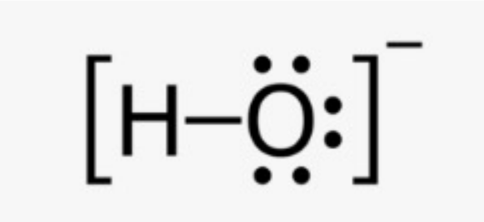


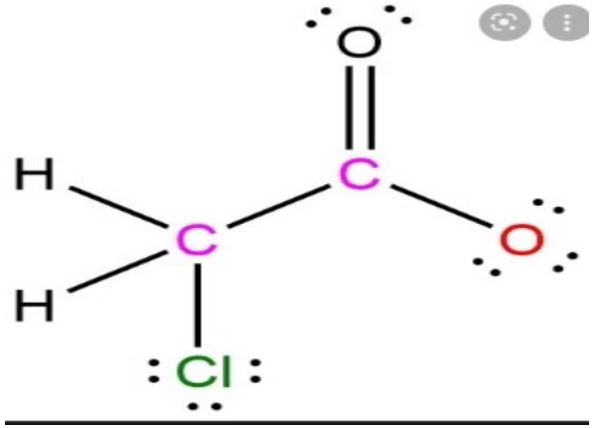
9.44 Write Lewis structures for the following molecules and ions: (a) OF_2, (b) N_2F_2, (c) Si_2H_6, (d) OH^-, (e) CH_2ClCOO^-, (f) $CH_3NH_3^+$.

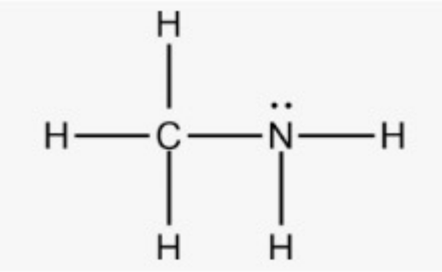




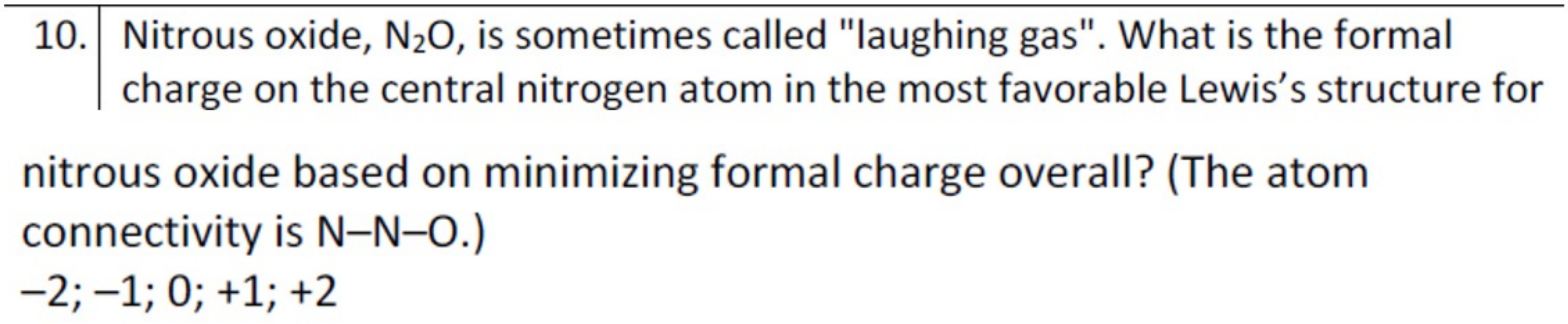








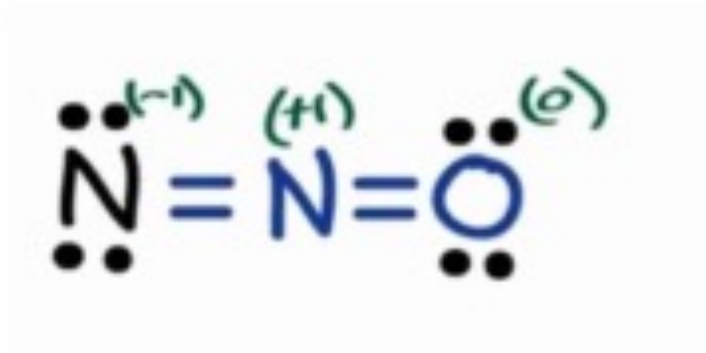
Objective: EOT Coverage:



10. Nitrous oxide, N_2O, is sometimes called "laughing gas". What is the formal charge on the central nitrogen atom in the most favorable Lewis's structure for nitrous oxide based on minimizing formal charge overall? (The atom connectivity is N–N–O.)

–2; –1; 0; +1; +2

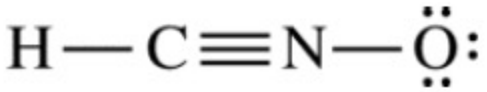
Formal charge = valence electrons— non bonded electrons—1/2 of bonded electrons.



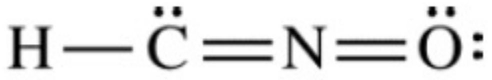
AP Question 2017

Answer the following questions about the isomers fulminic acid and isocyanic acid.

Two possible Lewis electron-dot diagrams for fulminic acid, HCNO, are shown below.



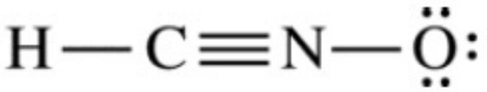
$H-C\equiv N-\underset{\cdot\cdot}{\overset{\cdot\cdot}{O}}:$



$H-\overset{\cdot\cdot}{C}=N=\overset{\cdot\cdot}{O}:$

Explain why the diagram on the left is the better representation for the bonding in fulminic acid. Justify your choice based on formal charges.

