MONDAY, Dec. 8: COVALENT NOMENCLATURE

Name the following covalent compounds.

1) P_4S_5 ____

2) O_2 ____

3) SeF_6 ____

4) Si_2Br_6 ____

5) SCl_4 ____

6) CH_4 ____

Write the formulas for the following covalent compounds.

1) antimony tribromide ____________________

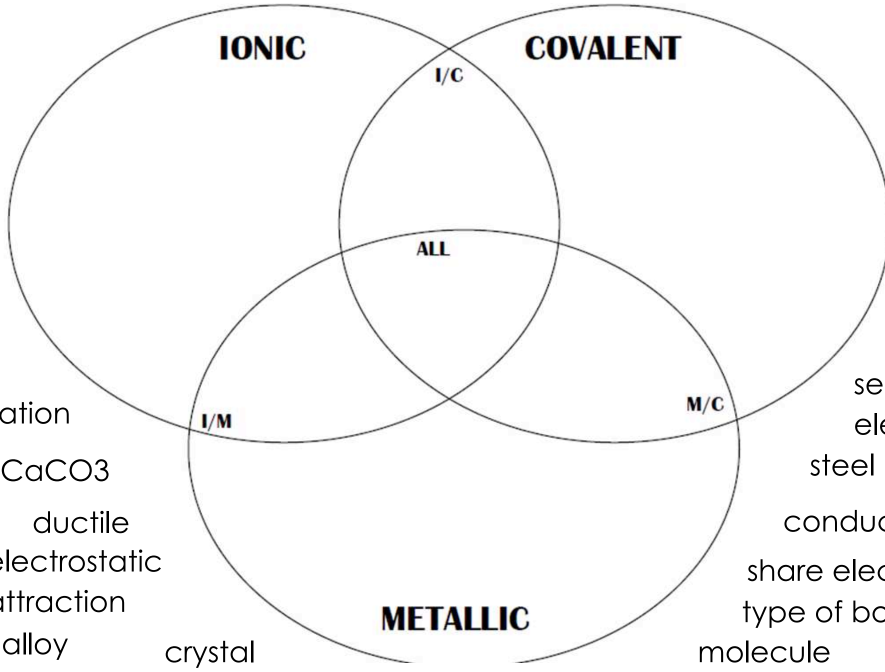
2) hexaboron silicide _____________________

3) chlorine dioxide _________________________

4) hydrogen iodide __________________________

5) iodine pentafluoride _____________________

6) dinitrogen trioxide _______________________



cation
CaCO3
ductile
electrostatic attraction
alloy
crystal
high melting pt
metal
low melting pt
transfer electrons
sea of electrons
steel
conducts
share electrons
type of bond
molecule
doesn't conduct
nonmetal
type of bond
malleable

Lewis Structures (NASL Method) VSPER

December 8, 2014

- Central atom = the atom with the lowest electronegativity (usually)
  - > Central atom is the usually the single atom of the molecule
  - > *Resonance: Sometimes there are alternate ways to draw a molecule that are both considered correct, these are resonance structures*
- ________ is always a terminal atom

Example: CF_4	

N

- For all atoms to fulfill the octet rule
  - > For most atoms, they need 8!!
- There are a few exceptions:
  - > Hydrogen =2
  - > Beryllium =4
  - > Boron =6

Example: 1 Carbon and 4 Fluorine

CF_4

A

- **For "A" groups, the number of valence electrons is equal to the group number.**
- **For groups 13-18, the number of valence electrons is equal to (group number – 10)**
  - > ***For anions, add the number of electrons equal to the charge of the anion***
  - > ***For cations, subtract the number of electrons equal to the charge of the cation***

