

Moles, formulae and equations: AS Chemistry practice questions with full answers

Cambridge International AS Level Chemistry 9701, topic: atoms, molecules and stoichiometry. Use these values throughout: A_r : H 1.0, C 12.0, N 14.0, O 16.0, Na 23.0, Mg 24.3, S 32.1, Cl 35.5, K 39.1, Ca 40.1, Fe 55.8, Cu 63.5. Avogadro constant $L = 6.02 \times 10^{23} \text{ mol}^{-1}$. Molar gas volume at room temperature and pressure (r.t.p.) = $24.0 \text{ dm}^3 \text{ mol}^{-1}$. Give answers to three significant figures unless the data justify fewer.

What you need to know

Relative masses

Relative atomic mass, A_r : the weighted mean mass of one atom of an element compared with one twelfth of the mass of one atom of carbon-12.

Relative molecular mass, M_r : the weighted mean mass of one molecule compared with one twelfth of the mass of one atom of carbon-12. For ionic compounds the same quantity is called the **relative formula mass**.

M_r is found by adding the A_r values of every atom in the formula. Water of crystallisation counts.

Worked example. M_r of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} = 63.5 + 32.1 + (4 \times 16.0) + 5 \times (2 \times 1.0 + 16.0) = 63.5 + 32.1 + 64.0 + 90.0 = 249.6$.

The mole and the Avogadro constant

One mole is the amount of substance that contains 6.02×10^{23} particles (atoms, molecules, ions or electrons). This number is the **Avogadro constant, L** .

Number of particles = amount in moles $\times L$. Be clear which particle you are counting: one mole of H_2O is one mole of molecules but three moles of atoms.

Worked example. Number of oxygen atoms in 0.250 mol of CO_2 : molecules = $0.250 \times 6.02 \times 10^{23} = 1.505 \times 10^{23}$. Each molecule has two O atoms, so O atoms = $2 \times 1.505 \times 10^{23} = 3.01 \times 10^{23}$.

Molar mass

The **molar mass, M** , is the mass of one mole of a substance in g mol^{-1} ; it is numerically equal to A_r or M_r . The key relationship is $n = m / M$, where n is the amount in mol and m is the mass in g.

Worked example. Amount of CaCO_3 in 5.00 g: $M = 40.1 + 12.0 + 48.0 = 100.1 \text{ g mol}^{-1}$; $n = 5.00 / 100.1 = 0.0500 \text{ mol}$ (3 s.f.).



Empirical and molecular formulae

Empirical formula: the simplest whole-number ratio of atoms of each element in a compound.

Molecular formula: the actual number of atoms of each element in one molecule of the compound. It is a whole-number multiple of the empirical formula.

Method: divide the mass (or percentage by mass) of each element by its A_r ; divide every result by the smallest; multiply up if the ratio contains a clear fraction such as 1.5. For the molecular formula, divide M_r by the empirical formula mass and multiply the empirical formula by that factor.

Worked example. A compound contains 40.0% C, 6.7% H and 53.3% O by mass, $M_r = 60.0$. C: $40.0 / 12.0 = 3.33$; H: $6.7 / 1.0 = 6.7$; O: $53.3 / 16.0 = 3.33$. Divide by 3.33: C 1 : H 2 : O 1, so the empirical formula is CH_2O (mass 30.0). $60.0 / 30.0 = 2$, so the molecular formula is $\text{C}_2\text{H}_4\text{O}_2$.

Balancing equations, including ionic equations

Every element must have the same number of atoms on each side. Balance by changing the numbers in front of formulae, never the subscripts. For an ionic equation, write soluble ionic compounds and strong acids as separate aqueous ions, then cancel the **spectator ions** (those unchanged on both sides). Charge must balance as well as atoms. Include state symbols: (s), (l), (g), (aq).

