Stoichiometry: the numerical relationship between chemical quantities in a balanced chemical equation.

Ex. \[ \text{4 NH}_3 + \text{5 O}_2 \rightarrow \text{4 NO} + \text{6 H}_2\text{O} \]

The reaction above can mean: 4 molecules of \( \text{NH}_3 \) reacts with 5 molecules of \( \text{O}_2 \) to produce 4 molecules of \( \text{NO} \) and 6 molecules of \( \text{H}_2\text{O} \).

It also can be interpreted as: 4 moles of \( \text{NH}_3 \) reacts with 5 moles of \( \text{O}_2 \) to produce 4 moles of \( \text{NO} \) and 6 moles of \( \text{H}_2\text{O} \).

How many moles of water can be produced from 4 moles of \( \text{NH}_3 \) in the chemical reaction above? (assume excess \( \text{O}_2 \))

How many moles of water can be produced from 5.6 moles of \( \text{NH}_3 \) in the chemical reaction above?

To solve this problem we need to use a mole ratio derived from the balanced chemical equation (text uses term "equivalence")

Mole ratios relating amounts of \( \text{NH}_3 \) and \( \text{H}_2\text{O} \) in equation above:

<table>
<thead>
<tr>
<th>4 moles ( \text{NH}_3 )</th>
<th>or</th>
<th>6 moles ( \text{H}_2\text{O} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>6 mole ( \text{H}_2\text{O} )</td>
<td>or</td>
<td>4 moles ( \text{NH}_3 )</td>
</tr>
</tbody>
</table>

?? Moles \( \text{H}_2\text{O} \) = 5.6 moles \( \text{NH}_3 \)
Mole to Mole Conversions

\[ 4 \text{NH}_3 + 5 \text{O}_2 \rightarrow 4 \text{NO} + 6 \text{H}_2\text{O} \]

How many moles of O\(_2\) are needed to produce 7.2 moles of NO? (assume there is excess NH\(_3\))

How many moles of ammonia (NH\(_3\)) is needed for complete reaction with 0.850 moles of oxygen gas?

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**Mass-to-Mass Conversions**

<table>
<thead>
<tr>
<th>Grams A</th>
<th>Moles A</th>
<th>Moles B</th>
<th>Grams B</th>
</tr>
</thead>
<tbody>
<tr>
<td>Conversion factor:</td>
<td>Use molar mass of A</td>
<td>Use mole ratio from the balanced equation</td>
<td>Use molar mass of B</td>
</tr>
</tbody>
</table>

Ex. How many grams of H\(_2\)O\(_2\) are needed to completely react with 50.0 g of N\(_2\)H\(_4\)?

(molar mass N\(_2\)H\(_4\) = 32.05 g/mol; H\(_2\)O\(_2\) = 34.02 g/mol)

\[ \text{N}_2\text{H}_4 (l) + 2 \text{H}_2\text{O}_2(l) \rightarrow \text{N}_2(g) + 4 \text{H}_2\text{O}_2(g) \]
Mass-to-Mass Conversions

![Diagram](image)

\[
\text{Conversion factor:} \quad \text{Use molar mass of A} \quad \text{Use mole ratio from the balanced equation} \quad \text{Use molar mass of B}
\]

\[
\begin{align*}
\text{N}_2\text{H}_4 (l) & \quad + \quad 2 \text{H}_2\text{O}_2(l) \quad \rightarrow \quad \text{N}_2(g) & \quad + \quad 4 \text{H}_2\text{O}(g) \\
\end{align*}
\]

How many grams of H\textsubscript{2}O will be produced if 50.0 g of N\textsubscript{2}H\textsubscript{4} completely react?

How many grams of MgCl\textsubscript{2} will be formed if 1.00 g of HCl is consumed?

\[
\begin{align*}
\text{Mg(OH)}_2 (s) & \quad + \quad 2 \text{HCl}(_{aq}) \quad \rightarrow \quad \text{MgCl}_2(_{aq}) & \quad + \quad 2 \text{H}_2\text{O}(l) \\
\end{align*}
\]
Mass-to-Mass Conversions

How many moles of H$_2$O are needed to completely react with 5.00g of NO$_2$?

$$3 \text{ NO}_2 (g) + \text{ H}_2\text{O}(l) \rightarrow 2 \text{ HNO}_3(l) + \text{ NO}(g)$$

**Limiting Reactant**

Analogy: There are 3 assembly lines in a bicycle factory. On a single day, the 3 assembly lines produced the following quantities:

- Handle bars - 34
- Tires - 62
- Frames - 37

How many bicycles can be made?

**Limiting reactant** - the reactant that is USED UP first (limits the amount of product that can be made).
Suppose you have 500.0 g Pb and 500.0 g H₂S. Which reactant will be used up first? (which is the limiting reactant?)

1) Which reactant is used up first? (which is the limiting reactant?)

and/or

1) How much product can be formed?

Suppose you have 500.0 g Pb and 500.0 g H₂S. Which reactant will be used up first?

\[ \text{Pb} + \text{H}_2\text{S} \rightarrow \text{PbS} + \text{H}_2 \]

To solve this problem: Need to do stoichiometric calculations with BOTH given starting materials and convert to amount of ONE chosen product in both sets of calculations....
Limiting Reactant Calculations

If 100.0 g Fe₂O₃ is reacted with 45.00 g Al:

\[
\text{Fe}_2\text{O}_3(s) + 2 \text{Al}(s) \rightarrow \text{Al}_2\text{O}_3(s) + 2 \text{Fe}(l)
\]

(159.70 g/mol) (26.98 g/mol) (101.96 g/mol) (55.85 g/mol)

a) Which reactant is used up first (is the limiting reactant)?
b) How many grams of iron will be formed?

If 1.00 mole of Cr is reacted with 1.00 mole of O₂, how much Cr₂O₃ can be produced?

\[
4 \text{Cr}(s) + 3 \text{O}_2(g) \rightarrow 2 \text{Cr}_2\text{O}_3(s)
\]
Theoretical Yield & Percent Yield

**Theoretical Yield** - the CALCULATED amount of a product that can be obtained from given amounts of reactants in a chemical reaction.

- It is the **maximum** amount of product that can be formed under ideal conditions.

Ex. \[4 \text{ NH}_3 + 3 \text{ O}_2 \rightarrow 2 \text{ N}_2 + 6 \text{ H}_2\text{O}\]

- Theoretically, starting with 4 moles of NH\(_3\) a maximum of 2 moles N\(_2\) can be produced.

**Actual Yield** - the amount of product **actually** obtained

(Not the actual yield is always smaller than theoretical yield)

- Starting with 4 moles of NH\(_3\), only 1 mole of N\(_2\) was actually obtained when the experiment was carried out. (The actual yield will be given - it is NOT calculated)

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

What is the percent yield of N\(_2\) formed in the reaction above?
Limiting Reactant/Percent Yield Practice

4 NH₃ + 3 O₂ → 2 N₂ + 6 H₂O
(17.03 g/mol) (32.00 g/mol) (28.02 g/mol) (18.02 g/mol)

a) If 115.0 g NH₃ is reacted with 200.0 g of O₂, which is the limiting reactant?

b) What is the theoretical yield of N₂ (in grams)?

c) If only 67.0 g of N₂ is obtained, what is the percent yield?