

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

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

b) Patterns of Chemical Reactivity

- Synthesis
 - Two or more substances \rightarrow single product
- Decomposition
 - Single reactant \rightarrow two or more products
- Single Replacement
 - element + compound \rightarrow element + compound
- Burning (Simple Combustion): rapid reactions that produce a flame
 - Simple hydrocarbon + oxygen \rightarrow carbon dioxide and water (for complete combustion)

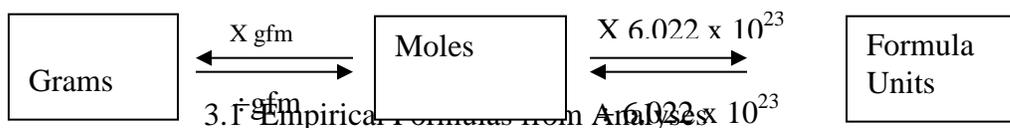
c) Formula Weights (aka formula mass or gram formula mass)

- FW = gfm = molar mass = sum of the atomic weights of each atom in its chemical formula
- Suggestion: keep at least two decimal places; better yet, keep all decimal places from periodic table
- Percent composition from formulas
- $\% \text{ element} = \frac{(\text{number of atoms of that element})(\text{atomic weight of element})}{\text{formula weight of compound}} \times 100\%$

d) Avogadro's Number and the Mole

- Mole = the amount of matter that contains as many objects as the number of atoms in exactly 12 grams of ^{12}C
- Avogadro's number = 6.022×10^{23} particles = 1 mole
- Molar Mass = the mass in grams of 1 mole of a substance
- Molar Mass (g/mol) = formula weight of substance (in amu)
- We use dimensional analysis to convert from masses to number of particles

(1) "Mole Bridge"



- i) From % composition data
 - (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: $whole\ number\ multiple = \frac{molecular\ weight}{empirical\ formula\ weight}$
- 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 \rightarrow moles reactant \rightarrow moles product \rightarrow 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 2nd reactant. If you have more of the 2nd reactant than you need, it is in XS; if you don't have enough of the 2nd 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)
$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$