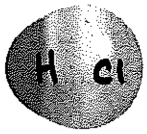
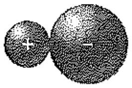
Key

1. The following electrostatic potential diagrams represent H_2, HCL, or NaCl. Label each and explain your choices.



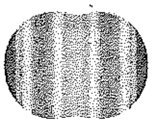
a

HCl This diagram represents a polar Covalent Bond



b

NaCl This diagram represents an ionic Bond. The Electronegicity differences Btwn NaCl are so great that valence e is transfued from Na to Cl



c

H_2 This diagram represents a pure covalent Bond

2. Describe the type of bonding that exists in the F_2(g) molecule. How does this type of bonding differ from that found in the HF(g) molecule? How is it similar?

F_2 is a pure Covalent Bonding. In HF there is also shared pair of Bonding electrons, But the shared pair is drawn more closely to the F atom. HF is polar Covalent.

3. When comparing the size of different ions, the general radii trend is usually not very useful. What do you concentrate on when comparing sizes of ions to each other or when comparing the size of an ion to its neutral atom?

For Ion's concentrate on the # of protons & # of Electrons present.

Atoms whose Nucleus has a greater amount of P to E holds the E more tightly will Be smallest.
Anion are larger than the neutral Atom.
Cation are Smaller than the neutral Atom

4. Predict which bond in each of the following groups will be post polar.
   a. C-F, Si-F, Ge-F
   b. P-Cl or S-Cl
   c. S-F, S-Cl, S-Br
   d. Ti-Cl, Si-Cl, Ge-Cl

* The most polar Bond will have the greatest difference in electronegativity Btwn the 2 Atoms.

* Using Position in Periodic table - farest Apart ⇒ most polar

Key

5. Predict the type of bond one would expect to form between the following pairs of elements.

   a. Rb and Cl

   Ionic

   b. S and S

   Covalent

   c. C and F

   Polar Covalent

   d. Ba and S

   Ionic

   e. N and P

   3.0 - 2.1 = .9

   Polar Covalent

   f. B and H

   2.0 - 2.1 = .1

   Covalent

6. Hydrogen has an electronegativity value between boron and carbon and identical to phosphorous. With this is mind, rank the following bonds in order of decreasing polarity: P-H, O-H, N-H, F-H, C-H

   F-H > OH > N-H > C-H > P-H

Ionic Bonds

- Any compound that conducts an electric current when melted will be classified as ionic.
- Composed of elements with very large differences in electronegativity
- Usually made up of a metal and a non metal OR metal and polytomic ion
- **Lattice energy** – energy that is released when two ions of an ionic compound come together to form a crystal
- Higher the charges the greater the attractive energy
- Greater the distance, the smaller the attractive energy
  - The change in energy that takes place when separated gaseous ions are packed together to form an ionic solid.

$$\text{Lattice energy} = k\left(\frac{Q_1Q_2}{r}\right)$$

k = proportionality constant
Q_1 and Q_2 = charges on the ions
r = shortest distance between the centers of the cations and anion

7. Which compound in each of the following pairs of ionic substances has the most exothermic lattice energy? Justify your answer.

   a. NaCl, KCl NaCl Na^+ is smaller than K^+

   b. LiF, LiCl LiF, F^- is smaller than Cl^-

   c. $Mg(OH)_2$, MgO MgO, O^{-2} has a greater charge than OH^-

   d. $Fe(OH)_2$, $Fe(OH)_3$ $Fe(OH)_3$, Fe^{+3} has a greater charge than Fe^{+2}

   e. NaCl, Na_2O Na_2O, O^{-2} has a greater charge than Cl^-

   f. MgO, BaS MgO, Both ions are smaller in MgO

Key

8. The lattice energies of $FeCl_3$, $FeCl_2$, and Fe_2O_3 are (in no particular order) -2631, -5359, -14,774 kJ/mol. Match the appropriate formula to each lattice energy. Explain.

