

3.2 Moles and reacting masses

Name: _____ Class: _____ Date: _____

Total: 31 marks**KEY WORDS** mole • Avogadro constant • limiting reactant

Objective

Convert mass and moles, calculate reacting masses, balance from masses, use the limiting reactant and concentration. HT. Original questions for spec statements 4.3.2.1–4.3.2.5.

1 Key knowledge and worked examples

Moles = mass (g) $\div$ Mr. Mass of one mole in grams = Mr. The **Avogadro constant** = 6.02×10^{23} particles per mole.

Equation $\rightarrow$ ratio $\rightarrow$ mass chain:

1. write the balanced equation;
2. convert the given mass to moles ($\div$ Mr);
3. multiply by the equation ratio;
4. convert back to mass ($\times$ Mr).

Balance from masses: each mass $\div$ Mr $\rightarrow$ divide all by the smallest $\rightarrow$ clear fractions.

Limiting reactant: the one fully used up; moles of product are limited by it. The other is in **excess**.

Concentration: $\text{g/dm}^3 = \text{mass} \div \text{volume}(\text{dm}^3)$; (HT) **mol/dm³:** moles = $C \times V(\text{dm}^3)$.

Worked example (4 marks). What mass of magnesium oxide forms when 2.4 g of magnesium burns in excess oxygen? (Mg = 24, O = 16.)

$2 \text{ Mg} + \text{O}_2 \rightarrow 2 \text{ MgO}$; moles Mg = $2.4 \div 24 = 0.1$; ratio 2:2, so moles MgO = 0.1; mass = $0.1 \times 40 = 4.0 \text{ g}$.

2 Practice

2.1 How many moles are in 8.0 g of sodium hydroxide, NaOH (Na = 23, O = 16, H = 1)? [2]

2.2 What mass is 0.25 mol of CaCO_3 ($M_r = 100$)? [1]

2.3 0.2 mol of HCl is mixed with 0.2 mol of NaOH . Why does neither remain in excess? [1]

3 Exam-style questions

3.1 Calculate the mass of carbon dioxide produced when 10.0 g of calcium carbonate decomposes: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ ($\text{Ca} = 40$, $\text{C} = 12$, $\text{O} = 16$). [4]

3.2 6.0 g of magnesium reacts with 4.0 g of oxygen: $2 \text{Mg} + \text{O}_2 \rightarrow 2 \text{MgO}$. Identify the limiting reactant, if any, and calculate the mass of MgO formed ($\text{Mg} = 24$, $\text{O} = 16$). [5]

3.3 (*HT*) 250 cm^3 of sodium chloride solution contains 3.5 g of solute. Calculate the concentration in g/dm^3 and in mol/dm^3 ($\text{Na} = 23$, $\text{Cl} = 35.5$). [4]

3.4 A reaction uses 4.8 g Mg and 6.4 g O₂. Use $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$, Mg = 24 and O = 16. Identify the limiting reactant, calculate the mass of MgO and the mass of oxygen remaining. [6]

3.5 11.2 g iron reacts with 4.8 g oxygen to make 16.0 g iron(III) oxide. Use Fe = 56 and O = 16; reactant formulae Fe and O₂, product Fe₂O₃. Convert all three masses to moles and deduce the balanced equation. [4]

3.6 A sample contains 0.25 mol O₂ molecules. Use 6.02×10^{23} particles per mole and O = 16. Calculate the number of O₂ molecules, the amount and number of oxygen atoms, and the sample mass. [4]

Solutions

2.1 $M_r(\text{NaOH}) = 40$; moles $= 8.0 \div 40 = 0.2$.

2.2 $0.25 \times 100 = 25 \text{ g}$.

2.3 The equation ratio $\text{NaOH} : \text{HCl}$ is $1 : 1$, and the moles are equal, so both are fully used.

3.1 $M_r(\text{CaCO}_3) = 100$, moles $= 10.0 \div 100 = 0.10$; ratio $1 : 1$, so 0.10 mol CO_2 ; $M_r(\text{CO}_2) = 44$; mass $= 0.10 \times 44 = 4.4 \text{ g}$.

3.2 Moles $\text{Mg} = 6.0 \div 24 = 0.25$; moles $\text{O}_2 = 4.0 \div 32 = 0.125$; ratio needs 2 Mg per 1 O_2 , so 0.125 mol O_2 requires 0.25 mol Mg – exactly what is present: neither is limiting / both fully react; moles $\text{MgO} = 0.25$; mass $= 0.25 \times 40 = 10.0 \text{ g}$.

3.3 g/dm^3 : 3.5 g in $0.250 \text{ dm}^3 \rightarrow 3.5 \div 0.25 = 14 \text{ g/dm}^3$; $M_r(\text{NaCl}) = 58.5$; $\text{mol/dm}^3 = 14 \div 58.5 = 0.24 \text{ mol/dm}^3$.

3.4 $n(\text{Mg}) = 4.8/24 = 0.20 \text{ mol}$; $n(\text{O}_2) = 6.4/32 = 0.20 \text{ mol}$. The $2:1$ equation ratio means Mg is limiting: 0.20 mol Mg needs only 0.10 mol O_2 . MgO amount $= 0.20 \text{ mol}$, so mass $= 0.20 \times 40 = 8.0 \text{ g}$. O_2 left $= (0.20 - 0.10) \times 32 = 3.2 \text{ g}$.

3.5 Amounts: $\text{Fe} = 11.2/56 = 0.200 \text{ mol}$; $\text{O}_2 = 4.8/32 = 0.150 \text{ mol}$; $\text{Fe}_2\text{O}_3 = 16.0/160 = 0.100 \text{ mol}$. Ratio $2:1.5:1$ becomes $4:3:2$, hence $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$; all atoms and total mass balance.

3.6 O_2 molecules $= 0.25 \times 6.02 \times 10^{23} = 1.505 \times 10^{23}$. Each molecule contains two O atoms: amount of O atoms $= 2 \times 0.25 = 0.50 \text{ mol}$, atom count $= 0.50 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$. O_2 molar mass $= 32 \text{ g/mol}$, so sample mass $= 0.25 \times 32 = 8.0 \text{ g}$. Accept consistent sensible rounding.