4. Stoichiometry

1. Stoichiometric Equations
2. Limiting Reagent Problems
3. Percent Yield
4. Limiting Reagent Problems
5. Concentrations of Solutes
6. Chemical Analysis
4.1 Stoichiometry

“Chemical Arithmetic”

Objective: Determine Quantitative Relationships Between Reactants and Products

Two “Ways” to Quantify Matter

1. Count
2. Measure
Stoichiometry

- Chemical Equations Relate the Numbers of Various Species to Each Other (Counted Quantities)
- We Observe Measurable Quantities Like Mass and Volume
- We Need to be Able to Connect the Two
3 Basic Steps

1. **Mass to Mole Conversion:** Determine # of Moles of all Species Present (Start with Mass Reactants)

2. **Mole to Mole Conversion:** Use Stoichiometric Ratios to Relate Various Species to Each Other

3. **Mole to Mass Conversion:** Multiple Moles Product by Molar Mass (End with Mass Products)
Butane Lighter

Consider a Butane Lighter that weighs 16.0g full and 12.0g empty. What information can we get from this?

\[2\text{C}_4\text{H}_{10} + 13\text{O}_2 \rightarrow 10\text{H}_2\text{O} + 8\text{CO}_2\]

1. Quantity (Mass or Moles) of Butane in Lighter
2. Quantity of Oxygen Consumed by Lighter
3. Quantity of CO\(_2\) Produced by Lighter
4. Quantity of Water Produced by Lighter
\[ 2C_4H_{10} + 13O_2 \rightarrow 10H_2O + 8CO_2 \]

1. Quantity of Butane in Lighter

\[ (16.0 - 12.0 \text{ g} \cdot C_4H_{10}) \left( \frac{\text{mol} \cdot C_4H_{10}}{58 \text{ g}} \right) = 0.069 \text{ mol} \cdot C_4H_{10} \]

2. Quantity of Oxygen Consumed by Lighter

\[ (0.069 \text{ mol} \cdot C_4H_{10}) \left( \frac{13 \text{ mol} \cdot O_2}{2 \text{ mol} \cdot C_4H_{10}} \right) = 0.45 \text{ mol} \cdot O_2 \]

or

\[ (0.45 \text{ mol} \cdot O_2) \left( \frac{32 \text{ g}O_2}{\text{mol}} \right) = 14 \text{ g} \cdot O_2 \]
In 1 Equation:

\[ 2\text{C}_4\text{H}_{10} + 13\text{O}_2 \rightarrow 10\text{H}_2\text{O} + 8\text{CO}_2 \]

1. Mass to Mole

\[ \frac{4\text{g} \cdot \text{C}_4\text{H}_{10}}{58\text{g}} \]

2. Mole to Mole

\[ \frac{13\text{mol} \cdot \text{O}_2}{2\text{mol} \cdot \text{C}_4\text{H}_{10}} \]

3. Mole to Mass

\[ \frac{32\text{g} \cdot \text{O}_2}{\text{mol}} = 14\text{g} \cdot \text{O}_2 \]
Take Home Problems:

\[ 2C_4H_{10} + 13O_2 \rightarrow 10H_2O + 8CO_2 \]

For the 4 grams of Butane, Determine:

1. Moles Water Produced = 0.35 mol \( H_2O \)
2. Grams Water Produced = 6.2 g \( H_2O \)
3. Moles Carbon Dioxide Produced = 0.28 mol \( CO_2 \)
4. Grams Carbon Dioxide Produced = 12 g \( CO_2 \)
4.2 Limiting Reagent Problems

• *Stoichiometric Proportions*: Reactants mixed in ratios given by stoichiometric coefficients. Everything is completely consumed.

• *Nonstoichiometric Proportions*: Reactants are mixed in ratios different than stoichiometric coefficients. One species is completely consumed *(Limiting Reagent)* while another is not completely consumed *(Excess Reagent)*.
Theoretical Yield

The Maximum Amount of Product Which Could be Produced by the Complete Consumption of the Limiting Reagent
Silver Tarnishes in the Presence of Hydrogen Sulfide and Oxygen due to the Reaction:

$$4\text{Ag} + 2\text{H}_2\text{S} + \text{O}_2 \rightarrow 2\text{Ag}_2\text{S} + 2\text{H}_2\text{O}$$

How many grams of Silver Sulfide Would be Formed from 2.4g Ag, 0.48g H$_2$S & 0.16g O$_2$?

