Chapter 4: Chemical and Solution Stoichiometry
(Sections 4.1-4.4)

Reaction Stoichiometry

➢ The coefficients in a balanced chemical equation specify the relative amounts in moles of each of the substances involved in the reaction

$$2 \text{C}_4\text{H}_{10}(g) + 13 \text{O}_2(g) \rightarrow 8 \text{CO}_2(g) + 10 \text{H}_2\text{O}(g)$$

➢ 2 molecules of \(\text{C}_4\text{H}_{10}\) react with 13 molecules of \(\text{O}_2\) to form 8 molecules of \(\text{CO}_2\) and 10 molecules of \(\text{H}_2\text{O}\)

➢ 2 moles of \(\text{C}_4\text{H}_{10}\) react with 13 moles of \(\text{O}_2\) to form 8 moles of \(\text{CO}_2\) and 10 moles of \(\text{H}_2\text{O}\)

**Mole ratio**

\(2 \text{ mol } \text{C}_4\text{H}_{10} : 13 \text{ mol } \text{O}_2 : 8 \text{ mol } \text{CO}_2 : 10 \text{ mol } \text{H}_2\text{O}\)
Predicting Amounts from Stoichiometry

- The amount of any other substance in a chemical reaction can be determined from the amount of just one substance.

- How much CO$_2$ can be made from 22.0 moles of C$_4$H$_{10}$ in the combustion reaction of C$_4$H$_{10}$?

$$2 \text{C}_4\text{H}_{10} (g) + 13 \text{O}_2 (g) \rightarrow 8 \text{CO}_2 (g) + 10 \text{H}_2\text{O} (g)$$

$$\frac{22 \text{ moles} \text{C}_4\text{H}_{10}}{2 \text{ moles} \text{C}_4\text{H}_{10}} \times \frac{8 \text{ moles} \text{CO}_2}{2 \text{ moles} \text{C}_4\text{H}_{10}} = 88 \text{ moles} \text{CO}_2$$

Practice

- According to the following equation, how many moles of water are made in the combustion of 0.10 moles of glucose?

$$\text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{O}_2 \rightarrow 6 \text{CO}_2 + 6 \text{H}_2\text{O}$$

*Answer: 0.60 mol H$_2$O*
Stoichiometry and Chemical Reactions

The most common stoichiometric problem will present you with a certain *mass of a reactant* and then ask the amount or *mass of product* that can be formed.

- This is called **mass-to-mass** stoichiometry problem

Predicting Amounts from Stoichiometry – *Cont.*

- What if the mass of a substance is given, how do we know how much of another substance is needed (reactant) or is produced (product)?

  For the balanced equation:
  
  \[ a \text{ A} + b \text{ B} \rightarrow c \text{ C} + d \text{ D} \]

1. You cannot convert mass (g) of one substance directly to mass (g) of another substance in a given reaction.

   \[ \text{Grams of A} \rightarrow \text{X} \rightarrow \text{Grams of B} \]

2. However, you can convert mass to moles, then use their mole ratio to convert moles to grams of another substance.
Predicting Amounts from Stoichiometry – Cont.
Mass-to-Mass Conversions

Case 1: **Known mass** of one of the reactants, calculate the mass of product(s)

1. *Balance* the chemical equation.
2. Convert the known *mass* of a chemical species *to moles* using MM as a conversion factor.
3. Use the ratio of the appropriate equation coefficients (i.e. *mole* or “stoichiometric” *ratio*) to convert *moles* of one species *to moles* of another species.
4. Finally, use the MM of the latter species to convert *moles to mass.*
Example: Estimate the mass of $\text{CO}_2$ produced in 2007 by the combustion of $3.5 \times 10^{15}$ g gasoline.

- Assuming that gasoline is octane, $\text{C}_8\text{H}_{18}$, the equation for the reaction is:

$$2 \text{C}_8\text{H}_{18} (l) + 25 \text{O}_2 (g) \rightarrow 16 \text{CO}_2 (g) + 18 \text{H}_2\text{O} (g)$$

- Since we cannot convert mass of A directly to mass of B, we follow the process:

1. Use MM
2. Use mole ratio in balanced equation
3. Use MM

<table>
<thead>
<tr>
<th>Given:</th>
<th>Find:</th>
</tr>
</thead>
<tbody>
<tr>
<td>3.4 $\times$ 10$^{15}$ g $\text{C}<em>8\text{H}</em>{18}$</td>
<td>g $\text{CO}_2$</td>
</tr>
</tbody>
</table>

**Conceptual Plan:**

<table>
<thead>
<tr>
<th>1 mol</th>
<th>16 mol $\text{CO}_2$</th>
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</thead>
<tbody>
<tr>
<td>114.22 g</td>
<td>44.01 g</td>
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</table>

**Relationships:**

$$2 \text{C}_8\text{H}_{18} (l) + 25 \text{O}_2 (g) \rightarrow 16 \text{CO}_2 (g) + 18 \text{H}_2\text{O} (g)$$