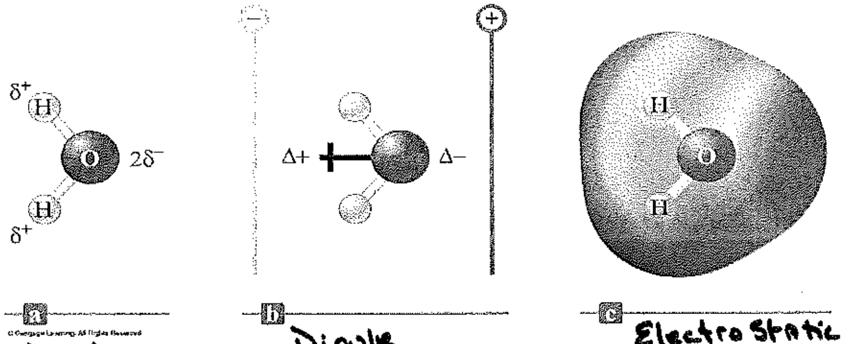
	QQ	Lattice Energy
$FeCl_2$	$(+2)(-1) = -2$	-2631 KJ/mol
$FeCl_3$	$(+3)(-1) = -3$	-5359 KJ/mol
Fe_2O_3	$(+3)(-2) = -6$	-14 774 KJ/mol

more Negative Lattice energy ⇒ more stable compound

Dipole Moment

- Property of a molecule whose charge distribution can be represented by a center of positive charge and a center of negative charge.
- Use an arrow to represent a dipole moment.
  - Point to the negative charge center with the tail of the arrow indicating the positive center of charge.

Examples



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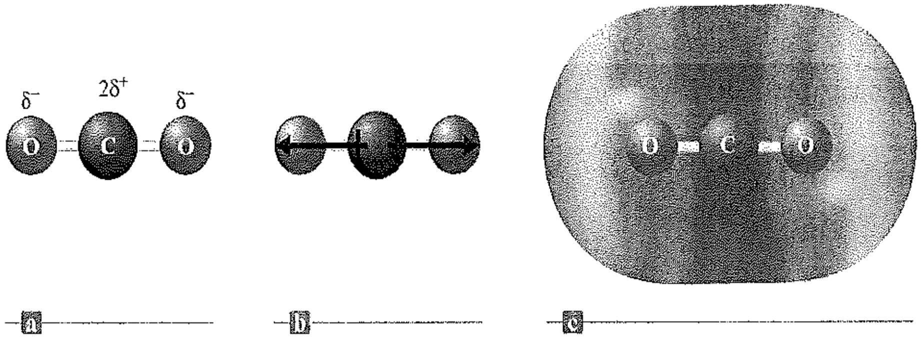
Molecular Structure

Dipole movement

ElectroStatic Potential Map



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Key

9. Which of the following incorrectly shows the bond polarity? Show the correct bond polarity for those that are incorrect.

   a. $^{\delta+}H\text{-}Cl^{\delta-}$ Correct
   2.1 3.0

   b. $^{\delta+}Cl\text{-}I^{\delta-}$ $Cl^{\delta-}\text{-}I^{\delta+}$ $Cl \leftarrow I$
   3.0 2.5

   c. $^{\delta+}Si\text{-}S^{\delta-}$ correct
   1.8 2.5

   d. $^{\delta+}Br\text{-}Br^{\delta-}$ pure covalent, equal sharing of Bond e's & no dipole movement exists

   e. $^{\delta+}O\text{-}P^{\delta-}$ $O^{\delta-}\text{-}P^{\delta+}$ $O \leftarrow P$
   3.5 2.1

* Use Electronegativity trend to predict partial neg end & the partial positive end of the Bond dipole

Key

Bond Energies

- To break bonds, energy must be *added* to the system (endothermic, energy term carries a positive sign).
- To form bonds, energy is *released* (exothermic, energy term carries a negative sign).

$$\Delta H = \Sigma\, \Delta H \text{ (bonds broken)} - \Sigma\, \Delta H \text{ (bonds formed)}$$

ΔH in this case represents the BOND ENTHALPY per mole of bonds (always has a positive sign).

