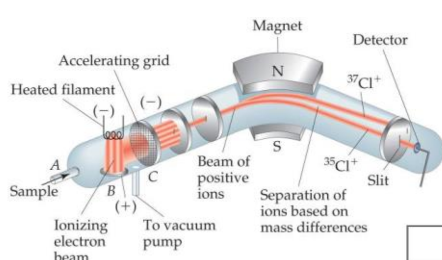
AP CHEM TOPIC #3
STOICHIOMETRY NOTES

Table of Contents

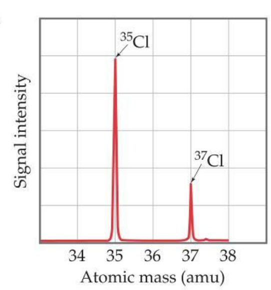
Objectives

- Moles, mass, representative particles (atoms, molecules, formula units), molar mass (MM), and Avogadro's number.
- Molarity; preparation of solutions.
- The percent composition of an element in a compound.
- Balanced chemical equations; for example, for a given mass of reactant, calculate the amount of product produced.
- Limiting reactants: calculate the amount of product formed when given the amounts of all the reactants present.
- Reactions in solution: given the molarity and the volume of the reactants, calculate the amount of product produced or the amount of reactant required to react.
- The percent yield of a reaction.

Atomic Masses Section#1



- Chemical stoichiometry
  - The area of study of the quantities of materials consumed and produced in chemical reactions.
- Atomic Mass (Average atomic mass. atomic weight)
  - Based on the carbon-12 atom
    - Assigned the mass of exactly 12 atomic mass units (amu)
  - Mass spectroscopy
    - Most accurate method for comparing the masses of atoms
  - A weighted average of all the isotopes in a given sample
    - Carbon has a atomic mass of 12.01
      - 98.89%(12 amu) + 1.11%(13.0034amu) = 0.9889(12amu) + 0.111(13.0034amu) = 12.01amu



        - No atom of carbon has an exact mass of 12.01 amu.
    - **<u>Sample Exercise 3.1</u>** – The mass of an Element
      When a sample of natural copper is vaporized and injected into a mass spectrometer, the results shown in Fig. 3.3 are obtained. Use these data to compute the average mass of natural copper. The mass values for ^{63}Cu and ^{65}Cu are 62.93amu and 64.93amu, respectively.) Ans: 63.55amu

The Mole Section#2

- Mole (mol)
    - As the number equal to the number of carbon atoms in exactly 12 grams of pure ^{12}C.
        - Avogadro's number
            - $6.022x10^{23}$ units (parts) of that substance
                - 1 mole of marbles would cover the Earth's surface to a depth of 50 miles.
    - Thus the mole is defined such that a sample of a natural element with a mass equal to the element's atomic mass expressed in grams contains 1 mol of atoms.

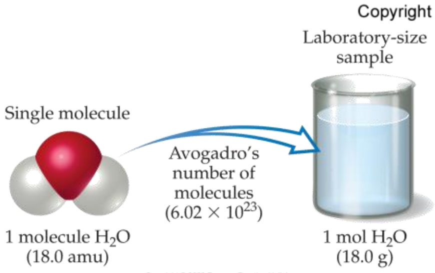
Comparison of 1 Mole Samples of Various Elements

Element	Number of Atoms Present	Mass of Sample(g)
aluminum	$6.022x10^{23}$	26.98
copper	$6.022x10^{23}$	63.55
iron	$6.022x10^{23}$	55.85
sulfur	$6.022x10^{23}$	32.07
iodine	$6.022x10^{23}$	126.9
mercury	$6.022x10^{23}$	200.6

TABLE 3.2 ▪ Mole Relationships

Name of Substance	Formula	Formula Weight (amu)	Molar Mass (g/mol)	Number and Kind of Particles in One Mole
Atomic nitrogen	N	14.0	14.0	6.02×10^{23} N atoms
Molecular nitrogen	N_2	28.0	28.0	6.02×10^{23} N_2 molecules; $2(6.02 \times 10^{23})$ N atoms
Silver	Ag	107.9	107.9	6.02×10^{23} Ag atoms
Silver ions	Ag^+	107.9[a]	107.9	6.02×10^{23} Ag^+ ions
Barium chloride	$BaCl_2$	208.2	208.2	6.02×10^{23} $BaCl_2$ units; 6.02×10^{23} Ba^{2+} ions; $2(6.02 \times 10^{23})$ Cl^- ions

[a] Recall that the electron has negligible mass; thus, ions and atoms have essentially the same mass.