2.4g Ag ($\frac{1\text{mol Ag}}{107.9\text{g}}$) ($\frac{2\text{mol } \text{Ag}_2\text{S}}{4\text{mol Ag}}$) = 0.011mol Ag$_2$S

0.48g H$_2$S ($\frac{1\text{mol H}_2\text{S}}{34\text{g}}$) ($\frac{2\text{mol } \text{Ag}_2\text{S}}{2\text{mol H}_2\text{S}}$) = 0.014mol Ag$_2$S

0.16g O$_2$ ($\frac{1\text{mol O}_2}{32\text{g}}$) ($\frac{2\text{mol } \text{Ag}_2\text{S}}{\text{mol O}_2}$) = 0.010mol Ag$_2$S

Oxygen is the Limiting Reagent
Mass Silver Sulfide Produced:

\[ 4\text{Ag} + 2\text{H}_2\text{S} + \text{O}_2 \rightarrow 2\text{Ag}_2\text{S} + 2\text{H}_2\text{O} \]

\[(0.010\text{mol Ag}_2\text{S}) \times (247.9\text{Ag}_2\text{S/mol}) = 2.5\text{g Ag}_2\text{S}\]
How Do We Quickly Identify Limiting Reactants in Stoichiometric Problems?

1. Determine the Initial Number of Moles for Each Reactant
2. Divide Each by Its Stoichiometric Coefficient.
3. Limiting Reagent Has Smallest Value (Reactants are in Stoichiometric Proportions if all Values are Equal)
A Closer Look at the Last Problem

\[ 4\text{Ag} + 2\text{H}_2\text{S} + \text{O}_2 \rightarrow 2\text{Ag}_2\text{S} + 2\text{H}_2\text{O} \]

Determine Moles Each Species

Divide by Coefficient

Note: Next Step Multiples all by the Same Value

2.4g Ag (1mol Ag/107.9g) \( (2\text{mol } \text{Ag}_2\text{S}/4\text{mol Ag}) = 0.011\text{mol } \text{Ag}_2\text{S} \)

0.48g H\(_2\)S (1mol H\(_2\)S/34g) \( (2\text{mol } \text{Ag}_2\text{S}/2\text{mol } \text{H}_2\text{S}) = 0.014\text{mol } \text{Ag}_2\text{S} \)

0.16g O\(_2\) (1mol O\(_2\)/32g) \( (2\text{mol } \text{Ag}_2\text{S}/\text{mol } \text{O}_2) = 0.010\text{mol } \text{Ag}_2\text{S} \)
A Closer Look at Short Cut Method

4Ag + 2H₂S + O₂ ----> 2Ag₂S + 2H₂O

2.4g Ag (1mol Ag/107.9g) / 4 mol Ag = 0.00556
0.48g H₂S (1mol H₂S/34g) / 2 mol H₂S = 0.00706
0.16g O₂ (1mol O₂/32g) / mol O₂ = 0.005

Possible Questions:

1. Theoretical Mass Yield Ag₂S
2. Theoretical Mass Yield H₂O
3. Excess Ag
4. Excess H₂S
Theoretical Mass Yield $\text{Ag}_2\text{S}$

$$4\text{Ag} + 2\text{H}_2\text{S} + \text{O}_2 \longrightarrow 2\text{Ag}_2\text{S} + 2\text{H}_2\text{O}$$

0.16g $\text{O}_2$ ($\text{mol O}_2/32g$)($2\text{molAg}_2\text{S }/\text{mol O}_2$)(247.9g/mol$\text{Ag}_2\text{S}$) = 2.5g

or

$$[0.005](2\text{molAg}_2\text{S })(247.9\text{g/molAg}_2\text{S})= 2.5\text{g}$$
Theoretical Mass Yield $\text{H}_2\text{O}$

