Chapter 6: Chemical Bonding

Section 6.1 – Introduction to Chemical Bonding

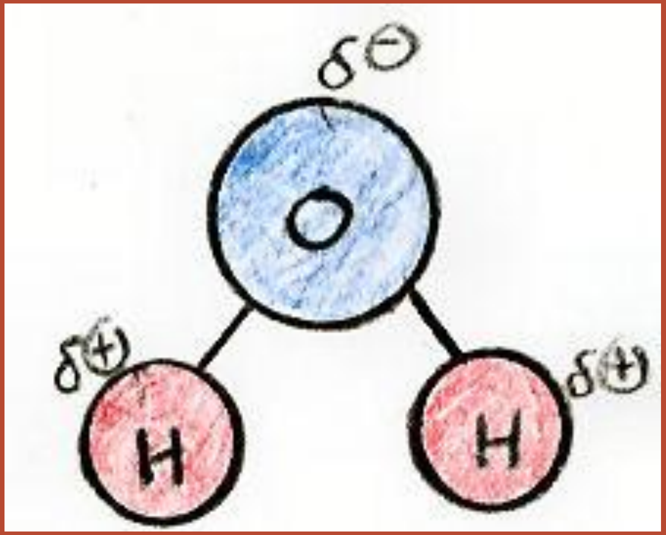
- A **chemical bond** is a mutual electrical attraction between the nuclei and valence electrons of different atoms that binds the atoms together.
- There are two types of bonding:
  - **Ionic bonding** is bonding that results from the electrical attraction between anions and cations
  - **Covalent bonding** results from the sharing of electron pairs between two atoms

Ionic or Covalent?

- Bonding between atoms of different elements is rarely purely ionic or covalent.
- The degree of ionic or covalent bonding is determined by the differences in the electronegativity of the elements.
  - Polar covalent bonds
  - Nonpolar covalent bonds
  - Ionic bonds

Covalent Bonding

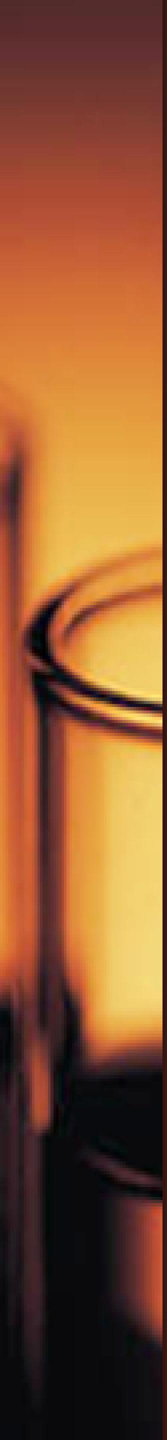
Polar Covalent – difference in electronegativities of 0.4-1.7



Non-Polar Covalent – slight to no difference in electronegativities 0-0.4 difference

6.2 – Covalent Bonding & Molecular Compounds

- **Molecules** – a neutral group of atoms held together by covalent bonds
  - Consist of nonmetal-nonmetal bonds
- **Molecular compound** – a chemical compound whose simplest units are molecules.
- **Molecular formula-** shows the types and numbers of atoms combined in a single molecule of a molecular compound.



Special Molecules

Diatomic Elements – a group of elements that naturally exist as two atoms covalently bonded together

- H_2
- N_2
- O_2
- F_2
- Cl_2
- Br_2
- I_2

Bond Length & Energy

- Bond length is the average distance between two bonded atoms
- Bond energy – energy required to break a chemical bond and form neutral isolated atoms
- Measured in KJ/mol
- Ex) H-H bonds take 436 KJ to break
- Table 1 page 182 has Bond Energies

Characteristics of Covalent Bonds

The Octet Rule – all atoms want 8 valence e-

- Exceptions: Hydrogen and Helium:
  - Only want 2 e^-
- Valence e^-: outermost electrons

Lewis Dot Structures

- Use the name of the representative element group to determine the # of **valence e^-'s**

Valence electrons

Grps	1	2	13	14	15	16	17	18
Per. 2	Li	B	Be	C	N	O	F	Ne
e^- config	s^1	s^2	s^2p^1	s^2p^2	s^2p^3	s^2p^4	s^2p^5	s^2p^6
Lewis Dot	1	2	3	4	5	6	7	8

Drawing Lewis Dot Structures for Covalent Compounds

1. Determine the number of shared electrons. (How many electrons do they need to obtain an octet?) This is how many bonds must be formed.
2. Place 1 pair of electrons in each bond.
3. Decide where any leftover bonding electrons should go.
4. Fill in the molecule with the rest of the electrons to give all atoms an octet.

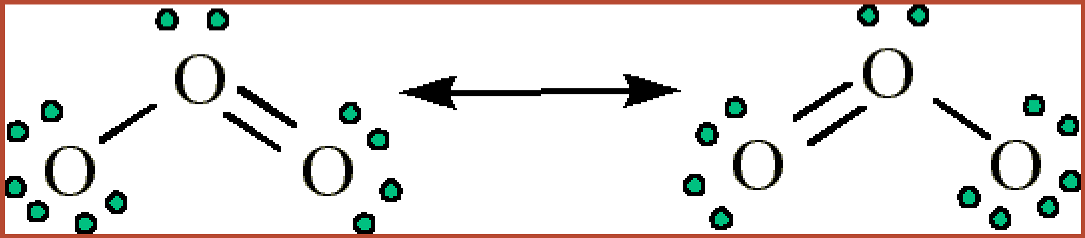
Structural formulas

Electron pairs in dot structures can be replaced by lines to make a structural formula.

