VSEPR & Molecular Geometry

Covalent Bonds

- The forces of attraction are much weaker than those of ionic bonds.
- Molecules melt at low temperatures.
- **Can not** conduct electricity in solution.
- Some have the ability to share more than one pair of electrons, forming multiple bonds.
  - *Double bond:* two pair of electrons (4 total) are shared between the two atoms (ex. O_2)
  - *Triple bond:* three pair of electrons (6 total) are shared between the two atoms (ex. N_2)

- The distance between the nuclei of two bonded atoms is called the *bond length*.
- Energy is released (*exothermic*) when a bond forms and is absorbed (*endothermic*) when a bond breaks
- The amount of energy required to break a covalent bond is called the *bond dissociation energy*.
- The stronger the bond, the greater the bond dissociation energy and, therefore, the more difficult it is to break the bond.
- Shorter bonds have greater bond dissociation energies than longer bonds.

Single bonds < double bonds < triple bonds

Lewis Structures: Uses electron dots to show the arrangement of electrons in a molecule

Steps:

1. Predict the location of the atoms
   a) Hydrogen is always a terminal atom
   b) Center atom is least electronegative element
   c) Carbon (if present) is always a center atom

2. Count the total number of valence electrons in the elements to be combined

3. Determine the number of "pairs" of electrons in the molecule by dividing the total number of valence electrons by 2

4. Place a single line ("bonding pair") between the center atom and the terminal atoms

5. Subtract the number of pairs used from the total number of pairs available

6. Starting with the terminal atoms, add unshared pairs so that each atom is surrounded by eight electrons (remember hydrogen only shares one pair)

7. If the center atom does not have an octet, one or two of the lone pairs around the terminal atoms must be converted to form multiple bonds

   In general, carbon, nitrogen, oxygen and sulfur can form double or triple bonds.

Resonance Structures

- Occurs when more than one valid Lewis Structure can be written for a molecule or ion.
- Only differ in the position of electron pairs, never in the atoms position.

Example:
- O_3 (ozone)
- NO_2^{-1} (nitrate)
- CO_3^{-2} (carbonate)

Exceptions to the Octet Rule

1) Fewer than eight electrons around the atom (hydrogen and boron containing compounds such as BH_3)

2) Odd number of total valence electrons (These compounds usually form polyatomic ions to "make-up" the difference) ex. ClO_2

2) Expanded Octets: central atom contains more than 8 electrons Usually occur with non-metals beyond period 3 when bound to highly electronegative elements fluorine, oxygen, and chlorine. ex. SF_6

Molecular Geometry (Shape)

- *VSEPR Theory*: "**v**alence-**s**hell **e**lectron-**p**air **r**epulsion."
- The electron pairs are oriented as far away from each other as possible to minimize the repulsion around the center atom.
- The shape of a molecule refers to the *positions of atoms only*.

5 main shapes (based on the octet rule)

1) Linear
2) Bent
3) Trigonal planar
4) Trigonal pyramid
5) Tetrahedral

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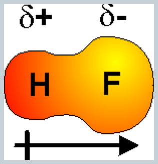
6) Trigonal bipyramid
7) Octahedral

Electronegativity and Polarity

- *Electronegativity*: a measure of the tendency of an atom to attract electrons in a chemical bond

- *Polarity*: the uneven distribution of electrons (molecule is asymmetrical around the center atom)

For polar covalent bonds, a dipole is established.



Forces

- The forces of attraction **within** a compound are known as *intra molecular forces.* (holds together the atoms making up a compound)
  - Ionic
  - Covalent
  - Metallic

- The forces of attraction **between molecules** are known as *inter molecular forces.*

1) London Dispersion Forces

- Very weak forces of attraction between *non-polar* molecules
- Result from the **temporary** dipole occurring as molecules approach one another
- The more electrons that are present, the stronger the dispersion forces will be.

- Polar molecules have *dipoles* (partial positive and partial negative regions.)

2) Dipole-dipole Forces

- Occur between *polar* molecules,
- The partial negative region in one molecule attracts the partial positive region in a neighboring molecules. There is an electrostatic attraction between the molecules.

3) Hydrogen Bonding

- Hydrogen bound to an atom that has lone pairs of electrons
- The hydrogen atom is attracted to an unshared pair of electrons on a neighboring molecule.
- Are the strongest intermolecular force