Example: 1 Carbon and 4 Fluorine

CF_4

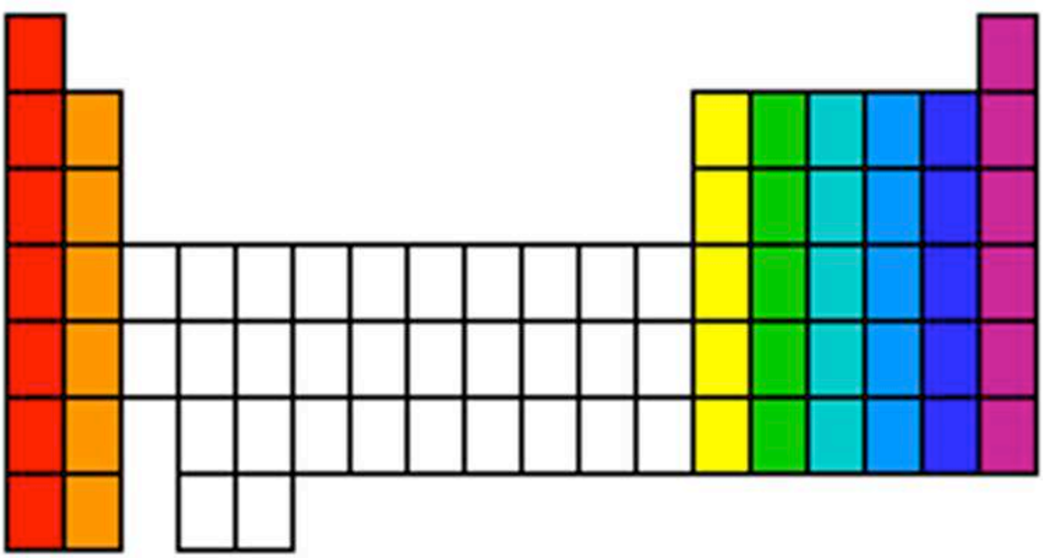
Group 2 = 2 electrons

Group 1 = 1 electron

3, 4, 5, 6, 7

Group 8 = 8 electrons

Except for He, it has 2 electrons



S

- **Calculate S (Shared) as the difference between N (needed electrons) – A (available)**
- **Divide S by 2 to obtain the number of bonds to be extended from the central atom**
  - > ***Multiple Bonds are possible: DOUBLE or TRIPLE***
    - **Do not add electrons.**
    - **"Borrow" them from surrounding atoms.**
    - **Only C, N, O, P, and S form multiple bonds.**

Example: Needed: Available:

CF_4 SHARED:

BONDS:

F

F C F

F

L

- **Calculate L (Lone-pair electrons or simply "dots") as the difference between A (available) – S (shared)**
- **Place lone pairs of electrons around each terminal atom to complete their octets**
- **An octet = 4 electrons pairs around an atom (eight total electrons)**
  - > **Remember: hydrogen can only have one bond and NO Lone Pairs (dots)**

Example: Available - Shared:

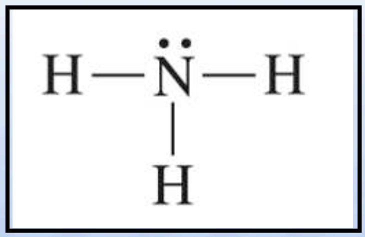
CF_4

F

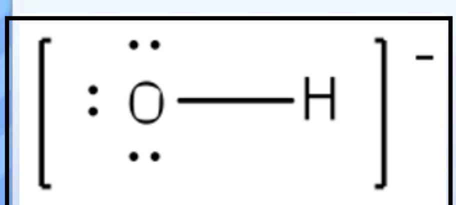
F C F

F

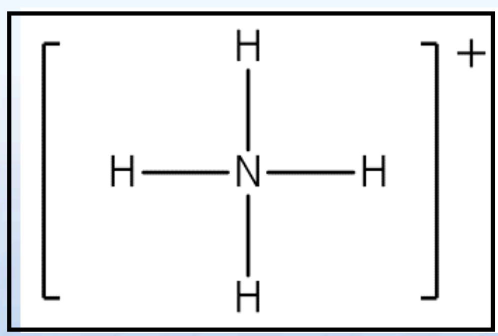
- **If there are more electrons left, place them as lone pairs on the central atom**
  - > **This will sometimes lead to an "expanded octet" around the central atom**
  - > **Expanded octet = five or six electron pairs around an atom**
  - > **Only central atoms from the third period and below can have expanded octets**