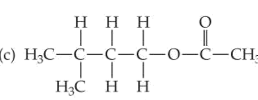
Notice that both masses have the same number but have different units (18.0 amu compared to 18.0 g) representing the huge difference in mass.

One mole of NaCl, the solid, has a mass of 58.45 g. One mole of H_2O, the liquid, has a mass of 18.0 g and occupies a volume of 18.0 mL. One mole of O_2, the gas, has a mass of 32.0 g and occupies a balloon whose diameter is 35 cm.

- **<u>Sample Exercise 3.2</u>** – Determining the Mass of a Sample of Atoms
  Americium is an element that does not occur naturally. It can be made in very small amounts in a devise known as a particle accelerator. Compute the mass in grams of a sample of americium containing six atoms. Ans: 2.42×10^{-21}g
- **<u>Sample Exercise 3.3</u>** – Determining Moles of Atoms
  Aluminum (Al) is a metal with a strength-to-mass ratio and a high resistance to corrosion; thus it is often used for structural purposes. Compute both the number of moles and the number of atoms in a 10.0g sample of aluminum. Ans: 0.371mol/2.23×10^{23}atoms
- **<u>Sample Exercise 3.4</u>** – Calculating Numbers of Atoms
  A silicon chip used in an integrated circuit of a microcomputer has a mass of 5.68mg. How many silicon (Si) atoms are present in the chip? Ans: 1.22×10^{20}atoms
- **<u>Sample Exercise 3.5</u>** – Calculating the Number of Moles and Mass
  Cobalt (Co) is a metal that is added to steel to improve its resistance to corrosion. Calculate both the number of moles in a sample of cobalt containing 5.00×10^{20}atoms and the mass of the sample.
  Ans: 8.30×10^{-4}mole Co/4.89×10^{-2}g Co

Molar Mass Section#3

- Molar mass (*MM*)
  - The mass in grams of one mole of the compound
    - Also called molecular weight (*MW*)
  - The sum of all the masses of the component atoms
  - **<u>Sample Exercise 3.6</u>** – Calculating Molar Mass
    Juglone, a dye known for centuries, is produced from the husks of black walnuts. It is also a natural herbicide (weed killer) that kills off competitive plants around the black walnut tree but does not affect grass and other noncompetitive plants. The formula for juglone is $C_{10}H_6O_3$.
    (a) Calculate the MM of juglone.
    (b) A sample of 1.56×10^{-2}g of pure juglone was extracted from black walnut husks. How many moles of juglone does this sample represent? Ans: 174.1g/8.96×10^{-5}mol
  - **<u>Sample Exercise 3.7</u>** – Calculating Molar Mass II
    Calcium carbonate ($CaCO_3$), also called calcite, is the principal mineral found in limestone, marble, pearls, and the shells of marine animals such as clams.
    (a) Calculate the *MM* of calcium carbonate.
    (b) A certain sample of calcium carbonate contains 4.86 moles. What is the mass in grams of this sample? What is the mass of the CO_3^{2-} ions present? Ans: 100.09g/mol, 292g CO_3^{2-}
  - **<u>Sample Exercise 3.8</u>** – Molar Mass and Number of Molecules.



