Chapter 9 – Stoichiometry
Section 9.1 – Introduction to Stoichiometry

**Standard 3.e.:** Students know how to calculate the masses of reactant and products in a chemical reaction from the mass of one of the reactants or products and the relevant atomic masses.

**Language Objective:** We will calculate the moles of a reactant or product using the moles of the reactant or product.

**Content Objective:** I will show through writing how to calculate the moles of a reactant or product using the moles of the reactant or product.

**Types of Stoichiometry Problems**
- **Given is in moles and unknown is in moles.**
  - moles of given $\rightarrow$ moles of unknown
- **Given is in moles and unknown is in mass (grams).**
  - moles of given $\rightarrow$ moles of unknown $\rightarrow$ mass of unknown
- **Given is in mass and unknown is in moles.**
  - mass of given $\rightarrow$ moles of given $\rightarrow$ moles of unknown
- **Given is in mass and unknown is in mass.**
  - mass of given $\rightarrow$ moles of given $\rightarrow$ moles of unknown $\rightarrow$ mass of unknown

**Writing and Using Mole Ratios**
- **Mole Ratio:** a conversion factor derived from the coefficients of a balanced chemical equation interpreted in terms of moles.
  - In chemical calculations, mole ratios are used to convert between:
    - Moles of reactants
    - Moles of reactants and moles of products
    - Moles of products

- $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
  - Three mole ratios can be derived from the balanced equation above:
    - $1 \text{ mol N}_2 \quad 2 \text{ mol NH}_3 \quad 3 \text{ mol H}_2$
    - $3 \text{ mol H}_2 \quad 1 \text{ mol N}_2 \quad 2 \text{ mol NH}_3$

- **Example:** $4 \text{ Li} + \text{O}_2 \rightarrow 2 \text{ Li}_2\text{O}$
  - $4 \text{ mol Li} \quad 4 \text{ mol Li} \quad 1 \text{ mol O}_2$
  - $1 \text{ mol O}_2 \quad 2 \text{ mol LiO}_2 \quad 2 \text{ mol Li}_2\text{O}$

**continued on next page**
Section 9.2 – Ideal Stoichiometric Calculations

Mole-Mole Calculations

- How many moles of ammonia (NH₃) are produced when 0.60 mol of N₂ reacts with H₂
  - **Step 1:** Write out equation and balance it
  - **Step 2:** Determine mole ratio of known (N₂) to unknown (NH₃).
  - **Step 3:** Multiply given amount of known by the correct mole ratio that will cancel out the known units, leaving you the unknown units

\[
\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3
\]

\[
\frac{1 \text{ mol N}_2}{2 \text{ mol NH}_3}
\]

\[
\left(\frac{0.60 \text{ mol } N_2}{1}\right) \cdot \left(\frac{2 \text{ mol NH}_3}{1 \text{ mol } N_2}\right) = 1.2 \text{ mol NH}_3
\]

- Example: How many moles of Chlorine are produced when 3.2 moles sodium chloride decomposes?

\[
2\text{NaCl} \rightarrow 2\text{Na} + \text{Cl}_2
\]

\[
\frac{2 \text{ mol NaCl}}{1 \text{ mol Cl}}
\]

\[
\left(\frac{3.2 \text{ mol NaCl}}{1}\right) \cdot \left(\frac{1 \text{ mol Cl}_2}{2 \text{ mol NaCl}}\right) = 1.6 \text{ mol Cl}_2
\]

Homework: page 301 #2 and page 311 #1 a, b (balance, and then a and b each have 3 answers)
Section 9.2 – Ideal Stoichiometric Calculations

Standard 3.e.: Students know how to calculate the masses of reactant and products in a chemical reaction from the mass of one of the reactants or products and the relevant atomic masses.

Language Objective: We will calculate the moles or mass of a reactant or product using the moles or mass of the reactant or product.

Content Objective: I will show through writing how to calculate the moles or mass of a reactant or product using the moles or mass of the reactant or product.

