Chemistry 3A

Introductory General Chemistry

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- Stoichiometry: mass, solution, gas, acid-base titration
- Limiting Reagent/Reactant
- Theoretical and Percent Yield
- Changes in Enthalpy

General Stoichiometry

- Stoichiometry: study of quantities in chemical reactions

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

- "two moles hydrogen react with one mole oxygen to produce two moles of water"
- Ratios we can deduce

$$\frac{2\ \text{mol}\ H_2}{1\ \text{mol}\ O_2}, \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2}, \frac{2\ \text{mol}\ H_2O}{1\ \text{mol}\ O_2}, \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O}, \frac{2\ \text{mol}\ H_2}{2\ \text{mol}\ H_2O}, \frac{2\ \text{mol}\ H_2O}{2\ \text{mol}\ H_2}$$

- These will be part of calculations that have been already done in this course and to estimate yields

Moles Reactant / Moles Product

- How many moles of oxygen (O_2) will react with hydrogen (H_2) to produce 27.6 mol H_2O?
- To solve: moles O_2
- Determine the BALANCED reaction:
  - Skeletal reaction: $H_2 + O_2 \rightarrow H_2O$
  - Balanced reaction: $2\ H_2 + O_2 \rightarrow 2\ H_2O$
- Identify needed quantities: 27.6 mol H_2O product
- Conversion factor: 1 mol O_2 / 2 mol H_2O
- Substitute values & solve:

$$27.6 \text{ mol } H_2O \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} = 13.8 \text{ mol } O_2$$

Stoichiometry Map

- To reinforce understanding of how certain conversions are done in quantitative chemistry, this visual "map" is found useful to some

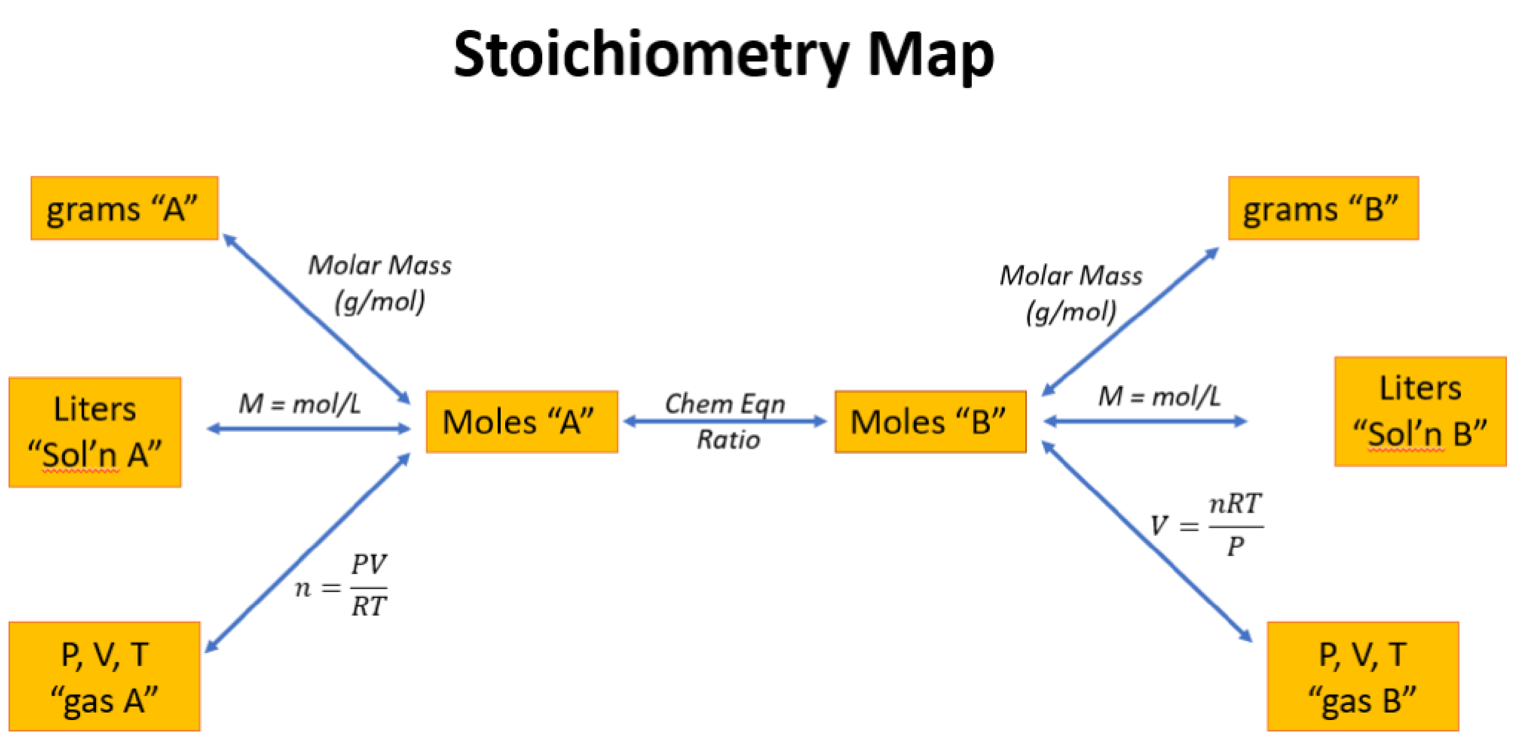
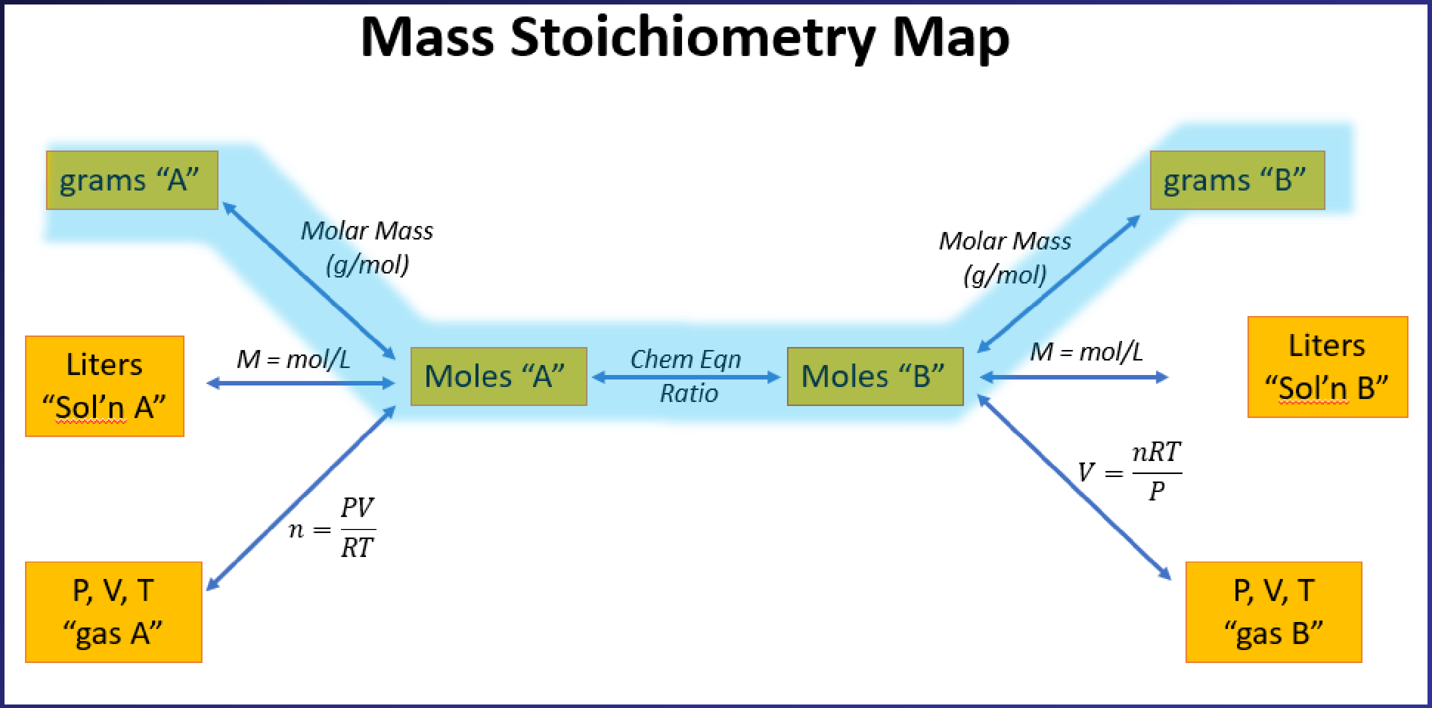