Worked example. Magnesium with aqueous copper(II) sulfate. Full equation: $\text{Mg(s)} + \text{CuSO}_4\text{(aq)} \rightarrow \text{MgSO}_4\text{(aq)} + \text{Cu(s)}$. Written in ions: $\text{Mg(s)} + \text{Cu}^{2+}\text{(aq)} + \text{SO}_4^{2-}\text{(aq)} \rightarrow \text{Mg}^{2+}\text{(aq)} + \text{SO}_4^{2-}\text{(aq)} + \text{Cu(s)}$. Cancel the sulfate spectator ions: $\text{Mg(s)} + \text{Cu}^{2+}\text{(aq)} \rightarrow \text{Mg}^{2+}\text{(aq)} + \text{Cu(s)}$. Check: charge 2+ on each side.

Reacting masses

Three steps every time: convert the given mass to moles, use the balanced equation to find the moles of the substance asked for, convert back to mass.

Worked example. Mass of calcium oxide produced when 25.0 g of calcium carbonate is fully decomposed: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$. $n(\text{CaCO}_3) = 25.0 / 100.1 = 0.2498$ mol. Ratio 1 : 1, so $n(\text{CaO}) = 0.2498$ mol. Mass = $0.2498 \times 56.1 = 14.0$ g.

Gas volumes at r.t.p.

One mole of any gas occupies 24.0 dm^3 ($24\,000 \text{ cm}^3$) at r.t.p. So $n = V / 24.0$ with V in dm^3 . Because the volume does not depend on which gas it is, the volume ratio of gases in a reaction equals the mole ratio in the equation.

Worked example. Volume of hydrogen at r.t.p. from 0.486 g of magnesium in excess acid: $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$. $n(\text{Mg}) = 0.486 / 24.3 = 0.0200$ mol; $n(\text{H}_2) = 0.0200$ mol; $V = 0.0200 \times 24.0 = 0.480 \text{ dm}^3 = 480 \text{ cm}^3$.

Concentrations and titrations

Concentration c is measured in mol dm^{-3} : $n = c \times V$, with V in dm^3 (divide cm^3 by 1000).

Concentration in $\text{g dm}^{-3} = c \times M$. In a titration, find the moles of the solution whose concentration and

volume are both known, use the equation to get the moles of the other reagent, then divide by its volume.

Worked example. 25.0 cm³ of sodium hydroxide needs 22.50 cm³ of 0.100 mol dm⁻³ hydrochloric acid. $n(\text{HCl}) = 0.100 \times 22.50 / 1000 = 2.25 \times 10^{-3}$ mol. $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$, ratio 1 : 1, so $n(\text{NaOH}) = 2.25 \times 10^{-3}$ mol. $c(\text{NaOH}) = 2.25 \times 10^{-3} / 0.0250 = 0.0900$ mol dm⁻³.

Percentage yield and atom economy

Percentage yield = (actual mass of product obtained / theoretical maximum mass of product) × 100. The theoretical mass is calculated from the limiting reagent.

Atom economy = (M_r of the desired product / sum of M_r of all products, each multiplied by its coefficient) × 100. It is a property of the equation, not of the experiment, and can also be written as M_r of desired product over total M_r of reactants.

Yields fall below 100% because reactions are reversible or incomplete, side reactions occur, and product is lost during separation and purification. Addition reactions have 100% atom economy; substitution and elimination reactions form by-products.

Worked example. 10.0 g of ethanol reacts with excess ethanoic acid to give 12.5 g of ethyl ethanoate: $\text{C}_2\text{H}_5\text{OH} + \text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O}$. $n(\text{ethanol}) = 10.0 / 46.0 = 0.2174$ mol, so theoretical $n(\text{ester}) = 0.2174$ mol and theoretical mass = $0.2174 \times 88.0 = 19.13$ g. Percentage yield = $12.5 / 19.13 \times 100 = 65.3\%$. Atom economy of the ester: $88.0 / (88.0 + 18.0) \times 100 = 83.0\%$.

