Chem 1025 General Chemistry by Zumdahl Decoste, 6th ed
Prof George W.J. Kenney, Jr **Last Update: 28-Nov-2009** *Draft*

Chapter 12: Chemical Bonding

These Notes are to SUPPLIMENT the Text, They do NOT Replace reading the Text Material. Additional material that is in the Text will be on your tests!

To get the most information, READ THE CHAPTER prior to the Lecture, bring in these lecture notes and make comments on these notes. These notes alone are NOT enough to pass any test!

Bond: a force that holds groups of two or more atoms together and makes them function as a unit

Bond Energy: the energy required to break the bond

Ionic Bonding: Ionic substances are formed when an atom loses an electron relatively easily [$Na \rightarrow Na^+ + 1e^-$] and reacts with an atom that has a high affinity for electrons [$Cl + 1e^- \rightarrow Cl^-$] and this forms an **Ionic Compound [$Na^+ + Cl^- \rightarrow NaCl$].**

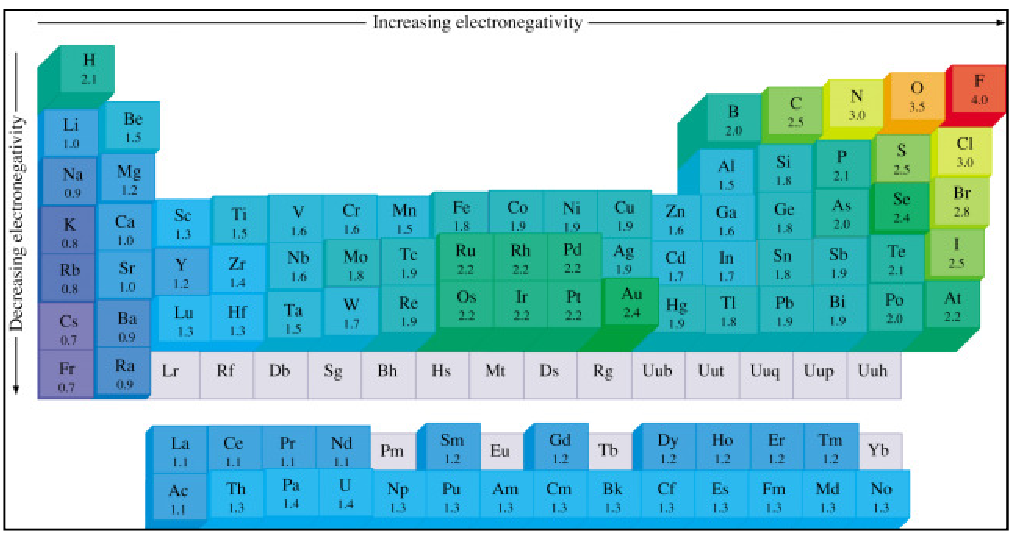
Covalent Bonding: Electrons are shared by the nuclei of two atoms. Hydrogen forms H2 which is H : H where the electrons reside in the space between the two nuclei and around the two atoms.

Pic of Hydrogen and electrons

Polar Covalent Bond: An unequal sharing of electrons due to the difference in electronegativity of each atom.

Pic of HF

Electronegativity: is the relative ability of an atom in a molecule to attract shared electron to itself.



The higher the Electronegativity value, the more that atom attracts an electron.

Left side – Cesium (Cs) has a low Electronegativity of 0.7 – it gives up electrons

Right Side – Fluroine (F) has a high Electronegativity of 4.0 – it sucks up electrons

Bond Polarity is related to the difference between the Electronegativity values of the different atoms

Difference	Bond type	Example
Greater than 1.8	Ionic Bonds	NaCl $MgBr_2$
Greater than 0.5	Polar / Polar Covalent Bonds	H-Cl H-F H-O-H
Less than 0.5	Covalent Bonds which are Non Polar.	H:H CH_4 $H_3C\text{-}CH_3$

Dipole Moment: a molecule where the electrons reside more towards one atom than the other due to differences in Electronegativity.