$4\text{Ag} + 2\text{H}_2\text{S} + \text{O}_2 \longrightarrow 2\text{Ag}_2\text{S} + 2\text{H}_2\text{O}$

$0.16\text{ g} \ 	ext{O}_2 (\text{mol O}_2/32\text{g})(2\text{mol H}_2\text{S}/\text{mol O}_2)(18\text{g/mol H}_2\text{O}) = 0.18\text{g}$

or

$[0.005] (2\text{mol H}_2\text{O})(18\text{g/mol H}_2\text{O}) = 0.18\text{g}$
Excess Mass of Ag

\[ 4Ag + 2H_2S + O_2 \rightarrow 2Ag_2S + 2H_2O \]

**Mass Ag Left Over** = Initial Mass - Mass Consumed

\[
2.4g - 0.16gO_2 \text{ (mol } O_2/32g)(4\text{ mol } Ag /\text{mol } O_2)(107.9g/\text{mol } Ag) = 0.24g
\]

or

\[
2.4 - [0.005](4 \text{ mol } Ag)(107.9g/\text{mol } Ag) = 0.24g
\]
Excess Mass of $\text{H}_2\text{S}$

$$4\text{Ag} + 2\text{H}_2\text{S} + \text{O}_2 \rightarrow 2\text{Ag}_2\text{S} + 2\text{H}_2\text{O}$$

Mass $\text{H}_2\text{S}$ Left Over $=$ Initial Mass - Mass Consumed

$$0.48g - 0.16g\text{O}_2 \left(\frac{\text{mol} \text{O}_2}{32g}\right) \left(2\text{molH}_2\text{S}/\text{molO}_2\right) \left(\frac{34g/\text{molH}_2\text{S}}{}\right) = 0.24g$$

or

$$0.48 - [0.005](2 \text{ molH}_2\text{S})(34g/\text{mol H}_2\text{S}) = 0.24g$$
How Much Oxygen is Left Over?

None:

It is the limiting reagent
Phosphoric Acid can be synthesized from phosphorus, oxygen and water according to the following reactions:

$$4P + 5O_2 \rightarrow P_4O_{10}$$

$$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$$

Starting with 20.0g P, 30.0g O\textsubscript{2} and 15.0 g H\textsubscript{2}O, how many grams of Phosphoric acid can be Produced?
Combining Equations gives:

\[4P + 5O_2 \rightarrow P_4O_{10}\]

\[P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4\]

\[4P + 5O_2 + 6H_2O \rightarrow 4H_3PO_4\]

(Note, the \(P_4O_{10}\) could be considered as a reaction intermediate)
\[4P + 5O_2 + 6H_2O \rightarrow 4H_3PO_4\]

\[
\begin{align*}
20gP(\text{mol P/31g})(1/4 \text{ mol P}) &= 0.1613 \\
30gO_2(\text{molO}_2/32g)(1/5 \text{ mol O}_2) &= 0.1875 \\
15gH_2O(\text{molH}_2O/18g)(1/6 \text{ mol H}_2O) &= 0.1389 \\
\end{align*}
\]

\[
\begin{align*}
(15gH_2O)(\text{molH}_2O/18g)(4\text{molH}_3PO_4/6\text{mol H}_2O)(98g \text{H}_3PO_4/\text{mol}) &= 54.5g \text{H}_3PO_4 \\
0.1389(4\text{molH}_3PO_4)(98g \text{H}_3PO_4/\text{mol}) &= 54.5g \text{H}_3PO_4 \\
\end{align*}
\]
4.3 Percent Yields

Often Times the Actual Yield of a Chemical Reaction is Less Than That Predicted by the Complete Consumption of the Limiting Reagent

\[
\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100
\]

Note: Theoretical Yield Is Based on the Complete Consumption of the Limiting Reagent
Why Would the Actual Yield Be Less Than the Theoretical Yield?