- 1 mol $\text{C}_8\text{H}_{18}$ = 114.22 g;
- 1 mol $\text{CO}_2$ = 44.01 g;
- 2 mol $\text{C}_8\text{H}_{18}$: 16 mol $\text{CO}_2$
Solution:

\[
\begin{align*}
3.5 \times 10^8 \text{ g } C_8H_{18} & \times \frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}} = 3.0643 \times 10^{13} \text{ mol } C_8H_{18} \\
3.0643 \times 10^{13} \text{ mol } C_8H_{18} & \times \frac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}} = 2.4514 \times 10^{14} \text{ mol } CO_2 \\
2.4514 \times 10^{14} \text{ mol } CO_2 & \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 1.0789 \times 10^{16} \text{ g } CO_2 \\
& = 1.1 \times 10^{16} \text{ g } CO_2
\end{align*}
\]

Alternate Solution: One step with multiple conversion factors

\[
\begin{align*}
\text{Given:} & \quad 3.4 \times 10^{15} \text{ g } C_8H_{18} \\
\text{Find:} & \quad \text{g } CO_2
\end{align*}
\]

\[
\begin{align*}
\text{Conceptual Plan:} & \quad \text{g } C_8H_{18} \rightarrow \text{mol } C_8H_{18} \rightarrow \text{mol } CO_2 \rightarrow \text{g } CO_2 \\
\text{Relationships:} & \quad \text{MM } C_8H_{18} = 114.22 \text{ g/mol} \quad \text{MM } CO_2 = 44.01 \text{ g/mol} \quad 2 \text{ mol } C_8H_{18} : 16 \text{ mol } CO_2 \text{ (fr. the balanced eq'n)}
\end{align*}
\]

\[
\begin{align*}
\text{Solution:} & \quad 3.5 \times 10^8 \text{ g } C_8H_{18} \times \frac{1 \text{ mol } C_8H_{18}}{114.22 \text{ g } C_8H_{18}} \times \frac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 1.1 \times 10^{16} \text{ g } CO_2
\end{align*}
\]
Your turn: (1) How many grams of chlorine gas can be liberated from the decomposition of 64.0 g of AuCl₃ by the reaction:

\[ 2 \text{AuCl}_3 (s) \rightarrow 2 \text{Au} (s) + 3 \text{Cl}_2 (g) \]

Answer: 22.4 g Cl₂

Relationships:

\[
\begin{align*}
1 \text{ mol AuCl}_3 & \rightarrow \frac{303.329 \text{ g AuCl}_3}{1 \text{ mol AuCl}_3} \\
3 \text{ mol Cl}_2 & \rightarrow \frac{70.906 \text{ g Cl}_2}{2 \text{ mol AuCl}_3} \quad \frac{70.906 \text{ g Cl}_2}{1 \text{ mol Cl}_2}
\end{align*}
\]

http://dbhs.wvusd.k12.ca.us/webdocs/Stoichiometry/Mass-Mass.html

Limiting reactant, theoretical yield, and percent yield
Limiting Reactant and Theoretical Yield

- In real life, we don’t use the exact mole ratio of reactants as shown in the balanced equation.

- In practice, an excess of one reactant is used for two reasons:
  1. To drive the reaction to completion
  2. To maximize the yield of products

- The reactant that is added in excess amount is called the excess reactant, while the one in lower or limiting amount is called the limiting reactant or limiting reagent.

Limiting Reactant and Theoretical Yield - Cont.

NOTE: The limiting reactant is:
  1. completely consumed in a chemical reaction. Thus, it
  2. determines (or limits) the amount of product formed.

- The maximum amount of product that can form when all of the limiting reactant is used up is called the theoretical yield.

Calculating theoretical yield:
  1. Calculate yield assuming the first reactant is limiting.
  2. Calculate yield assuming the 2nd reactant is limiting.
  3. Choose the smaller of the two amounts.

  The reactant that produces the smaller yield is the limiting reactant.
Calculating Theoretical Yield - Cont.

**Example:** 31.84 g of aluminum and 73.15 g of sulfur are combined to form aluminum sulfide according to the equation:

\[
\text{Al} \,(s) \quad + \quad \text{S} \,(s) \quad \rightarrow \quad \text{Al}_2\text{S}_3 \,(s)
\]

(a) Balance the equation.

(b) Determine the limiting reactant.

(c) Calculate the theoretical yield of \( \text{Al}_2\text{S}_3 \) in grams.

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**Theoretical vs. Actual Yield**

The amount of product *actually obtained* is called the **actual yield**.

Actual yield < Theoretical yield  
*WHY?*

(1) Not all the reactants may react
(2) Presence of significant side reactions
(3) Physical recovery of 100% of the sample may be impossible  
- like getting all the peanut butter out of the jar

**To calculate percent yield:**

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100
\]
More Mass-Mass Stoichiometry Problems

Another example: **Limiting reactant calculation**:
Consider the reaction: \[ 2 \text{Al} + 3 \text{I}_2 \rightarrow 2 \text{AlI}_3 \]

Determine the limiting reagent and the theoretical yield of the product if one starts with:

- a) 1.20 mol Al and 2.40 mol iodine.
- b) 1.20 g Al and 2.40 g iodine
- c) How many grams of Al are left over in part b?