Figure 11.1.1: Stoichiometry Map

The Mass-to-Mass Pathway

- The mass (grams) of compound A to mass (grams) of compound B shows that it is first necessary to convert to moles in order to relate atom-for-atom reaction in a chemical reaction equation



Example: Mass-to-Mass

- Ammonium nitrate decomposes to dinitrogen monoxide and water. **45.7 g** NH_4NO_3 is decomposed: find mass of each product
- To solve: mass, i.e. grams, of products N_2O and H_2O from given data
- Determine the BALANCED reaction:
  - Skeletal reaction: $NH_4NO_3 \rightarrow N_2O + H_2O$
  - Balanced reaction: $NH_4NO_3\ (s) \rightarrow N_2O\ (g) + 2\ H_2O\ (l)$
- Identify given data and needed quantities:
  - provided = 45.7 g NH_4NO_3, molar mass NH_4NO_3 = 80.06 g/mol, molar mass N_2O = 44.02 g/mol, molar mass H_2O = 18.02 g/mol
- Conversion factor: 2 mol H_2O / 1 mol NH_4NO_3, 1 mol N_2O / 1 mol NH_4NO_3
- Substitute values & solve:

$$45.7\ \mathrm{g\ NH_4NO_3} \times \frac{1\ \mathrm{mol\ NH_4NO_3}}{80.06\ \mathrm{g\ NH_4NO_3}} \times \frac{1\ \mathrm{mol\ N_2O}}{1\ \mathrm{mol\ NH_4NO_3}} \times \frac{44.02\ \mathrm{g\ N_2O}}{1\ \mathrm{mol\ N_2O}} = 25.1\ \mathrm{g\ N_2O}$$

$$45.7\ \mathrm{g\ NH_4NO_3} \times \frac{1\ \mathrm{mol\ NH_4NO_3}}{80.06\ \mathrm{g\ NH_4NO_3}} \times \frac{2\ \mathrm{mol\ H_2O}}{1\ \mathrm{mol\ NH_4NO_3}} \times \frac{18.02\ \mathrm{g\ H_2O}}{1\ \mathrm{mol\ H_2O}} = 20.6\ \mathrm{g\ H_2O}$$