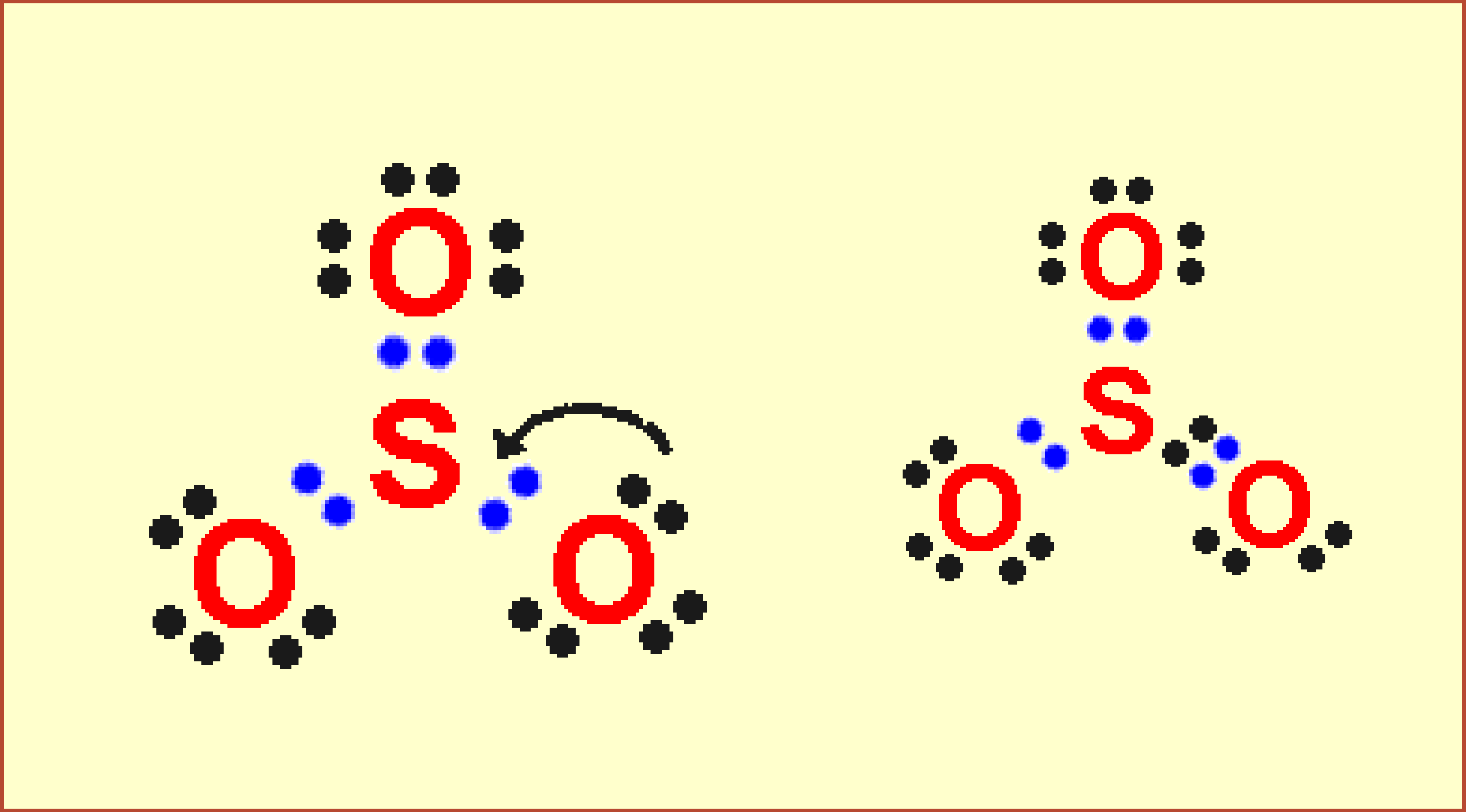
Single bond = 1 line	$Cl{-}Cl$
Double bond = 2 lines	$O{=}O$
Triple bond = 3 lines	$N{\equiv}N$

Covalent Bonds and Resonance Structures

- Resonance: when 2 or more equally valid electron dot structures can be written for a molecule
  - Ex) Ozone: O_3
- Proof: bond lengths are the same, there is no clear side for the single and the double bond



SO_3 Resonance Structures



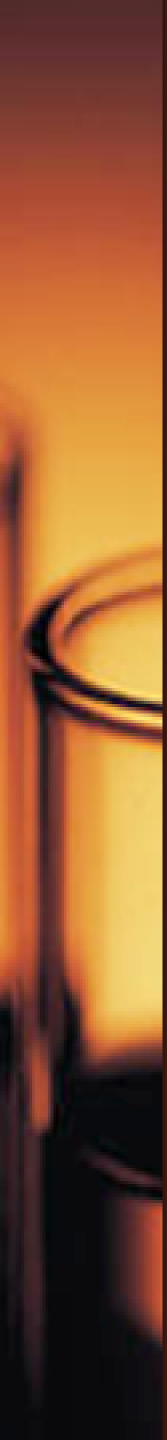
Section 6.3: Ionic Bonding

An **ionic compound** is composed of anions and cations combined so that the numbers of positive and negative charges are equal.

- A **formula unit** shows the lowest whole number ratio of atoms in an ionic compound.
- Ionic compounds exist as 3-dimensional arrays of ions held together by the force of attraction between the oppositely charged ions.
- Compounds comprised of metal-nonmetal bonding

Other Ions

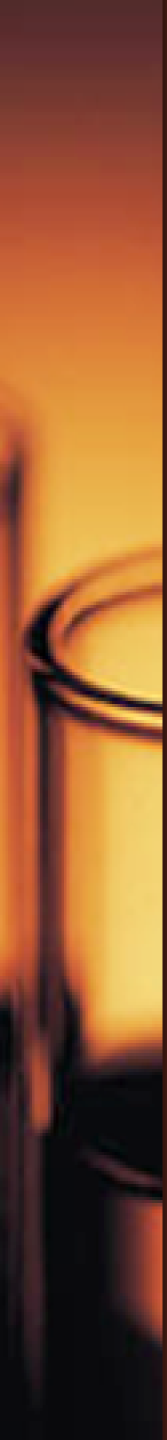
- Polyatomic ions – ions that are made up of more then one atom
  - Behave like atoms
  - Very common & stable in nature
  - Have special names:
  - Ammonium = NH_4^+