(c) Isopentyl acetate (banana flavor)

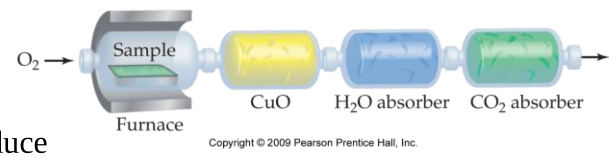
Isopentyl acetate is the compound responsible for the scent of bananas. A molecular model of isopentyl acetate is shown below. Interestingly, bees release about µg (1×10^{-6}g) of this compound when they sting. The resulting scent attracts other bees to join the attack. How many molecules of isopentyl acetate are released in a typical bee sting? How many atoms of carbon are present? Ans: 5×10^{15}molecules/4×10^{16} C atoms

Percentage Composition Section#4

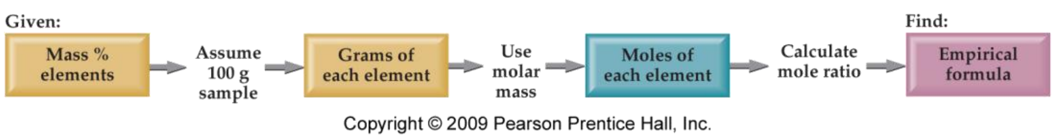
- Mass percent (weight percent)
  - The percentage by mass of the constituent atoms of a molecule or formula unit (ionic).
  - **<u>Sample Exercise 3.9</u>** – Calculating Mass Percent I
    Carvone is a substance that occurs in two forms having different arrangements of the atoms but the same molecular formula ($C_{10}H_{14}O$) and mass. One type of carvone gives caraway seeds their characteristic smell, and the other type is responsible for the smell of spearmint oil. Compute the mass percent of each element in carvone. Ans: C-79.96%, H – 9.394%, O – 10.65%
  - **<u>Sample Exercise 3.10</u>** – Calculating Mass Percent II
    Penicillin, the first of a now large number of antibiotics (antibacterial agents), was discovered accidentally by Scottish Alexander Fleming in 1928, but he was never able to isolate it as a pure compound. This and similar antibiotics have saved millions of lives that might have been lost to infections. Penicillin F has the formula $C_{14}H_{20}N_2SO_4$. Compute the mass percent of each element.
    Ans: C – 53.81%, H – 6.453%, N – 8.969%, S – 10.27%, O – 20.49%

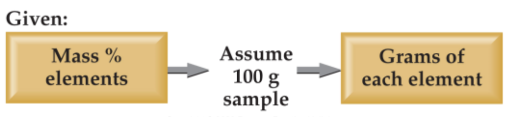
Determining the Formula of a Compound Section#5

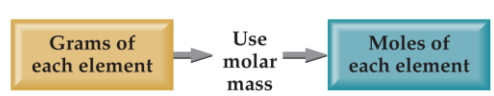
- New compound
  - Decompose it into its constituent elements
    -or-
  - Combustion analysis
    - React organic compound with oxygen to produce CO_2, H_2O, N_2, etc.
      - The products are then collected and massed.



  - Empirical formula (EF)
    - Simplest whole number ratio of elements in a compound
      - Use for all ionic compounds
    - Empirical mass (EM) the mass of the simplest ratio in a formula
      - Mass of empirical formula
  - Molecular formula (MF)
    - True ratio of elements in a compound
      - Use for all molecular (covalent compounds)
    - MF = $(EF)_n$
      - Where n is an integer
      - Found by dividing MM by EM
        - n = MM/EM







Home

- **Sample Exercise 3.11** – Determining Empirical and Molecular Formulas I
  Determine the empirical and molecular formulas for a compound that gives the following percentages upon analysis (in mass percents): 71.65% Cl 24.27%C 4.07% H
  The molar mass (MM) is known to be 98.96g/mol Ans: EF = $ClCH_2$, MF = $Cl_2C_2H_4$
- **Sample Exercise 3.12** – Determining Empirical and Molecular Formulas II
  A white powder is analyzed and found to contain 43.64% phosphorus and 56.36% oxygen by mass. The compound has a MM of 283.88g/mol. What are the compound's empirical and molecular formula? Ans: P_2O_5 and P_4O_{10}
- **Sample Exercise 3.13** – Determining Molecular Formula
  Caffeine, a stimulant found in coffee, tea, and chocolate, contains 49.48% carbon, 5.15% hydrogen, 28.87% nitrogen, and 16.49% oxygen by mass and has a MM of 194.2g/mol. Determine the molecular formula of caffeine. Ans: $C_8H_{10}N_4O_2$

Empirical Formula (*EF*) Determination

- Since mass percentage gives the number of grams of a particular element per 100 grams of compound, base calculation on 100 grams of compound. Each percent will then represent the mass in grams of that element.
- Determine the number of moles of each element present in 100 grams of compound using the atomic masses of the elements present.
- Divide each value of the number of moles by the smallest of the values. If each resulting number is a whole number (after appropriate rounding), these numbers represent the subscripts of the elements in the empirical formula.
- If the numbers obtained in the previous step are not whole numbers, multiply each number by an integer so that the results are all whole numbers.