Formal Charge = Valence electrons - nonbonding electrons - bonds to the atom



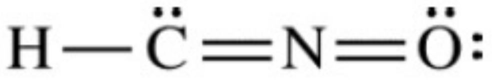
$H-C\equiv N-\underset{\cdot\cdot}{\overset{\cdot\cdot}{O}}:$

H: 1 - 0 - 1 = 0

C: 4 - 0 - 4 = 0

N: 5 - 0 - 4 = +1

$\boxed{\text{O: 6 - 6 - 1 = -1}}$



$H-\overset{\cdot\cdot}{C}=N=\overset{\cdot\cdot}{O}:$

H: 1 - 0 - 1 = 0

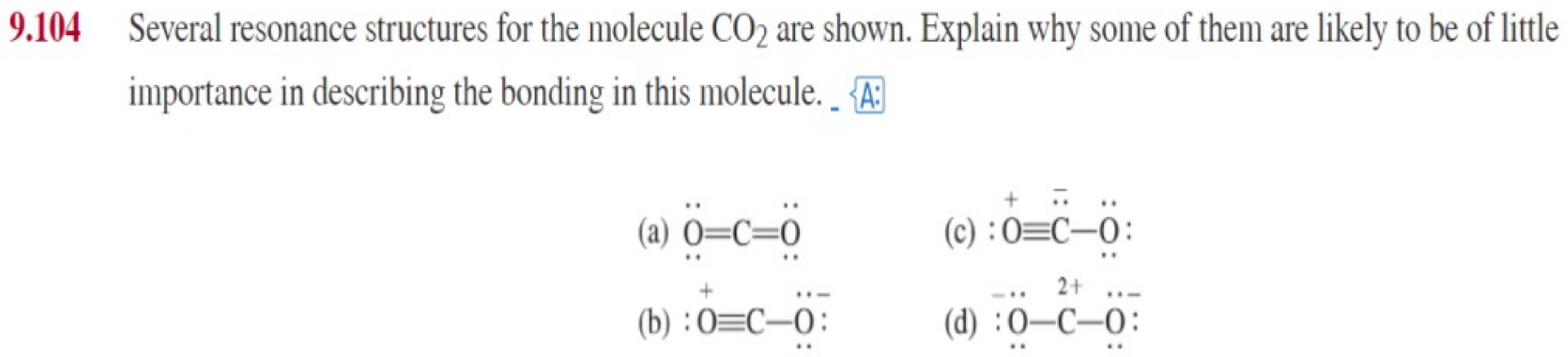
$\boxed{\text{C: 4 - 2 - 3 = -1}}$

N: 5 - 0 - 4 = +1

O: 6 - 4 - 2 = 0

The diagram on the left is the better representation of because it puts negative formal charge on O, which is more electronegative than C.

11.



9.104 Several resonance structures for the molecule CO_2 are shown. Explain why some of them are likely to be of little importance in describing the bonding in this molecule.

LO: Employ formal charges to identify the most likely structure of a compound when multiple Lewis structures are possible. (Sec. 9.7)

Calculate the formal charge.

Diagram (c) is unlikely because negative charge is on C which less electronegative than O.

Diagram (d) has large formal charges.

9.58 Draw two resonance structures for diazomethane, CH_2N_2. Show formal charges. The skeletal structure of the molecule is

LO: Evaluate the concept of resonance and draw resonance structures of a given compound or polyatomic ion. (Sec. 9.8)

(A)

(B)

C: 4 - 0 - 4 = 0

N: 5 - 0 - 4 = +1

N: 5 - 4 - 2 = -1

H: 1 - 0 - 0 = 0

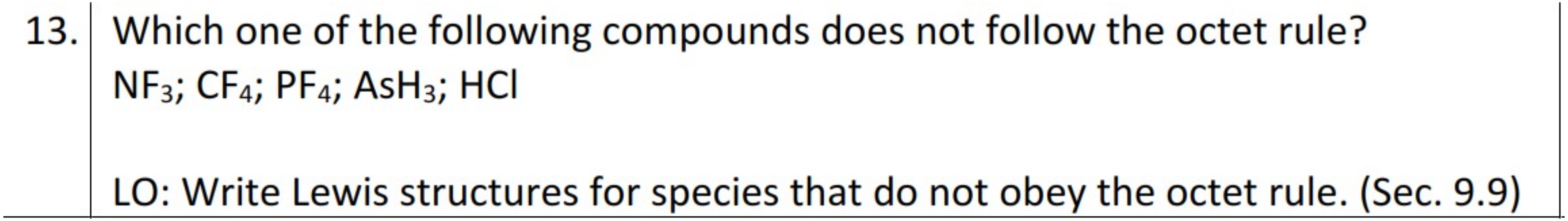
C: 4 - 2 - 3 = -1

N: 5 - 0 - 4 = +1

N: 5 - 2 - 3 = 0

H: 1 - 0 - 0 = 0

The diagram A is the better representation of because it puts negative formal charge on O, which is more electronegative than C.



