AP Chemistry
A. Allan

Chapter 3 Notes - Stoichiometry

3.1 Counting by Weighing

- A. Average Mass
  1. When a particle (or object) has a characteristic AVERAGE mass, then counting large numbers can be done by weighing
- B. Assumptions
  1. Large sample size
  2. A representative sample (it represents the assumed average)

3.2 Atomic Masses

- A. C-12, the Relative Standard
  1. C-12 is assigned a mass of exactly 12 atomic mass units (amu)
  2. Masses of all elements are determined in comparison to the carbon - 12 atom (^{12}C) the most common isotope of carbon
  3. Comparisons are made using a mass spectrometer
- B. Atomic Mass (Average atomic mass, atomic weight)
  1. Atomic masses are the average of the naturally occurring isotopes of an element
  2. Atomic mass does not represent the mass of any actual atom
  3. Atomic mass can be used to "weigh out" large numbers of atoms

3.3 The Mole

- A. Avogadro's number
  1. 6.022×10^{23} units = 1 mole
  2. Named in honor of Avogadro (he did NOT discover it)
- B. Measuring moles
  1. An element's atomic mass expressed in grams contains 1 mole of atoms of that element
     - a. 12.01 grams of carbon is 1 mole of carbon
     - b. 12 grams of carbon-12 is 1 mole of carbon-12

3.4 Molar Mass

- A. Molar Mass (Gram molecular weight)
  1. The mass in grams of one mole of a compound
  2. The sum of the masses of the component atoms in a compound
     - a. Molar mass of ethane (C_2H_6):

$$\begin{array}{lr} \text{Mass of 2 moles of C} = & 2(12.01\text{ g}) \\ \text{Mass of 6 moles of H} = & \underline{6(1.008\text{ g})} \\ & 30.07\text{ g} \end{array}$$

3.5 Learning to Solve Problems

No notes – but PLEASE read this section of the chapter thoroughly. A few points:

1. The ability to solve complex problems separates the successful student from the unsuccessful student
2. Complex problems often do not have a single, repeatable algorithm
3. There may be many paths to a correct answer
4. Not all pathways to an answer are equally EFFICIENT

3.6 Percent Composition of Compounds

A. Calculating any percentage
    1. "The part, divided by the whole, multiplied by 100"
B. Percentage Composition
    1. Calculate the percent of each element in the total mass of the compound

$$\frac{\text{(\#atoms of the element)(atomic mass of element)}}{\text{(molar mass of the compound)}} \times 100$$

C. The importance of the concept
    1. Compounds have a characteristic percent by mass
    2. % composition cannot always identify a compound (why not?)

3.7 Determining the Formula of a Compound

A. Determining the empirical formula
    1. Determine the percentage of each element in your compound
    2. Treat % as grams, and convert grams of each element to moles of each element
    3. Find the smallest whole number ratio of atoms
    4. If the ratio is not all whole number, multiply each by an integer so that all elements are in whole number ratio
B. Determining the molecular formula
    1. Find the empirical formula mass
    2. Divide the known molecular mass by the empirical formula mass, deriving a whole number, n
    3. Multiply the empirical formula by n to derive the molecular formula

3.8 Chemical Equations

A. Chemical reactions
    1. Reactants are listed on the left hand side
    2. Products are listed on the right hand side
    3. Atoms are neither created nor destroyed
        a. All atoms present in the reactants must be accounted for among the products, in the same number
        b. No new atoms may appear in the products that were not present in the reactants

B. The Meaning of a Chemical Reaction
    1. Physical States
        a. Solid - (*s*)
        b. Liquid - (*l*)
        c. Gas - (*g*)
        d. Dissolved in water (aqueous solution) - (*aq*)
    2. Relative numbers of reactants and products
        a. Coefficients give atomic/molecular/mole ratios

3.9 Balancing Chemical Equations

A. Determine what reaction is occurring
    1. It is sometimes helpful to write this in word form:

       Hydrogen + oxygen → water

B. Write the unbalanced equation
    1. Focus on writing correctly atomic and compound formulas

$$H_2 + O_2 \rightarrow H_2O$$

C. Balance the equation by inspection
    1. It is often helpful to work systematically from left to right

$$2H_2 + O_2 \rightarrow 2H_2O$$

D. Include phase information

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(l)$$

3.10 Stoichiometric Calculations: Amounts of Reactants and Products

A. Balance the chemical equation
B. Convert grams of reactant or product to moles
C. Compare moles of the known to moles of the desired substance
    1. A ratio derived from the coefficients in the balanced equation
D. Convert from moles back to grams if required

3.11 Calculations Involving a Limiting Reactant

A. Concept of limiting reactant (limiting reagent):

" I want to make chocolate chip cookies. I look around my kitchen (I have a BIG kitchen!) and find 40 lbs. of butter, two lbs. of salt, 1 gallon of vanilla extract, 80 lbs. of chocolate chips, 200 lbs. of flour, 150 lbs. of sugar, 150 lbs. of brown sugar, ten lbs. of baking soda and TWO eggs. It should be clear that it is the number of eggs that will determine the number of cookies that I can make."

    1. The limiting reactant controls the amount of product that can form

B. Solving limiting reactant problems
    1. Convert grams of reactants to moles
    2. Use stoichiometric ratios to determine the limiting reactant
    3. Solve as before, beginning the stoichiometric calculation with the grams of the limiting reactant

C. Calculating Percent Yield

1. Actual yield - what you got by actually performing the reaction
2. Theoretical yield - what stoichiometric calculation says the reaction SHOULD have produced

$$\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% = \text{percent yield}$$