Naming Covalent Compounds

- When it is all NONMETALS

Compounds vs Molecules

- A Compound is any substance composed of two or more DIFFERENT elements.
- A Molecule is any substance composed of two or more atoms COVALENTLY BONDED.

Properties of Covalent Compounds

- Generally Low Melting and Boiling Points
- Generally Soft and Flexible
- Tend to be Flammable
- Don’t conduct electricity
- Normally won’t dissolve in water

Types of Covalent Bonds

- **Formed between two nonmetals in 14, 15, 16, and 17**
- **Nonmetals have high electronegativity values**
- **Electrons are shared**

single bond shares one pair electrons

double bond shares two pairs electrons

triple bond shares three pairs electrons

Diatomic Elements

Gases that exist as diatomic molecules are H_2, F_2, N_2, O_2, Cl_2, Br_2, I_2

- They are simply given their elements name.
- Exist this way only when not in compounds

Learning Check

Use the name of the element to name the following diatomic molecules.

H_2 hydrogen

N_2 nitrogen

Cl_2 ____________________

O_2 ____________________

I_2 ____________________

Naming Covalent Compounds

Two nonmetals

- **Name each element**
- **End the last element in -ide**
- **Add prefixes to show more than 1 atom**

Prefixes

mon1 hexa6

di2hepta7

tri3octa8

tetra4nona9

penta5deca10

Learning check

Fill in the blanks to complete the following names of covalent compounds.

CO carbon _______oxide

CO_2 carbon __________________

PCl_3 phosphorus ________chloride

CCl_4 carbon _________chloride

N_2O ______nitrogen ______oxide

Learning Check

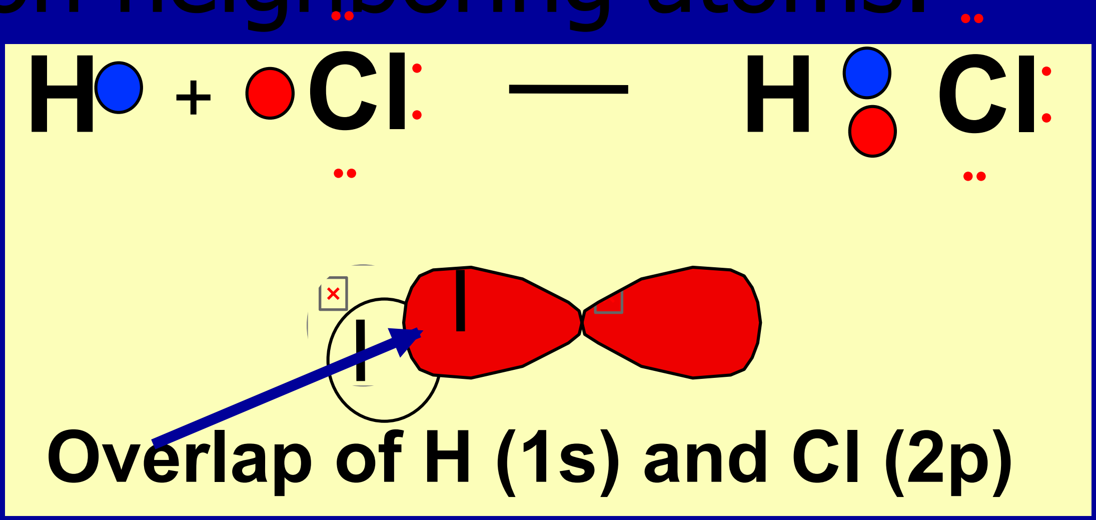
A. P_2O_5 1) **phosphorus oxide**
2) **phosphorus pentoxide**
3) **diphosphorus pentoxide**

B. Cl_2O_7 1) **dichlorine heptoxide**
2) **dichlorine oxide**
3) **chlorine heptoxide**

C. Cl_2 1) **chlorine**
2) **dichlorine**
3) **dichloride**

Bond Formation

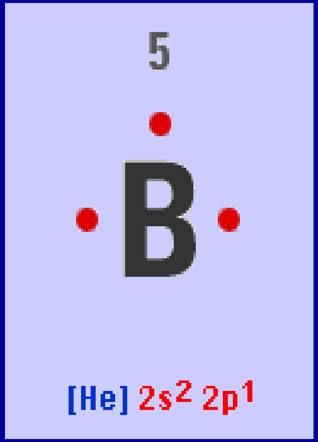
A bond can result from an **overlap** of atomic orbitals on neighboring atoms.