Average Bond Energies (kJ/mol)

Bond	Energy	Bond	Energy	Bond	Energy	Bond	Energy
Single Bonds							
H—H	432	N—H	391	Si—H	323	S—H	347
H—F	565	N—N	160	Si—Si	226	S—S	266
H—Cl	427	N—P	209	Si—O	368	S—F	327
H—Br	363	N—O	201	Si—S	226	S—Cl	271
H—I	295	N—F	272	Si—F	565	S—Br	218
		N—Cl	200	Si—Cl	381	S—I	~170
C—H	413	N—Br	243	Si—Br	310		
C—C	347	N—I	159	Si—I	234	F—F	159
C—Si	301					F—Cl	193
C—N	305	O—H	467	P—H	320	F—Br	212
C—O	358	O—P	351	P—Si	213	F—I	263
C—P	264	O—O	204	P—P	200	Cl—Cl	243
C—S	259	O—S	265	P—F	490	Cl—Br	215
C—F	453	O—F	190	P—Cl	331	Cl—I	208
C—Cl	339	O—Cl	203	P—Br	272	Br—Br	193
C—Br	276	O—Br	234	P—I	184	Br—I	175
C—I	216	O—I	234			I—I	151
Multiple Bonds							
C=C	614	N=N	418	C≡C	839	N≡N	945
C=N	615	N=O	607	C≡N	891		
C=O	745 (799 in CO_2)	O_2	498	C≡O	1070		

10. Use the bond energies to estimate ΔH for each of the following reactions in the gas phase:

a. $H_2 + Cl_2 \rightarrow 2HCl$

$$H-H + Cl-Cl \longrightarrow 2\, H-Cl$$

432 239 427 KJ/mol

$$\Delta H = \Sigma H_R - \Sigma H_P$$

$$= 1\text{mole}(432 \tfrac{KJ}{mol}) + 1\text{mole}(239 \tfrac{KJ}{mol}) - [2^{\text{mole}}(427 \tfrac{KJ}{mol})]$$

$$\boxed{\Delta H = -183\ KJ}$$

b. $N_2 + 3H_2 \rightarrow 2NH_3$

$$N \equiv N + 3\, H-H \longrightarrow 2\, H-\overset{H}{N}-H \qquad \Delta H = \Sigma H_B - \Sigma H_F$$

945 432 $3(391)\ \frac{KJ}{mol}$

$$= 945 + 3(432) - [2(3)(391)]$$

$$\boxed{\Delta H = -105\ KJ}$$

Key

11. Use bond energies (In chart above) to predict ΔH for the following reaction:
$H_2S(g) + 3F_2(g) \rightarrow SF_4(g) + 2HF(g)$

H-S-H + 3 F-F → F-S(F)(F)-F + 2 H-F $\Delta H = \Sigma H_B - \Sigma H_F$

$2(347 \frac{KJ}{mol}) + 3(159)$ $-[4(327) + 2(565)]$

$\boxed{\Delta H = -1267\ KJ}$

12. The major industrial source of hydrogen gas is by the following reaction. Use bond energies to predict the ΔH for this reaction: $CH_4(g) + H_2O(g) \rightarrow CO(g) + 3\ H_2(g)$

H-C(H)(H)-H + H-O-H → C≡O + 3 H-H $\Delta H = \Sigma H_B - \Sigma H_F$

$4(413\ KJ/mole) + 2(467)$ $-[\ 1070 + 3(432)]$

$= 2586 - 2366$

$\boxed{\Delta H = 220\ KJ}$

13. Consider the following reaction: $A_2 + B_2 \rightarrow 2AB$ $\Delta H = -285\ kJ$
The bond energy for A_2 is one-half the amount of the AB bond energy. The bond energy of B_2 = 432 kJ/mol. What is the bond energy of A_2?

$A_2 + B_2 \rightarrow 2AB$

$A_2 = \frac{1}{2} AB$

$\frac{1}{2} AB + B_2 \rightarrow 2AB$

$\Delta H = \Sigma H_B - \Sigma H_F$

$-285\ KJ = \frac{1}{2} AB + 432 KJ - [2AB]$

$\frac{1}{2} AB - \frac{4}{2} AB = -717\ KJ$

$-\frac{3}{2} AB = -717\ KJ$

$AB = 478\ KJ$

$A_2 = \frac{1}{2} AB$

$= \frac{1}{2}(478\ KJ)$

$\boxed{A_2 = 239\ KJ}$

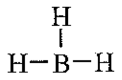
Lewis Structures

- Shows how valence electrons are arranged among atoms in a molecule.
- Reflects central idea that stability of a compound relates to noble gas electron configuration.
- Used primarily in drawing COVALENT compounds

Exceptions to Octet Rule

- Boron tends to form compounds in which the boron atom has fewer than eight electrons around it (it does not have a complete octet).