13. Which one of the following compounds does not follow the octet rule?

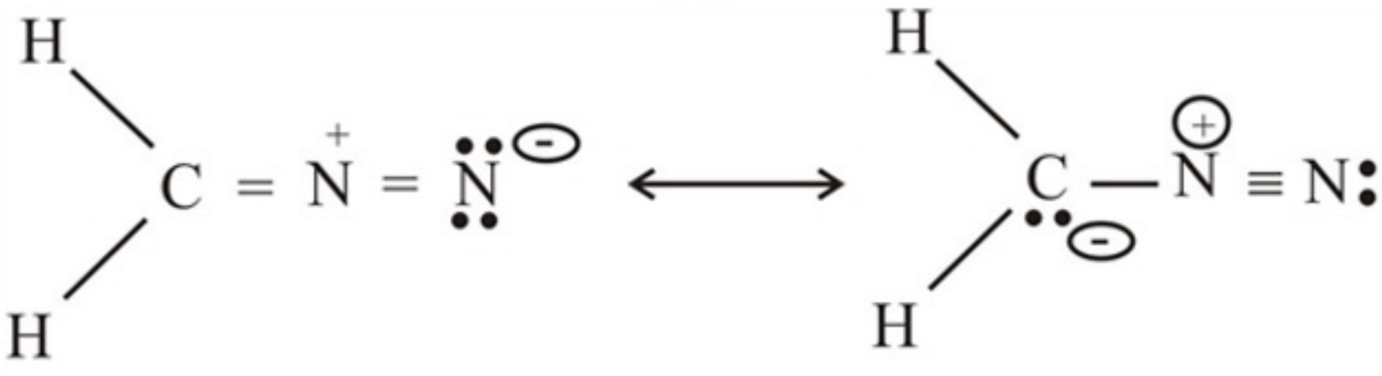
NF_3; CF_4; PF_4; AsH_3; HCl

LO: Write Lewis structures for species that do not obey the octet rule. (Sec. 9.9)

PF_4 does not follow octet rule, the central atom (P) has more than 8 valence electrons

Objective: EOT Coverage:

9.58 Draw two resonance structures for diazomethane, CH_2N_2. Show formal charges. The skeletal structure of the molecule is



(A)

(B)

Objective: EOT Coverage:

13. Which one of the following compounds does not follow the octet rule?

NF_3; CF_4; PF_4; AsH_3; HCl

PF4

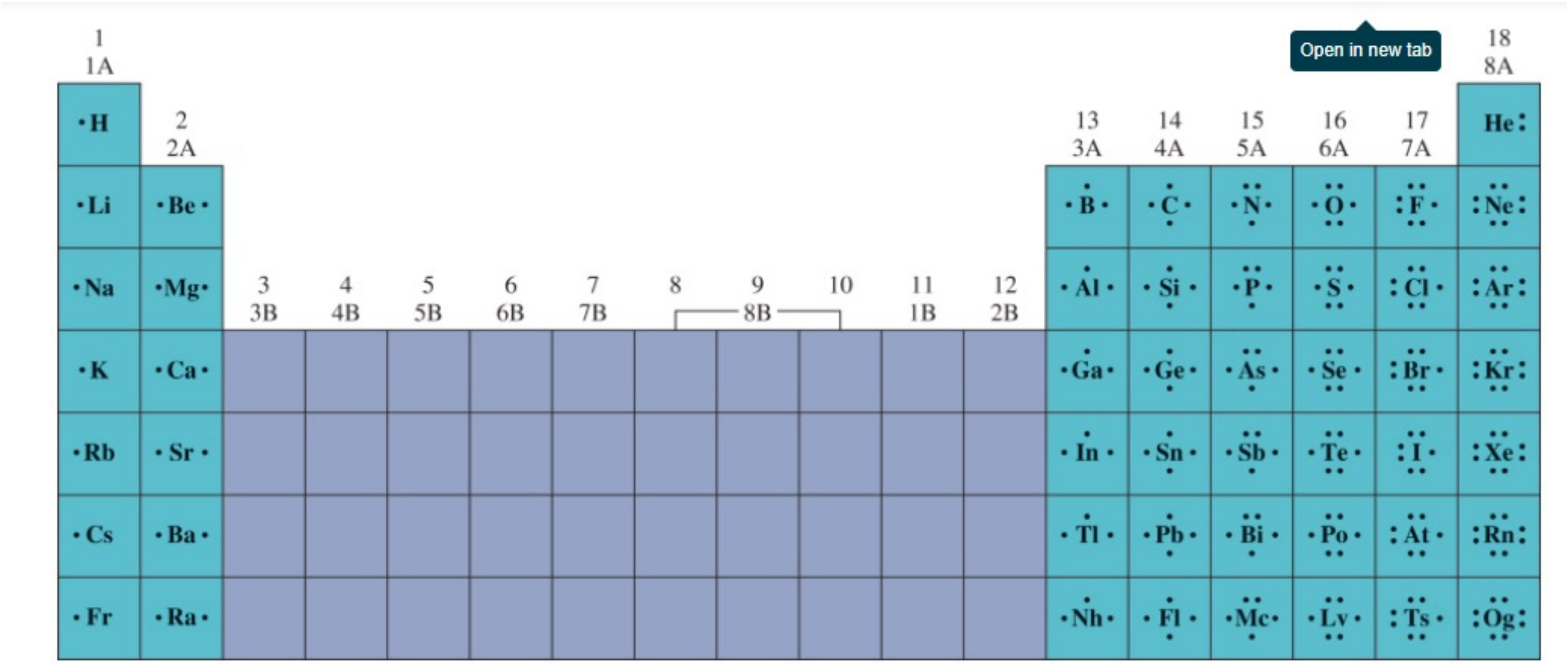
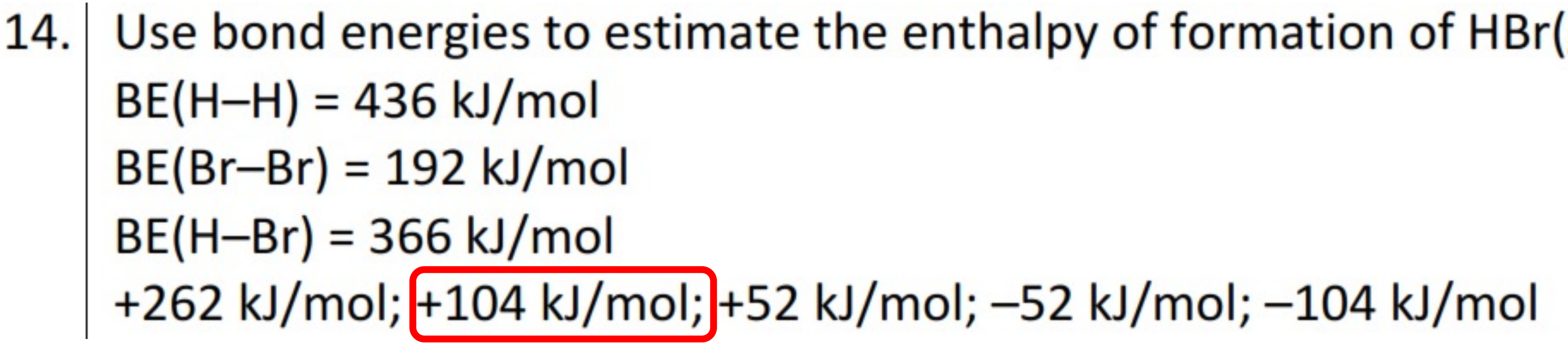