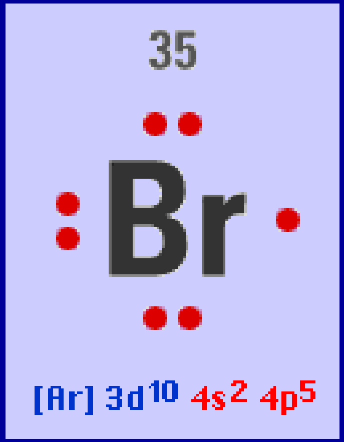


Note that each atom has a single, unpaired electron.

Review of Valence Electrons

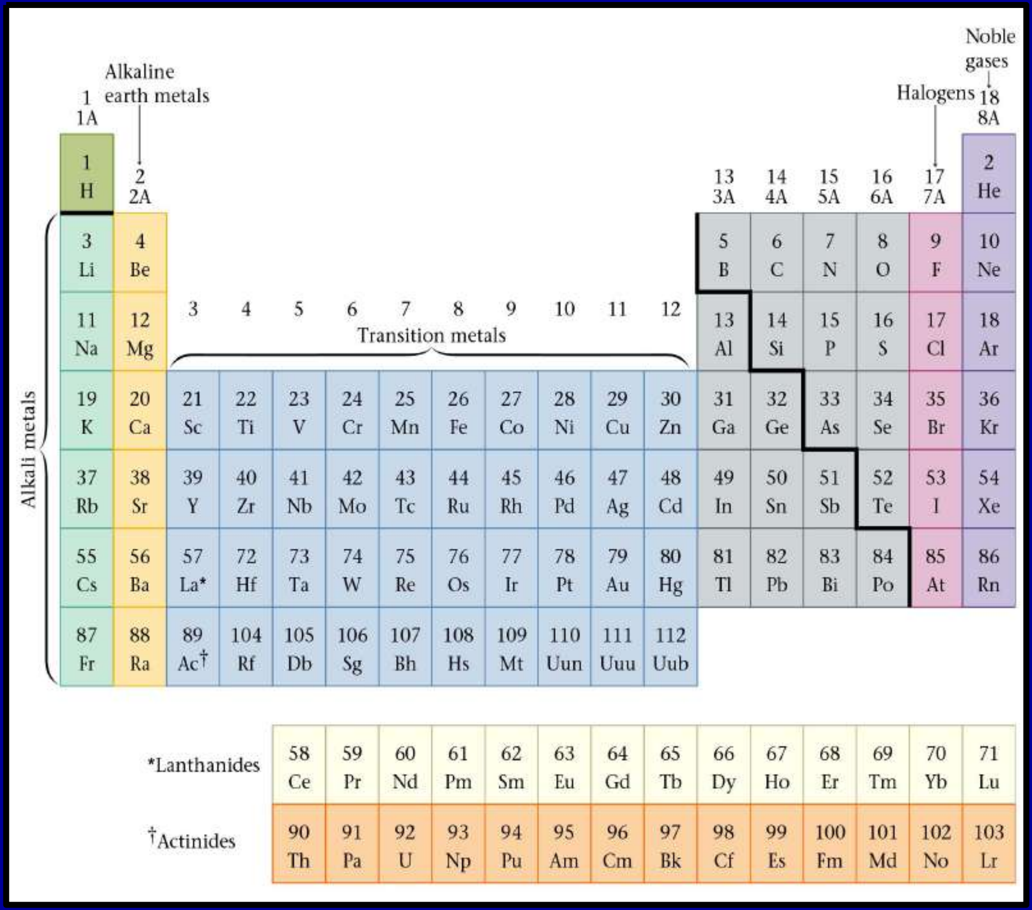
- Remember from the electron chapter that valence electrons are the electrons in the OUTERMOST energy level… that's why we did all those electron configurations!
- B is $1s^2\ 2s^2\ 2p^1$; so the outer energy level is 2, and there are 2+1 = 3 electrons in level 2. These are the valence electrons!
- Br is [Ar] $4s^2\ 3d^{10}\ 4p^5$
  How many valence electrons are present?

