### Reaction Stoichiometry

**Stoichiometry**

Relationships Derived from a Balanced Chemical Equation

**Mole Ratios**

Basic Steps Used in Solving Stoichiometry Problems

**Moles → Moles**

**Moles → Mass**

**Mass → Moles**

**Mass → Mass**
Relationships Derived from a Balanced Chemical Equation

Chemical equations can be interpreted to describe the relationships occurring in the reaction.

Example:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Coefficient</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Iron (Fe)</td>
<td>4</td>
<td>Fe(s)</td>
</tr>
<tr>
<td>Oxygen (O₂)</td>
<td>3</td>
<td>3O₂(g)</td>
</tr>
<tr>
<td>Iron (III) oxide (Fe₂O₃)</td>
<td>2</td>
<td>2Fe₂O₃(s)</td>
</tr>
</tbody>
</table>

4 Fe(s) + 3O₂(g) → 2Fe₂O₃(s)

Stoichiometry

Stoichiometry is the study of quantitative relationships between amounts of reactants used and products formed by a chemical reaction.

In order to do stoichiometric calculations, you must have a balanced chemical equation.

Basic Steps Used in Solving Stoichiometry Problems

1. Write a balanced equation.
2. Convert the given substance to moles.
3. Use the mole ratio from the balanced equation to calculate the moles of the second substance needed to react with the given substance.
4. Convert the moles of the second substance to the appropriate quantity.

Mole Ratios

A mole ratio is a ratio between the numbers of moles of any two substances in a balanced chemical equation.

Ex. What mole ratios can be written relating the moles of aluminum to the moles of bromine based upon the following equation?

2Al(s) + 3Br₂(l) → 2AlBr₃(s)

\[
\begin{align*}
\text{Al} : \text{Br}_2 &= 2 : 3 \\
\text{Al} : \text{AlBr}_3 &= 1 : 1 \\
\text{Br}_2 : \text{AlBr}_3 &= 1 : 1
\end{align*}
\]

Moles → Mass

\[
\text{mass of given (g)} \times \frac{1 \text{ mol of given}}{\text{molar mass of given (g/mol)}} \times \frac{\text{mol ratio (from equation)}}{\text{molar mass of unknown (g/mol)}} = \text{mass of unknown (g)}
\]

Example:

When magnesium burns in air, it combines with oxygen to form magnesium oxide according to the following equation:

2Mg(s) + O₂(g) → 2MgO(s)

What mass in grams of oxygen combines with 2.00 mol of magnesium?

\[
2.00 \text{ mol Mg} \times \frac{1 \text{ mol O₂}}{2 \text{ mol Mg}} \times \frac{32.00 \text{ g O₂}}{1 \text{ mol O₂}} = 32.00 \text{ g O₂}
\]

Moles → Moles

\[
\text{mol of given} \times \frac{1 \text{ mol of unknown}}{\text{molar ratio (from equation) of unknown}} = \text{mol of unknown}
\]

Example: The decomposition of potassium chlorate is used as a source of oxygen in the laboratory. How many moles of potassium chlorate are needed to produce 15 mol of oxygen?

\[
\begin{align*}
2\text{KClO}_3 & \rightarrow 2\text{KCl} + 3\text{O}_2 \\
15 \text{ mol O}_2 & \times \frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2} = 10 \text{ mol KClO}_3
\end{align*}
\]

Mass → Mass

\[
\text{mass of given (g)} \times \frac{1 \text{ mol unknown}}{\text{molar mass of unknown (g/mol)}} \times \frac{\text{mol ratio (from equation) of unknown}}{\text{molar ratio (from equation) of given}} = \text{mass of unknown (g)}
\]

Example:

Tin(II) fluoride, SnF₂, is used in some tooth pastes. It is made by the reaction of tin with hydrogen fluoride according to the following equation.

Sn(s) + 2HF(g) → SnF₂(s) + H₂(g)

How many grams of SnF₂ are produced from the reaction of 30.00 g of HF with Sn?

\[
30.00 \text{ g HF} \times \frac{1 \text{ mol HF}}{20.01 \text{ g HF}} \times \frac{1 \text{ mol SnF}_2}{2 \text{ mol HF}} \times \frac{157.71 \text{ g SnF}_2}{1 \text{ mol SnF}_2} = 117.5 \text{ g SnF}_2
\]

Mass → Moles

\[
\text{mass of given (g)} \times \frac{1 \text{ mol unknown}}{\text{molar mass of unknown (g/mol)}} \times \frac{\text{mol ratio (from equation) of unknown}}{\text{mol ratio (from equation) of given}} = \text{mol of unknown (g)}
\]

Example:

Oxygen was discovered by Joseph Priestley in 1774 when he heated mercury(II) oxide to decompose it to form its constituent elements.