Figure 9.1 *Lewis dot symbols for the representative elements and the noble gases. The number of unpaired dots corresponds to the*



14. Use bond energies to estimate the enthalpy of formation of HBr(g).

BE(H–H) = 436 kJ/mol

BE(Br–Br) = 192 kJ/mol

BE(H–Br) = 366 kJ/mol

+262 kJ/mol; +104 kJ/mol; +52 kJ/mol; –52 kJ/mol; –104 kJ/mol

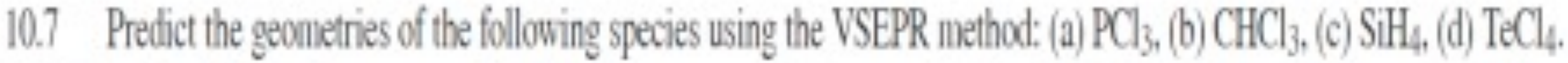
$$H_2(g) \quad + \quad Br_2(g) \rightarrow 2HBr(g)$$

$$\mathbf{\Delta H_r =} [\ 2(\text{H—Br})] - [(\text{H—H}) + (\text{Br—Br})]$$

$$\mathbf{=} \ [2(366)] - [(436) + (192)]$$

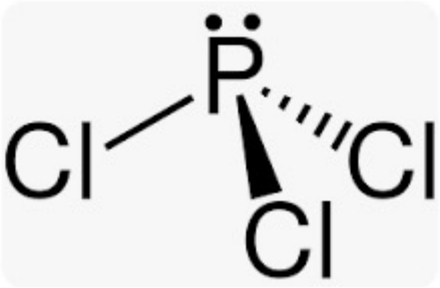
$$\mathbf{= \ +104\ kJ/mol}$$

15.


10.7 Predict the geometries of the following species using the VSEPR method: (a) PCl_3, (b) $CHCl_3$, (c) SiH_4, (d) $TeCl_4$.

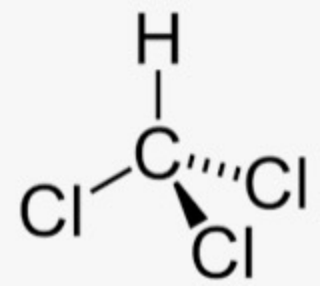
LO: Apply the valence-shell electron-pair repulsion (VSEPR) model to predict the shape of a molecule or polyatomic ion. (Sec 10.1)

p.454
Questions &
Problems 10.7

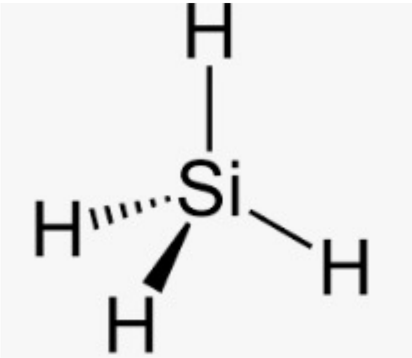


Trigonal Pyramidal

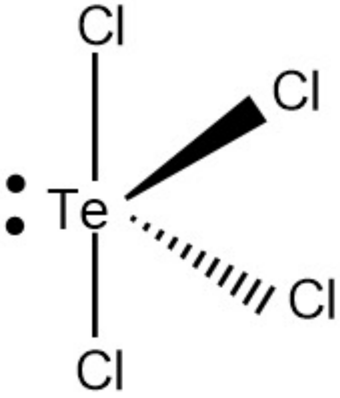
P= 5 ve
Bonding with 3 Cl
So 5+3=8/2= 4pairs
3 bonding pairs and 1 lone pair
For 4 pair of electron the regular geometry is tetrahedral, since one lone pair is present the shape changes to trigonal.



Tetrahedral
C= 4ve
Bonding with 4 other atom.
So, 4+4 = 8/2=4 pairs
All four pairs are bonding.
So its having a regular tetrahedral geometry.



Tetrahedral
Si= 4ve
Bonding with 4 other atom.
So, 4+4 = 8/2=4 pairs
All four pairs are bonding.
So its having a regular tetrahedral geometry.



Sea Saw
Te= 6ve
So, 6+4= 10/2=5 pairs
4 bonding and 1 lone pair. For a molecule with 5 pair of electron is trigonal bipyramidal since a lone pair is present it changes into Sea Saw

16.	**10.14** Which of the following species are tetrahedral? $SiCl_4$, SeF_4, XeF_4, CI_4, $CdCl_4^{2-}$ LO: Predict deviations from ideal bond angles in structures based on presence of lone pairs on a central atom. (Sec 10.1).	p.454 Questions & Problems 10.14

$SiCl_4$

$Si = 4ve + 4 = \frac{8}{2} = 4$ pairs

All are bonding pairs.