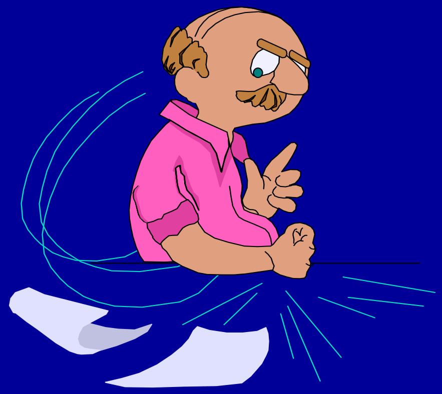




Review of Valence Electrons

Number of valence electrons of a main (A) group atom = Group





Lewis Structures

- Drawings of covalent compounds
- Dots are used for nonbonding electrons
- Lines are used for bonding pairs of electrons
- All atoms must have 8 electrons in some combination except H which has 2

Lone pairs versus bonding pairs

- Lone pairs are pairs of electrons not in bonding
- Bonding pairs are pairs of electrons involved in bonding

Steps Building a Lewis Structure

1. Decide on the central atom; never H. Why?

If there is a choice, the central atom is atom of lowest affinity for electrons. (Most of the time, this is the least electronegative atom…the single atom is normally the lowest or the one further to the left on the Periodic Table.)

2. Add up the number of valence electrons that can be used.

Building a Dot Structure

- ➓ Form a single bond between the central atom and each surrounding atom (each bond takes 2 electrons!)

4. Remaining electrons form LONE PAIRS to complete the octet as needed (or duet in the case of H).

Building a Dot Structure

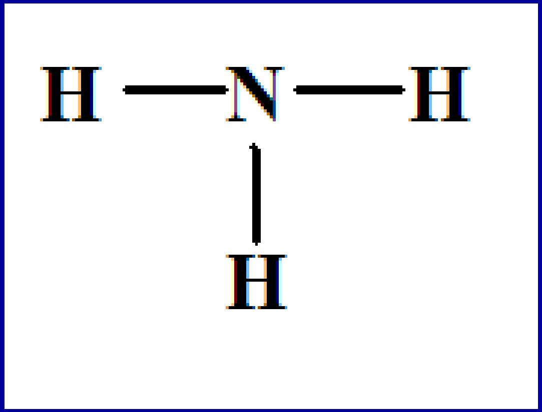
- 5. Check to make sure there are 8 electrons around each atom except H. H should only have 2 electrons. This includes SHARED pairs.
- 6. Move nonbonding pairs if necessary to make double or triple bonds to insure the octet rule

Example – NH_3

- Central Atom. In this case it would have to be N, because H NEVER can be
- Count Electrons.
  - N = 5
  - H = 3 (1) = 3
  - 5 + 3 = 8