Molecular Formula (*MF*) Determination

Method One

- Obtain the empirical formula.
- Compute the mass corresponding to the empirical formula (*EF*).
- Calculate the ratio $n = MM/EM$ (*MM* – molar mass and *EM* – empirical mass)
- The integer from the previous step represents the number of empirical formula units in one molecule. When the empirical formula subscripts are multiplied by this integer, the molecular formula results. This procedure is summarized by the equation:

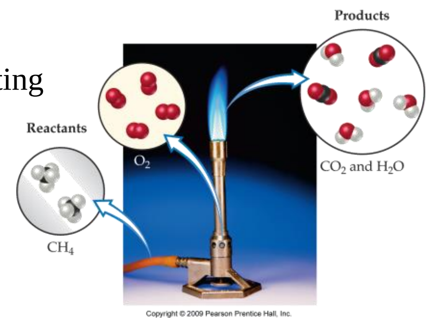
$$MF = (EF) \times \frac{MM}{EM}$$

Method Two

- Using the mass percentages and the MM, determine the mass of each element present in 1 mole of compound.
- Determine the number of moles of each element present in 1 mole of compound.
- The integers from the previous step represent the subscripts in the molecular formula.

Chemical Equations Section#6

- Chemical Reactions
  - Chemical equation
    - Reactants on left, products on right and an arrow (→) separating them



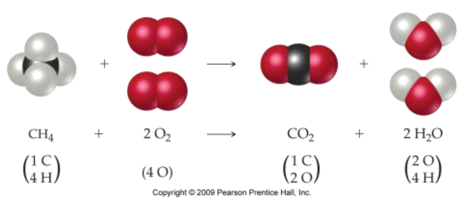
$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

methane and oxygen yields carbon dioxide and water

  - Balancing Chemical equations
    - Obeys the Law of Conservation of Mass (Dalton)
    - Process where the atoms of one element are made equal on the reactant and product side of a chemical equation using whole number coefficients.

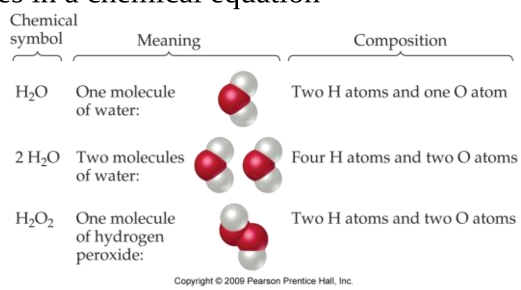
$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

1C 4H 4O 1C 2O 4H 2O



- The meaning of a Chemical Equation
  - Physical states
    - Represent as (*s*) – solid, (*l*) – liquid, (*g*) – gas, and (*aq*) – aqueous solution
  - Coefficients/Subscripts
    - Coefficients are used to balance atoms/molecules in a chemical equation
    - Subscripts are within a molecule/formula unit

Notice how adding the coefficient 2 in front of the formula (line 2) has a different effect on the implied composition than adding the subscript 2 to the formula (in line 3). The number of atoms of each type (listed under composition) is obtained by multiplying the coefficient and the subscript associated with each element in the formula.



Information Conveyed by the Balanced Equation for the Combustion of Methane

Reactants		Products
$CH_4(g) + 2O_2(g)$	→	$CO_2(g) + 2H_2O(g)$
1 molecule + 2 molecule	→	1 molecule + 2 molecule
1 mole + 2 moles	→	1 mole + 2 moles
$6.022\text{x}10^{23}$molecules + 2($6.022\text{x}10^{23}$molecules)	→	$6.022\text{x}10^{23}$molecules + 2($6.022\text{x}10^{23}$molecules)
16g + 2(32g)	→	44g + 2(18g)
80g reactants	→	80g products