mass conservation note: 25.1 g + 20.6 g = 45.7 g

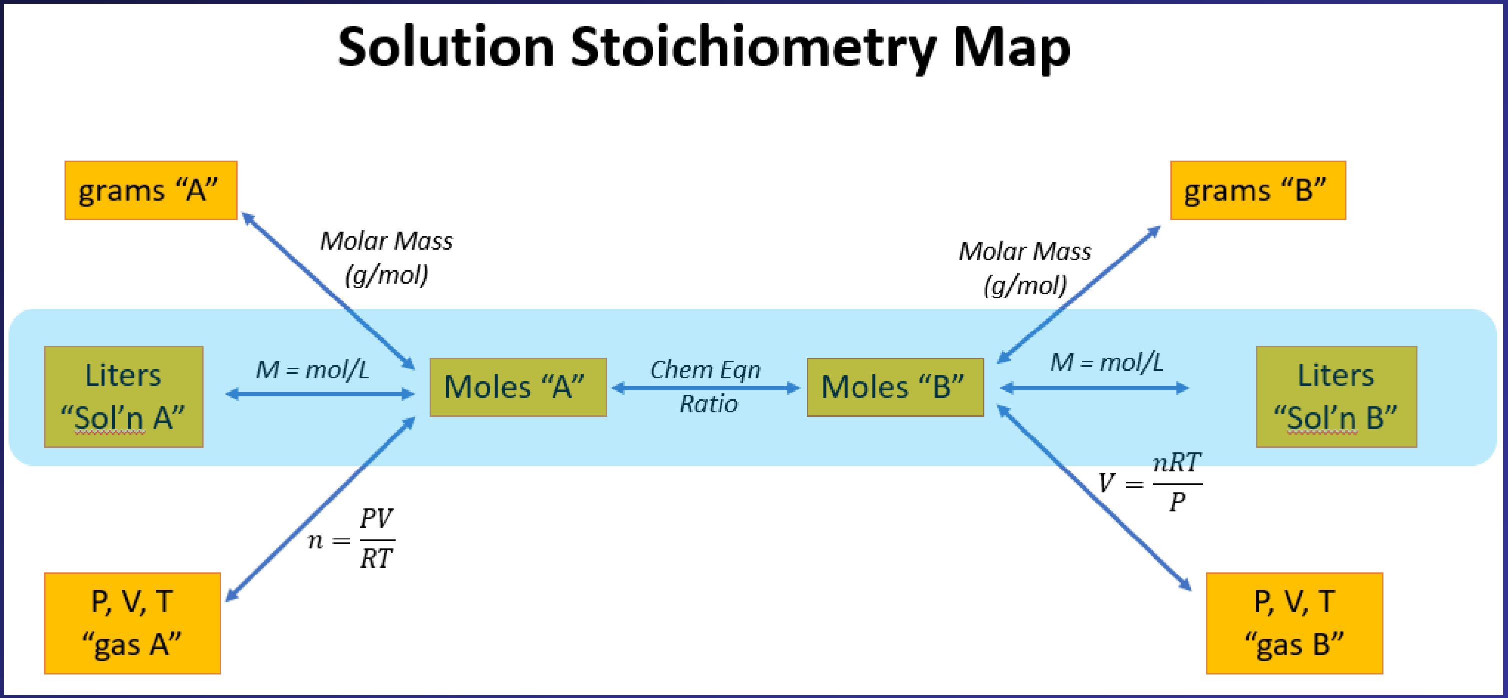
Practice: Mass-to-Mass

- Methane reacts with elemental chlorine to make carbon tetrachloride. How many grams HCl are produced with **100.0 g** methane?
- To solve: mass, i.e. grams, of product HCl
- Determine the BALANCED reaction:
  - Skeletal reaction: $CH_4 + Cl_2 \rightarrow CCl_4 + HCl$
  - Balanced reaction: CH_4 (*g*) + 4 Cl_2 (*g*) → CCl_4 (*l*) + 4 HCl (*g*)
- Identify given data and needed quantities:
  - provided=100.0 g CH_4, molar mass CH_4=80.06 g/mol, molar mass HCl = 44.02 g/mol
- Conversion factor: 4 mol HCl / 1 mol CH_4
- Substitute values & solve:

$$100.0\ \mathrm{g\ CH_4} \times \frac{1\ \mathrm{mol\ CH_4}}{16.04\ \mathrm{g\ CH_4}} \times \frac{4\ \mathrm{mol\ HCl}}{1\ \mathrm{mol\ CH_4}} \times \frac{36.46\ \mathrm{g\ HCl}}{1\ \mathrm{mol\ HCl}} = 909.2\ \mathrm{g\ HCl}$$

Solution Stoichiometry

- This conversion pathway uses a volume to get the amount in moles of a substance. It requires knowing a concentration—molarity—for the solution where the problem is to be solved



Example: Volume Calculations

- How many liters of 0.500 M sodium sulfate in necessary to precipitate all the barium in a 275 mL Ammonium nitrate decomposes to dinitrogen monoxide and water. **45.7 g** NH_4NO_3 is decomposed: find mass of each product
- To solve: mass, i.e. grams, of products N_2O and H_2O from given data
- Determine the BALANCED reaction:
  - Skeletal reaction: $NH_4NO_3 \rightarrow N_2O + H_2O$
  - Balanced reaction: $NH_4NO_3\ (s) \rightarrow N_2O\ (g) + 2\ H_2O\ (l)$
- Identify given data and needed quantities:
  - provided = 45.7 g NH_4NO_3, molar mass NH_4NO_3 = 80.06 g/mol, molar mass N_2O = 44.02 g/mol, molar mass H_2O = 18.02 g/mol
- Conversion factor: 2 mol H_2O / 1 mol NH_4NO_3, 1 mol N_2O / 1 mol NH_4NO_3
- Substitute values & solve:

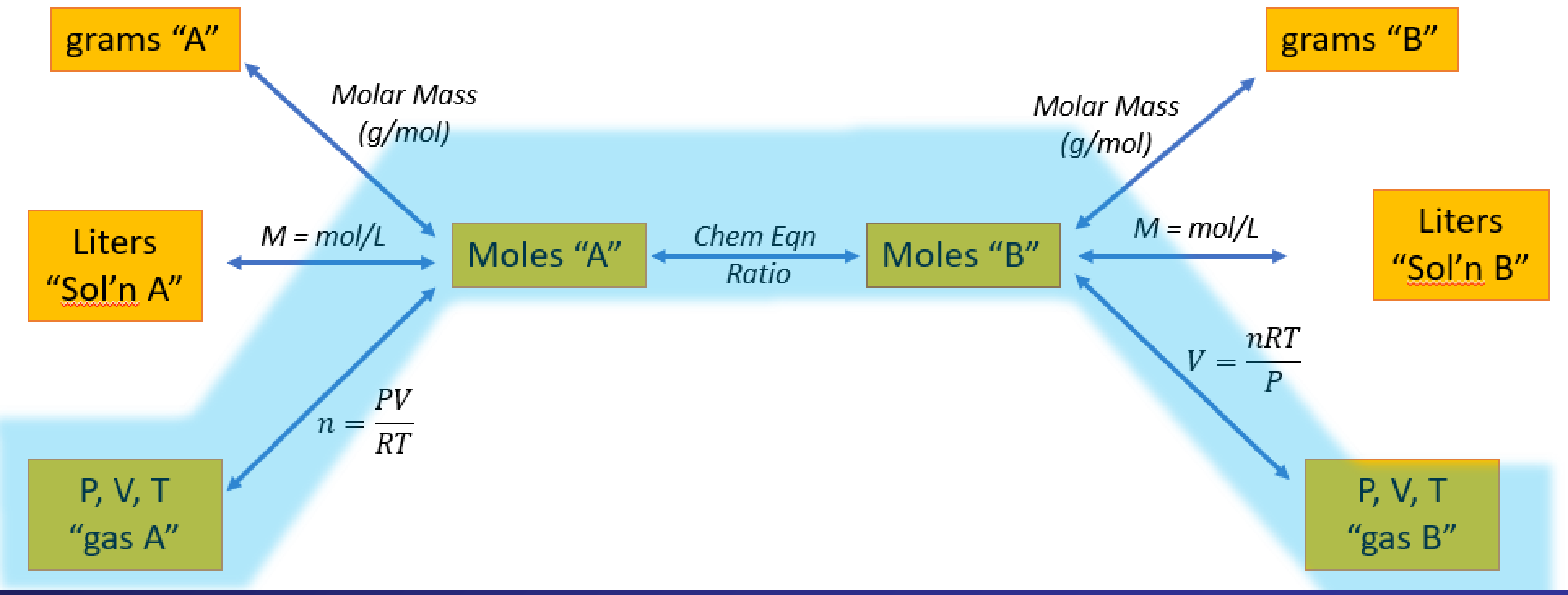
$$45.7\ \mathrm{g\ NH_4NO_3} \times \frac{1\ \mathrm{mol\ NH_4NO_3}}{80.06\ \mathrm{g\ NH_4NO_3}} \times \frac{1\ \mathrm{mol\ N_2O}}{1\ \mathrm{mol\ NH_4NO_3}} \times \frac{44.02\ \mathrm{g\ N_2O}}{1\ \mathrm{mol\ N_2O}} = 25.1\ \mathrm{g\ N_2O}$$

$$45.7\ \mathrm{g\ NH_4NO_3} \times \frac{1\ \mathrm{mol\ NH_4NO_3}}{80.06\ \mathrm{g\ NH_4NO_3}} \times \frac{2\ \mathrm{mol\ H_2O}}{1\ \mathrm{mol\ NH_4NO_3}} \times \frac{18.02\ \mathrm{g\ H_2O}}{1\ \mathrm{mol\ H_2O}} = 20.6\ \mathrm{g\ H_2O}$$

mass conservation note: 25.1 g + 20.6 g = 45.7 g

Moles Reactant / Moles Product

Gas Stoichiometry Map



Example: Gas Calculations

- The production of sulfuric acid (H_2SO_4) is achieved in two step by combustion of sulfur solid with oxygen (O_2) to sulfur dioxide (SO_2), further oxidized to sulfur trioxide (SO_3):

$$2\ S\ (s) + 3\ O_2\ (g) + 2\ H_2O\ (l) \rightarrow 2\ H_2SO_4\ (aq)$$

- To make **1.00 ton** (**907.18 kg**) H_2SO_4, what volume in **L** of O_2 at **22°C** with **745 mmHg** pressure is required?
- To solve: volume of reactant O_2 from moles O_2 from moles H_2SO_4 product from mass H_2SO_4
- From the reaction: 3 moles O_2 → 2 moles H_2SO_4: $\frac{2\ \text{mol}\ H_2SO_4}{3\ \text{mol}\ O_2}$
- Identify given data and needed quantities:
  - provided: 907.18 kg H_2SO_4, T = 22°C, P = 745 mmHg, molar mass H_2SO_4 = 98.03 g/mol, molar mass O_2 = 32.00 g/mol; needed: R = 62.36 L mmHg/mol K
- Substitute values & solve:

$$907.18\ \text{kg}\ H_2SO_4 \times \frac{10^3\ \text{mol}\ H_2SO_4}{98.03\ \text{kg}\ H_2SO_4} \times \frac{3\ \text{mol}\ O_2}{2\ \text{mol}\ H_2SO_4} = 13881.16\ \text{mol}\ O_2$$

$$V = \frac{nRT}{P}: \frac{13881.16\ \text{mol}}{1} \times \frac{62.36\ \text{L mmHg}}{\text{mol K}} \times \frac{(22+273)\text{K}}{1} \times \frac{1}{745\ \text{mmHg}} = 342765.9\ \text{L}\ O_2$$

$$= 342.8\ \text{kL}\ O_2$$

Example: Gas Calculations

- To produce H_2 gas for a balloon in 1783 with volume 31,150 L, metallic iron was placed in hydrochloric acid

$$Fe\ (s) + 2\ HCl\ (aq) \rightarrow H_2\ (g) + FeCl_2\ (aq)$$

- How many kilograms of iron metal to make that volume at **30°C** with **745 mmHg** pressure is needed?
- To solve: kg Fe reactant from mol Fe from moles H_2 from L H_2
- From the reaction: 1 mole H_2 → 1 moles Fe: $\frac{1\ \text{mol H}_2}{1\ \text{mol Fe}}$
- Identify given data and needed quantities:
  - provided: V = 31150 L, T = 30°C, P = 745 mmHg, molar mass H_2 = 2.016 g/mol, molar mass Fe = 55.845 g/mol; needed: R = 62.36 L mmHg/mol K
- Substitute values & solve:

$$n = \frac{PV}{RT}: \frac{745\ \text{mmHg}}{1} \times \frac{\text{mol K}}{62.36\ \text{L mmHg}} \times \frac{1}{(30 + 273)\text{K}} \times \frac{31150\ \text{L}}{1} = 1228.19\ \text{mol H}_2$$

$$1228.19\ \text{mol H}_2 \times \frac{1\ \text{mol Fe}}{1\ \text{mol H}_2} \times \frac{55.845\ \text{kg Fe}}{10^3\ \text{mol Fe}} = 68.588\ \text{kg Fe}\ (\sim 150\ \text{lb})$$