**Solution:**

a) We already have moles (instead of grams) so we use those numbers directly. To find the limiting reagent, we use their mole ratio from the balanced equation (theor. mole ratio) and compare it with the actual ratio from the given moles.

\[
\begin{align*}
\text{Al} & : \text{I}_2 \\
1.20 \text{ mol Al} & : 2.40 \text{ mol I}_2
\end{align*}
\]

Obviously, 489 g is less than 652, so **Al is the limiting reagent** (lower yield from Al), making **I\(_2\)** the excess reagent.

a) Calculate yield of AlI\(_3\) from 1.20 mol Al and 2.40 mol iodine.

MM Al = 26.982; MM I\(_2\) = 253.80 and MM AlI\(_3\) = 407.68 g/mol

Assume Al is limiting:

\[
1.20 \text{ mol Al} \times \frac{2 \text{ mol AlI}_3}{2 \text{ mol Al}} \times \frac{407.78 \text{ g AlI}_3}{1 \text{ mol AlI}_3} = 489 \text{ g AlI}_3
\]

Now assume I\(_2\) is limiting:

\[
2.40 \text{ mol I}_2 \times \frac{2 \text{ mol AlI}_3}{3 \text{ mol I}_2} \times \frac{407.78 \text{ g AlI}_3}{1 \text{ mol AlI}_3} = 652 \text{ g AlI}_3
\]

Obviously, 489 g is less than 652, so **Al is the limiting reagent** (lower yield from Al), making **I\(_2\)** the excess reagent.
Solution – Cont.
(b) Determine the yield of the $\text{AlI}_3$ if one starts with 1.20 g Al and 2.40 g $\text{I}_2$.  

\[
\text{MM Al} = 26.982; \text{MM I}_2 = 253.80 \text{ and MM AlI}_3 = 407.68 \text{ g/mol}
\]

\[
2 \text{Al} + 3 \text{I}_2 \rightarrow 2 \text{AlI}_3
\]

WORK:

(i) Calc. yield of the $\text{AlI}_3$ assuming Al is limiting

\[
1.20 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.982 \text{ g Al}} \times \frac{2 \text{ mol AlI}_3}{2 \text{ mol Al}} \times \frac{407.68 \text{ g AlI}_3}{1 \text{ mol AlI}_3} = 18.1 \text{ g AlI}_3
\]

(ii) Calc. yield of the $\text{AlI}_3$ assuming $\text{I}_2$ is limiting

\[
2.40 \text{ g I}_2 \times \frac{1 \text{ mol I}_2}{253.80 \text{ g I}_2} \times \frac{2 \text{ mol AlI}_3}{3 \text{ mol I}_2} \times \frac{407.68 \text{ g AlI}_3}{1 \text{ mol AlI}_3} = 2.57 \text{ g AlI}_3
\]

(iii) Pick the lower of two yields because it came from the limiting reactant. This is the theoretical yield of product.

Thus:
- limiting reactant is $\text{I}_2$ and
- theoretical yield of $\text{AlI}_3$ is 2.57 g
Another HOMEWORK problem:
A 2.00 g sample of ammonia is mixed with 4.00 g of oxygen.

\[ \text{NH}_3 \text{(g)} + \text{O}_2 \text{(g)} \rightarrow \text{NO} \text{(g)} + \text{H}_2\text{O} \text{(g)} \]

Which is the limiting reactant? How much NO is produced?

First, balance the equation:

\[ \text{NH}_3 \text{(g)} + \text{O}_2 \text{(g)} \rightarrow \text{NO} \text{(g)} + \text{H}_2\text{O} \text{(g)} \]

Next we can use stoichiometry to calculate how much NO product is produced by each reactant. NOTE: It does not matter which product is chosen, but the same product must be used for both reactants so that the amounts can be compared.

Given:

\[ \begin{align*}
4 \text{NH}_3 \text{(g)} & \rightarrow 5 \text{O}_2 \text{(g)} \\
2.00 \text{g} & \rightarrow 4.00 \text{g}
\end{align*} \]

\[ \begin{align*}
2.00 \text{g} \text{NH}_3 \times \frac{1 \text{ mol NH}_3}{17.0 \text{ g NH}_3} \times \frac{4 \text{ mol NO}}{4 \text{ mol NH}_3} \times \frac{30.0 \text{ g NO}}{1 \text{ mol NO}} &= 3.33 \text{ g NO} \quad \text{from NH}_3 \\
4.00 \text{ g} \text{O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \times \frac{4 \text{ mol NO}}{5 \text{ mol O}_2} \times \frac{30.0 \text{ g NO}}{1 \text{ mol NO}} &= 3.00 \text{ g NO} \quad \text{from O}_2
\end{align*} \]

The reactant that produces the lesser amount of product, in this case the oxygen, is the limiting reactant.