TABLE 3.3 Information from a Balanced Equation

Equation:	$2\ H_2(g)$	+	$O_2(g)$	$\longrightarrow$	$2\ H_2O(l)$
Molecules:	2 molecules H_2	+	1 molecule O_2	$\longrightarrow$	2 molecules H_2O
					
Mass (amu):	4.0 amu H_2	+	32.0 amu O_2	$\longrightarrow$	36.0 amu H_2O
Amount (mol):	2 mol H_2	+	1 mol O_2	$\longrightarrow$	2 mol H_2O
Mass (g):	4.0 g H_2	+	32.0 g O_2	$\longrightarrow$	36.0 g H_2O

Balancing Chemical Equations Section#7

- The formulas of the compounds (once formulated) must NEVER be changed in balancing a chemical equation.
  - Use only coefficients (in front of formula) to balance the equation.

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

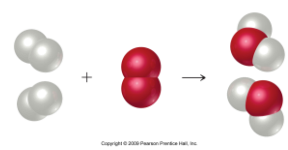
2C   2C

(5+1)H   (3x2)H

1O + (3x2)O   (2x2)O + 3O

2C, 6H, and 7O   2C, 6H, and 7O

> Because there are four H atoms and two O atoms on each side of the equation, the equation is balanced. We can represent the balanced equation by these molecular models, which illustrate that the number of atoms of each kind is the same on both sides of the arrow.



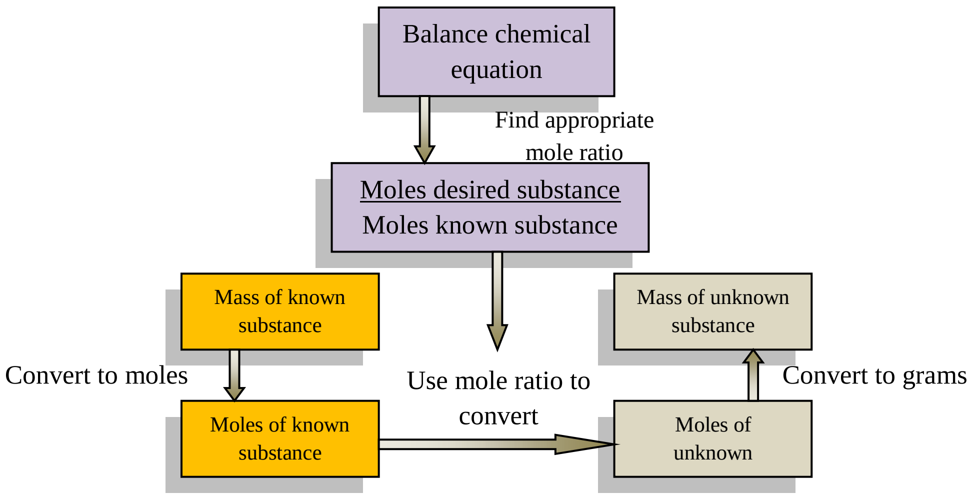
> Writing and Balancing the Equation for a Chemical Reaction
> 1. Determine what reaction is occurring. What are the reactants, the products, and the physical states involved?
> 2. Write an unbalanced equation that summarizes the reaction described in step 1.
> 3. Balance the equation by inspection, starting with the most complicated molecule(s). Determine what coefficients are necessary so that the same number of each type of atom appears on both reactant and product sides. Do not change the identities (formulas) of any of the reactants or products.

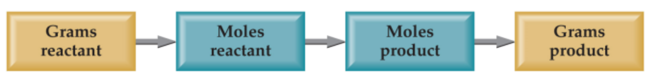
- **Sample Exercise 3.14** – Balancing a Chemical Equation I
  Chromium compounds exhibit a variety of bright colors. When solid ammonium dichromate, $(NH_4)_2Cr_2O_7$, a vivid orange compound, is ignited, a spectacular reaction occurs. Although the reaction is actually somewhat more complex, let's assume here that the products are solid chromium (III) oxide, nitrogen gas, and water vapor. Balance the equation for this reaction.
- **Sample Exercise 3.15** – Balancing a Chemical Equation II
  At 1000°C, ammonia gas, $NH_3(g)$, reacts with oxygen gas to form gaseous nitric oxide, NO(g), and water vapor. This reaction is the first step in the commercial production of nitric acid by the Ostwald process. Balance the equation for this reaction.