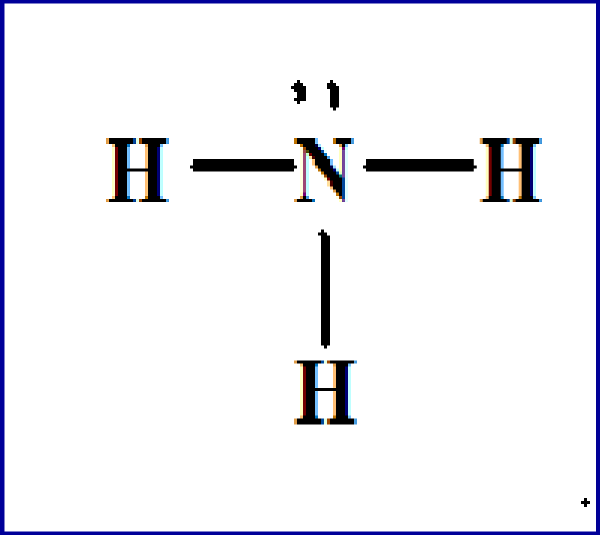
Example – NH_3

3. Place the N in the middle and attach the H's



Example – NH_3

4. Place the unshared pairs



Example – NH_3

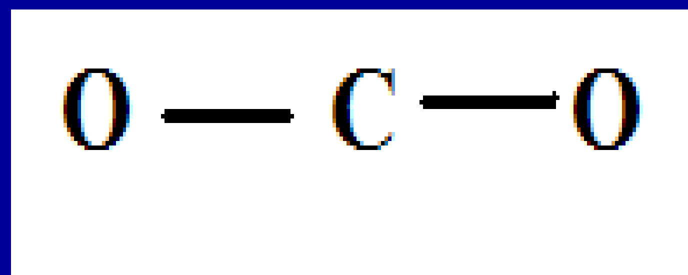
- Count your electrons and check your work
  - $3 (2) + 2 = 8$
- Add double or triple bonds if needed
  - None needed here so you are done

Example – CO_2

- Central Atom: C is the single atom and farthest to the right – SO "C" it is.
- Count the electrons
  - C = 4 and O = 6 x 2
  - 16 electrons total

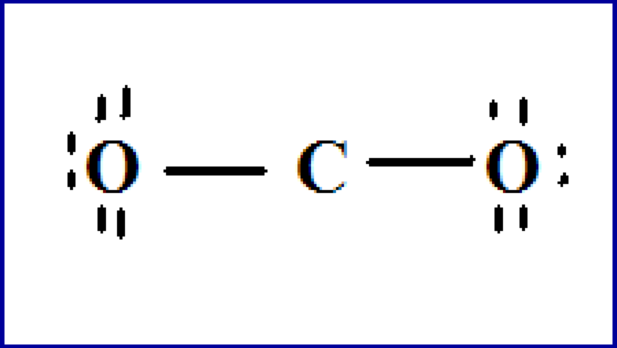
Example – CO_2

❿ Place the C in the middle and attach the O’s



Example – CO_2

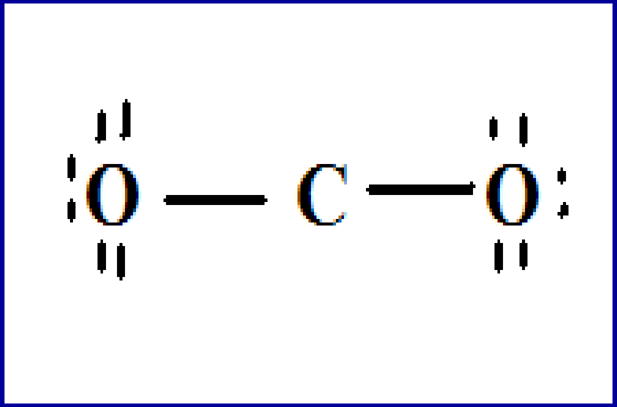
- ❿ Place the unshared pairs



Example – CO_2

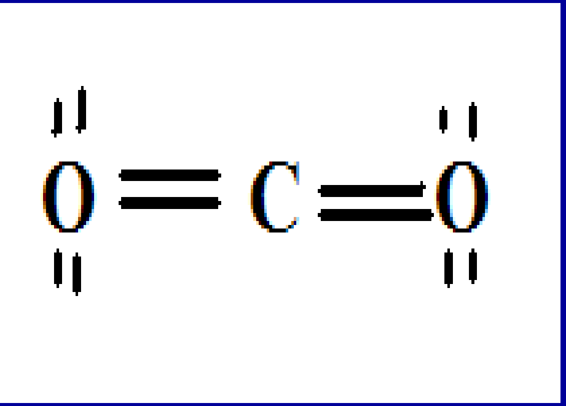
- ⑩ Count your electrons – insure all except H have 8