Questions

- Calculate the relative formula mass of (a) H_2SO_4 , (b) $\text{Ca}(\text{OH})_2$, (c) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. [3]
- Calculate the amount, in mol, in (a) 8.00 g of sodium hydroxide, NaOH , (b) 12.6 g of iron, (c) 2.50 g of calcium carbonate, CaCO_3 . [3]
- Calculate the mass of (a) 0.350 mol of magnesium chloride, MgCl_2 , (b) 2.50×10^{-3} mol of potassium nitrate, KNO_3 , (c) 1.20×10^{22} molecules of carbon dioxide. [4]
- Calculate (a) the number of molecules in 0.400 mol of ammonia, NH_3 , (b) the number of hydrogen atoms in 0.400 mol of ammonia, (c) the total number of ions in 5.85 g of sodium chloride. [3]
- Balance the following equations. (a) $\text{C}_3\text{H}_8 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$; (b) $\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$; (c) $\text{KOH} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + \text{H}_2\text{O}$; (d) $\text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2$. [4]
- Write an ionic equation, with state symbols, for each reaction. (a) Magnesium ribbon placed in aqueous copper(II) sulfate. (b) Aqueous potassium hydroxide neutralised by dilute hydrochloric acid. (c) Aqueous calcium chloride mixed with aqueous sodium carbonate, giving a white precipitate of calcium carbonate. [3]
- Two fertilisers are ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$, and ammonium nitrate, NH_4NO_3 . Calculate the percentage by mass of nitrogen in each and state which supplies more nitrogen per gram of fertiliser. [3]

8. Iron is extracted from iron(III) oxide in the blast furnace: $\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$. Calculate the maximum mass of iron that can be obtained from 1.00 kg of iron(III) oxide. [3]
9. Determine the empirical formula of (a) a compound of magnesium and nitrogen containing 72.2% magnesium by mass, (b) a compound containing 31.9% potassium, 29.0% chlorine and 39.1% oxygen by mass. [5]
10. A gaseous hydrocarbon contains 85.7% carbon by mass and has a relative molecular mass of 56.0. Determine its empirical formula and its molecular formula. [4]
11. A 4.99 g sample of hydrated copper(II) sulfate, $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$, was heated to constant mass. The mass of the white anhydrous solid remaining was 3.19 g. Calculate the value of x. [4]
12. 0.460 g of an organic compound containing carbon, hydrogen and oxygen only was burned completely in oxygen. 0.880 g of carbon dioxide and 0.540 g of water were formed. (a) Calculate the mass of carbon, hydrogen and oxygen in the sample. (b) Determine the empirical formula. (c) The relative molecular mass of the compound is 46.0. Deduce its molecular formula. [5]
13. (a) 5.30 g of sodium carbonate, Na_2CO_3 , is dissolved in water and made up to 250 cm^3 . Calculate the concentration in mol dm^{-3} . (b) Calculate the mass of potassium hydroxide needed to prepare 500 cm^3 of a 0.150 mol dm^{-3} solution. (c) Express the concentration of the sodium carbonate solution in (a) in g dm^{-3} . [4]
14. In a titration, 25.0 cm^3 of 0.100 mol dm^{-3} sodium hydroxide was exactly neutralised by 20.00 cm^3 of dilute sulfuric acid. Calculate the concentration of the sulfuric acid in mol dm^{-3} . [3]
15. 1.50 g of impure sodium carbonate was dissolved in water and the solution made up to 250 cm^3 . 25.0 cm^3 portions of this solution were titrated with 0.105 mol dm^{-3} hydrochloric acid using methyl orange indicator. The mean titre was 24.60 cm^3 . The impurities do not react with the acid. Calculate the percentage purity of the sodium carbonate. [5]
16. 5.00 g of magnesium ribbon is added to 100 cm^3 of 2.00 mol dm^{-3} hydrochloric acid. (a) Write the balanced equation. (b) Show by calculation which reagent is in excess. (c) Calculate the volume of hydrogen formed at r.t.p. (d) Calculate the mass of magnesium chloride formed. [5]
17. 10.0 g of iron filings and 10.0 g of powdered sulfur are mixed and heated strongly: $\text{Fe} + \text{S} \rightarrow \text{FeS}$. (a) Identify the limiting reagent. (b) Calculate the maximum mass of iron(II) sulfide that can form. (c) Calculate the mass of the reagent left unreacted. [4]
18. Calculate the volume of carbon dioxide, measured at r.t.p., produced when 10.0 g of calcium carbonate reacts completely with an excess of dilute hydrochloric acid. [3]
19. Propane burns completely in oxygen according to $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$. Calculate the volume of oxygen needed to burn 500 cm^3 of propane completely and the volume of carbon dioxide formed, both measured at r.t.p. Explain why you do not need the molar gas volume to answer this. [3]
20. 0.560 g of a gas occupies 480 cm^3 at r.t.p. Calculate the relative molecular mass of the gas and suggest two possible identities for it. [3]
21. A student prepared hydrated copper(II) sulfate by reacting 4.00 g of copper(II) oxide with an excess of warm dilute sulfuric acid, filtering, and crystallising the solution: $\text{CuO} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}$, followed by $\text{CuSO}_4 + 5\text{H}_2\text{O} \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. The student obtained 10.2 g of blue crystals. Calculate the percentage yield. [4]
22. Ethanol can be made by the direct hydration of ethene, $\text{C}_2\text{H}_4 + \text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_5\text{OH}$, or by fermentation of glucose, $\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow 2\text{C}_2\text{H}_5\text{OH} + 2\text{CO}_2$. (a) Calculate the atom economy of