- **For ions, put brackets around the entire the entire Lewis structure and make sure to indicate its charge**
  - > ***Also, REMEMBER to add electrons for anions (negative charge) and subtract electrons for cations (positive charge)***



OH^{-1} Hydroxide



NH^{+4} Ammonium

Example #1: Simple Structures

-The central atom is the least electronegative.
-Make sure all atoms have 8 valence electrons except hydrogen, which should have 2

a) PCl_3

Needed e:

Available e:

Shared:

- Bonds:

Lone pair e:

Example #2: Polyatomic Ions

-Add number of electrons equal to charge for negatively charged ions.
-Subtract number of electrons for positively charged ions.
-Place brackets around the entire structure with the correct charge on the outside

a) NH_4^+

Needed e:

Available e:

Shared:

- Bonds:

Lone pair e:

Example #1: Simple Structures

-The central atom is the least electronegative.
-Make sure all atoms have 8 valence electrons except hydrogen, which should have 2

b) SiF_4

Needed e:

Available e:

Shared:

- Bonds:

Lone pair e:

Example #2: Polyatomic Ions

-Add number of electrons equal to charge for negatively charged ions.
-Subtract number of electrons for positively charged ions.
-Place brackets around the entire structure with the correct charge on the outside

b) SO_4^{-2}

Needed e:

Available e:

Shared:

- Bonds:

Lone pair e:

Example #3: Multiple Bonds
-Use double or triple bonds when the central atom does NOT have 8 valence electrons
-Borrow lone pairs of electrons from atoms on the terminal ends of structure

b) N_2
Needed e:
Available e:
Shared:
- Bonds:
Lone pair e:

WEDNESDAY, Dec. 10
1. Draw the Lewis Diagrams for the following compounds:

a) Water, H_2O
Needed e:
Available e:
Shared:
- Bonds:
Lone pairs:

b) Oxygen, O_2
Needed e:
Available e:
Shared:
- Bonds:
Lone pairs:

2. Directions: Read the following and use the electronegativity chart below to answer the following questions.

Covalent bonds involve atoms sharing electrons (pulled in a tug of war). When electrons are pulled equally, then a nonpolar covalent bond forms. When electrons are pulled unequally, then a polar covalent bond. The magnitude of attraction for electrons is called "Electronegativity". The more electronegative an atom is, the more it wants the electrons. So, in covalent bonds, the more electronegative atom, pulls the electron more. We can use the chart below to tell if a molecule is nonpolar covalent, polar covalent or ionic. Subtract the values of electronegativity of the atoms and determine where the value falls.

ΔE_n	Bond type
0.0 - 0.3	nonpolar covalent
0.4 - 1.9	polar covalent
2.0 - 3.3	ionic

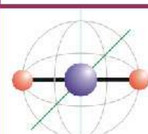
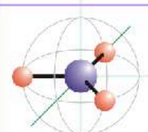
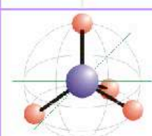
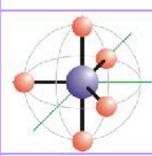
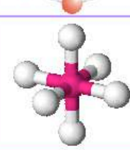
EXAMPLE: H_2O

Electroneg. of Hydrogen: 2.1

Electroneg. of Oxygen: 3.5

What type of bond forms between hydrogen and oxygen?