Gas Stoichiometry at STP

- Avogadro's Law is about the amount, ***n***, (in moles) of gas related to its volume ***V*** when pressure ***P*** and temperature ***T*** are held constant
- But the ***V=kn*** relationship can also be affected by what is set as ***P*** and ***T***. It is for this reason scientists decided to study Avogadro's Law from a reference ***T*** and ***P***, which is ***STP***
- After setting STP as a reference, it was found that when ***n*** = **1 mol** gas, ***V*** = **22.4 L**
- When "at STP" is used, it is a given that "1 mol gas = 22.4 L"

Titration

- Titration is a technique to quantitatively determine an amount of a substance in a chemical reaction using a substance of known amount/concentration
- In an acid-base titration as example, a base solution of known concentration is added (from a volumetric instrument like a buret) to an acidic solution of unknown concentration to determine the concentration and amount of acid
- Presence of a pH-sensitive color indicator like phenolphthalein can help determine the end point or equivalence point for when acid and base are equal (neutralizing)

Titration Example

25.66 mL of **0.1078 M** HCl was used to titrate solution of NaOH of unknown amount and mass

- To solve: amount (moles) and mass (grams) of NaOH in solution
- What principle of chemistry: Concentration times volume = amount (*CV=m*), Acid-base neutralization: 1 mol HCl neutralizes 1 mol NaOH

  $$HCl + NaOH \rightarrow H_2O + NaCl$$

- Identify given data and needed quantities:
  - provided: *C* = 0.1078 *M* = 0.1078 *mol/L*, *V* = 25.66 mL = 0.02566 L, molar mass NaOH = 40.00 g/mol
- Substitute values & solve:

$$\frac{0.02566\text{ L solution}}{1} \times \frac{0.1078\text{ mol HCl}}{1\text{ L solution}} \times \frac{1\text{ mol NaOH}}{1\text{ mol HCl}} = 0.002766\text{ mol NaOH}$$

$$0.002766\text{ mol NaOH} \times \frac{40.00\text{ g NaOH}}{1\text{ mol NaOH}} = 0.1106\text{ g NaOH}$$

Titration Indicators

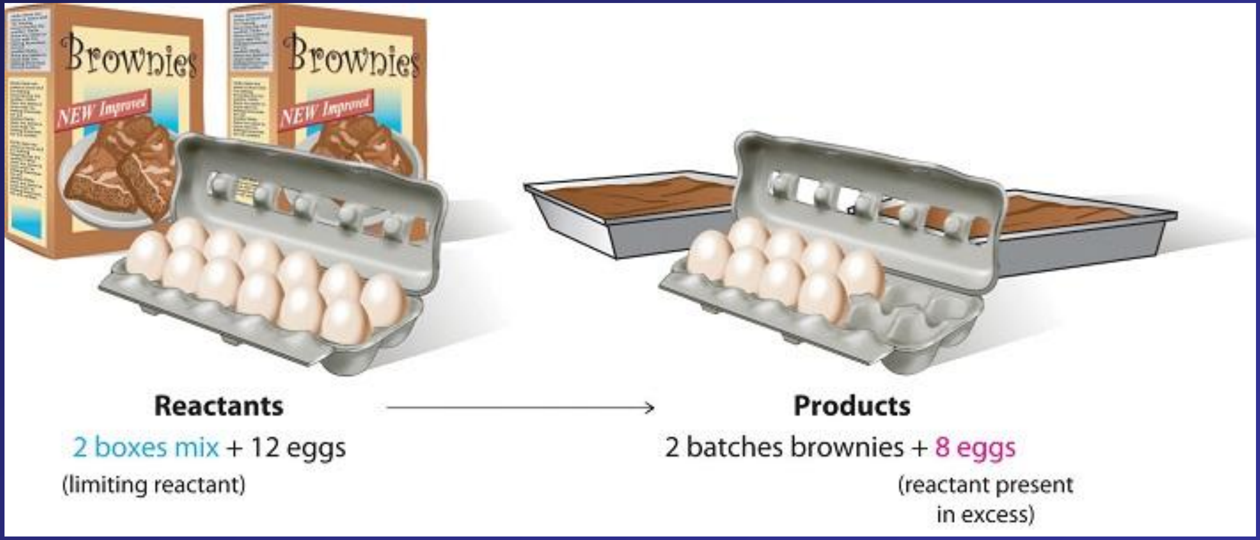
Titrations would usually be done using strong acids or bases in a buret to quantify levels of a weak acid or weak base with a pH indicator to indicate a transition from one delineating pH to another (such from acidic to neutral to basic or the other way around

Indicator will used in the Determination of Citric Acid Levels in a Soft Drink laboratory experiment

Titration between . . .	Indicator	Explanation
strong acid and strong base	any	
strong acid and weak base	methyl orange	changes color in the acidic range (3.2 - 4.4)
weak acid and strong base	phenolphthalein	changes color in the basic range (8.2 - 10.6)

Limiting Reactant (Reagent)

- Chemical reactions done in a laboratory do not have perfect amounts of reactants in a ratio for the chemical reaction equation
- Usually one reactant may be present in fewer/smaller mole amounts compared to another
- The reactant that is consumed first compared to other reactants is called the **limiting reactant** or limiting reagent. All other reactants are **excess reactants**



Identifying the Limiting Reactant

1. Reactant Mole Ratio Method

a. Balance the chemical reaction equation

b. Make sure everything is in moles: if masses (grams) given, use molar mass to convert

c. Use info to calculate mole ratio: compare calculated ratio to actual ratio

d. Now compute/estimate product amount from limiting reactant

e. Determine as well amounts not reactant (left over) in the excess reactants/reagents