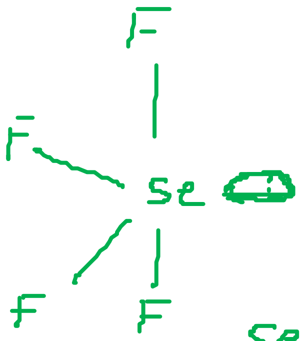
Shape – $\boxed{\text{Tetrahedral}}$

SeF_4

$Se = 6 + 4 = \frac{10}{2} = 5$ pairs

4 - bonding

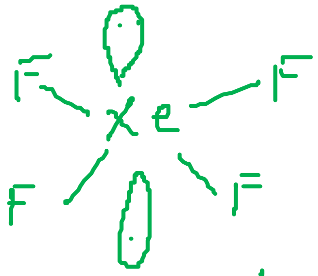
1 - lone pair



Sea saw

XeF_4

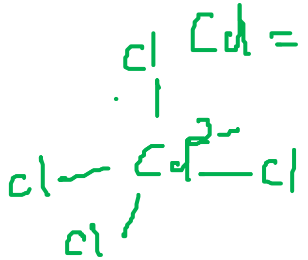
$8 + 4 = \frac{12}{2} = 6$



Square planar

$CdCl_4^{2-}$

$Cd = 2+4+2 = \frac{8}{2} = 4P$



$\boxed{\text{Tetrahedral.}}$

$\boxed{SiCl_4 \text{ and } CdCl_4^{2-}}$

17.	The bond angle in Cl_2O is expected to be approximately 90°; 109.5°; 120°; 145°; 180°. LO: Predict deviations from ideal bond angles in structures based on presence of lone pairs on a central atom. (Sec 10.1).	Ch.10 Questions bank #38

Cl_2O

Central atom 'O'

$O = 6 + 2 = \frac{8}{2} = 4$ pairs

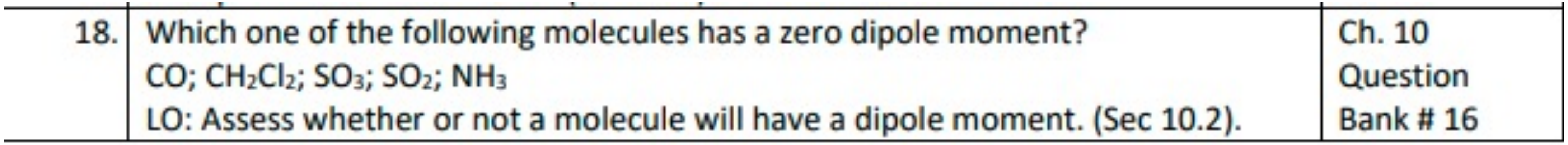
2 → bonding

2 → lone pair

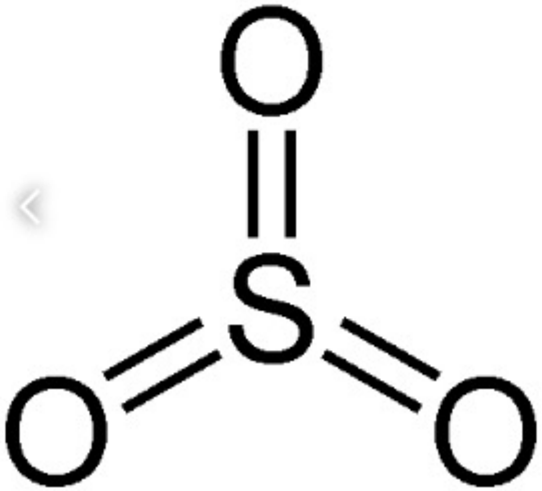
Since 4 pairs of $\bar{e}$ are present its having tetrahedral geometry with 109.5° [regular angle]

For one lone pair subtract 2.5 from regular angle. In reality the angle will be $109.5 - 5$

$= 104.5$



18.	Which one of the following molecules has a zero dipole moment? CO; CH_2Cl_2; SO_3; SO_2; NH_3 LO: Assess whether or not a molecule will have a dipole moment. (Sec 10.2).	Ch. 10 Question Bank # 16



All three S—O bonds are polar, but because the molecule is **symmetrical** the dipoles cancel out one another. The molecule is **non-polar**.

Predicting the polarity of a molecule

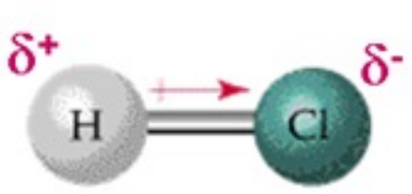
1. Draw the Lewis structure and determine the molecular geometry (shape).

2. Determine whether the bonds in the molecule are polar.

    - If there are no polar bonds, the molecule is non-polar.

3. Determine whether the molecule is symmetrical about all axis. If so, the molecule is non-polar. If not, it is polar.

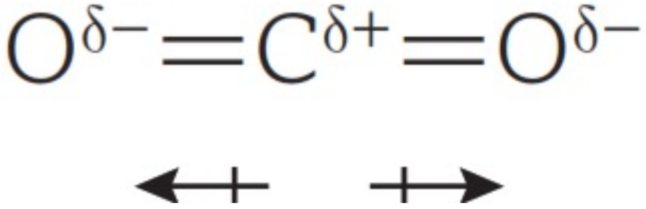
Example 1: Hydrogen chloride



The following symbol is used to represent dipole: +→

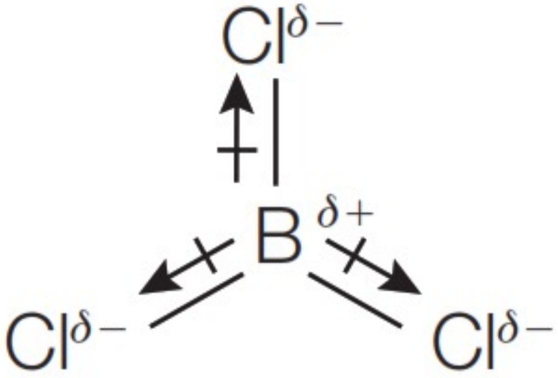
Since this is the only polar bond in the molecule, the molecule itself is polar.