1. **Equilibrium State** (where Reactants exist together with Products)

2. **Existence of Intermediates** (two step process like formation of Phosphoric Acid)

3. **Kinetic Factors** (the reaction is slow, maybe not all of the reactants have “diffused” together).
Determine Percent Yield if 2.3g of Silver Sulfide form when 2.4g Ag, 0.28g H₂S & 0.16g O₂ react.

From Previous Problem, the Theoretical Yield is 2.5g Ag₂S

\[
\% \text{ Yield (Ag}_2\text{S)} = \frac{2.3\text{g}}{2.5\text{g}} \times 100 = 92\%
\]
Class Problem

If, 20.0g P, 30.0g O₂ and 15.0 g H₂O; Yields 50.3 g H₃PO₄, what is the Percent Yield?

Note in Previous Problem the Theoretical Yield was determined to be 54.5 g phosphoric acid.
Class Problem:

\[
\left( \frac{50.3 \text{g } H_3PO_4}{54.5 \text{g } H_3PO_4} \right)(100) = 92.3\% \text{ Yield } H_3PO_4
\]
4.4 Chemical Equations & Chemical Analysis
Acidum Formicum – Formic Acid
What is the empirical formula of Formic Acid?

Obtained through the distillation of Formica rufa
**Empirical Formula of Formic Acid from Combustion Data**

Sample $(C_x H_y O_n)$

- 2.3482 g formic acid

Step 1: Determine Mass of each element

$$m_H = 0.9606 g \times \left( \frac{2gH}{18gH_2O} \right) = 0.1067g \text{ H}$$

$$m_C = 2.3482 g \times \left( \frac{12gC}{44gCO_2} \right) = 0.6404g \text{ C}$$

$$m_O = 2.2482g - 0.1067g - 0.6404g = 1.6011g \text{ O}$$
Step 2: Determine Mole of Each Element

\[
\begin{align*}
\text{moles } H &= 0.1067 \text{ gH} \left( \frac{1 \text{ mol}}{1.008 \text{ g}} \right) = 0.1067 \text{ mol } H \\
\text{moles } C &= 0.6404 \text{ gC} \left( \frac{1 \text{ mol}}{12.01 \text{ g}} \right) = 0.0533 \text{ mol } C \\
\text{moles } O &= 1.6011 \text{ gO} \left( \frac{1 \text{ mol}}{16.00 \text{ g}} \right) = 0.1000 \text{ mol } O
\end{align*}
\]

Step 3: Divide by Smallest Number

\[
\begin{align*}
H : & \quad 0.1067 \left( \frac{1}{0.0533} \right) = 2 \\
C : & \quad 0.0533 \left( \frac{1}{0.0533} \right) = 1 \\
H : & \quad 0.1000 \left( \frac{1}{0.0533} \right) = 2
\end{align*}
\]

\(\text{CH}_2\text{O}_2\)
4.5 Concentration of a Solute

1. Mass Concentration:
   Typical Units: g/L

\[
\text{mass solute} \quad \frac{m}{\text{vol solution}} = \frac{m}{V}
\]

2. Mole Concentration:
   (Molarity)
   Typical Units: M=mol/L

\[
\text{moles solute} \quad \frac{n}{\text{vol solution}} = \frac{n}{V} = M
\]

Conventions
- \( m \) = mass
- \( n \) = moles
- \( M \) = molarity
\[ M = \text{molarity} = \frac{\text{moles solute}}{\text{Liter solution}} \]

**Why is Molarity so Important in Solution Stoichiometric Calculations?**

Because solutes are typically reactants and we can measure the solution volume, this gives us the moles solute. i.e., we can “\text{Count moles solute}” by measuring the volume of a solution of known Molarity.
Calculate the Molarity of a Solution if 2.00 g NaCl is diluted to 250.0 ml with water.

\[
(2.00 \text{ g NaCl})(\frac{\text{mol NaCl}}{58.5 \text{ g}})(\frac{1}{.250 \text{ L}}) = .137M
\]

How much NaOH would you need to make 500. mL of .70M NaOH?

\[
.500L(\frac{.70 \text{ mol NaOH}}{L})(\frac{40 \text{ g NaOH}}{\text{mol}})=14 \text{ g}
\]
How do You Make 500.0 ml of .1560M CuSO₄?