Identifying the Limiting Reactant

2. Product Method
   a. Balance the chemical reaction equation
   b. If necessary, convert all to moles
   c. Use stoichiometry for each reactant to find mass of product that will be produced
   d. The reactant producing lesser amount of product is the limiting reactant
   e. Reactant(s) producing larger amounts of product are the excess reactants
   f. If needed (perhaps to guard against waste), subtract excess reactant consumed from total excess reactant to be utilized

Moles Reactant / Moles Product

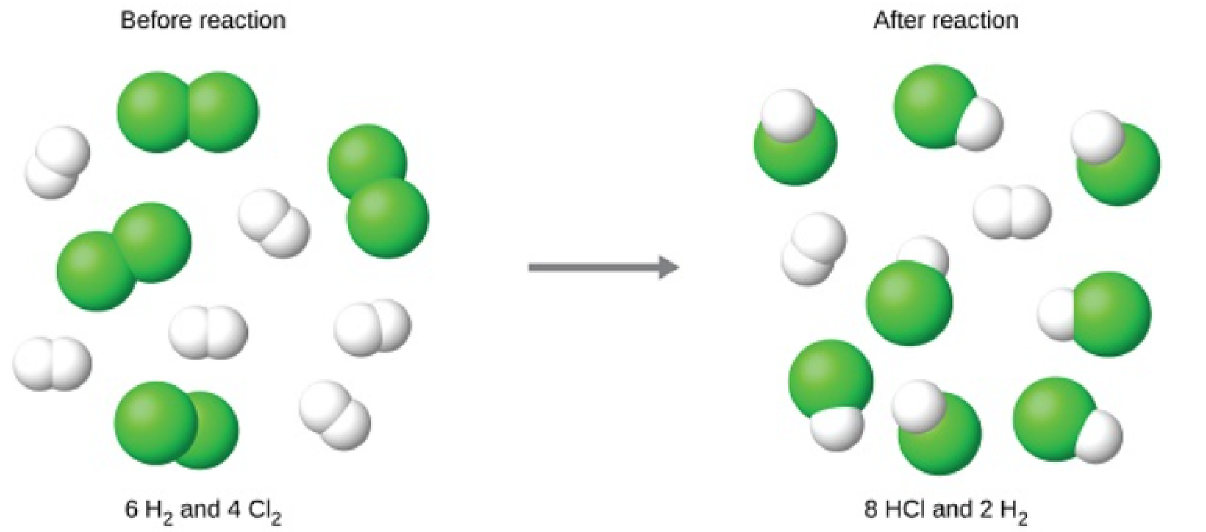


Figure 11.2.2: When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant. The figure shows a space-filling molecular models reacting. There is a reaction arrow pointing to the right in the middle. To the left of the reaction arrow there are three molecules each consisting of two green spheres bonded together. There are also five molecules each consisting of two smaller, white spheres bonded together. Above these molecules is the label, "Before reaction," and below these molecules is the label, "6 H subscript 2 and 4 C l subscript 2." To the right of the reaction arrow, there are eight molecules each consisting of one green sphere bonded to a smaller white sphere. There are also two molecules each consisting of two white spheres bonded together. Above these molecules is the label, "After reaction," and below these molecules is the label, "8 H C l and 2 H subscript 2."

Practice: Reactant Limit/Excess

76.4 g tribromoethane reacts with **49.1 g** oxygen in the reaction below. Which is limiting?

$$4\ C_2H_3Br_3 + 11\ O_2 \rightarrow 8\ CO_2 + 6\ H_2O + 6\ Br_2$$

- To solve: convert reactants actual moles, compare one reactant to another in consumption
- Identify given data and needed quantities:
  - provided: masses: $C_2H_3Br_3$ = 76.4 g, O_2 = 49.1 g; reaction ratio: 11 mol O_2 / 4 mol $C_2H_3Br_3$; molar masses: $C_2H_3Br_3$ = 266.72 g/mol, O_2 = 32.00 g/mol,
- Substitute values & solve:

$$76.4\ \text{g}\ C_2H_3Br_3 \times \frac{1\ \text{mol}\ C_2H_3Br_3}{266.72\ \text{g}\ C_2H_3Br_3} \times \frac{11\ \text{mol}\ O_2}{4\ \text{mol}\ C_2H_3Br_3} = 0.788\ \text{mol}\ O_2\ \text{needed/consumed}$$

$$49.1\ \text{g}\ O_2 \times \frac{1\ \text{mol}\ O_2}{32.00\ \text{g}\ O_2} = 1.53\ \text{mol}\ O_2\ \text{available}$$

$$49.1\ \text{g}\ O_2 \times \frac{1\ \text{mol}\ O_2}{32.00\ \text{g}\ O_2} \times \frac{4\ \text{mol}\ C_2H_3Br_3}{11\ \text{mol}\ O_2} = 0.558\ \text{mol}\ C_2H_3Br_3\ \text{needed/consumed}$$

$$76.4\ \text{g}\ C_2H_3Br_3 \times \frac{1\ \text{mol}\ C_2H_3Br_3}{266.72\ \text{g}\ C_2H_3Br_3} = 0.286\ \text{mol}\ C_2H_3Br_3\ \text{available}$$

Practice: Reactant Limit/Excess

2.40 g magnesium reacts with **10.0 g** oxygen to make MgO. Which is limiting?

$$2\ Mg\ (s) + O_2\ (g) \rightarrow 2\ MgO\ (s)$$

- Identify given data and needed quantities:
  - provided: Mg: mass = 76.4 g, molar mass = 24.31 g/mol; O_2: mass = 10.0 g; molar mass = 32.00 g/mol, MgO: molar mass = 40.31 g/mol
- Substitute values & solve:

$$2.40\ \text{g Mg} \times \frac{1\ \text{mol Mg}}{24.31\ \text{g Mg}} \times \frac{1\ \text{mol O}_2}{2\ \text{mol Mg}} \times \frac{32.00\ \text{g O}_2}{1\ \text{mol O}_2} = 1.58\ \text{g O}_2\ \text{needed/consumed}$$