Ionic Compounds

- Ionic Compounds = neutral
- Combine a cation and an anion
- Cation + Anion → <u>Neutral</u> ionic compound

$$A^{+} \quad + \quad X^{-} \rightarrow AX$$

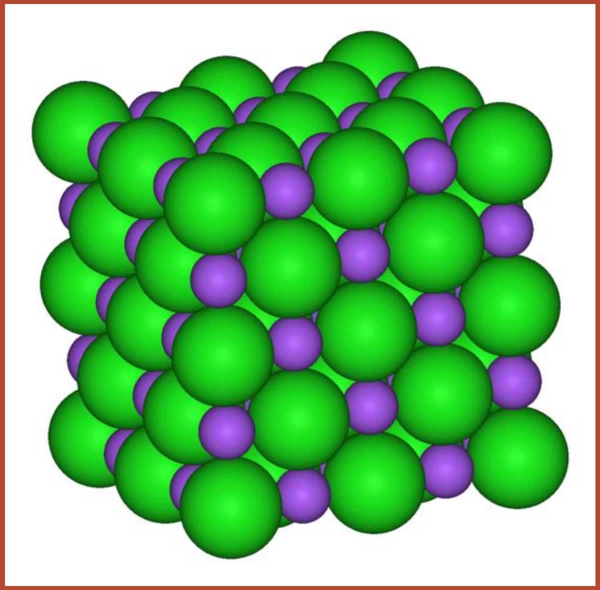


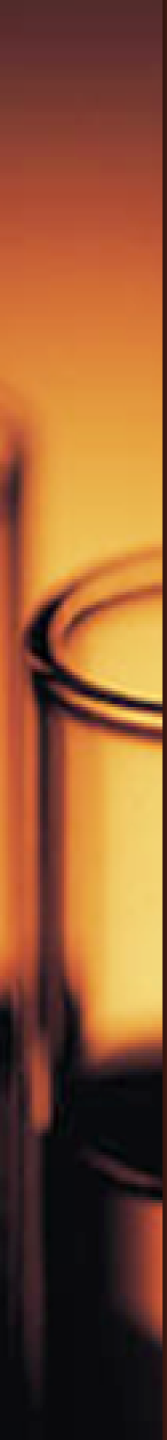
Properties of Ionic Compounds

- Repeating 3D patterns
- Most are crystalline solids at room temp
- Characteristics
  - High melting points
  - Conductivity in molten states
  - Existence in crystalline form
  - Tendency to dissolve in water
  - Produce electrical conductivity when dissolved in water

Coordination Number

- Definition – gives the number of ions of opposite charge that surround each ion in a crystal
- Ex) NaCl
- Coord # = 6





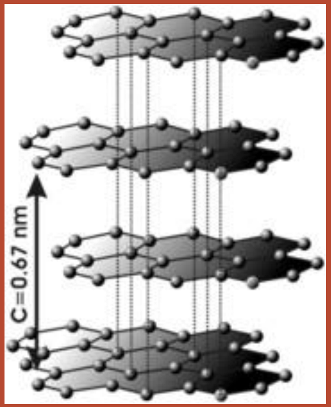
Properties of Molecular Substances

- Exist in all states of matter
- Melting points and boiling points are low compared to ionic compounds
- Some exceptions:
  - Network Solids – stable substances in which all of the atoms are covalently bonded to each other
  - All atoms are interconnected
  - Ex) Diamond & Silicon carbide

Allotropes

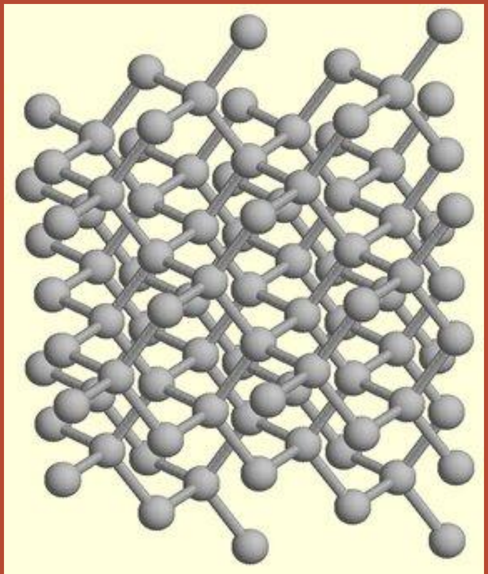
Definition – different forms of carbon with different types of bonding