1. Calculate amount of Copper(II) Sulfate you will need.

\[
0.5000L \left( \frac{0.1560 \text{mol} \ CuSO_4}{L} \right) \left( \frac{159.6 \text{g} \ CuSO_4}{\text{mol}} \right) = 12.48 \text{g}
\]
2. Weight the Copper(II) Sulfate
3. Quantitatively Transfer to 500 ml Volumetric Flask
4. Dilute to Volume
Concentration of Ions

The concentration of ions are related to the salt via its formula.

\[
\text{Al}_2(\text{SO}_4)_3 \rightarrow 2 \text{Al}^{+3} + 3 \text{SO}_4^{-2}
\]

What are the ion concentrations for a 1M Al\textsubscript{2}(SO\textsubscript{4}) solution?

2M in Al\textsuperscript{+3} and 3M in SO\textsubscript{4}^{-2}

(5M in Total Ion Concentration)
Dilution Problems

Many Reagents Come as Stock Solutions
Adding Solvent Does Not Change The Moles of Solute

Moles Before Dilution = Moles After Dilution

\[ n_{\text{initial}} = n_{\text{final}} \]

\[ M_i V_i = M_f V_f \]
How would we make 100. mL of 4.0 M Sulfuric Acid from Stock 16 M Sulfuric Acid?

1. Determine Amount of 16 M Sulfuric Acid needed.

\[ M_i V_i = M_f V_f \]

\[ V_i = \left( \frac{M_f}{M_i} \right) V_f = \left( \frac{4.0M}{16M} \right) 100\text{mL} = 25\text{mL} \]
2. Accurately Measure 25mL with Volumetric Pipet (TD)
3. Transfer to 100ml Volumetric Flask
4. Dilute to Volume
Solution Reaction Stoichiometry

1. Identify species present and possible rxns.

2. Write Balanced Net Ionic Eq.

3. Calculate Moles of Each Reactant

4. Determine Limiting Reagent

5. Calculate Moles of Other Reactants and Products as Desired

6. Convert Results to Desired Units (Moles, Molarity, Mass... )
Solution Stoichiometry Problem

Consider mixing 30.0mL of 0.750M sodium chromate with 40.0 mL of 0.500 M aluminum chloride.

1. Predict the mass and identity of any precipitate formed.

2. Predict the concentrations of any ions remaining in solution.
1. Write Balanced Eq.

\[ 3\text{Na}_2\text{CrO}_4(\text{aq}) + 2\text{AlCl}_3(\text{aq}) \rightarrow 6\text{NaCl} + \text{Al}_2(\text{CrO}_4)_3 \]

2. Use Solubility Rules to Predict States of Products

\[ 3\text{Na}_2\text{CrO}_4(\text{aq}) + 2\text{AlCl}_3(\text{aq}) \rightarrow 6\text{NaCl}(\text{aq}) + \text{Al}_2(\text{CrO}_4)_3(\text{s}) \]

3. Write Net Ionic Equation

\[ 3\text{CrO}_4^{2-} + 2\text{Al}^{3+} \rightarrow \text{Al}_2(\text{CrO}_4)_3(\text{s}) \]
3. Identify Initial Moles of all Species

**Na⁺:**

\[
0.0300 \ell \left( \frac{0.750 \text{ mol} \ Na_2\text{CrO}_4}{\ell} \right) \left( \frac{2 \text{ mol} \ Na^+}{\text{mol} \ Na_2\text{CrO}_4} \right) = 0.0450 \text{ mol} \ Na^+
\]

**CrO₄²⁻:**

\[
0.0300 \ell \left( \frac{0.750 \text{ mol} \ Na_2\text{CrO}_4}{\ell} \right) \left( \frac{\text{mol} \ CrO_4^{2-}}{\text{mol} \ Na_2\text{CrO}_4} \right) = 0.0225 \text{ mol} \ CrO_4^{2-}
\]

