Chapter 7 – Chemical Formulas and Chemical Compounds

7-1 Chemical Names and Formulas

I. Significance of a Chemical Formula

A. Molecular formulas

1. Number of atoms of each element in one molecule of a compound

C_2H_6 = ethane (2 carbon atoms, 6 hydrogen atoms)

B. Ionic Compounds

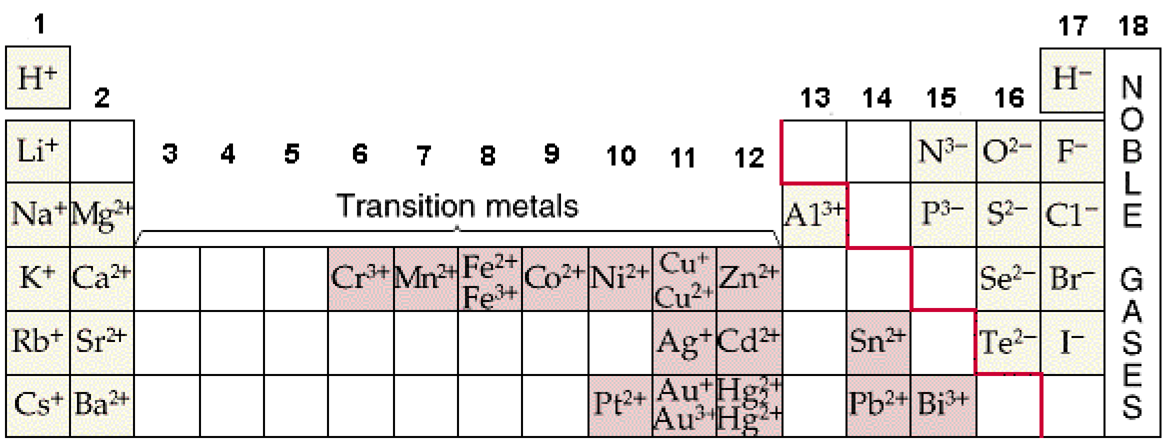
1. Represents the simplest whole number ratio of the compounds cations and anions

$Al_2(SO_4)_3$ = aluminum sulfate (2 aluminum ions, 3 sulfate ions)

II. Monatomic Ions

A. Monatomic ions

1. Ions formed from a single atom



B. Naming Monatomic Ions

1. Monatomic cations are
   a. Identified by the element's name
2. Monatomic anions
   a. Drop the ending of the element name
   b. Add an "–ide" ending

III. Binary Ionic Compounds

A. Binary Compounds

1. Compounds composed of two different elements

B. Writing Formulas for Binary Ionic Compounds

1. Write the symbols for the ions side by side. ALWAYS write the cation first!
2. Cross over the charges by using the absolute value of each ion's charge as the subscript for the other ion
3. Check that the subscripts are in smallest whole number ratio

C. Naming Binary Ionic Compounds from Their Formulas

1. Name the cation
2. Name the anion

D. The Stock System of Nomenclature
    1. Roman numerals are used to denote the charge of metals that can form two or more cations.
    2. The numeral is enclosed in parentheses and placed immediately after the metal name
        a. Iron(II) and Iron(III), pronounced "iron two" and "iron three"
    3. Roman numerals are never used:
        a. For anions
        b. For metals that form only one ion

E. Compounds Containing Polyatomic Ions
    1. Oxyanions
        a. Polyatomic anions that contain oxygen
    2. Naming a series of similar polyatomic ions

ClO^-	ClO_2^-	ClO_3^-	ClO_4^-
Hypochlorite	Chlorite	Chlorate	Perchlorate

    3. Naming compounds containing polyatomic ions
        a. Same as for monatomic ions
    4. Writing formulas including polyatomic ions
        a. Use parentheses when you need MORE THAN one of a polyatomic ion
        b. Parentheses are NEVER used for monatomic ions, regardless of how many are in the formula

IV. <u>Naming Binary Molecular Compounds</u>

A. Binary Molecular Compounds
    1. Covalently bonded molecules containing only two elements, both nonmetals

B. Naming
    1. Least electronegative element is named first
    2. First element gets a prefix if there is more than 1 atom of that element
    3. Second element ALWAYS gets a prefix, and an "-ide" ending

Examples: N_2O_3 = dinitrogen trioxide
CO = carbon monoxide, **not** monocarbon monoxide

Table 7-3	**Numerical Prefixes**									
Number	1	2	3	4	5	6	7	8	9	10
Prefix	mono	di	tri	tetra	penta	hexa	hepta	octa	nona	deca