$10.0\ \text{g O}_2$ available, O_2 in excess, Mg in limiting

$$2.40\ \text{g Mg} \times \frac{1\ \text{mol Mg}}{24.31\ \text{g Mg}} \times \frac{1\ \text{mol MgO}}{1\ \text{mol Mg}} \times \frac{40.31\ \text{g MgO}}{1\ \text{mol MgO}} = 3.98\ \text{g MgO to be produced}$$

Practice: Reactant Limit/Excess

22.7 g MgO reacts with **17.9 g** H_2S to make MgS and H_2O. What is mass of excess reactant that will remain?

$$MgO\ (s) + H_2S\ (g) \rightarrow MgS\ (s) + H_2O\ (l)$$

- Identify given data and needed quantities:
  - provided: MgO: mass = 22.7 g, molar mass = 40.304 g/mol; H_2S: mass = 17.9 g molar mass = 34.082 g/mol
- Substitute values & solve:

$$22.7\ \text{g MgO} \times \frac{1\ \text{mol MgO}}{40.304\ \text{g MgO}} \times \frac{1\ \text{mol H}_2\text{S}}{1\ \text{mol MgO}} \times \frac{34.082\ \text{g H}_2\text{S}}{1\ \text{mol H}_2\text{S}} = 19.2\ \text{g H}_2\text{S needed/consumed}$$

$17.9\ \text{g H}_2\text{S available}$, MgO in excess, H_2S in limiting

$$17.9\ \text{g H}_2\text{S} \times \frac{1\ \text{mol H}_2\text{S}}{34.082\ \text{g H}_2\text{S}} \times \frac{1\ \text{mol MgO}}{1\ \text{mol H}_2\text{S}} \times \frac{40.304\ \text{g MgO}}{1\ \text{mol MgO}} = 21.2\ \text{g MgO}\ \frac{\text{needed}}{\text{consumed}}$$

$$22.7\ \text{g MgO available} - 21.2\ \text{g MgO needed} = 1.5\ \text{g excess}$$

Yield: Theoretical, Actual, Percent

- When doing a synthesis from reactants to products, there is a concern about amount of product obtained for reactant used. This is the **yield**
- The **theoretical yield** is the possible yield based on the amount of reactants used keeping in mind the limiting reactant
- The **actual yield** is the amount of product obtained during the processing.
- The **percent yield** is the ratio of actual yield to theoretical yield

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Does it matter if the amounts are in moles or grams?

Practice: Yields

40.0 g potassium chlorate is decomposed in Bunsen burner and the oxygen gas is collected mass = **14.9 g** What is theoretical, % yield

$$2\ KClO_3\ (s) \rightarrow 2\ KCl\ (s) + 3\ O_2\ (g)$$

- Identify given data and needed quantities:
  - provided: $KClO_3$: mass = 40.0 g, molar mass = 122.55 g/mol; O_2: mass = 14.9 g; molar mass = 32.00 g/mol
- Substitute values & solve:

$$40.0\ \text{g KClO}_3 \times \frac{1\ \text{mol KClO}_3}{122.55\ \text{g KClO}_3} \times \frac{3\ \text{mol O}_2}{2\ \text{mol KClO}_3} \times \frac{32.00\ \text{g O}_2}{1\ \text{mol O}_2} = 15.7\ \text{g O}_2\ \text{theoretical mass}$$

$$\frac{14.9\ \text{g O}_2\ \text{(actual mass)}}{15.7\ \text{g O}_2\ \text{(theoretical mass)}} \times 100\% = 94.9\%$$

Practice: Yields

32.9 g carbon tetrachloride is reacted in excess hydrofluoric acid and it makes **12.5 g** of Freon (CF_2Cl_2). What is the percent yield?

$$CCl_4\ (l) + 2\ HF\ (l) \rightarrow CF_2Cl_2\ (g) + 2\ HCl\ (g)$$

- Identify given data and needed quantities:
  - provided: CCl_4: mass = 32.9 g, molar mass = 153.81 g/mol; CF_2Cl_2: mass = 12.5 g; molar mass = 120.91 g/mol
- Substitute values & solve:

$$32.9\text{ g }CCl_4 \times \frac{1\text{ mol }CCl_4}{153.81\text{ g }CCl_4} \times \frac{1\text{ mol }CF_2Cl_2}{1\text{ mol }CCl_4} \times \frac{120.91\text{ g }CF_2Cl_2}{1\text{ mol }CF_2Cl_2} = 25.9\text{ g }CF_2Cl_2\text{ theoretical mass}$$

$$\frac{12.5\text{ g }CF_2Cl_2\text{ (actual mass)}}{25.9\text{ g }CF_2Cl_2\text{ (theoretical mass)}} \times 100\% = 48.3\%$$

Causes for Less Than 100% Percent Yield

- Incomplete reaction: some reactions may be slow or even reversible (equilibrium between reactants and products), so limiting reactant not converted to product
- Side reaction: there is competing, undesired side reaction
- Losses in transfer/workup/handling: these are physical losses related to loss of intermediates or products between containers, purification, filtration, extraction, spillage
- Impure reactants: what was thought to be a certain amount of limiting reactant was actually less because impurities added to the total mass

Causes for Greater Than 100% Percent Yield

- Experimental or measurement error in determining amounts/masses of reactant(s): otherwise it violates conservation of mass principle
- Incomplete drying or impurity: most common reason is product that is contaminated with substance adding mass:
  - Solvent (water): used for product washing and product mass obtained before full drying
  - Unreacted starting material or byproducts not removed during purification
- Measurement error: errors getting mass of final product (forgetting to "tare" or zero balance), calculation errors (wrong molar mass)

Energy in Reactions

Previously addressed

- Law of Conservation of Energy
- System and Surroundings
- Endothermic and Exothermic