2HgO(s) → 2Hg(l) + O₂(g)

How many moles of HgO are needed to produce 125 g O₂?

\[
125 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \times \frac{2 \text{ mol HgO}}{2 \text{ mol O}_2} = 7.81 \text{ mol HgO}
\]
Reaction Stoichiometry

Volume of a Gas
(at STP)

Volume of a Gas
(Not at STP)

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Reaction Stoichiometry

Volume of a Gas
(Not at STP)

Limiting and Excess Reagents

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Reaction Stoichiometry

Limiting Reagents
(Identifying the Limiting Reagent)

Limiting Reagents
(Determining the Maximum Amount of Product)

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Reaction Stoichiometry

Limiting Reagents
(Calculating the amount of excess)

Percent Yield

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**Volume of a Gas (Not at STP)**

If a gas is not at STP, you will have to use the Ideal Gas Equation.

\[ PV = nRT \]

- \( P \) = pressure
- \( V \) = volume in L
- \( n \) = number of moles
- \( T \) = temperature in Kelvin
- \( R \) = ideal gas constant

At STP, 1 mole of a gas occupies 22.4 L. This information can be used when converting to or from liters.

**Example:**

Water can be decomposed by electrolysis according to the following equation:

\[ 2\text{H}_2\text{O} \xrightarrow{\text{electrolysis}} 2\text{H}_2(g) + \text{O}_2(g) \]

What volume of oxygen is produced when a sample of 29.2 g of water is decomposed by electrolysis according to this balanced equation?

\[ n = \frac{m}{M} \]

\[ n_{\text{H}_2\text{O}} = \frac{29.2 \text{ g H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 1.62 \text{ mol H}_2\text{O} \]

\[ n_{\text{H}_2} = \frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \times 1.62 \text{ mol H}_2\text{O} = 1.62 \text{ mol H}_2 \]

\[ n_{\text{O}_2} = \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2} \times 1.62 \text{ mol H}_2 = 0.81 \text{ mol O}_2 \]

**Limiting and Excess Reagents**

Often an excess of one or more reactants is used in a chemical reaction. The reactant which is not in excess is the limiting reagent, and determines the amount of product formed. The reagent in excess is called the excess reagent.

**Calculating the Maximum Amount of Product**

Hydrogen gas can be produced in the laboratory by the reaction of magnesium metal with hydrochloric acid.

\[ \text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(g) \]

The limiting reagent when 6.00 g of HCl reacts with 5.00 g Mg is HCl (see previous card). What is the maximum amount of hydrogen gas that can be produced?

\[ 0.0165 \text{ mol HCl} \times \frac{1 \text{ mol H}_2}{2 \text{ mol HCl}} \times \frac{2.02 \text{ g H}_2}{1 \text{ mol H}_2} = 0.166 \text{ g H}_2 \]

**Identifying the Limiting Reagent**

Hydrogen gas can be produced in the laboratory by the reaction of magnesium metal with hydrochloric acid.

\[ \text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(g) \]

Identify the limiting reagent when 6.00 g of HCl reacts with 5.00 g Mg.

\[ 6.00 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} = 0.165 \text{ mol HCl} \]

\[ 5.00 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 0.206 \text{ mol Mg} \]

\[ 0.165 \text{ mol HCl} \times \frac{1 \text{ mol Mg}}{2 \text{ mol HCl}} = 0.0825 \text{ mol Mg needed} \]

The HCl is limiting.

**Calculating the Amount of Excess**

Hydrogen gas can be produced in the laboratory by the reaction of magnesium metal with hydrochloric acid.

\[ \text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(g) \]

Calculate the amount of excess Mg when 6.00 g of HCl reacts with 5.00 g Mg. It has been previously determined that 5.00 g Mg is equal to 0.206 mol Mg and that only 0.0825 mol Mg will be used (Card 23).

\[ 0.206 \text{ mol Mg} - 0.0825 \text{ mol Mg} = 0.1235 \text{ mol Mg} \]

\[ 0.1235 \text{ mol Mg} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 3.00 \text{ g Mg} \]
Reaction Stoichiometry

Calculating Percent Yield

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Calculating Percent Yield

When 1.77 g H₂ (molar mass – 2.02 g/mol) reacts with an excess of N₂, 8.85 g NH₃ (molar mass – 17.04 g/mol) is produced. What is the percent yield of the reaction?

N₂(g) + 3H₂(g) → 2NH₃(g)

First calculate the theoretical yield.

\[1.77 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.02 \text{ g H}_2} \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} \times \frac{17.04 \text{ g NH}_3}{1 \text{ mol NH}_3} = 9.95 \text{ g NH}_3\]

Then calculate the percent yield.

Percent Yield = \(\frac{\text{actual}}{\text{theoretical}}\) \times 100 = \(\frac{8.85 \text{ g}}{9.95 \text{ g}}\) \times 100 = 88.9%