  C only has 4 so it is short, we will need more bonds



Example – CO_2

❿ Move electrons around and make double bonds, then recount



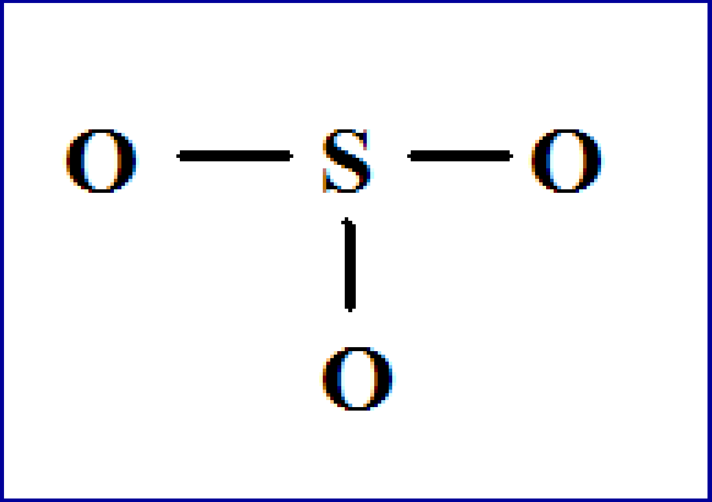
7. Now all have 8, so we are done

Example SO_3^{-2}

- Central Atom here is "S"
- Count electrons:
  - S = 6
  - O = 6 x 3 = 18
  - -2 means add two more
  - 6 + 18 + 2 = 26 total