Example 2: carbon dioxide, CO_2



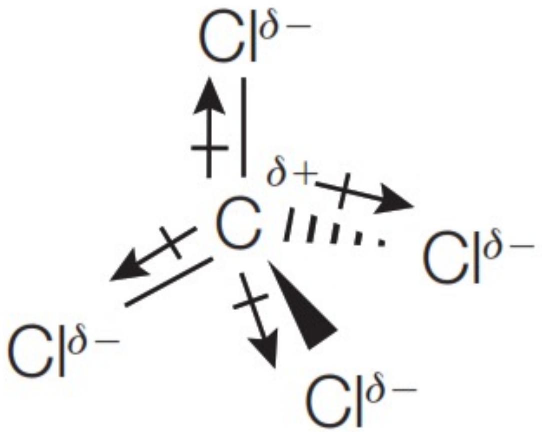
Both bonds in the carbon dioxide molecule are polar, but because the molecule is **symmetrical** the dipoles cancel out one another. The molecule is **non-polar**.

Example 3: Boron trichloride, BCl_3



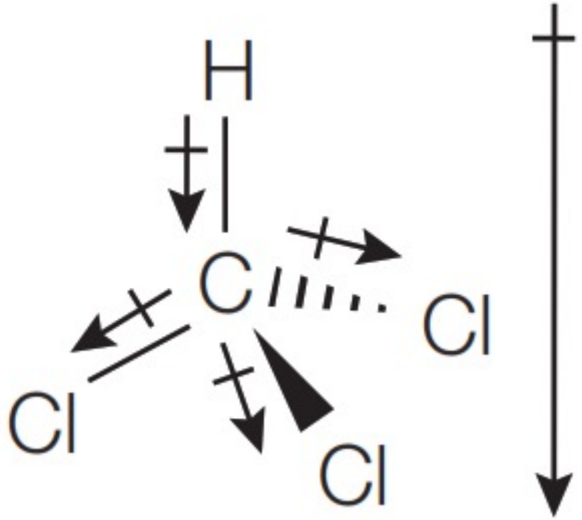
All three B—Cl bonds are polar, but because the molecule is **symmetrical** the dipoles cancel out one another. The molecule is **non-polar**.

Example 4: carbon tetrachloride, CCl_4



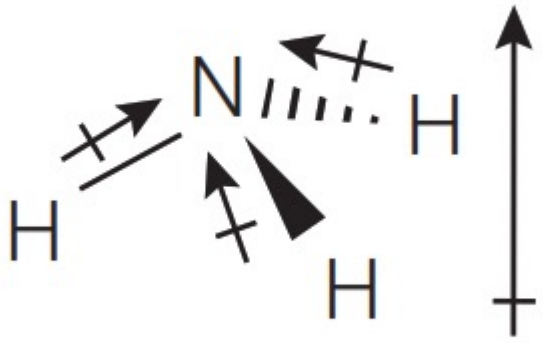
All four C—Cl bonds are polar, but because the molecule is **symmetrical** the dipoles cancel out one another. The molecule is **non-polar**.

Example 5: trichlormethane, CH_3Cl



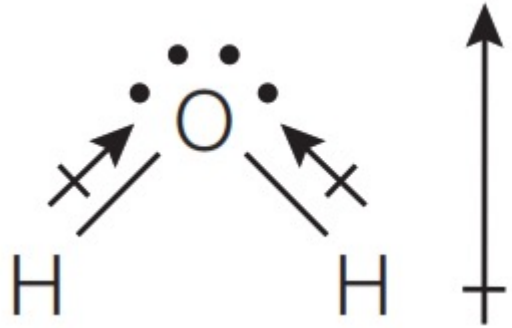
The dipoles reinforce one another and so the molecule is **polar**

Example 6: Ammonia, NH_3



All three N—H bonds are polar and the dipoles reinforce one another. The molecule is **polar**.