**Al³⁺:**

\[
0.0400 \ell \left( \frac{0.5 \text{ mol} \ Al\text{Cl}_3}{\ell} \right) \left( \frac{\text{mol} \ Al^{3+}}{\text{mol} \ Al\text{Cl}_3} \right) = 0.0200 \text{ mol} \ Al^{3+}
\]

**Cl⁻:**

\[
0.0400 \ell \left( \frac{0.5 \text{ mol} \ Al\text{Cl}_3}{\ell} \right) \left( \frac{3 \text{ mol} \ Cl^-}{\text{mol} \ Al\text{Cl}_3} \right) = 0.0600 \text{ mol} \ Cl^-
\]
4. Determine Precipitate Based on Complete Consumption of Limiting Reagent

$$3\text{CrO}_4^{-2} + 2\text{Al}^{+3} \rightarrow \text{Al}_2(\text{CrO}_4)_3(s)$$

\[
0.0225 \text{ mol CrO}_4^{-2} \left( \frac{\text{mol Al}_2(\text{CrO}_4)_3}{3 \text{ mol CrO}_4^{-2}} \right) \left( \frac{401.96 \text{ g Al}_2(\text{CrO}_4)_3}{\text{mol}} \right) = 3.01 \text{ g Al}_2(\text{CrO}_4)_3
\]

\[
0.0200 \text{ mol Al}^{+3} \left( \frac{\text{mol Al}_2(\text{CrO}_4)_3}{2 \text{ mol Al}^{+3}} \right) \left( \frac{401.96 \text{ g Al}_2(\text{CrO}_4)_3}{\text{mol}} \right) = 4.02 \text{ g Al}_2(\text{CrO}_4)_3
\]
5. Determine Concentration of Excess Reagent

\[
\frac{\text{moles}_{\text{Initial}} - \text{moles}_{\text{Consumed}}}{\text{vol}_{\text{Total}}} = 0.714 M \text{ Al}^{+3}
\]

\[
0.0200 \text{ mol Al}^{+3} - 0.0225 \text{ mol CrO}_4^{-2} \left( \frac{2 \text{ mol Al}^{+3}}{3 \text{ mol CrO}_4^{-2}} \right) = 0.0400 l + 0.0300 l
\]
6. Determine Concentration of Spectator Ions

Na\(^+\): \[
\frac{0.0450 \text{ mol Na}^+}{0.0700 \text{ L}} = 0.643 \text{M}
\]

Cl\(-\): \[
\frac{0.0600 \text{ mol Cl}^-}{0.0700 \text{ L}} = 0.857 \text{M}
\]
pH Scale

\[ pX = -\log[X] \]

\[ \text{pH} = -\log[H_3O^+] \]

\[ [H^+] = 10^{-\text{pH}} = \frac{1}{10^{\text{pH}}} \]
What is pH if $[H^+] = 3.50 \times 10^{-6}$?

Exact # tells position of decimal

Shows # of significant figures

$5.456$
pH of Strong Acids & Bases

Consider 100% Dissociation

What is the pH of a 0.00836M HCl solution?

\[
pH = -\log(0.00836) = 2.078
\]
What is the hydronium ion concentration for a solution with a pH of 4.041?

\[ [H_3O^+] = 10^{-pH} = 10^{-4.041} = 9.099 \times 10^{-5} \]

= 9.10 \times 10^{-5}
Acid-Base Titrations

Analytical Technique to Determine the Quantity of an Unknown Acid or Base by Neutralizing With an Acid or Base of Known Concentration.
1. Add Acid to Erlenmeyer Flask, then add indicator
2. Add Base with Buret
3. Quit When Indicator Turns pink
What is the concentration of an unknown Sulfuric Acid solution if it takes 35mL of .40 M NaOH to neutralize 50.0 mL of the acid?

\[
2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}
\]

\[
0.035L \text{ NaOH} \times \left(\frac{\text{4mol NaOH}}{\text{L}}\right) \times \left(\frac{1\text{mol } \text{H}_2\text{SO}_4}{2\text{mol NaOH}}\right) \times \left(\frac{1}{0.05L}\right) = 0.14M \text{ H}_2\text{SO}_4
\]

-Note, we are using the initial volume of the acid as we want to know it’s initial strength