AP Chemistry

Chapter 3 Outline

Stoichiometry = the study of the quantities of substances consumed and produced in chemical reactions

- Compositional stoichiometry
- Reaction stoichiometry

a) Chemical Equations
   i) Reactants → products
   ii) Balanced equations: equal numbers of atoms of each element on each side of the arrow
      (1) Place coefficients in front of formulas to achieve this, by trial and error
   iii) States of matter can be indicated by adding symbols: (g), (l), (s), (aq)
      (1) Reaction conditions can be written over the arrow
         (a) Δ indicates addition of heat

b) Patterns of Chemical Reactivity
   i) Synthesis
      (1) Two or more substances → single product
   ii) Decomposition
      (1) Single reactant → two or more products
   iii) Single Replacement
      (1) element + compound → element + compound
   iv) Burning (Simple Combustion): rapid reactions that produce a flame
      (1) Simple hydrocarbon + oxygen → carbon dioxide and water (for complete combustion)

c) Formula Weights (aka formula mass or gram formula mass)
   i) FW = gfm = molar mass = sum of the atomic weights of each atom in its chemical formula
   ii) Suggestion: keep at least two decimal places; better yet, keep all decimal places from periodic table
   iii) Percent composition from formulas
   iv) $\%\ element = \dfrac{(number\ of\ atoms\ of\ that\ element)(atomic\ weight\ of\ element)}{formula\ weight\ of\ compound} \quad x \quad 100\%$

d) Avogadro's Number and the Mole
   i) Mole = the amount of matter that contains as many objects as the number of atoms in exactly 12 grams of ^{12}C
   ii) Avogadro's number = $6.022x\ 10^{23}$ particles = 1 mole
   iii) Molar Mass = the mass in grams of 1 mole of a substance
   iv) Molar Mass (g/mol) = formula weight of substance (in amu)
   v) We use dimensional analysis to convert from masses to number of particles

(1) “Mole Bridge”

$$\boxed{\text{Grams}} \;\underset{\div\ \text{gfm}}{\overset{\text{X gfm}}{\leftrightarrows}}\; \boxed{\text{Moles}} \;\underset{\div\ 6.022 \times 10^{23}}{\overset{\text{X } 6.022 \times 10^{23}}{\rightleftarrows}}\; \boxed{\text{Formula Units}}$$

3.1 Empirical Formulas from Analyses

i) From % composition data eActivity
   (1) “percent to mass, mass to mole, divide by small, multiply ‘til whole”
ii) Finding the Molecular Formula
   (1) The subscripts in the molecular formula are some whole number multiple of the empirical formula
   (2) To find multiple: $\textit{whole number multiple} = \dfrac{\textit{molecular weight}}{\textit{empirical formula weight}}$
i) Combustion Analysis VERY IMPORTANT!
   (1) Used for hydrocarbons, other organic substances
   (2) Assume all C in original substance is converted to $CO_2 \Rightarrow$ moles or grams of C
   (3) Assume all H in original substance is converted to $H_2O \Rightarrow$ moles or grams of H
   (4) If a third element is present, its mass can be found by difference
   (5) Use data from calculations to find empirical formula as outlined above

e) Quantitative Information from Balanced Equations
   i) Coefficients from balanced equations indicate both relative numbers of molecules (or formula units) and the relative numbers of moles involved in the equation.
   ii) Multiple problem types can be solved: mole-mole, mole-mass, mass-mass, etc.
   iii) Several different problem-solving strategies can be used
      (1) Dimensional analysis
         (a) grams reactant → moles reactant → moles product → grams product

f) Limiting Reactants (or reagents) Animation,
   i) Limiting reactant = the reactant that is completely consumed in a chemical reaction
   ii) Excess reagent = the reactant(s) left over when reaction stops
   iii) Several different problem solving strategies can be used. These problems can be identified
      (1) when the mass of more than one reactant is given in the problem.
      (2) Use one reactant to solve for the amount needed of the 2^{nd} reactant. If you have more of the 2^{nd} reactant than you need, it is in XS; if you don’t have enough of the 2^{nd} reactant, it is limiting
      (3) Solve for the amount of product needed using the mass of each reactant given. The reactant that results in less product is the LR.

iv) Summarizing table:

	A	B	C
Initial quantities			
Change (reaction)			
Final quantities			

v) Theoretical yield = the quantity of product that is calculated to form when all of the limiting reactant reacts

vi) Actual yield = the amount of product actually obtained in a reaction; usually less than the theoretical yield

(1) Side reactions, unreacted reactants, or loss of product can lower the percent yield.

(2) $$\%\, yield = \frac{actual\ yield}{theoretical\ yield}\ x\, 100\%$$