V. <u>Covalent Network Compounds</u>

A. Naming
    1. Use the same system as binary molecular compounds (prefixes)

VI. Acids and Salts

A. Binary Acids

1. Acids that consist of two elements, usually hydrogen and one of the halogens

B. Oxyacids

1. Acids that contain hydrogen, oxygen and a third element (usually a nonmetal)

C. Naming Acids

1. Refer to the "Naming Acids" worksheet

7-2 Oxidation Numbers

Oxidation Number – numbers assigned to atoms composing a compound or ion that indicate the general distribution of electrons among bonded atoms

I. Assigning Oxidation Numbers

A. Rules

Rule	Example
1. The atoms of a pure element have an ox. # of zero	$\overset{0}{Au}\quad \overset{0}{F}_2$
2. The more electronegative element in a binary compound is assigned the # equal to the charge it would have as an anion. The less electronegative is assigned the # equal to the charge it would have as a cation	$\overset{+5}{As}_2\,\overset{-2}{S}_5$
3. Fluorine has an ox. # of –1 in all of its compounds	$\overset{+4}{C}\,\overset{-1}{F}_4$
4. Oxygen has an ox. # of –2 in almost all compounds. It is a –1 in peroxides, and a +2 in compounds with fluorine	$\overset{+4}{S}\,\overset{-2}{O}_2$
5. Hydrogen has an ox. # of +1 in all compounds containing elements that are more electronegative than it; it has an ox. # of –1 in compounds with metals	$\overset{+1}{H}_2\,\overset{-2}{O}$
6. The algebraic sum of the ox. #'s of all atoms in a neutral compound is equal to zero	$\overset{-4}{C}\,\overset{+1}{H}_4$
7. The algebraic sum of the ox. #'s of all atoms in a polyatomic ion is equal to its charge	$\overset{+6}{S}\,\overset{-2}{O}_4{}^{2-}$
8. For monatomic ions, the ox. # is equal to the charge	$\overset{+2}{Mg}\,\overset{-1}{Cl}_2$

II. Using Oxidation Numbers for Formulas and Names

Be aware that the stock number (Roman numerals) system may be used for molecular compounds as well as ionic compounds. In this course, we will use the traditional (prefix) method of naming binary molecular compounds.

7-3 Using Chemical Formulas

I. Formula Masses

A. Formula Mass

1. The sum of the average atomic masses of all the atoms represented in the formula of a molecule, formula unit, or ion

Formula Mass of glucose, $C_6H_{12}O_6$:

C = 12.01 amu	6 x 12.01 amu = 72.06 amu
H = 1.01 amu	12 x 1.01 amu = 12.12 amu
O = 16.00 amu	6 x 16.00 amu = 96.00 amu
	Formula Mass = 180.18 amu

B. Molar Masses

1. A compound's molar mass is numerically equal to it formula mass, but expressed in units of grams/mole (g/mol)

Molar Mass of glucose, $C_6H_{12}O_6$ = 180.18 g/mol

II. Molar Mass as a Conversion Factor

A. Converting moles of compound to grams

$$Amount\ \ in\ moles\ \ x\ \ molar\ \ mass(g/mol) = Mass\ \ in\ \ grams$$

B. Converting grams of compound to mass

$$Mass\ \ in\ \ grams\ \ x\frac{1}{molar\ \ mass(g/mol)} = Amount\ \ of\ \ moles$$

III. Percentage Composition

A. Percentage Composition

1. The percentage by mass of each element in a compound

$$\frac{Mass\ \ of\ \ element\ \ in\ 1\ mol\ \ of\ \ compound}{molar\ \ mass\ \ of\ \ compound}\ x\ 100 = \%\ \ element\ \ in\ \ compound$$

B. Hydrates

1. Crystalline compounds in which water molecules are bound in the crystal structure

Copper (II) sulfate pentahydrate

$CuSO_4 \bullet 5H_20$

a. The raised dot means "Water is loosely attached" It does **NOT** mean multiply when determining formula weight

7-4 Determining Chemical Formulas

Empirical Formula - the symbols for the elements combined in a compound, with subscripts showing the smallest whole-number ratio of the different atoms in the compound

I. Calculation of Empirical Formula
   A. Assume a 100 g sample of the compound
      1. Treat % as grams
   B. Convert grams to moles using molar mass of each element
   C. Place each mole quantity in ratio to the smallest number of moles
      1. Construct element ratios from the nearest resulting whole numbers

II. Calculation of Molecular Formula
   A. Necessary Information
      1. Empirical Formula
      2. Molecular weight
   B. Calculations
      1. $(\text{empirical formula wt.})_x$ = molecular weight
      2. $(\text{empirical formula})_x$ = molecular formula
   C. Example (empirical formula = HO molecular wt. = 34.02)
      1. $(\text{HO weight})_x = 34.02$
         a. HO = 17.01 (1.01 + 16.00)
      2. $(17.01)_x = 34.02$
         a. x = 2
      3. Molecular formula is $(HO)_2$
         a. Molecular formula is H_2O_2