$BH_3 = 6e^-$

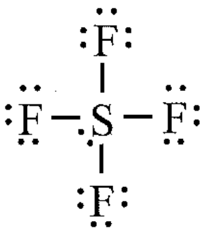


BH_3

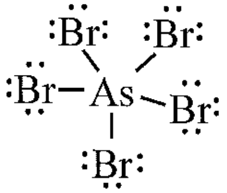
- When it is necessary to exceed the octet rule for one of several third-row (or higher) elements, place the extra electrons on the central atom.

Key

$SF_4 = 34e^-$



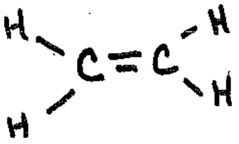
$AsBr_5 = 40e^-$



14. Draw a Lewis structure for each of the following molecules:

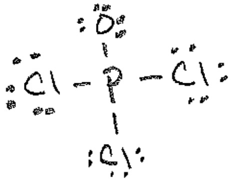
a. C_2H_4

C 2x4=8
H 4x1=4
12



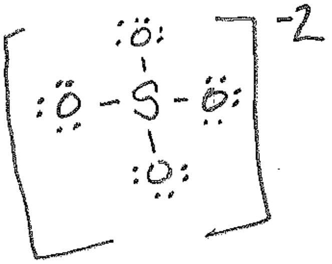
b. $POCl_3$

P 1x5=5
O 1x6=6
Cl 3x7=21
32VE



c. SO_4^{-2}

S 1x6=6
O 4x6=24
-2e = 2
32Ve



d. SF4

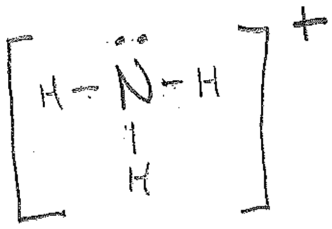
S 1x6=6
F 4x7=28
34Ve

e. HCN

H 1x1=1
C 1x4=4
N 1x5=5
10Ve

f. NH_4^+

N 1x5=5
H 1x4=4
+ = -1
8Ve



Review

- C, N, O, and F should always be assumed to obey the octet rule.
- B and Be often have fewer than 8 electrons around them in their compounds.
- Second-row elements never exceed the octet rule.
- Third-row and heavier elements often satisfy the octet rule but can exceed the octet rule by using their empty valence *d* orbitals.
- When writing the Lewis structure for a molecule, satisfy the octet rule for the atoms first. If electrons remain after the octet rule has been satisfied, then place them on the elements having available *d* orbitals (elements in Period 3 or beyond).

Resonance

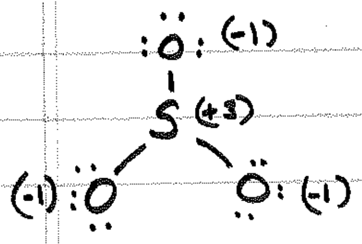
- More than one valid Lewis structure can be written for a particular molecule.
- Actual structure is an average of the resonance structures.
- Electrons are really delocalized – they can move around the entire molecule.

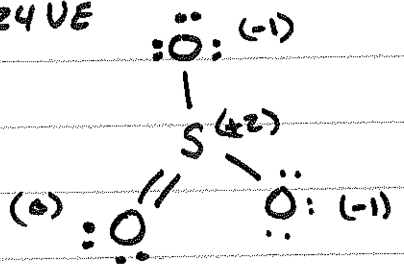
AP Chem - Unit 5 - Homework Packet

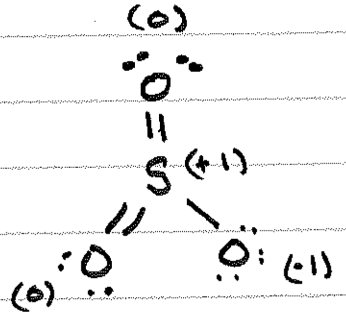
15)

