Topic 3 – Quantitative chemistry

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3.1 Chemical measurements, conservation of mass and the quantitative interpretation of chemical equations

3.1.1 Conservation of mass and balanced chemical equations

Law of conservation of mass

- No atoms are lost or made during a chemical reaction
- So mass of products = mass of reactants

The student's results are shown in the table below.

	Mass in g
Beaker **A** and contents before mixing	127.60
Beaker **B** and contents before mixing	126.86
Beaker **A** and contents after mixing	153.09
Beaker **B** after mixing	101.37

Use the data from the table to show that the law of conservation of mass is true.

[3 marks]

- Total mass before = 127.6 + 126.86 = 254.46
- Total mass after = 153.09 + 101.37 = 243.46
- So mass of products = mass of reactants

A fellow student also tests the law of conservation of mass and decides to use the same method but performs the experiment with a different reaction.

The equation for the reaction is:

$$Na_2CO_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$$

This students' results appear to fail to support the law of conservation of mass.

Explain why this is so.

[3 marks]

- CO_2 is a gas
- Which escapes during the reaction
- So the mass at the end of the experiment is less expected as mass has been lost

Explain why an unbalanced chemical equation cannot correctly describe a chemical reaction (2)

- Must end up with the same no of atoms as at the start
- Otherwise matter is shown to be lost / gained
- Won't show correct amount of each element / compound

3.1.2 Relative formula mass

Relative formula mass (Mr)

Calculate the relative formula mass (M_r) of the compound lead nitrate $Pb(NO_3)_2$

Relative atomic masses (A_r): N = 14; O = 16; Pb = 207

[2 marks]

- 207 + 2 × [14 + (3 × 16)] = 331

In a balanced chemical equation

- Sum of Mr of reactants in the quantities = sum of Mr of products in the quantities

Why is there a change in mass?

- A reactant or product is a gas & its mass has not been taken into account

Percentage composition

$$Percentage\ (\%) = \frac{mass\ or\ Mr\ of\ element}{mass\ or\ Mr\ of\ compound} \times 100\%$$

How much Fe is there in $FeSO_4$?

Mr(Fe) = 56, Mr($FeSO_4$) = 152

$$Percentage\ (\%) = \frac{Mr\ of\ element}{Mr\ of\ compound} \times 100\% = \frac{56}{152} \times 100\% = 36.8\%$$

3.1.3 Mass changes when a reactant or product is a gas

When a metal reacts with O_2

- Mass of the oxide produced is greater than the mass of the metal
- In thermal decompositions of metal carbonates
  - $CO_2$is produced
  - Which escapes into the atmosphere
  - Leaving the metal oxide as the only solid product

The equation for the reaction is 2HCl(aq) + $CaCO_3$(s) → $CaCl_2$(aq) + H_2O(l) + CO_2 (g). Explain why there is a loss in mass in this investigation (2)

- A gas is produced
- Which escapes from the flask

3.1.4 Chemical measurements

(Do practice questions)

3.2 Use of amount of substance in relation to masses of pure substances

3.2.1 Moles (HT only)

Moles

$$mole\ (mol) = \frac{mass\ (g)}{Mr}$$

Avogadro constant = 6.02×10^{23}
No of molecules = mole × 6.02×10^{23}

Mass of 1 mol of a substance
= relative atomic mass (Ar) in grams if substance is an element
= relative formula mass (Mr) in grams if substance is a compound

Reacting masses

Eg If I react 10g of $CaCO_3$ with excess HCl, what mass of $CaCl_2$ should I produce?

$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

$$n(CaCO_3) = \frac{\text{mass (g)}}{\text{Mr}} = \frac{10}{100} = 0.1\text{mol}$$

$$0.1 = \frac{\text{mass (CaCl2)}}{110}$$

mass($CaCl_2$) = 11g

3.2.2 Amounts of substances in equations (HT only)

(Linked to other sub-topics)

3.2.3 Using moles to balance equations (HT only)

(Linked to other sub-topics)

3.2.4 Limiting reactants (HT only)

0 8 . 7 In **Stage 2**, 40 kg of titanium chloride was added to 20 kg of sodium.

The equation for the reaction is:

$$TiCl_4 + 4\,Na \rightarrow Ti + 4\,NaCl$$

Relative atomic masses (A_r): Na = 23 Cl = 35.5 Ti = 48

Explain why titanium chloride is the limiting reactant.

You **must** show your working.

