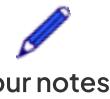




Cambridge (CIE) A Level Chemistry



Your notes

Reacting Masses & Volumes (of Solutions & Gases)

Contents

- * Reacting Masses & Volumes



Mole Calculations

- The number of moles of a substance can be found by using the following equation:

$$\text{number of moles} = \frac{\text{mass of a substance (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

- It is important to be clear about the type of particle you are referring to when dealing with moles
 - E.g. 1 mole of CaF2 contains one mole of CaF2 formula units, but one mole of Ca^{2+} and two moles of F^- ions

Reacting masses

- The **masses** of reactants are useful to determine how much of the reactants **exactly** react with each other to prevent waste
- To calculate the reacting masses, the chemical equation is required
- This equation shows the ratio of moles of all the reactants and products, also called the **stoichiometry**, of the equation
- To find the mass of products formed in a reaction the following pieces of information is needed:
 - The mass of the reactants
 - The molar mass of the reactants
 - The balanced equation

Percentage yield

- In a lot of reactions, not all reactants react to form products which can be due to several factors:
 - Other reactions take place simultaneously
 - The reaction does not go to **completion**
 - Reactants or products are **lost** to the atmosphere
- The **percentage yield** shows how much of a particular product you get from the reactants compared to the maximum theoretical amount that you can get:

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

- Where **actual yield** is the number of moles or mass of product obtained **experimentally**
- The **predicted yield** is the number of moles or mass obtained by calculation



Your notes

Worked Example

Mass calculation using moles

Calculate the mass of magnesium oxide that can be produced by completely burning 6 g of magnesium in oxygen.



Answer

- **Step 1:** The symbol equation is:
$$2\text{Mg (s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO (s)}$$
- **Step 2:** The relative atomic and formula masses are:
 - Magnesium: 24.3
 - Oxygen: 32
 - Magnesium oxide: 40.3
- **Step 3:** Calculate the moles of magnesium used in the reaction:

$$\text{Number of moles} = \frac{6.0 \text{ g}}{24.3 \text{ g mol}^{-1}} = 0.2469 \text{ moles}$$

- **Step 4:** Find the ratio of magnesium to magnesium oxide using the balanced chemical equation:

| | Magnesium | Magnesium oxide |
|-----------------|-----------|-----------------|
| Moles | 2 | 2 |
| Ratio | 1 | 1 |
| Change in moles | - 0.2469 | + 0.2469 |

- Therefore, **0.2469 mol of MgO** is formed
- **Step 5:** Find the mass of magnesium oxide
 - mass = mol $\times M_r$
 - mass = $0.2469 \text{ mol} \times 40.3 \text{ g mol}^{-1}$
 - mass = 9.95 g
- Therefore, the mass of magnesium oxide produced is **9.95 g**



Worked Example

Calculate % yield using moles

In an experiment to displace copper from copper(II) sulfate, 6.54 g of zinc was added to an excess of copper(II) sulfate solution.

The copper was filtered off, washed and dried.

The mass of copper obtained was 4.80 g.



Your notes

Answer

- **Step 1:** The symbol equation is:
$$\text{Zn (s)} + \text{CuSO}_4 \text{ (aq)} \rightarrow \text{ZnSO}_4 \text{ (aq)} + \text{Cu (s)}$$
- **Step 2:** Calculate the amount of zinc reacted in moles

$$\text{Number of moles} = \frac{6.54 \text{ g}}{65.4 \text{ g mol}^{-1}} = 0.10 \text{ moles}$$

- **Step 3:** Calculate the maximum amount of copper that could be formed from the molar ratio:
 - Since the ratio of Zn(s) to Cu(s) is 1:1 a maximum of 0.10 moles can be produced
- **Step 4:** Calculate the maximum mass of copper that could be formed (**theoretical yield**)
 - Mass = mol $\times M_r$
 - Mass = $0.10 \text{ mol} \times 63.5 \text{ g mol}^{-1}$
 - Mass = **6.35 g**
- **Step 5:** Calculate the percentage yield of copper

$$\text{Percentage yield} = \frac{4.80 \text{ g}}{6.35 \text{ g}} \times 100 = 75.6\%$$