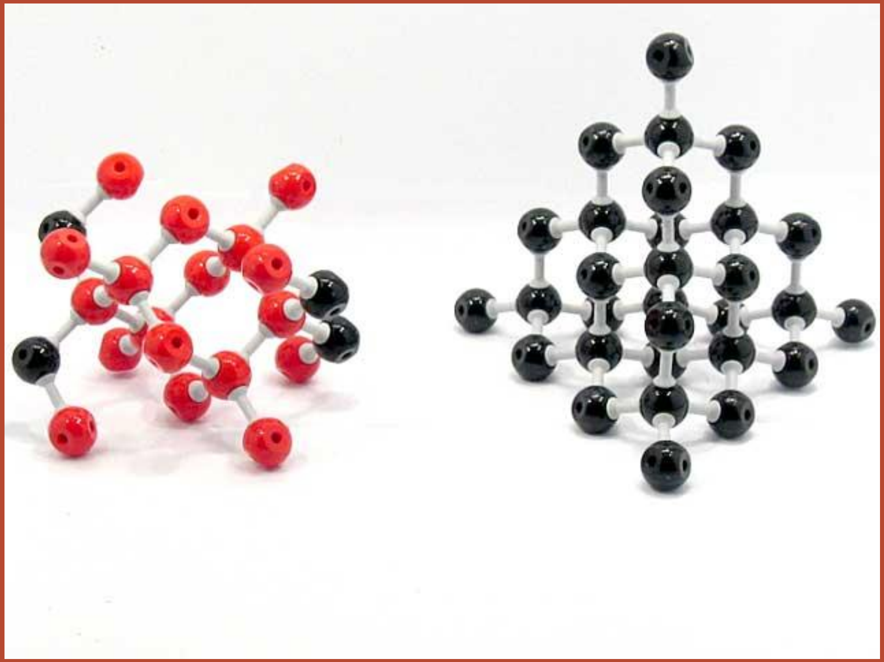
- Carbon – 3 allotropes
  1. Graphite



Allotropes, cont'd

2. Diamond

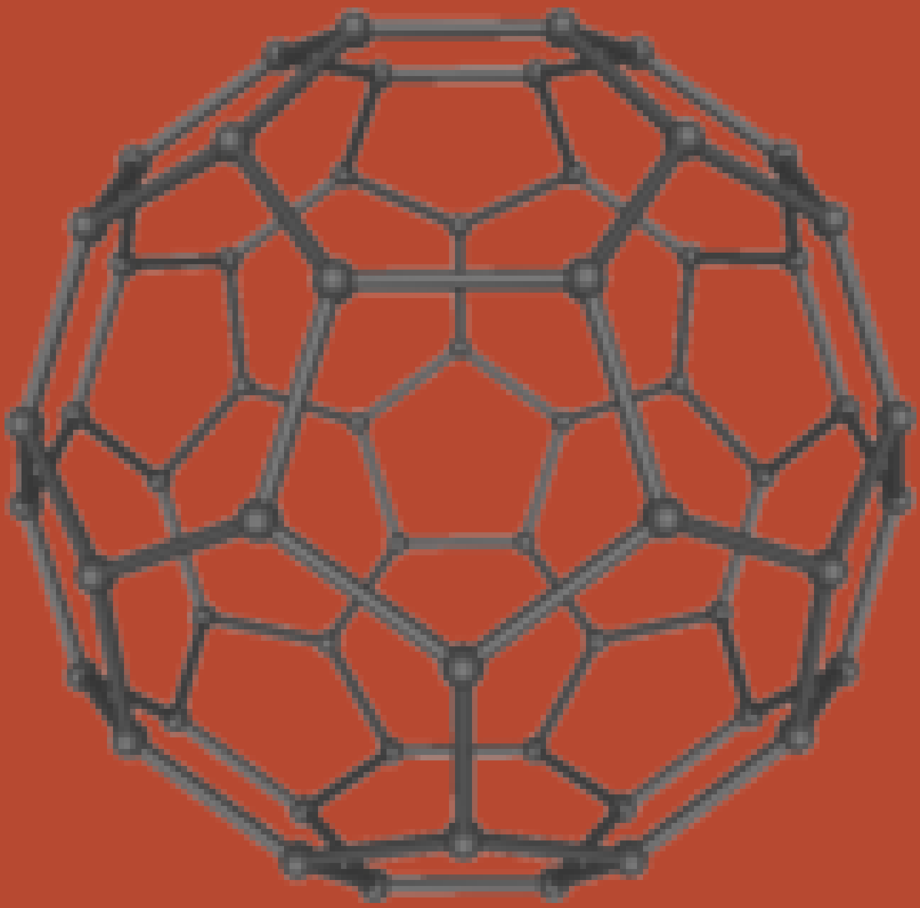




Allotropes, cont'd

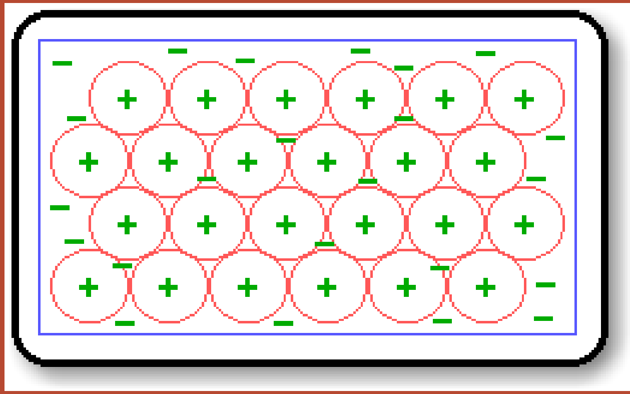
3. Buckminsterfullerene

   C_{60}



Section 6.4 – Metallic Bonds

- Definition – consist of the attraction of the free floating electrons for the positive charged metal ions



  - Metals are believed to be composed of closely packed cations
  - The cations are surrounded by mobile valence electrons

Requirements

1. Must have vacant valence electrons
2. Must have low ionization energies so that loosely held electrons are available for bonding

Properties of Metallic Bonded Substances

1. Good conductors of electricity
   - Explained by this model
   - Electrons enter one end of the metal and leave the other
2. Malleable – hammered into different shapes
3. Ductile – drawn into wires

Section 6.5 Molecular Geometry

- Unpaired electrons around a central atom play a large role in determining a molecule's 3-D shape
- Negatively charged electrons repel one another
  - electron pairs in different orbital stay as far apart as possible

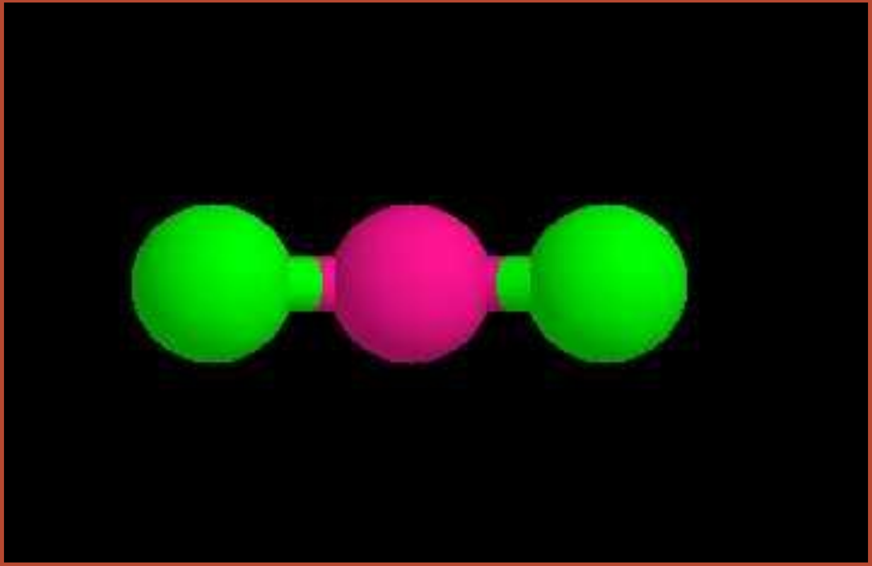
Valence-Shell Electron-Pair Repulsion Theory (VSEPR)