[4 marks]

Mr of TiCl4 = 190

$$n(Na) = \frac{\text{mass (g)}}{\text{Mr}} = \frac{20000}{23} = 870\text{mol}$$

$$n(TiCl4) = \frac{\text{mass (g)}}{\text{Mr}} = \frac{40000}{190} = 211\text{mol}$$

Na is in excess as n(Na) = 870 mol is more than 844 mol needed or n($TiCl_4$) = 211 mol is less than 217.5 mol needed

3.2.5 Concentration of solutions

(See 3.4 Using concentrations of solutions in mol/dm3 (chemistry only) (HT only))

3.3 Yield and atom economy of chemical reactions (chemistry only)

3.3.1 Percentage yield

Percentage yield

$$Percentage\ yield\ (\%) = \frac{actual\ mass\ of\ product}{\text{max}\ mass\ of\ product} \times 100\%$$

For a Stage 2 reaction the percentage yield was 92.3% The theoretical maximum mass of titanium produced in this batch was 13.5 kg. Calculate the actual mass of titanium produced. (2)

$$Percentage\ yield\ (\%) = \frac{actual\ mass\ of\ product}{\text{max}\ mass\ of\ product} \times 100\%$$

$$92.3\% = \frac{actual\ mass\ of\ titanium}{13.5} \times 100\%$$

$$Actual\ mass\ of\ titanium = 12.5kg$$

Factors that affect % yield

- Not all reactant is reacted
- Some product is lost when separated from reaction mixture
- Unexpected reaction

3.3.2 Atom economy

$$Atom\ economy = \frac{Mr\ of\ desired\ product}{Total\ Mr\ of\ all\ reactants} \times 100\%$$

0 2 . 5 An equation for the reaction is:

NiO + C → Ni + CO

Calculate the percentage atom economy for the reaction to produce nickel.

Relative atomic masses (A_r): C = 12 Ni = 59

Relative formula mass (M_r): NiO = 75

Give your answer to 3 significant figures.

[3 marks]

$$Atom\ economy = \frac{Mr\ of\ desired\ product}{Total\ Mr\ of\ all\ reactants} \times 100\% = \frac{59}{87} \times 100\% = 67.8\%$$

Why is it important for percentage of atom economy of a reaction to be as high as possible (2)

- Important for sustainable development
- Economic reasons
- Waste products may be pollutants

Different results – additional product is made e.g. CO_2, H_2O

3.4 Using concentrations of solutions in mol/dm3 (chemistry only) (HT only)

$$mole = concentration \times volume$$
$$n = cv$$

n = number of moles (mol)
c = concentration of solution (mol/dm^{-3})
v = volume of solution (dm^{-3})

$1dm^3 = 1000cm^3$

A student titrated 25.0 cm^3 portions of dilute sulfuric acid with a 0.105 mol/dm^3 sodium hydroxide solution.

0 9 . 3 **Table 4** shows the student's results.

Table 4

	Titration 1	Titration 2	Titration 3	Titration 4	Titration 5
Volume of sodium hydroxide solution in cm^3	23.50	21.10	22.10	22.15	22.15

The equation for the reaction is:

$$2\,NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2\,H_2O$$

Calculate the concentration of the sulfuric acid in mol/dm^3

Use only the student's concordant results.

Concordant results are those within 0.10 cm^3 of each other.

[5 marks]

$$Average\ titre = \frac{22.10+22.15+22.15}{3} = 22.13cm^3 = 0.02213dm^{-3}$$

$$n(NaOH) = cv = 0.105 \times 0.02213 = 0.002324$$

$$n(H2SO4) = \frac{1}{2} \times 0.002324 = 0.001162$$

$$0.001162 = c(H2SO4) \times 0.025 = 0.0465mol/dm^{-3}$$

3.5 Use of amount of substance in relation to volumes of gases (chemistry only) (HT only)

1 mole of gas = 24 dm^3 1 dm^3 = 1000 cm^3 1 m^3 = 1000 cm^3

$$moles = \frac{volume\ (dm^3)}{24\ dm^3}$$

A helium balloon has a volume of 48000 cm^3. Calculate the moles of helium in the balloon. (2)

$$n(He) = \frac{volume\ (dm^3)}{24\ dm^3} = \frac{48}{24} = 2mol$$

All the formulae

$$mole\ (mol) = \frac{mass\ (g)}{Mr}$$

$$moles = \frac{volume\ (dm^3)}{24\ dm^3}$$

$$Percentage\ (\%) = \frac{mass\ or\ Mr\ of\ element}{mass\ or\ Mr\ of\ compound} \times 100\%$$

$$mole = concentration \times volume$$
$$n = cv$$

$$Percentage\ yield\ (\%) = \frac{actual\ mass\ of\ product}{\max\ mass\ of\ product} \times 100\%$$

$$Atom\ economy = \frac{Mr\ of\ desired\ product}{Total\ Mr\ of\ all\ reactants} \times 100\%$$