Excess & limiting reagents

- Sometimes, there is an **excess** of one or more of the reactants (**excess reagent**)
- The reactant which is not in excess is called the **limiting reagent**
- To determine which reactant is limiting:
 - The number of moles of the reactants should be calculated



Worked Example

Excess & limiting reagent

9.2 g of sodium is reacted with 8.0 g of sulfur to produce sodium sulfide, Na₂S.

Which reactant is in excess and which is the limiting reagent?

Answer

- **Step 1:** Calculate the moles of each reactant
 - Number of moles (Na) = $\frac{9.2 \text{ g}}{23.0 \text{ g mol}^{-1}} = 0.40 \text{ mol}$



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- Number of moles (S) = $\frac{8.0 \text{ g}}{32.1 \text{ g mol}^{-1}} = 0.25 \text{ mol}$
- **Step 2:** Write the balanced equation and determine the molar ratio
 - $2\text{Na} + \text{S} \rightarrow \text{Na}_2\text{S}$
 - The molar ratio of Na: Na₂S is **2:1**
- **Step 3:** Compare the moles and determine the limiting reagent
 - To completely react 0.40 moles of Na requires 0.20 moles of S
 - Since there are 0.25 moles of S, then **S is in excess**
 - Therefore, **Na is the limiting reactant**

Volumes of gases

- Avogadro suggested that:

equal volumes of gases contain the same number of molecules

- This is also called **Avogadro's hypothesis**
- At room temperature and pressure **one mole** of any gas has a volume of 24.0 dm³
 - Room temperature is 20 °C
 - Room pressure is 1 atmosphere
- This **molar gas volume** of 24.0 dm³ can be used to find:
 - The volume of a given mass or number of moles of gas:

$$\text{volume of gas (dm}^3\text{)} = \text{amount of gas (mol)} \times 24.0$$

- The mass or number of moles of a given volume of gas:

$$\text{amount of gas (mol)} = \frac{\text{volume of gas (dm}^3\text{)}}{24.0}$$



Worked Example

Calculating the volume of gas using excess & limiting reagents

Calculate the volume that the following gases occupy:

1. Hydrogen (3 mol)
2. Carbon dioxide (0.25 mol)
3. Oxygen (5.4 mol)
4. Ammonia (0.02 mol)

Calculate the moles in the following volumes of gases:

1. Methane (225.6 dm³)
2. Carbon monoxide (7.2 dm³)
3. Sulfur dioxide (960 dm³)

Answer



Your notes

| Gas | Amount of gas (mol) | Volume of gas (dm ³) |
|-----------------|----------------------------|----------------------------------|
| Hydrogen | 3.0 | $3.0 \times 24.0 = 72.0$ |
| Carbon dioxide | 0.25 | $0.25 \times 24.0 = 6.0$ |
| Oxygen | 5.4 | $5.4 \times 24.0 = 129.6$ |
| Ammonia | 0.02 | $0.02 \times 24.0 = 0.48$ |
| Methane | $\frac{225.6}{24.0} = 9.4$ | 225.6 |
| Carbon monoxide | $\frac{7.2}{24.0} = 0.30$ | 7.2 |
| Sulfur dioxide | $\frac{960}{24.0} = 40$ | 960 |

Volumes & concentrations of solutions

- The **concentration** of a solution is the amount of **solute** dissolved in a **solvent** to make 1 dm³ of **solution**
 - The solute is the substance that dissolves in a solvent to form a solution
 - The solvent is often water

$$\text{concentration (mol dm}^{-3}\text{)} = \frac{\text{number of moles of solute (mol)}}{\text{volume of solution (dm}^3\text{)}}$$