Mole-Mass Calculations
- How many grams of aluminum oxide (Al₂O₃) are produced when 0.50 mol of Al reacts with O₂
  - Step 1: Write out equation and balance it
  - Step 2: Multiply by the mole ratio to get from known to unknown moles.
  - Step 3: Multiply the moles of unknown by the molar mass of the unknown to get grams.

\[ 4 \text{Al} + 3 \text{O}_2 \rightarrow 2 \text{Al}_2\text{O}_3 \]

\[
\begin{align*}
4 \text{ mol Al} & \quad \text{Molar Mass } \text{Al}_2\text{O}_3 = 101.96 \text{ g/mol} \\
2 \text{ mol } \text{Al}_2\text{O}_3 & \\
\left( \frac{0.50 \text{ mol Al}_2}{1} \right) \cdot \left( \frac{2 \text{ mol } \text{Al}_2\text{O}_3}{4 \text{ mol Al}} \right) \cdot \left( \frac{101.96 \text{ g Al}_2\text{O}_3}{1 \text{ mol Al}_2\text{O}_3} \right) & = 25 \text{ g Al}_2\text{O}_3
\end{align*}
\]

- Example: How many grams of Copper are needed to react with Sulfur to form 5.70 moles of copper sulfide?

\[ 2 \text{Cu} + \text{S} \rightarrow \text{Cu}_2\text{S} \]

\[
\left( \frac{5.70 \text{ mol Cu}_2\text{S}}{1} \right) \cdot \left( \frac{2 \text{ mol Cu}}{1 \text{ mol Cu}_2\text{S}} \right) \cdot \left( \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} \right) = 724 \text{ g Cu}
\]

Mass-Mole Calculations
- Calculate the number of moles of H₂O produced by the reaction of 5.40 g of O₂ with an excess of H₂. *excess* means that there is enough to react.
  - Step 1: Write out equation and balance it.
  - Step 2: Multiply the given amount by the molar mass of the known to get moles.
  - Step 3: Multiply by the mole ratio to get from known to unknown moles.

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

\[
\left( \frac{5.40 \text{ g O}_2}{1} \right) \cdot \left( \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) \cdot \left( \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \right) = 0.338 \text{ mol H}_2\text{O}
\]
• Example: How many moles of hydrogen are needed to react with nitrogen to form 10.5 grams of ammonia (NH₃)?

\[
\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3
\]

\[
\left( \frac{10.5 \text{ g} \text{NH}_3}{1} \right) \cdot \left( \frac{1 \text{ mol} \text{NH}_3}{17.04 \text{ g} \text{NH}_3} \right) \cdot \left( \frac{3 \text{ mol} \text{H}_2}{2 \text{ mol} \text{NH}_3} \right) = 0.924 \text{ mol H}_2
\]

• Example: How many moles of Lithium are formed by the decomposition of 3.40 grams of Lithium Oxide?

\[
4\text{Li} + \text{O}_2 \rightarrow 2\text{Li}_2\text{O}
\]

\[
\left( \frac{3.40 \text{ g} \text{Li}_2\text{O}}{1} \right) \cdot \left( \frac{1 \text{ mol} \text{Li}_2\text{O}}{29.88 \text{ g} \text{Li}_2\text{O}} \right) \cdot \left( \frac{4 \text{ mol Li}}{2 \text{ mol} \text{Li}_2\text{O}} \right) = 0.228 \text{ mol Li}
\]

Homework: page 311 #2-3
Section 9.2 – Ideal Stoichiometric Calculations

Standard 3.e.: Students know how to calculate the masses of reactant and products in a chemical reaction from the mass of one of the reactants or products and the relevant atomic masses.

Language Objective: We will calculate the mass of a reactant or product using the mass of the reactant or product.

Content Objective: I will show through writing how to calculate the mass of a reactant or product using the mass of the reactant or product.