- VSEPR (Valence Shell) Theory:
  - The tendency of electron pairs to adjust the orientation of their orbitals to maximize the distance between them
  - Depends on the number of electrons or atoms bonded to a central atom
  - Bond angle: shape characterized between the central atom and the atoms bonded to it

VSEPR Models

of atoms or electron pairs attached to central atom: 2

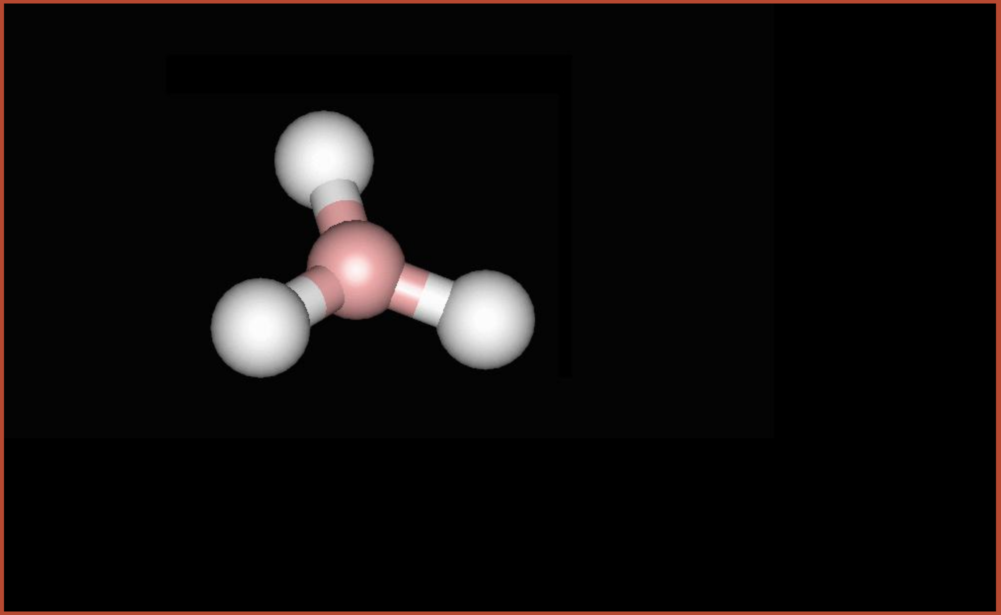
- # of unshared pairs: 0
- Bond Shape: Linear
- Bond Angle: 180º
- Examples
  - $BeCl_2$
  - CO_2
  - HCN



VSEPR Models

of atoms or electron pairs: 3

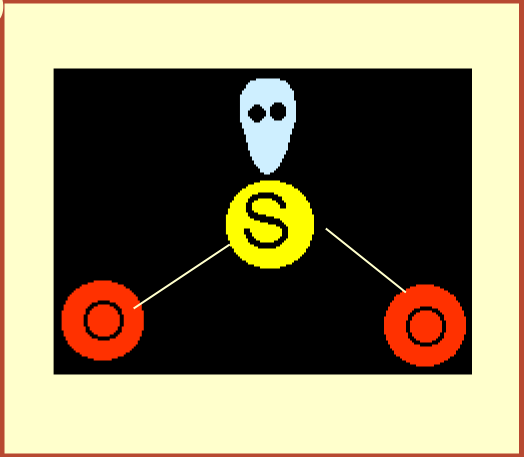
- # of unshared pairs: 0
- Bond Shape: Trigonal planar
- Bond Angle: 120º
- Examples:
  - BF_3
  - BH_3
  - SO_3

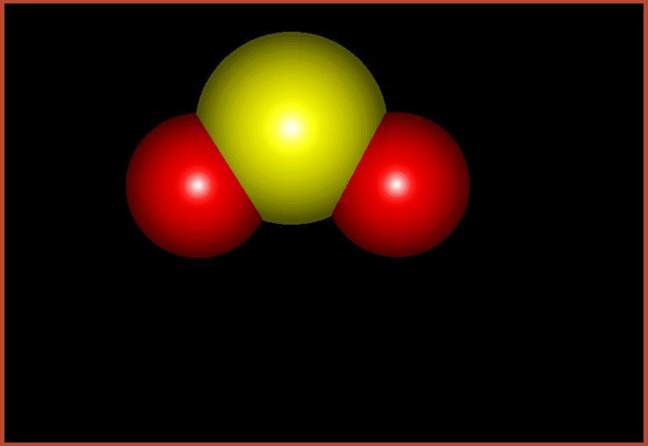


VSEPR Models

of atoms or electron pairs: 3

- # of unshared pairs: 1
- Bond Shape: Bent or Angular
- Bond Angle: < 120º
- Examples:
  - SO