A) SO_3

S 1×6 = 6
O 3×6 = 18
24 VE







FC

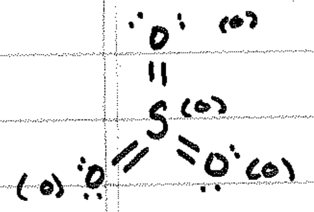
O	S
6	6
-7	-3
-1	+3

FC

O	S
6	6
-6	-4
0	+2

FC

S
6
-5
+1

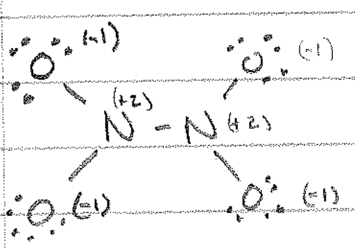


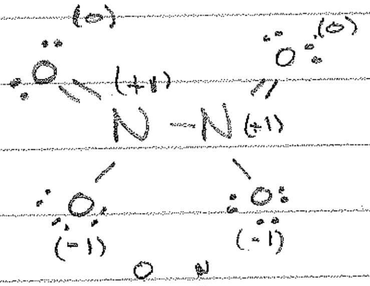
← Correct due to All Atoms having a FC = 0

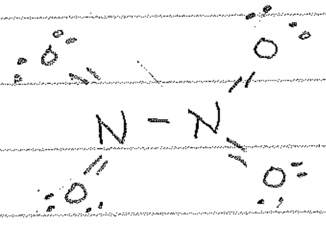
S
6
-6
0

B) N_2O_4

N 2×5 = 10
O 4×6 = 24
34 VE





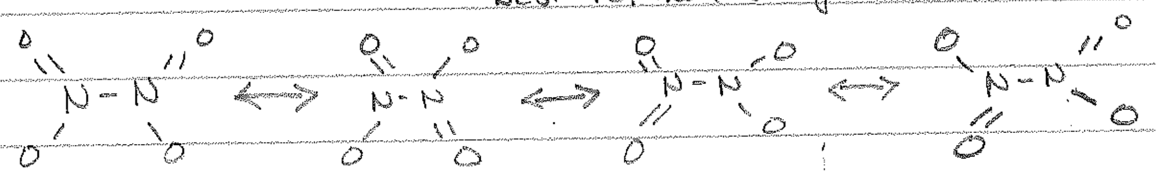


O	N
6	5
-7	-3
-1	+2

O	N
6	5
-6	-4
0	+1

N has to many Bonds

Best Formal Charge!!



15) conti

c) OCN$^-$

O 1×6 = 6
C 1×4 = 4
N 1×5 = 5
ė = 1
16VE

FC

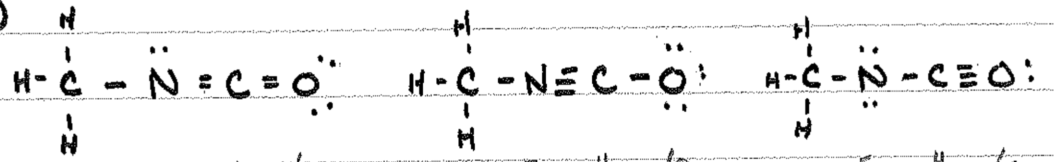
$\frac{6}{-6}\ 0$ $\quad \frac{4}{-4}\ 0$ $\quad \frac{5}{-6}\ -1$

$\frac{6}{-7}\ -1$ $\quad \frac{4}{-4}\ 0$ $\quad \frac{5}{-5}\ 0$

O is more Electronegative than N
∴ Above is correct Lewis Structure

16) CH_3NCO

C 2×4 = 8
H 3×1 = 3
N 1×5 = 5
O 1×6 = 6
22 Ve



FC

$\frac{4}{4}\ 0$ $\quad \frac{5}{-5}\ 0$ $\quad \frac{4}{-4}\ 0$ $\quad \frac{6}{-6}\ 0$