Mass-Mass Calculations

- Calculate the number of grams of CO produced by the reaction of 3.25 g of C with an excess of O₂.
  - Step 1: Write out equation and balance it.
  - Step 2: Multiply the given amount by the molar mass of the known to get moles.
  - Step 3: Multiply by the mole ratio to get from known to unknown moles.
  - Step 4: Multiply the moles of unknown by the molar mass of the unknown to get grams.

\[2C + O₂ \rightarrow 2CO\]

\[
\left(\frac{3.25 \text{ g C}}{1}\right) \cdot \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}}\right) \cdot \left(\frac{2 \text{ mol CO}}{2 \text{ mol C}}\right) \cdot \left(\frac{28.01 \text{ g CO}}{1 \text{ mol CO}}\right) = 7.58 \text{ g CO}
\]

- Example: How many grams of Oxygen are needed to make 20.0 grams of NO₂?

\[2NO + O₂ \rightarrow 2NO₂\]

\[
\left(\frac{20.0 \text{ g NO₂}}{1}\right) \cdot \left(\frac{1 \text{ mol NO₂}}{46.01 \text{ g NO₂}}\right) \cdot \left(\frac{1 \text{ mol O₂}}{2 \text{ mol NO₂}}\right) \cdot \left(\frac{32.00 \text{ g O₂}}{1 \text{ mol O₂}}\right) = 6.96 \text{ g O₂}
\]

- Example: How many grams of SO₃ are produced when 55.3 grams of SO₂ reacts with an excess of O₂.

\[2SO₂ + O₂ \rightarrow 2SO₃\]

\[
\left(\frac{55.3 \text{ g SO₂}}{1}\right) \cdot \left(\frac{1 \text{ mol SO₂}}{64.07 \text{ g SO₂}}\right) \cdot \left(\frac{2 \text{ mol SO₃}}{2 \text{ mol SO₂}}\right) \cdot \left(\frac{80.07 \text{ g SO₃}}{1 \text{ mol SO₃}}\right) = 69.1 \text{ g SO₃}
\]

Homework: page 311 #4 and practice problem #1 on page 311
Section 9.3 – Limiting Reactants and Percentage Yield

Standard 3.: Students know how to calculate percent yield in a chemical reaction.

Language Objective: We will find the limiting reactant in a chemical reaction.

Content Objective: I will show through writing how to find the limiting reactant in a chemical reaction.

- **Limiting reactant:** the substance that determines the amount of product that can be formed by a reaction.
- **Excess reactant:** the substance that is not completely used up in a reaction.
- Keep in mind that the reactant that is present in the smaller amount by mass or volume is not necessarily the limiting reactant.

Determining Limiting Reactant

- **What is the limiting reactant when 80.0g Cu reacts with 25.0g S?**

  \[2Cu(s) + S(s) \rightarrow Cu_2S(s)\]

  - Using the grams of reactant given, determine the moles of reactant.
  - Using the balanced chemical equation, determine the mole ratio.
  - Multiply the moles of one reactant by the mole ratio to get the needed amount of the other reactant and compare to the amount you have.

  \[
  \left(\frac{80.0 \text{ g Cu}}{1}\right) \cdot \left(\frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}}\right) = 1.26 \text{ mol Cu} \quad \text{Amount we have}
  \]

  \[
  \left(\frac{25.0 \text{ g S}}{1}\right) \cdot \left(\frac{1 \text{ mol S}}{32.07 \text{ g S}}\right) = 0.780 \text{ mol S} \quad \text{Amount we have}
  \]

  \[
  \left(\frac{1.26 \text{ mol Cu}}{1}\right) \cdot \left(\frac{1 \text{ mol S}}{2 \text{ mol Cu}}\right) = 0.630 \text{ mol S} \quad \text{Amount we need}
  \]

  \[
  \left(\frac{0.780 \text{ mol S}}{1}\right) \cdot \left(\frac{2 \text{ mol Cu}}{1 \text{ mol S}}\right) = 1.56 \text{ mol Cu} \quad \text{Amount we need}
  \]

  **Limiting reactant is Copper**

  *When you need more than you have, that is your limiting reactant.
  *When you have more than you need, that is your excess reactant.

