Percentage Yield

Many chemical reactions are carried out in order to make a desired product. The amount of the desired product that is produced is called the **yield**. If the amount of all the reactants in a chemical reaction are known, it is possible to accurately calculate what the yield of a product should be.

The maximum possible amount of a product that can be made in a chemical reaction is known as the **theoretical yield**.

The actual amount of product that is made in a chemical reaction is known as the **actual yield**.

Calculating the **percentage yield** of a reaction provides an indication of how successful the reaction has been.

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

A high percentage yield may indicate an efficient practical process.

The percentage yield of a reaction is **very** rarely 100%. With some chemical reactions, it would be impossible to achieve 100% yield even with the best practical skills and processes.

The actual yield of a chemical reaction is almost always less than the theoretical yield because:

- Reactants or products may have been 'lost' during a step in the process. Such as when reactants and products are transferred from one piece of equipment to another. For example, during filtration, some solution may be absorbed by the filter paper, therefore reducing the amount of product in the filtrate.
- A chemical reaction may not go to completion, especially if the reaction is reversible.
- Some reactants may form by-products in a side reaction.

Chemists modify reaction conditions and practical processes to maximise the percentage yield of a chemical reaction.

The overall percentage yield of a multi-step reaction is usually low. The overall percentage yield is found by multiplying the percentage yield of each step. So, for a reaction involving three steps where each step has a percentage yield of 85%, the overall percentage yield

$$\frac{85}{100} \times \frac{85}{100} \times \frac{85}{100} \times 100 = 61.425 = 61.4\% \text{ (3 s.f.)}$$

ADVANCED | CHEMISTRY

Examples

1. **Calculate the percentage yield of a reaction when 7.10 g of iron is extracted from 13.3 g of iron(III) oxide.**

$$2Fe_2O_3 + 3C \longrightarrow 4Fe + 3CO_2$$

moles of $Fe_2O_3 = \dfrac{\text{mass of substance (g)}}{A_r \text{ or } M_r \text{ of substance (g mol}^{-1}\text{)}}$

M_r of Fe_2O_3 = (55.8 × 2) + (16.0 × 3) = 159.6

$\dfrac{13.30}{159.6}$ = 0.083333... moles

stoichiometric ratio of Fe_2O_3 : Fe = 2 : 4 = 1 : 2

maximum possible moles of Fe = 0.083333... × 2 = 0.16666... moles

theoretical yield of Fe = number of moles × A_r

A_r of Fe = 55.8

0.16666... × 55.8 = 9.30 g

percentage yield = $\dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100$

$\dfrac{7.10}{9.30}$ × 100 = 76.34408602

= 76.3% (3 s.f.)

ADVANCED | CHEMISTRY

2. **In the Haber process, nitrogen and hydrogen react to make ammonia.**

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

At a pressure of 200 atmospheres and a temperature of 450 °C, 1.00 dm^3 of nitrogen and 3.00 dm^3 of hydrogen typically react to produce 300 cm^3 of ammonia.

a. **Calculate the percentage yield of the reaction under these conditions.**

One mole of any gas occupies a constant volume at a given temperature and pressure.

The stoichiometry of the equation tells us that 1.00 dm^3 of nitrogen and 3.00 dm^3 of hydrogen should produce a theoretical yield of 2.00 dm^3 of ammonia.

$$\frac{300 \text{ cm}^3}{1000} = 0.300 \text{ dm}^3$$

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

$$\frac{0.300}{2.00} \times 100 = 15$$

$$= 15.0\% \text{ (3 s.f.)}$$

b. **Explain why the percentage yield is not 100%.**

The reaction is reversible. Therefore, some of the ammonia will always decompose to form nitrogen and hydrogen.

ADVANCED | CHEMISTRY

3. 8.43 g of magnesium carbonate was heated for five minutes. After cooling, the mass of the remaining mixture was 5.68 g.

$$MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$$

Calculate the percentage of the magnesium carbonate that decomposed.

initial moles of $MgCO_3 = \dfrac{\text{mass of substance (g)}}{A_r \text{ or } M_r \text{ of substance (g mol}^{-1})}$

Mr of $MgCO_3 = (24.3 \times 1) + (12.0 \times 1) + (16.0 \times 3) = 84.3$

$\dfrac{8.43}{84.3} = 0.100$ moles

stoichiometric ratio of $MgCO_3 : CO_2 = 1 : 1$

number of moles of $MgCO_3$ decomposed = number of moles of CO_2 produced

mass of CO_2 = 8.43 g – 5.68 g = 2.75 g

moles of $CO_2 = \dfrac{\text{mass of substance (g)}}{A_r \text{ or } M_r \text{ of substance (g mol}^{-1})}$

Mr of $CO_2 = (12.0 \times 1) + (16.0 \times 2) = 44.0$

2.75/44.0 = 0.0625 moles

percentage decomposition $= \dfrac{\text{actual number of moles decomposed}}{\text{theoretical number of moles decomposed}} \times 100$

$0.0625/0.100 \times 100 = 62.5$

= 62.5%

ADVANCED | CHEMISTRY