0 $\quad \frac{5}{-4}\ +1$ $\quad \frac{4}{-4}\ 0$ $\quad \frac{6}{-7}\ -1$

0 $\quad \frac{5}{-6}\ -1$ $\quad \frac{4}{-4}\ 0$ $\quad \frac{6}{-5}\ +1$

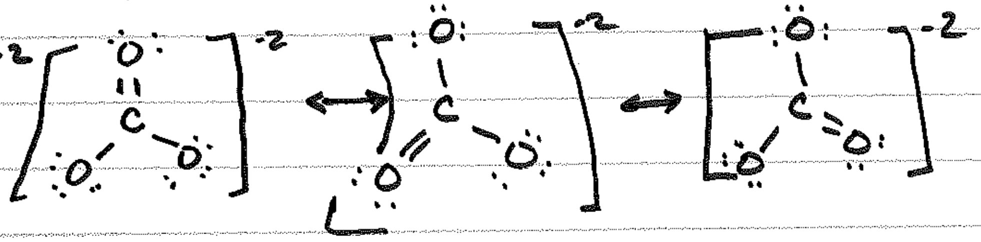
↑ lowest FC ∴ Correct Structure

17) CO, CO_2, CO_3^{-2}, CH_3OH

CO $:C\equiv O:$

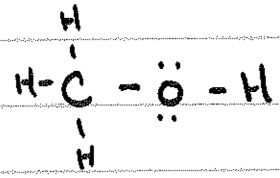
CO_2

CO_3^{-2}



17) conti:

CH_3OH



$H-\overset{H}{\underset{H}{C}}-\ddot{\underset{..}{O}}-H$

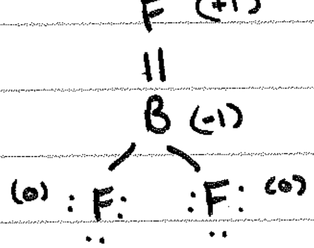
As # of Bonds increase Btwn 2 Atoms, Bond strength ↑ & Bond length ↓

∴ Longest O-C Bond to Shortest: $CH_3OH > CO_3^{-2} > CO_2 > CO$

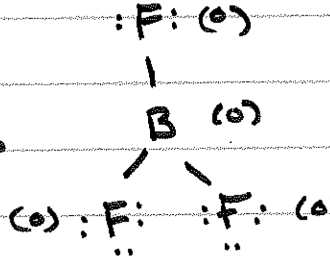
Weakest O-C Bond to strongest: $CH_3OH < CO_3^{-2} < CO_2 < CO$

18) BF_3

B 1x3 = 3
F 3x7 = 21
24VE



VS



FC	F=	F-	B
	7	7	3
	-6	-7	-4
	1	0	-1

B
3
-3
0

1st one obeys octet Rule But FC's not = to 0

Doesn't Follow octet Rule But FC is 0 on Each Atom ∴ Structures generally want to minimize FC this structure w/ Single Bonds is Best

19) CO

$:C\equiv O:$

↑

higher Electronegative

$\longmapsto$

δ^+ δ^-

$\overset{(-1)}{:C}\equiv\overset{(+1)}{O:}$

$$\begin{array}{lcc} \text{Fc} & 4 & 6 \\ & \underline{-5} & \underline{5} \\ & -1 & +1 \end{array}$$

FC is opposite of EN

The two opposing effects seem to partially cancel to give a much less polar molecule than expected.

20) OCN⁻

$$\begin{array}{l} O\ 1\times 6 \\ C\ 1\times 4 \\ N\ 1\times 5 \\ -e^-\ \underline{1} \\ \quad 16VE \end{array}$$

$[\,\ddot{\underset{..}{O}}=C=\ddot{\underset{..}{N}}\,]^- \leftrightarrow [\,:\ddot{\underset{..}{O}}-C\equiv N:\,]^- \leftrightarrow [\,:O\equiv C-\ddot{\underset{..}{N}}:\,]^{-1}$

$$\begin{array}{lccccccccc} \text{Fc} & 0 & 0 & -1 & -1 & 0 & 0 & +1 & 0 & -2 \end{array}$$