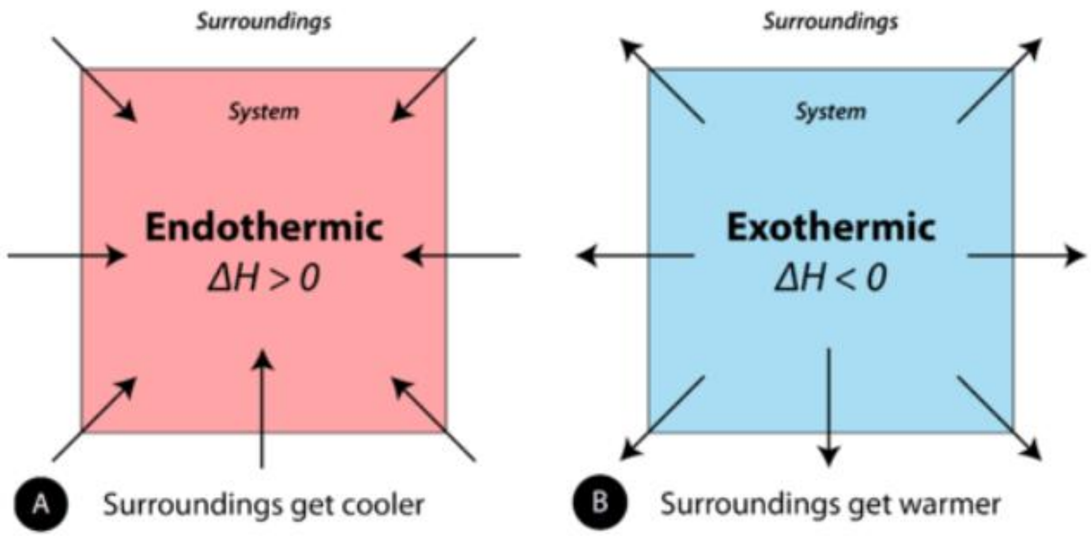


Figure 11.4.1: (A) Endothermic reaction. (B) Exothermic reaction.

Enthalpy

- **Enthalpy** refers to energy of a system under constant pressure (usually atmospheric)
- Its symbol is $\boldsymbol{H}$
- The change in enthalpy (ΔH) = heat (energy) absorbed or release in a chemical reaction at constant pressure
- $\Delta H = H_{products} - H_{reactants}$
- What energy? Energy that breaks and makes/forms chemical bonds in reactions
- Factors: amounts of and phase/state (solid, liquid, gas) determine enthalpy

Thermochemical Equation

- A thermochemical equation is a chemical reaction equation that includes the change in enthalpy of the reaction

$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (l) + \mathbf{890.4\ kJ}$

- The combustion of methane includes the release of 890.4 kJ as a product of the reaction
- In another typical expression of the reaction, the enthalpy is shown to the side:

$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (l) \quad \Delta H = -890.4\ kJ$

- Note that $\Delta H = \boldsymbol{H_{products} - H_{reactants}}$ is less than zero, that is the energy of the products is less than the energy of reactants, which makes sense, as combustion is an exothermic reaction

Thermochemical Equation

- Look at the thermochemical equation of an endothermic reaction

$CaCO_3$ (*s*) + **177.8 kJ** → CaO (*s*) + CO_2 (*g*)

- Heat is absorbed for this reaction to occur (you are likely getting this reaction to happen on a hot plate)
- The other way of expressing this is:

$CaCO_2$ (*s*) → CaO (*s*) + CO_2 (*g*) ΔH = 177.8 kJ

- A reaction like this can be reversed too: carbon dioxide gas in calcum oxide will form calcium carbonate:

CaO (*s*) + CO_2 (*g*) → $CaCO_2$ (*s*) ΔH = –177.8 kJ

Enthalpy Calculations

Energies in reactions under constant pressure of enthalpies. The energies also depend on amounts as well: if one mole of methane burned in oxygen releases 890.4 kJ energy, then two moles releases 2 x 890.4 kJ (= 1781 kJ) and a 0.5 mol releases 0.5 x 890.4 kJ (=445.2 kJ)

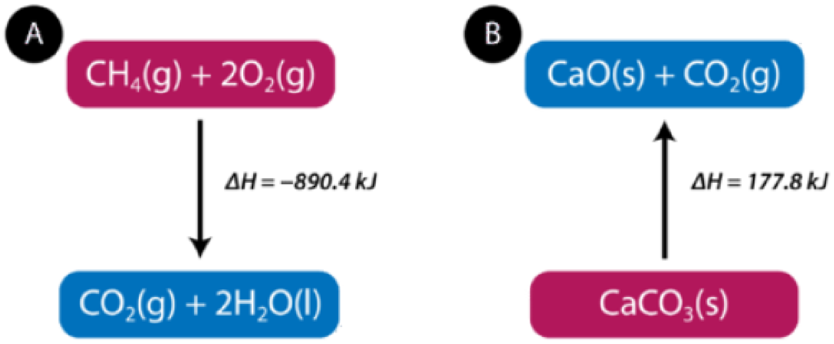


Figure 11.4.2: (A) As reactants are converted to products in an exothermic reaction, enthalpy is released into the surroundings. The enthalpy change of the reaction is negative. (B) As reactants are converted to products in an endothermic reaction, enthalpy is absorbed from the surroundings. The enthalpy change of the reaction is positive.

Practice: Enthalpy Calculations

Sulfur dioxide (SO_2) gas reacts with oxygen (O_2) to form sulfur trioxide (SO_3) in an exothermic reaction

$$2\ SO_2\ (g) + O_2\ (g) \rightarrow 2\ SO_3\ (g) + 198\ \text{kJ}$$

What is ΔH when **58.0 g** SO_2 reacts with excess O_2?

- Identify given data and needed quantities:
  - provided: SO_2: mass = 58.0 g, ΔH = 198 kJ for 2 mol SO_2; SO_2 molar mass = 64.07 g/mol
- Substitute values & solve:

$$\Delta H = 58.0\ \text{g SO}_2 \times \frac{1\ \text{mol SO}_2}{64.07\ \text{g SO}_2} \times \frac{-198\ \text{kJ}}{2\ \text{mol SO}_2} = -89.6\ \text{kJ}$$