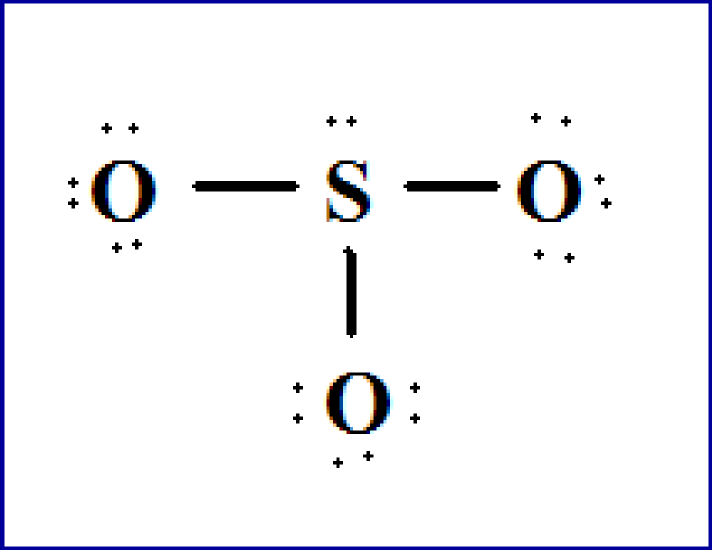
Example SO_3^{-2}

- Place the S in the middle and attach the O's



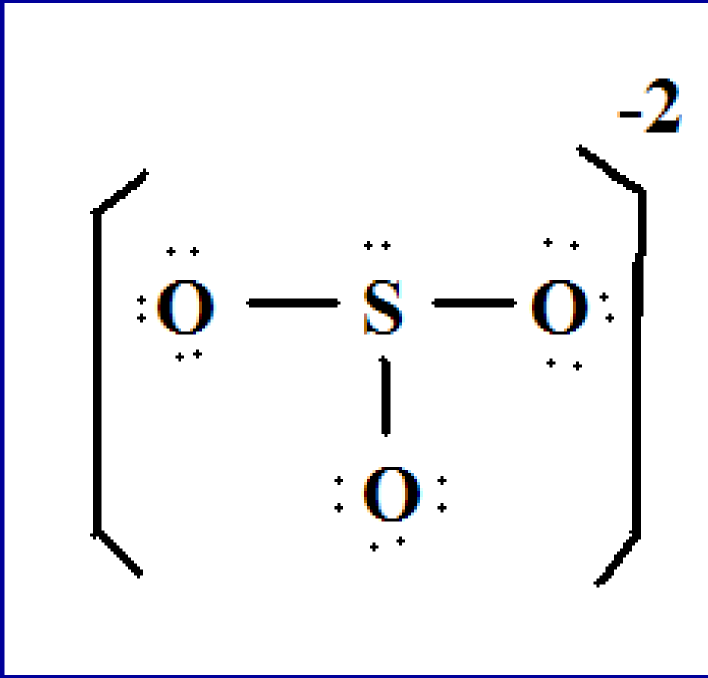
Example: SO_3^{-2}

- Place the extra electrons
- CHECK: All have 8 and I have placed 26 so I AM ALMOST DONE



Example: SO_3^{-2}

- One More little thing since it has a charge
- You have to put it in brackets and right the charge outside



VSEPR

- Valence Shell Electron Pair Repulsion
- Molecules take on shapes because electrons are all negative and attempt to repel as far away as possible.
- This Repulsion results in molecules taking on 3 dimensional Shapes

Shapes

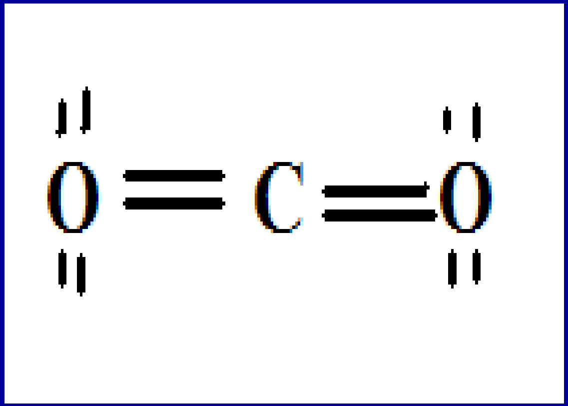
- The Shapes are dependent on the areas of electron concentration
  - A single pair is an area of concentration – NON Bonding area
  - A single bond counts as an area of concentration – BONDING AREA
  - A double bond counts as an area of concentration – BONDING AREA
  - A triple bond count as an area of concentration – BONDING AREA

SHAPES

SHAPE	BONDING AREAS	NON BONDING AREAS
LINEAR	2	0
BENT	2	1 or 2
TRIGONAL PLANAR	3	0
TRIGONAL PRYMIDAL	3	1
TETRAHEDRAL	4	0

Shapes - Examples

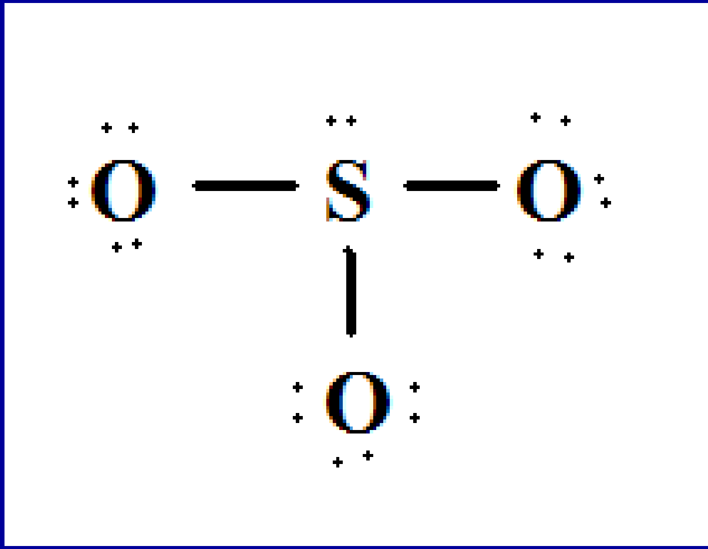
- Carbon Dioxide



- Two Bonding Areas, no Non bonding areas – Linear

Shapes - Example

- Sulfite (SO_3^{-2})



- Three Bonding areas, 1 Non bonding area = Trigonal Pyramidal