CHM 1020 Lecture Part 4

Octet Rule

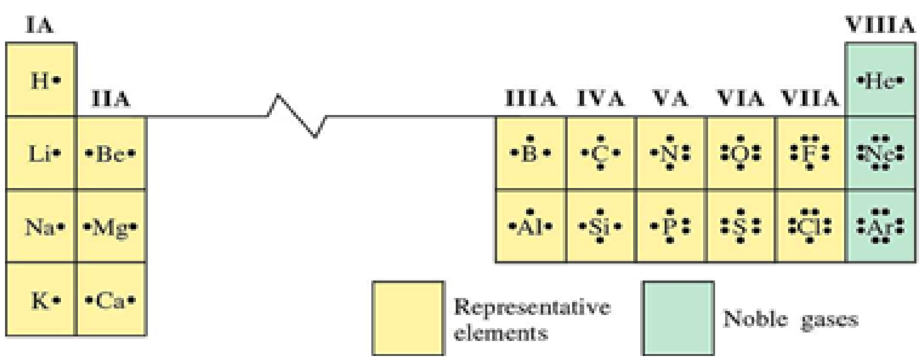
- **Octet rule:** Atoms in a compound will lose, gain or share electrons in order to achieve a stable noble gas configuration. (memorize this rule)
- It is the electrons in the outer shell that participate in these changes to create bonds.

Valence electrons

- The valence electrons of an atom are defined as the electrons in the outermost shell of the uncharged atom.
- The number of valence electrons of an uncharged atom is equal to the group number for main group elements.

Dot structures

- In Lewis dot structures, the valence electrons are represented by dots.
- Lewis dot structures play a more important role in covalent bonding than ionic bonding.
- Sodium, in group I has 1 valence electron
- Carbon in group IV has 4 valence electrons.



Types of bonding

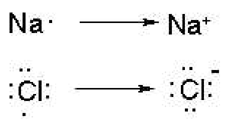
- metal + non-metal leads to ionic bonding.
  - the metal will lose electrons to become positively charged.
  - the non-metal will gain electrons to become negatively charged.
  - the ionic compound is held together by the electrostatic attraction between the positive and negative charges.
- Non- metal + non-metal leads to covalent bonding.

Table Salt

- Sodium is a metal, how many electrons will it lose? It will lose one electron which will give it the same noble gas configuration as Neon.
- Chlorine is a non-metal. It will gain 1 electron to achieve the same configuration as argon.
- Is Cl^- the same as argon? No! They have different numbers of protons.

Salt and Dot Structures

- The sodium and chlorine combine in a 1:1 ratio.



The Na^+ and the Cl^- are attracted to each other.

Charges of some Common Monatomic ions

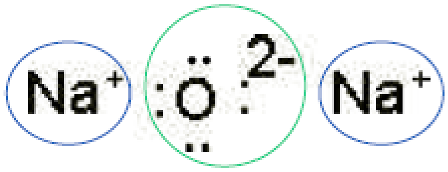
H 1+ 1-																	
Li 1+	Be 2+													N 3-	O 2-	F 1-	
Na 1+	Mg 2+											Al 3+				Cl 1-	
K 1+	Ca 2+	Sc 3+	Ti 3+ 4+	V 3+ 4+	Cr 2+ 3+	Mn 2+ 3+	Fe 2+ 3+	Co 2+ 3+	Ni 2+ 4+	Cu 1+ 2+	Zn 2+					Br 1-	
Rb 1+	Sr 2+								Pd 2+ 4+	Ag 1+	Cd 2+		Sn 2+ 4+			I 1-	
Cs 1+	Ba 2+								Pt 2+ 4+	Au 1+ 3+	Hg 2+ *		Pb 2+ 4+				
Fr 1+	Ra 2+																

Please note that many of the metals shown here can have more possibilities that I can show here. Vanadium, for example, can be 2+, 3+, 4+ or 5+. I have only shown the more common charges.

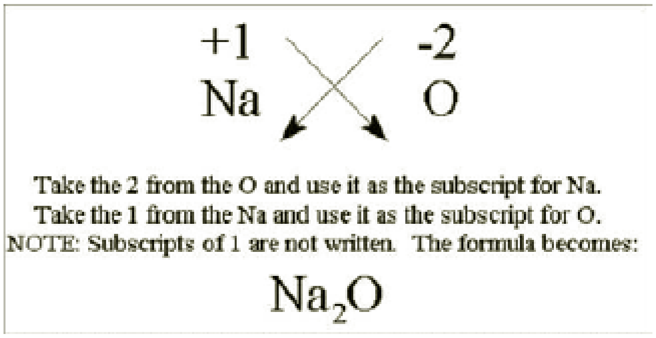
*Mercury can be 1+ in the polyatomic ion Hg_2^{2+}.

Does it have to be a 1:1 ratio? No!

- All ionic compounds must have no overall charge so positive charges must equal negative charges.
- Example sodium oxide: Na_2O



The Switcheroo Rule



Take the 2 from the O and use it as the subscript for Na.
Take the 1 from the Na and use it as the subscript for O.
NOTE: Subscripts of 1 are not written. The formula becomes:

Na_2O

Note that the cation is written first and the anion second.

A Caveat to the Switcheroo rule

- If you can divide by an integer greater than one, you must do so.
- Mg^{2+} and O^{2-} form MgO not Mg_2O_2

Polyatomic ions

- Polyatomic ions are groups of covalently bound atoms that act like a single ion.
- Example: nitrate NO_3^- combines with Mg^{2+} to form $Mg(NO_3)_2$.
- Note the use of ( ) to identify that it is 2 nitrates.

- Pb^{2+} and OH^- form:

Memorize these ten polyatomic ions

Formula	Name
NH_4^+	Ammonium
OH^-	Hydroxide
NO_3^-	Nitrate
$CH_3CO_2^-$	Acetate
CN^-	Cyanide
ClO_3^-	Chlorate
CO_3^{2-}	Carbonate
HCO_3^-	Bicarbonate
SO_4^{2-}	Sulfate
PO_4^{3-}	Phosphate

Rules for naming simple ionic compounds.
1. Name the metal by its elemental name.
2. Name the nonmetal by its elemental name and an -ide ending.
3. Name metals that can have different oxidation states using roman numerals to indicate positive charge. Example Fe^{2+} is Iron(II)
(See table "Charges of some Common Monatomic ions" to determine which metals can have more than one positive charge.)
4. Name polyatomic ions by their names.

Ionic Nomenclature Practice

- $CoCl_2$
- $Sn(ClO_2)_2$
- K_2S
- $NH_4C_2H_3O_2$
- $Mg(NO_2)_2$
- AgI

More Practice On the Web
- **Nomenclature Activity**
- **game**
- **Worksheet**
- **More Worksheets**