Stoichiometric Calculations: Amounts of Reactants and Products Section#8

Calculating Masses of Reactants/Products in Chemical Reactions

1) Balance the equation for the reaction.
2) Convert the known mass of the reactant/product into moles of that substance.
3) Use the balanced equation to set up the appropriate mole ratio.
4) Use the appropriate mole ratio to calculate the number of moles of the desired reactant/product.
5) Convert from moles back to grams if required by the problem.





- **Sample Exercise 3.16** – Chemical Stoichiometry I
  Solid lithium hydroxide is used in space vehicles to remove exhaled carbon dioxide from the living environment by forming solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide can be absorbed by 1.00kg of lithium hydroxide? Ans: 920.g of $CO_2(g)$
- **Sample Exercise 3.17** – Chemical Stoichiometry II
  Baking soda ($NaHCO_3$) is often used as an antacid. It neutralizes excess hydrochloric acid secreted by the stomach: $NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(aq)$. Milk of magnesia, which is an aqueous suspension of magnesium hydroxide, is also used as an antacid: $Mg(OH)_2(s) + 2HCl(aq) \rightarrow 2H_2O(l) + MgCl_2(aq)$. Which is the more effective antacid per gram, $NaHCO_3$ or $Mg(OH)_2$?
  Ans: 1.00g $Mg(OH)_2$ will neutralize 3.42×10^{-2}mol HCl

Calculations Involving a Limiting Reactant Section#9

- Often chemicals are mixed in exact quantities, so all of the reactants "run out" at the same time
  - Stoichiometric quantities
  - Haber Process
    - Ammonia a very important starting material for the production of fertilizer.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

- The hydrogen is made from the reaction of methane (CH_4) and water (H_2O):

$$CH_4(g) + H_2O(g) \rightarrow 3H_2(g) + CO(g)$$

Home

- Based on the above equation, 2.50×10^3kg of methane reacts with exactly 2.81×10^3kg H_2O.

- When the given quantities are inexact, then one of the reactants is a limiting reactant (will be used up in the reaction) and excess reactant (will NOT be used up in reaction)
  - LR → limiting reactant
  - ER → excess reactant
    - Sometimes the question will state one of the reactants as the ER by stating "given quantity is in excess."
    - The other reactant is then the LR (no need to calculate the LR).

	$2 H_2(g)$	+	$O_2(g)$	⟶	$2 H_2O(g)$
Initial quantities:	10 mol		7 mol		0 mol
Change (reaction):	−10 mol		−5 mol		+10 mol
Final quantities:	0 mol		2 mol		10 mol

	$N_2(g)$	+	$3 H_2(g)$	⟶	$2 NH_3(g)$
Initial quantities:	3.0 mol		6.0 mol		0 mol
Change (reaction):	−2.0 mol		−6.0 mol		+4.0 mol
Final quantities:	1.0 mol		0 mol		4.0 mol