O higher En
So most likely

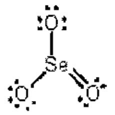
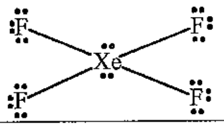
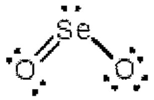
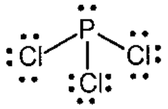
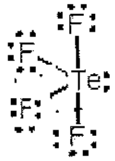
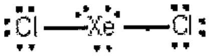
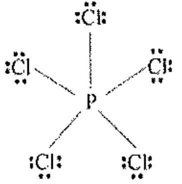
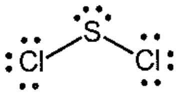
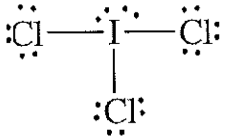
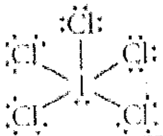
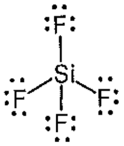
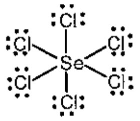
CNO⁻ Also 16VE

$[\,\ddot{\underset{..}{C}}=N=\ddot{\underset{..}{O}}\,]^- \leftrightarrow [\,:C\equiv N-\ddot{\underset{..}{O}}:\,]^- \leftrightarrow [\,:\ddot{\underset{..}{C}}-N\equiv O:\,]^-$

$$\begin{array}{ccccccccc} 4 & 5 & 6 & 4 & 5 & 6 & 4 & 5 & 6 \\ \underline{-6} & \underline{-4} & \underline{-6} & \underline{-5} & \underline{-4} & \underline{-7} & \underline{-7} & \underline{-4} & \underline{-5} \\ -2 & +1 & 0 & -1 & +1 & -1 & -3 & +1 & +1 \end{array}$$

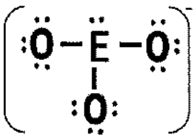
All the Resonance structures for CNO⁻ have greater FC than in OCN⁻, making fulminate (CNO⁻) more Reactive (less stable)

21. Complete the chart below:

Molecular Formula	**Molecular Structure**	**Electronic Structure (electronic)**	**Molecular Structure (geometry)**	**Bond Angles**	**Polarity**
SeO_3		Trigonal planar	Trigonal planar	120	nonpolar
XeF_4		Octahedral	Square planar	90	nonpolar
SeO_2		Trigonal planar	Bent	<120	polar
PCl_3		Tetrahedral	Trigonal pyramidal	<109.5	polar
TeF_4		Trigonal bipyramidal	See-saw	<90, <120	polar
$XeCl_2$		Trigonal bipyramidal	Linear	180	nonpolar
PCl_5		Trigonal bipyramidal	Trigonal bipyramidal	90, 120	nonpolar
SCl_2		Tetrahedral	Bent	<<109.5	polar
ICl_3		Trigonal bipyramidal	T shaped	<<90	polar
ICl_5		Octahedral	Square pyramidal	<90	polar
SiF_4		Tetrahedral	Tetrahedral	109.5	nonpolar
$SeCl_6$		Octahedral	Octahedral	90	nonpolar

22. Consider the following Lewis structure where E is an unknown element. What are some possible identities for element E? Predict the molecular structure and bond angles for this ion.

Lewis structure has 26 Ve



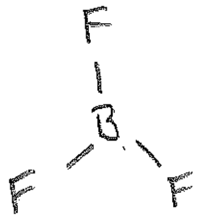
$26 = E + 3(6) + 1$

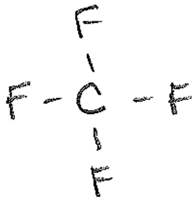
$\uparrow$ O $\quad \uparrow$ Ion

E = 7 valence Electrons ∴ its a halogen F, Cl, Br, I

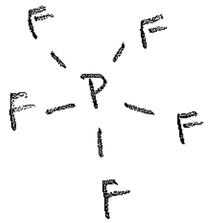
Trigonal pyramid so Bond Angles less than 109.5°

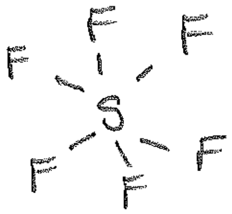
23. The molecules BF_3, CF_4, CO_2, PF_5, and SF_6 are all nonpolar, even though they all contain polar bonds. Explain why?





$O=C=O$





All these molecules have polar bonds that are Symmetrically arranged about the Central Atom.

In each molecule, the Individual Bond dipole cancel each other out to give No net overall dipole moment.