Example:

- If 4.00 mol C_2H_4 is reacted with 9.50 mol O_2, identify the limiting reactant and calculate the moles of excess reactant remaining.

  \[C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)\]

  *Since we already have moles, we know what we have, we just have to find what we need.

  \[
  \left(\frac{4.00 \text{ mol C}_2\text{H}_4}{1}\right) \cdot \left(\frac{3 \text{ mol O}_2}{1 \text{ mol C}_2\text{H}_4}\right) = 12.0 \text{ mol O}_2 \quad \text{Amount we need}
  \]

  **Limiting Reactant is O_2**

  \[
  \left(\frac{9.50 \text{ mol O}_2}{1}\right) \cdot \left(\frac{1 \text{ mol C}_2\text{H}_4}{3 \text{ mol O}_2}\right) = 3.17 \text{ mol C}_2\text{H}_4 \quad \text{Amount we need} \quad 4.00-3.17 = 0.83 \text{ moles}
  \]
Using a Limiting Reactant to Find Quantity of a Product

- What is the maximum number of grams of Cu₂S that can be formed when 45.0g Cu reacts with 24.0g S?

\[ 2\text{Cu(s)} + \text{S(s)} \rightarrow \text{Cu}_2\text{S(s)} \]

Find the limiting reactant and then use the moles of limiting reactant to calculate moles and then grams of Cu₂S.

\[
\left( \frac{45.0 \text{ g Cu}}{1} \right) \cdot \left( \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \right) = 0.708 \text{ mol Cu} \quad \text{Amount we have}
\]

\[
\left( \frac{24.0 \text{ g S}}{1} \right) \cdot \left( \frac{1 \text{ mol S}}{32.07 \text{ g S}} \right) = 0.748 \text{ mol S} \quad \text{Amount we have}
\]

\[
\left( \frac{0.708 \text{ mol Cu}}{1} \right) \cdot \left( \frac{1 \text{ mol S}}{2 \text{ mol Cu}} \right) = 0.354 \text{ mol S} \quad \text{Amount we need}
\]

\[
\left( \frac{0.748 \text{ mol S}}{1} \right) \cdot \left( \frac{2 \text{ mol Cu}}{1 \text{ mol S}} \right) = 1.50 \text{ mol Cu} \quad \text{Amount we need}
\]

Limiting Reactant is Copper, so we use the given amount of Copper to find grams of Cu₂S. We already have converted to moles in the first part, so start with moles.

\[
\left( \frac{0.708 \text{ mol Cu}}{1} \right) \cdot \left( \frac{1 \text{ mol Cu}_2\text{S}}{2 \text{ mol Cu}} \right) \cdot \left( \frac{159.17 \text{ g Cu}_2\text{S}}{1 \text{ mol Cu}_2\text{S}} \right) = 56.3 \text{ g Cu}_2\text{S}
\]

Example

- Calculate the moles of I₂ produced when 80.0g I₂O₅ reacts with 28.0g CO.

\[ \text{I}_2\text{O}_5(\text{g}) + 5\text{CO}(\text{g}) \rightarrow 5\text{CO}_2(\text{g}) + \text{I}_2(\text{g}) \]

\[
\left( \frac{80.0 \text{ g I}_2\text{O}_5}{1} \right) \cdot \left( \frac{1 \text{ mol I}_2\text{O}_5}{333.80 \text{ g I}_2\text{O}_5} \right) = 0.240 \text{ mol I}_2\text{O}_5 \quad \text{Amount we have}
\]

\[
\left( \frac{28.0 \text{ g CO}}{1} \right) \cdot \left( \frac{1 \text{ mol CO}}{28.01 \text{ g CO}} \right) = 1.00 \text{ mol CO} \quad \text{Amount we have}
\]