- A **concentrated** solution is a solution that has a **high** concentration of solute
- A **dilute** solution is a solution with a **low** concentration of solute
- When carrying out calculations involving concentrations in mol dm⁻³ the following points need to be considered:
 - Change mass in grams to **moles**
 - Change cm³ to dm³
- To calculate the **mass** of a substance present in solution of known **concentration and volume**:
 - Rearrange the concentration equation

$$\text{number of moles (mol)} = \text{concentration (mol dm}^{-3}\text{)} \times \text{volume (dm}^3\text{)}$$

- Multiply the moles of solute by its molar mass

$$\text{mass of solute (g)} = \text{number of moles (mol)} \times \text{molar mass (g mol}^{-1}\text{)}$$



Worked Example



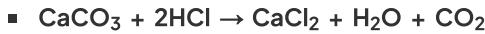
Your notes

Calculating volume from concentration

Calculate the volume of 1.0 mol dm^{-3} hydrochloric acid required to completely react with 2.5 g of calcium carbonate.

Answer

- **Step 1:** Write the balanced symbol equation



- **Step 2:** Calculate the amount, in moles, of calcium carbonate:

$$\text{Number of moles} (\text{CaCO}_3) = \frac{2.5 \text{ g}}{100 \text{ g mol}^{-1}} = 0.025 \text{ mol}$$

- **Step 3:** Calculate the moles of hydrochloric acid required using the reaction's stoichiometry:

- 1 mol of CaCO_3 requires 2 mol of HCl
- So 0.025 mol of CaCO_3 requires **0.05 mol of HCl**

- **Step 4:** Calculate the volume of HCl required:

$$\text{Volume (HCl)} = \frac{\text{amount (mol)}}{\text{concentration (mol dm}^{-3})} = \frac{0.05 \text{ mol}}{1.0 \text{ mol dm}^{-3}} = 0.05 \text{ dm}^3$$

- So, the volume of hydrochloric acid required is **0.05 dm³**



Worked Example

Neutralisation calculation

25.0 cm³ of 0.050 mol dm⁻³ sodium carbonate solution was completely neutralised by 20.0 cm³ of dilute hydrochloric acid.

Calculate the concentration, in mol dm⁻³, of the hydrochloric acid.

Answer

- **Step 1:** Write the balanced symbol equation:



- **Step 2:** Calculate the amount, in moles, of sodium carbonate reacted

- Rearrange the equation for the amount of substance (mol) and divide the volume by 1000 to convert cm³ to dm³

$$\text{Number of moles} (\text{Na}_2\text{CO}_3) = 0.025 \text{ dm}^3 \times 0.050 \text{ mol dm}^{-3} = 0.00125 \text{ mol}$$

- **Step 3:** Calculate the moles of hydrochloric acid required using the reaction's stoichiometry:

- 1 mol of Na_2CO_3 reacts with 2 mol of HCl, so the molar ratio is 1 : 2

- Therefore, 0.00125 moles of Na_2CO_3 react with **0.00250 moles of HCl**

- **Step 4:** Calculate the concentration, in mol dm⁻³, of hydrochloric acid:



Your notes

$$\text{Concentration (HCl)} = \frac{\text{amount (mol)}}{\text{volume (dm}^3\text{)}} = \frac{0.00250}{0.0200} = 0.125 \text{ mol dm}^{-3}$$

Stoichiometric relationships

- The stoichiometry of a reaction can be found if the **exact amounts** of reactants and products formed are known
- The **amounts** can be found by using the following equation:

$$\text{number of moles} = \frac{\text{mass of a substance (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

- The gas volumes can be used to deduce the **stoichiometry** of a reaction
 - E.g. in the **combustion** of 50 cm³ of propane reacting with 250 cm³ of oxygen, 150 cm³ of carbon dioxide is formed suggesting that the ratio of propane: oxygen : carbon dioxide is 1 : 5 : 3