VSEPR Theory

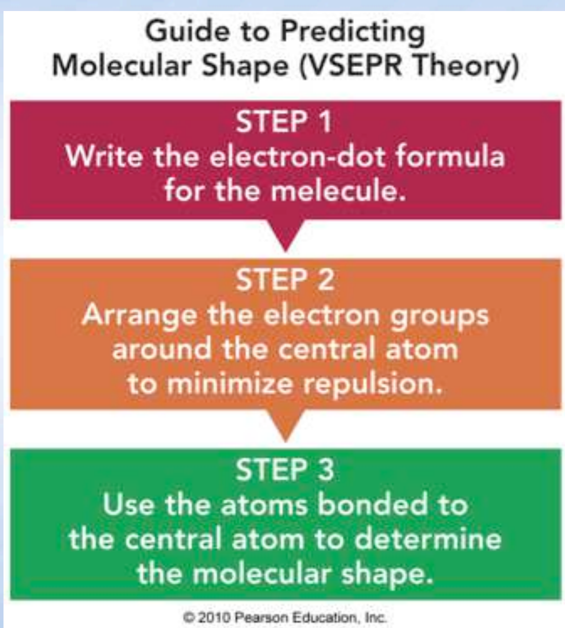
Number of Electron Domains	Arrangement of Electron Domains	Electron-Domain Geometry	Predicted Bond Angles
2		Linear	180°
3		Trigonal Planar	120°
4		Tetrahedral	109.5°
5		Trigonal Bipyramidal	120° 90°
6		Octahedral	90°

VSEPR Theory

ttps://www.youtube.com/watch?v=keHS-CASZfc

- **V**alence **S**hell **E**lectron **P**air **R**epulsion **Theory**
  - > Used to predict the 3-D shape of molecule
  - > **BIG Idea:** Structure around a given atom is determined by minimizing the repulsion between electron pairs
  - > Bonding and non-bonding electron pairs around the central atom are arranged as far apart as possible
    - – RULES: A) max separation between electron pairs
    - – B) Atom positions define molecule geometry
    - – C) Lone electron pairs squeeze bond angle (Actual angle < ideal angle)

VSEPR



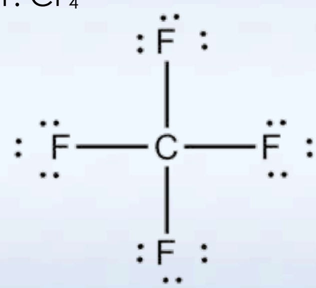
Steps to identifying the molecular shape of a compound

VSEPR

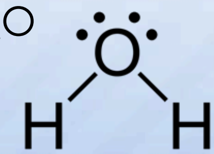
Pairs of Electrons	Pairs of Bonding Electrons	Pairs of Lone Electrons	Electron Distribution	Molecular Geometry	Bond Angle	Diagram
2	2	0	linear	linear	180	F—Be—F
3	3	0	trigonal planar	trigonal planar	120	H B H H
	2	1		bent	117	O O O
4	4	0	tetrahedral	tetrahedral	109.5	H C H H H
	3	1		trigonal pyramidal	107	H B H H
	2	2		bent	104⁺	H Ö H

Let's classify some examples:

1. CF_4



2. H_2O

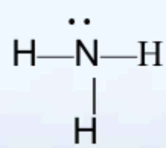


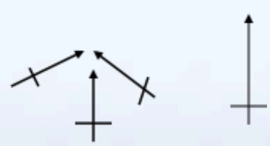
Bond Polarity

A **polar molecule**

- contains polar bonds (Use differences in electronegativity)
- has a separation of positive and negative charge called a dipole, indicated with δ^+ and δ^-
- has dipoles that do not cancel

δ^+ δ^-
H–Cl
dipole





dipoles do not cancel

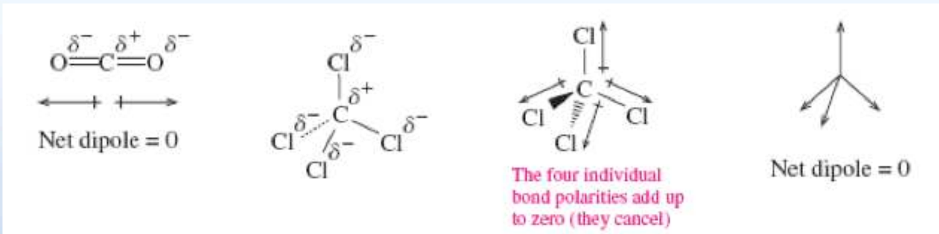
Bond Polarity

A **nonpolar molecule**

- contains nonpolar bonds

  Cl–Cl H–H

- or has a symmetrical arrangement of polar bonds



Determining Molecular Polarity

Determine the polarity of the H_2O molecule.