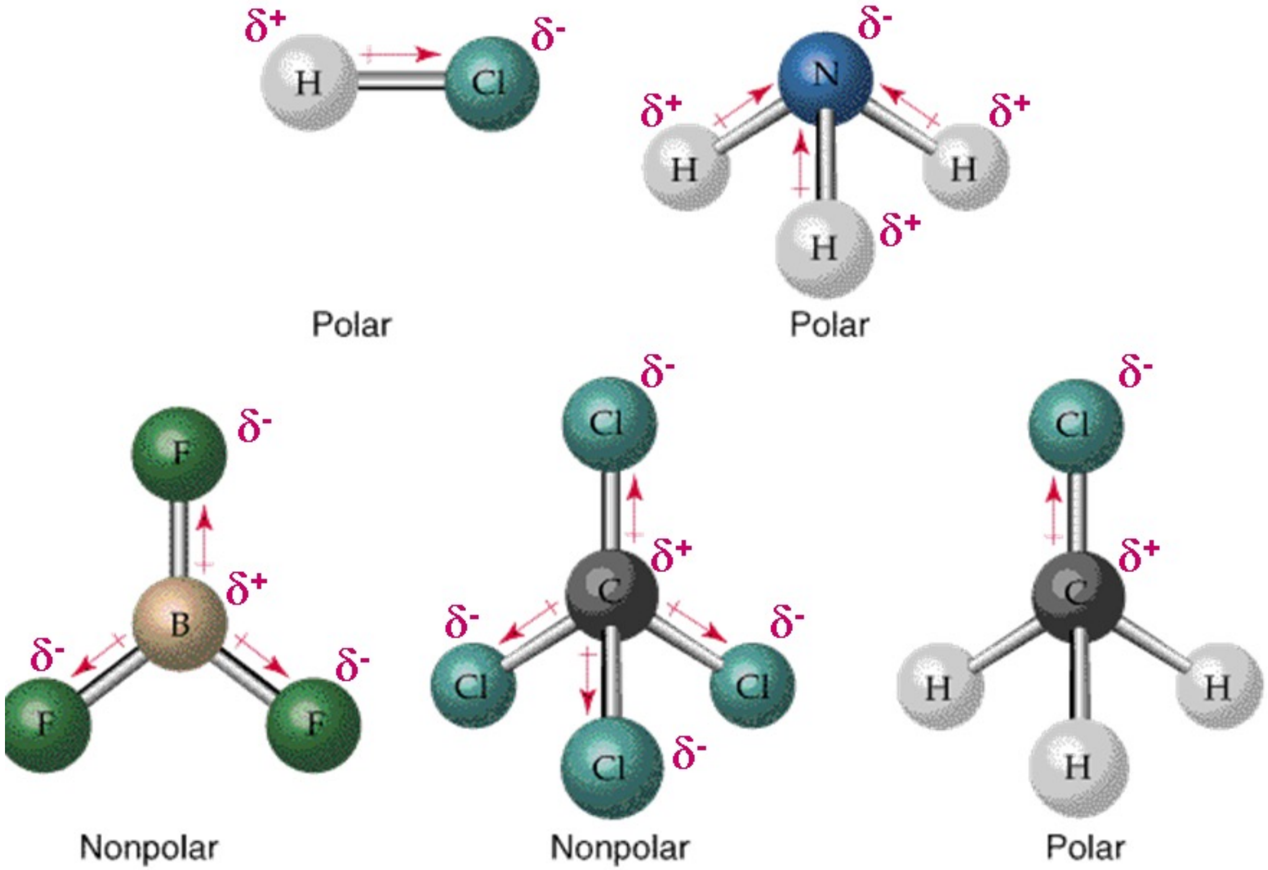
Example 7: Water, H_2O



Both O—H bonds are polar and the dipoles reinforce one another. The molecule is **polar**.

Polar molecules

- If the dipoles in a molecule **cancel** – usually when there is symmetry in the arrangement of the polar bonds, then the molecule will not be polar.



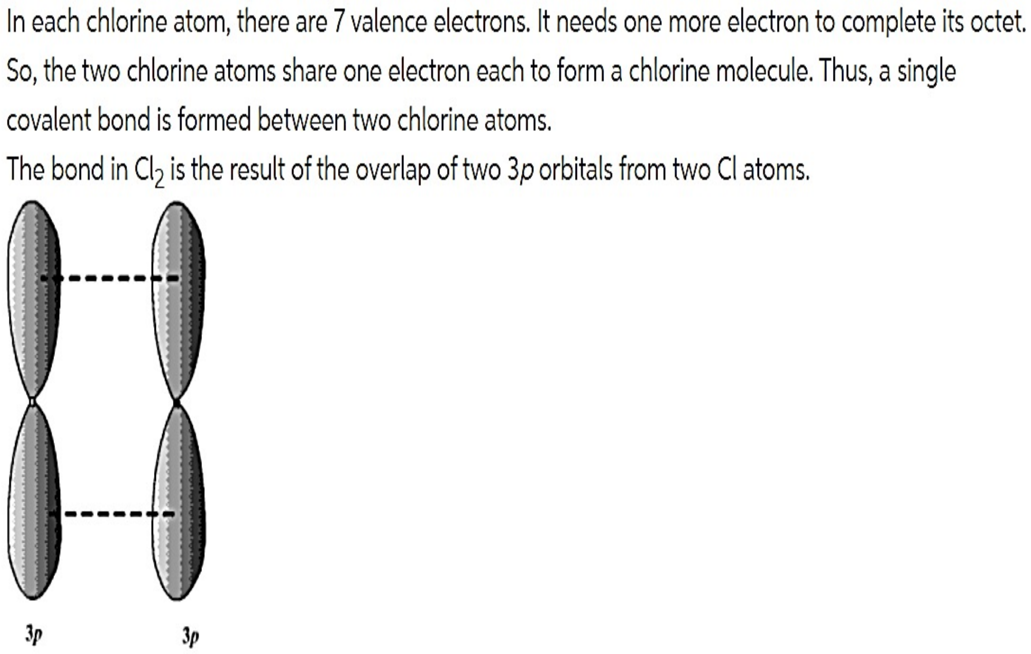
A polar molecule must have a defined dipole.

Does carbon dioxide have polar bonds, and if so, is it a polar molecule?

L.O: Employ valence bond theory to describe the bonding in covalent molecules. (Sec. 10.3)

Q. **Use valence bond theory to explain the bonding in Cl2 and HCl. Show how the atomic orbitals overlap when a bond is formed.**

A single covalent bond is formed sigma bond formed by the overlap of s orbital of H and p orbital of Cl



In each chlorine atom, there are 7 valence electrons. It needs one more electron to complete its octet. So, the two chlorine atoms share one electron each to form a chlorine molecule. Thus, a single covalent bond is formed between two chlorine atoms.

The bond in Cl_2 is the result of the overlap of two 3*p* orbitals from two Cl atoms.

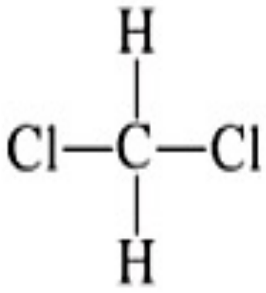
20. LO: Determine the hybridization of an atom in a molecule or polyatomic ion. (Sec. 10.4)

What is the hybridization of Si in SiH_4 and $H_3Si\text{-}SiH_3$?