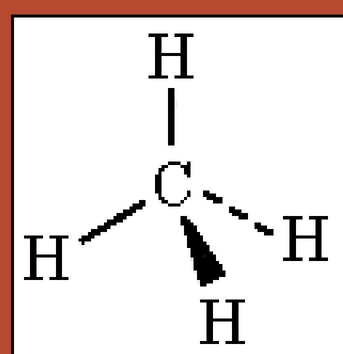


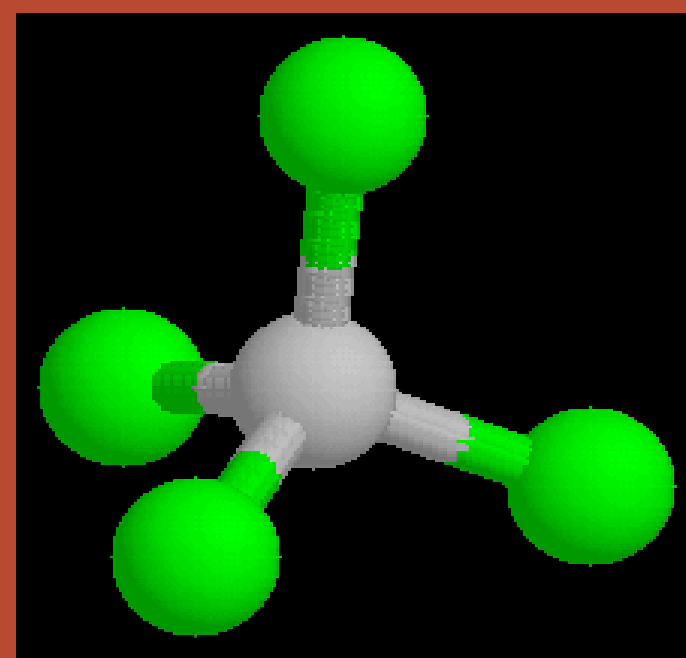


VSEPR Models

of atoms or electron pairs: 4

- # of unshared pairs: 0
- Bond Shape: Tetrahedral
- Bond Angle: 109.5º
- Examples:
  - CH_4
  - CH_2Cl_2
  - $SiCl_4$
  - $POCl_3$

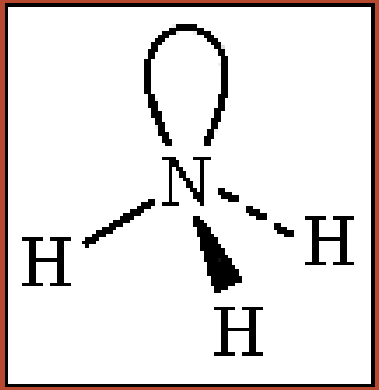


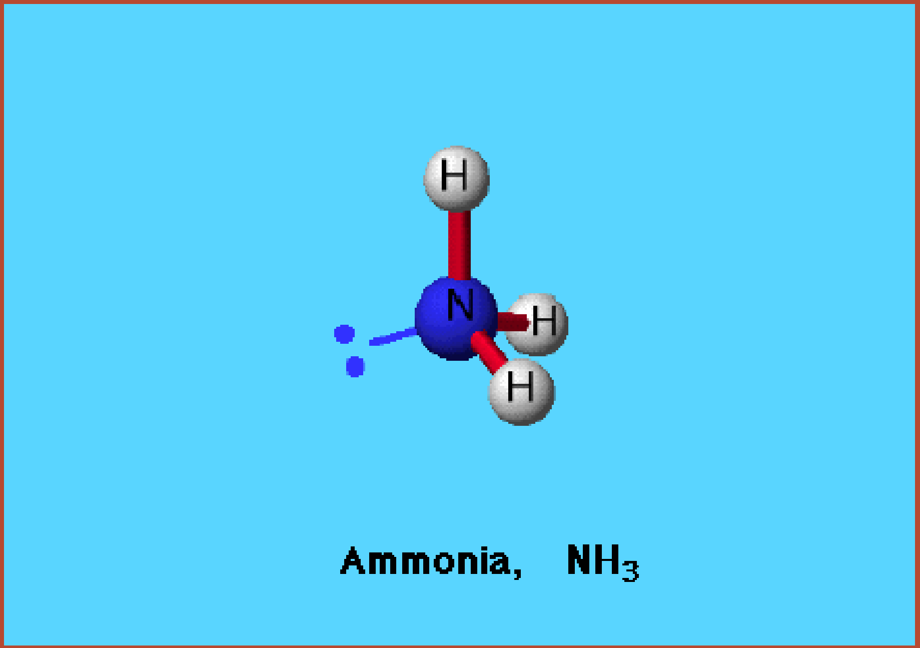


VSEPR Models

of atoms or electron pairs: 4

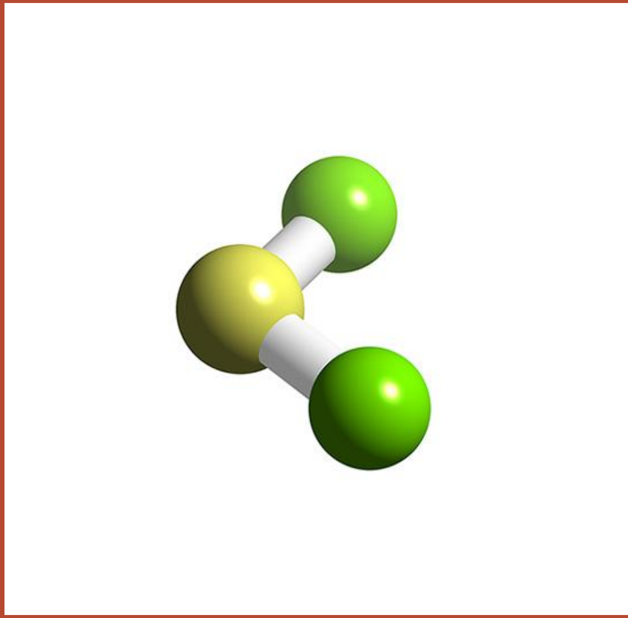
- # of unshared pairs: 1
- Bond Shape: Trigonal pyramidal
- Bond Angle: < 109.5º
- Examples:
  - NH_3
  - PF_3
  - NH_2Cl