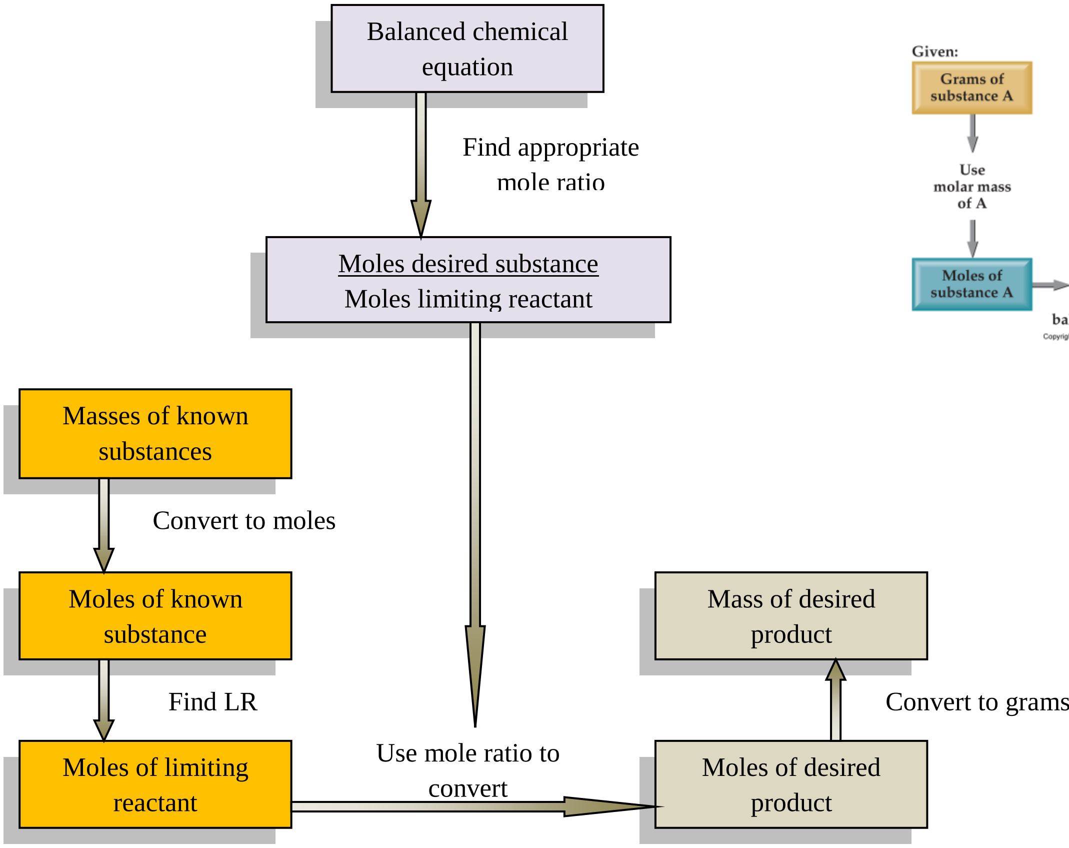
  - **Sample Exercise 3.18** – Stoichiometry: Limiting Reactant
    Nitrogen gas can be prepared by passing gaseous ammonia over solid copper (II) oxide at high temperatures. The other products of the reaction are solid copper and water vapor. If a sample containing 18.1g of NH_3 is reacted with 90.4g of CuO, which is the LR? How many grams of N_2 would be formed? Ans: LR is CuO; 10.6g N_2

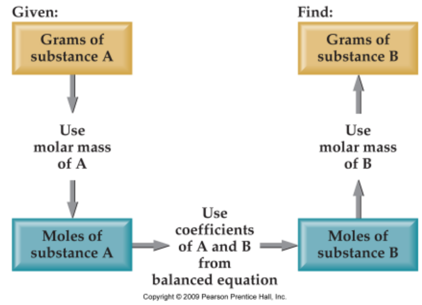
- Percent Yield
  - Theoretical yield
    - The amount of product formed when all of the LR is completely consumed.
      - The maximum amount of product formed (mathematically derived).
  - Actual yield
    - The amount of product formed from experimentation.
  - Percent yield (%yield)
    - The comparison between the experimental amount of product and the theoretical amount of product.
    - Written as a percent. $\dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \text{percent yield}$
  - **Sample Exercise 3.19** – Calculating Percent Yield
    Methanol (CH_3OH), also called methyl alcohol, is the simplest alcohol. It is used as a fuel in race cars and is a potential replacement for gasoline. Methanol can be manufactured by combination of gaseous carbon monoxide and hydrogen. Suppose 68.5kg CO(g) is reacted with 8.60kg $H_2(g)$. Calculate the theoretical yield of methanol. If 3.57×10^4g CH_3OH is actually produced, what is the percent yield of methanol? Ans: 6.86×10^4g CH_3OH; 52.0%

Solving a Stoichiometry Problem Involving Masses of Reactants and Products

1) Write/balance the equation for the reaction.
2) Convert the known masses of substances to moles.
3) Determine which reactant is limiting.
4) Using the amount of the limiting reactant and the appropriate mole ratios, compute the number of moles of the desired product.
5) Convert from moles to grams, using the molar mass.

This process is summarized in the following diagram:





[Home]