\[
\left( \frac{0.240 \text{ mol I}_2\text{O}_5}{1} \right) \cdot \left( \frac{5 \text{ mol CO}}{1 \text{ mol I}_2\text{O}_5} \right) = 1.2 \text{ mol CO} \quad \text{Amount we need}
\]

\[
\left( \frac{1.00 \text{ mol CO}}{1} \right) \cdot \left( \frac{1 \text{ mol I}_2\text{O}_5}{5 \text{ mol CO}} \right) = 0.200 \text{ mol I}_2\text{O}_5 \quad \text{Amount we need}
\]

Limiting Reactant is CO, so we use the given amount of CO to find moles of I₂. We already have converted to moles in the first part, so start with moles.

\[
\left( \frac{1.00 \text{ mol CO}}{1} \right) \cdot \left( \frac{1 \text{ mol I}_2}{5 \text{ mol CO}} \right) = 0.200 \text{ mol I}_2
\]

Homework: page 321 #22-23
Section 9.3 – Limiting Reactants and Percentage Yield

Standard 3.: Students know how to calculate percent yield in a chemical reaction.

Language Objective: We will calculate the percent yield in a chemical reaction.

Content Objective: I will show through writing how to calculate the percent yield in a chemical reaction.

Percent Yield

- **Theoretical Yield**: the maximum amount of product that could be formed from given amounts of reactants.
- **Actual Yield**: the amount of product that actually forms when the reaction is carried out in the laboratory.
- **Percent Yield**: the ratio of the actual yield to the theoretical yield expressed as a percent.

\[
\text{Percent Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%
\]

The percent yield is a measure of the efficiency of a reaction carried out in the laboratory.

Factors that can cause the actual yield to be less than the theoretical yield:

- Reactions do not always go to completion.
- Side reactions may occur.
- Poor laboratory technique.

Calculating Theoretical Yield

- What is the theoretical yield in grams of CaO if 24.8g CaCO₃ is heated?
  \[\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g)\]
  Calculate the theoretical yield by converting the mass of the reactant to the mass of the product, using molar mass and mole ratios.

\[
\left(\frac{24.8 \text{ g CaCO}_3}{1}\right) \cdot \left(\frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3}\right) \cdot \left(\frac{1 \text{ mol CaO}}{1 \text{ mol CaCO}_3}\right) \cdot \left(\frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}}\right) = 13.9 \text{ g CaO}
\]

Example

- What is the theoretical yield of AlCl₃ in moles when 3.00 mol Al reacts with an excess of Cl₂?
  \[2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3\]

\[
\left(\frac{3.00 \text{ mol Al}}{1}\right) \cdot \left(\frac{2 \text{ mol AlCl}_3}{2 \text{ mol Al}}\right) = 3.00 \text{ mol AlCl}_3
\]
Calculating Percent Yield

- What is the percent yield if 13.1g CaO is actually produced when 24.8g CaCO$_3$ is heated?

\[
\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})
\]

- Find the theoretical yield (13.9g CaO) and plug both the theoretical and actual yields into the percent yield equation.

\[
\text{Percent Yield} = \left(\frac{13.1 \text{ g CaO}}{13.9 \text{ g CaO}}\right) \times 100\% = 94.2\%
\]

Example

- What is the percent yield if 2.85mol AlCl$_3$ is actually produced when 3.00mol Al reacts with an excess of Cl$_2$?

\[
2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3
\]

\[
\text{Percent Yield} = \left(\frac{2.85 \text{ mol AlCl}_3}{13.00 \text{ mol AlCl}_3}\right) \times 100\% = 95.0\%
\]

Another Way To Look At It

- Think of percent yield as a grade on a test.
- The Theoretical Yield is the number of points possible: 70 pts
- The Actual Yield is the number of points you earned: 64 pts
- Percent yield = 64/70 x 100% = 91%

Homework: page 318 #1 & 3