SiH_4 has 4 atoms bonded to the central Si atom (4 single covalent Si-H bonds) and no lone pair in its Lewis structure . Since the electron group is ‘4’ silicon is sp3 hybridized.

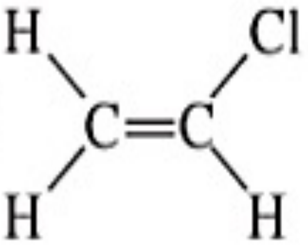
In $H_3Si\text{-}SiH_3$ each Si atom is bonded to 4 other atoms without any lone pairs so each Si atom is sp3 hybridized. One s and 3 p orbitals hybridize to form 4 new sp3 orbitals.

Q21. How many sigma and pi bonds are there in each of the following molecules?

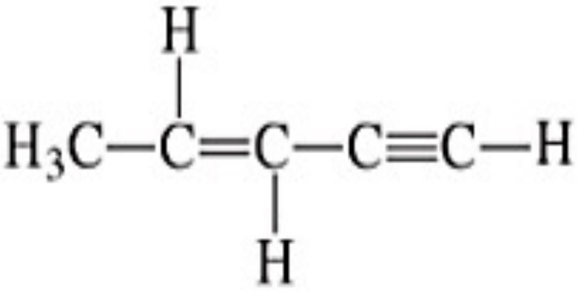


$\mathrm{Cl-\overset{H}{\underset{H}{C}}-Cl}$

(a)



(b)



$\mathrm{H_3C-\overset{H}{C}=\underset{H}{C}-C\equiv C-H}$

(c)

a) There are 4 single bonds . Hence, 4 sigma bonds only.

b) There are 5 single bonds and 1 double bond. Hence 5 sigma bonds and 1 pi bond.

c)There are 10 sigma bonds and 3 pi bonds

LO: Explain how atomic orbitals overlap to form molecular orbitals according to molecular orbital theory. (Sec. 10.6).

Q22. According to Molecular Orbital Theory, two separate 1s orbitals interact to form what molecular orbital(s)?

σ only

Option b

σ and σ*

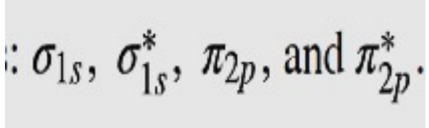
π only

π and π*

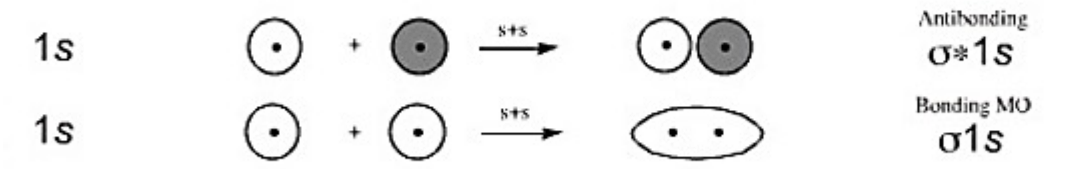
σ and π

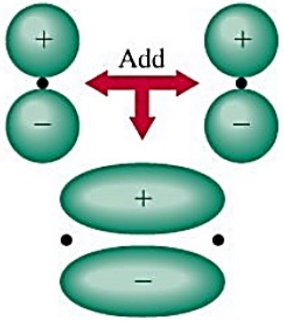
LO: Identify bonding and antibonding molecular orbitals.

- Q23. Sketch the shapes of the following molecular orbitals and how do their energies compare ?

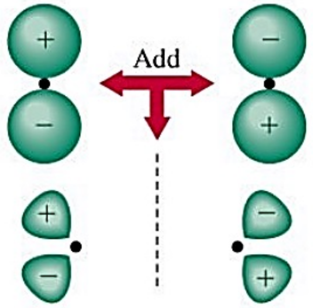


: σ_{1s}, σ^*_{1s}, π_{2p}, and π^*_{2p}.



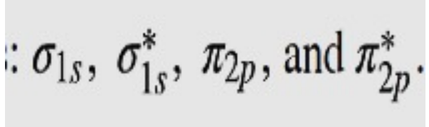


π_{2p}



π^*_{2p}

The energy is lowest for sigma 1s and highest for pi *2p.
Pi bond has higher energy and weaker than sigma bond. The order of increasing energies is



: σ_{1s}, σ^*_{1s}, π_{2p}, and π^*_{2p}.

LO: Draw Lewis structures with and without expanded octets for species where both are possible (Sec. 9.9).

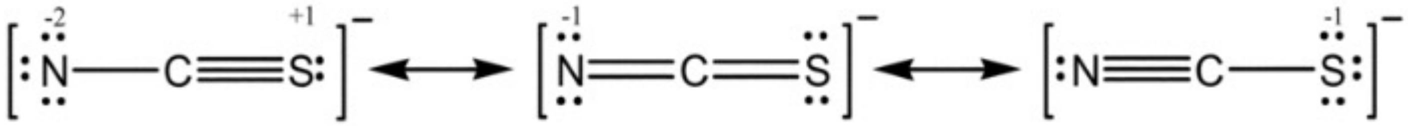
- Q24. Which of the elements listed below is most likely to exhibit an expanded octet in its compounds? O; S; Na; C; N

- Ans : S as it is in 3rd period and due to the availability of d orbitals it can accommodate more than 8 electrons.

LO: Evaluate the concept of resonance and draw resonance structures of a given compound or polyatomic ion. (Sec. 9.

- Q25. Define the terms resonance and resonance structure.
- Draw the resonance for NCS- and predict the stable structure.

- Ans : Whenever a single Lewis structure cannot describe all the properties of a molecule a number of structures with similar energy are written to describe all the properties. The actual structure is in between of all these contributing structures and is known as the resonance hybrid. The individual structures are resonance structures and the phenomenon is resonance.



II structure is most stable because N- can hold negative charge as it is more electronegative than Sulphur