Example: **$^{+}H - F^{-}$** **H_2O [show in 3d]**

Electron Configurations: [See Periodic Table]

Name	Symbol	# e^-	Config	Loss/Gain Equation	Final Config
Sodium	Na	11	$1s^2\ 2s^2\ 2p^6\ 3s^1$	$Na \rightarrow Na^{+1} + 1\ e^-$	$1s^2\ 2s^2\ 2p^6$
Magnesium	Mg	12	$1s^2\ 2s^2\ 2p^6\ 3s^2$	$Mg \rightarrow Mg^{+2} + 2\ e^-$	$1s^2\ 2s^2\ 2p^6$
Aluminum	Al	13	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	$Al \rightarrow Al^{+3} + 3\ e^-$	$1s^2\ 2s^2\ 2p^6$
Oxygen	O	8	$1s^2\ 2s^2\ 2p^4$	$O + 2\ e^- \rightarrow O^{2-}$	$1s^2\ 2s^2\ 2p^6$
Fluorine	F	9	$1s^2\ 2s^2\ 2p^5$	$F + 1\ e^- \rightarrow F^{1-}$	$1s^2\ 2s^2\ 2p^6$
Neon	Ne	10	$1s^2\ 2s^2\ 2p^6$	No Reaction	

Note: All of the above lose or gain electrons to form the Neon Noble Gas Configuration!
The Metals Group 1 [Na], Group 2 [Mg] and Group 3 [Al] lose electrons to form a noble gas configuration
The Nonmetals Group 6 [Oxygen] and Group 7 [Fluorine] gain electrons to form a noble gas configuration.

Predicting Formula of Ionic Compounds

1. Pick a Group 1, 2 or 3 — Ca
2. Pick a Group 8 or 9 — O
3. Show the gain or loss of electrons — $Ca \rightarrow Ca^{+2} + 2\ e^-$ $O + 2\ e^- \rightarrow O^{2-}$
4. Show the product formed! — $Ca + O \rightarrow Ca\,O$

<u>Common Ions and the Noble Gas Configurations</u>

Group 1	Group 2	Group 3	Group 6	Group 7	Noble Gas
Li^+	Be^{2+}				He
Na^+	Mg^{2+}	Al^{3+}	O^{2-}	F^-	Ne
K^+	Ca^{2+}		S^{2-}	Cl^-	Ar
Rb^+	Sr^{2+}		Se^{2-}	Br^-	Kr
Cs^+	Ba^{2+}		Te^{2-}	I^-	Xe

Ionic Bonding and Structures

Cation [$Li \rightarrow Li^+ + 1\ e^-$] is smaller than the parent ion as it loses an electron

Anion [$O + 2\ e^- \rightarrow O^{2-}$] is larger than the parent ion as it gains electrons

Lewis Structure

The Lewis Structure is a representation of a molecule that shows how the valenca electrons are arranged among the atoms in the molecule.

KBr Draw K^+ with no electrons and Br^- with 8 electrons

1. Hydrogen forms a stable molecule sharing two electrons: H:H
2. Helium [and the other noble gases] do not form bonds as their outer electron shell is filled
3. 2n Period Non-Metals [Carbon → Fluorine] form stable molecules when they absorb electrons to fill the valence orbitals [s^2 p^6].

Octet Rule: Atoms like to be surrounded by eight electrons
Bonding Pair: Electrons that are shared with another atom
Lone Pair: Electrons that are not involved in bonding Show example of H-F

Rules for Lewis Structure:

1. Calculate the sum of all of the valence electrons from all of the atoms
2. Use one pair [2 electrons] to form a bonding pair between each bounded atoms
3. Arrange the remaining electrons to satisfy the Octet Rule

Write Lewis Structures for: H_2O

Multiple Bonds

In step 3, if there are not enough electrons to fill the Octet Rule, share two pairs of electrons between atoms

Example: CO_2 2 Double Bonds

HCN 1 Triple Bond

Write Lewis Structures for:

HF	N_2	NH_3	CH_4
CF_4	NO^+	NO_3^-	
NF_3	O_2	CO	PH_3
H_2S	SO_4^{2-}	NH_4^+	SO_2

Molecular Structure or Geometric Structure represents the 3-D arrangement of atoms in space

Water: H-O-H does not exist in a straight line, the 3 D pic shows water at: DRAW ON BOARD ∟= 105^o

Carbon Dioxide: CO_2 is a linear structure.

Boron TriFluoride, BF_3 is Trigonal Planar with 120^o bond angles.

Methane, CH4 is tetrahedral

VSEPR: Valence Shell Electron Pair Repulsion: [I'll leave this for your 1045 class!]