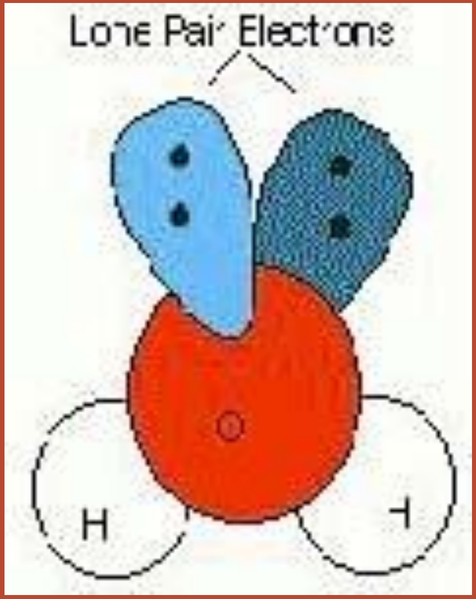
VSEPR Models

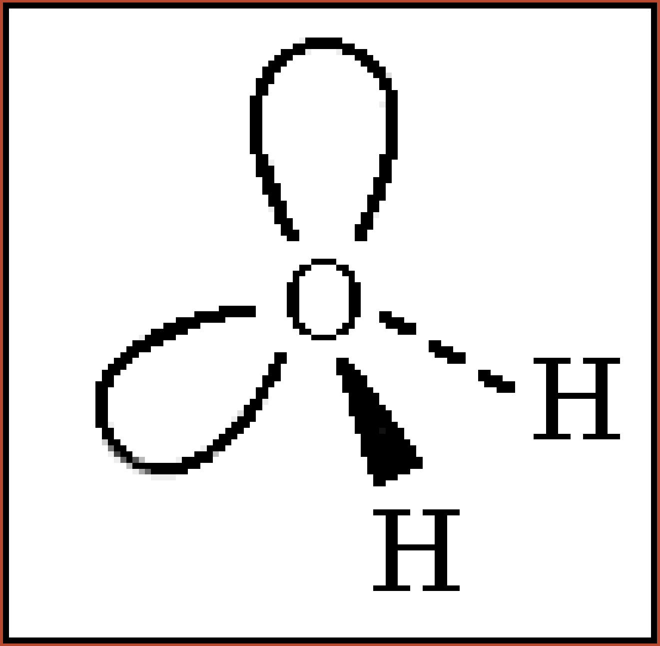
- # of atoms or electron pairs: 4
  - # of unshared pairs: 2
  - Bond Shape: Bent or Angular
  - Bond Angle: < 109.5°
  - Examples:
    - F_2O
    - BrO_2
    - SCl_2



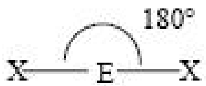
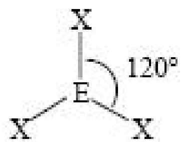
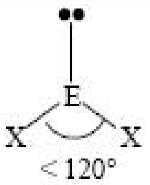
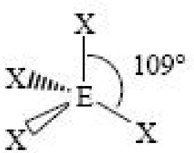
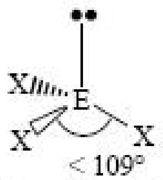
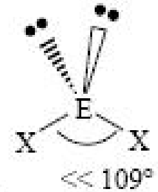
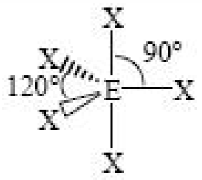
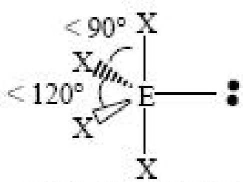
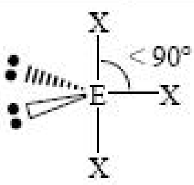
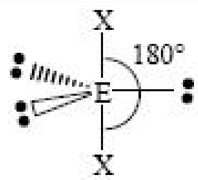
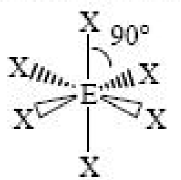
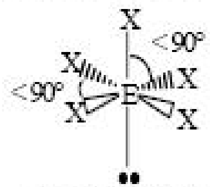
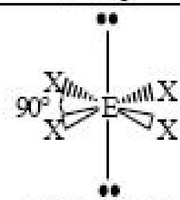
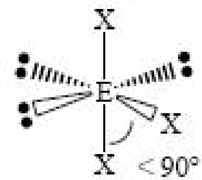
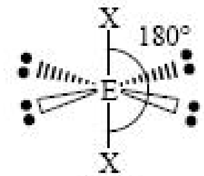
VSEPR Models

- # of atoms or electron pairs: 4
  - # of unshared pairs: 2 (Water)
  - Bond Shape: Bent or Angular
  - Bond Angle: 104.5º





VSEPR Geometries

Steric No.	Basic Geometry 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 **Linear**				
3	 **Trigonal Planar**	 **Bent or Angular**			
4	 **Tetrahedral**	 **Trigonal Pyramid**	 **Bent or Angular**		
5	 **Trigonal Bipyramid**	 **Sawhorse or Seesaw**	 **T-shape**	 **Linear**	
6	 **Octahedral**	 **Square Pyramid**	 **Square Planar**	 **T-shape**	 **Linear**

VSEPR

- To predict the shape of a molecule according to VSEPR theory

1. Draw a Lewis Formula
2. Count the number of atoms bonded to the central atom, and count the unshared electron pairs on the central atom

VSEPR

Predicting the shape of a molecule…

3. Add the numbers of atoms and the number of unshared electron pairs around the central atom. The total indicates the parent structure
4. The molecular shape is derived from the parent structure by considering only the positions in the structure occupied by bonded atoms

Hybridization

- Definition: the mixing of two or more atomic orbitals of similar energies in the same atom to give new orbitals of equal energies
  - Orbitals combine and rearrange

Hybridized Orbitals

- Definition – orbitals of equal energy produced by the combination of two or more orbitals of the same atom
- 3 Types
  - sp
  - sp^2
  - sp^3 (the major one we are going to look at)

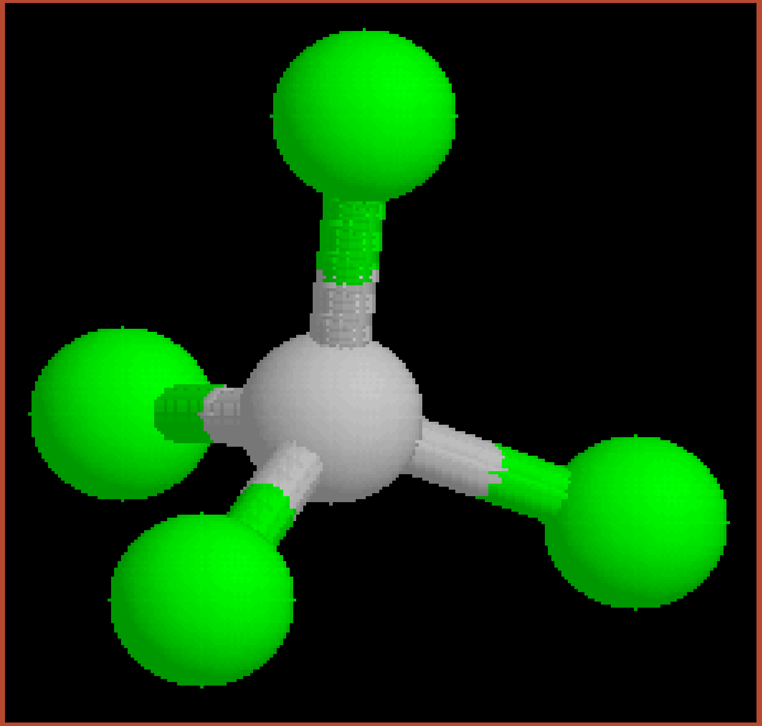
Hybridization

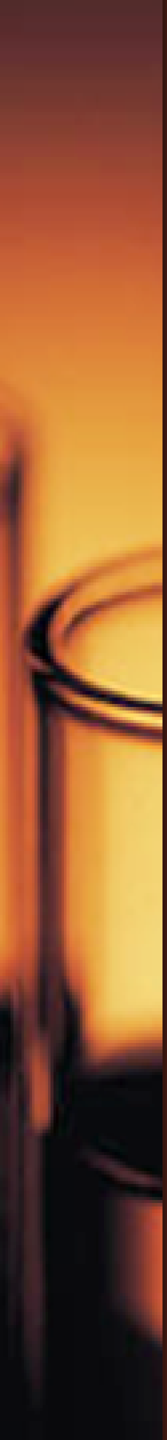
- Example: Carbon: $1s^2\ 2s^2\ 2p^2$
  - Has 4 valence electrons, two in 2s & two in 2p
  - Method: Hybridization
    - One s orbital: $\underline{\uparrow\downarrow}$
    - Two p orbitals: $\underline{\uparrow}\ \underline{\uparrow}\ \underline{\quad}$
    - They hybridize to make an sp^3: $\underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}$
      - **Four new identical bonds!**
  - <u>Now carbon can make 4 bonds!</u>

Hybridized Orbitals

sp^3 – contains 1 s and 3 p orbitals

- Resulting shape – tetrahedral
- Ex) CH_4





Intermolecular Forces

- Intermolecular Forces
  - An attractive force that operates between molecules
  - * DO NOT confuse with bonds! *
    - Bonds: attractive forces that hold atoms together in molecules
  - IMF are much weaker than bonding forces

Intermolecular Forces

van der Waals forces: collection of the weak interactions

Types:

1. London dispersion force
2. Dipole-dipole force (already covered)
3. Hydrogen-bonding force

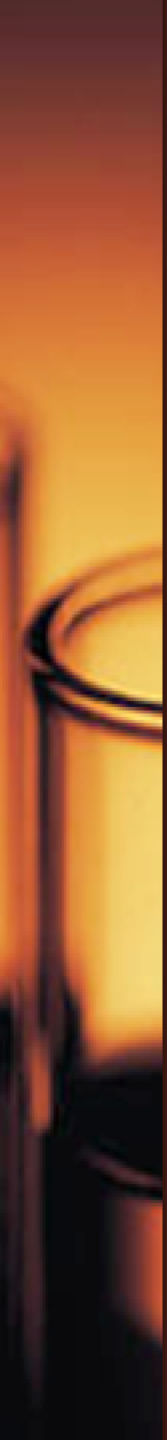
London Dispersion Forces

Electrons are in constant motion and aren’t always equally distributed

- Therefore they develop a temporary dipole, known as an induced dipole
- The effect passes onto other atoms, like a domino effect… and so on, and so on…

London Dispersion Forces

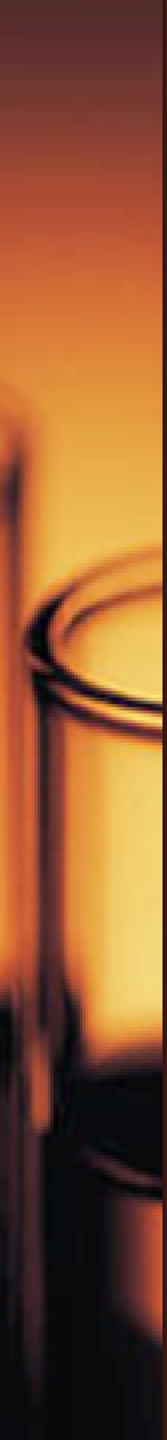
- – Attraction between temporary dipoles of molecules



London Dispersion Forces (L.D.F.)

- What do we know?

1. Occur between all atoms and molecules
2. The only intermolecular force at work in nonpolar substances
3. Relatively weak



London Dispersion Forces (L.D.F.)

- What do we know?
  - Tend to be <u>stronger</u> the <u>larger</u> the atom or molecule is
    - Therefore: at room temperature (Column 17-the halogens)
      - Cl_2 is a gas
      - Br_2 is a liquid
      - I_2 is a solid

Dipole-Dipole Forces

- Dipole-Dipole Force:
  - Attractions among ***polar molecules***
  - Electronegativity of atoms determines which part is the:
    - Partial positive (δ +)
    - Partial negative (δ -)
      - Positive and negative parts attract!

Hydrogen Bonding

- <u>Hydrogen Bonding</u>:
  - An especially strong dipole-dipole force between polar molecules that contain hydrogen attached to a highly electronegative element
    - Although, there is no bond between molecules in the "usual" sense
    - H-bond is a special type of dipole-dipole force