- each route. (b) Suggest why a manufacturer might still choose the route with the lower atom economy. [4]
23. Ammonia is manufactured by the Haber process: $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$. (a) State the atom economy of this reaction and explain your answer. (b) When 56.0 g of nitrogen reacts with an excess of hydrogen, 20.4 g of ammonia is collected. Calculate the percentage yield. (c) Suggest one reason why the yield is low. [4]
24. An indigestion tablet of mass 1.25 g contains calcium carbonate mixed with an inert binder. The tablet was crushed and added to 50.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid, which is an excess. After the reaction was complete, the unreacted acid required 28.40 cm³ of 1.00 mol dm⁻³ sodium hydroxide for neutralisation. (a) Write equations for the reaction of the acid with calcium carbonate and with sodium hydroxide. (b) Calculate the amount of hydrochloric acid added to the tablet. (c) Calculate the amount of hydrochloric acid left unreacted. (d) Calculate the amount of hydrochloric acid that reacted with the calcium carbonate and hence the amount of calcium carbonate in the tablet. (e) Calculate the percentage by mass of calcium carbonate in the tablet. (f) Calculate the volume of carbon dioxide, at r.t.p., released by the tablet. [9]
25. 2.46 g of hydrated magnesium sulfate, $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$, was heated until no further change in mass occurred. The loss in mass was 1.26 g. (a) Calculate the value of x. (b) The anhydrous residue was dissolved in water and an excess of aqueous sodium hydroxide was added, precipitating magnesium hydroxide. Write the ionic equation for the precipitation and calculate the maximum mass of magnesium hydroxide that could form. (c) The dried precipitate was heated strongly: $\text{Mg}(\text{OH})_2 \rightarrow \text{MgO} + \text{H}_2\text{O}$. Calculate the maximum mass of magnesium oxide obtainable. (d) The student actually obtained 0.350 g of magnesium oxide. Calculate the percentage yield over steps (b) and (c) combined. [9]

Answers

- (a) H_2SO_4 : $(2 \times 1.0) + 32.1 + (4 \times 16.0) = 98.1$. (b) $\text{Ca}(\text{OH})_2$: $40.1 + 2 \times (16.0 + 1.0) = 74.1$. (c) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$: $63.5 + 32.1 + 64.0 + (5 \times 18.0) = 249.6$.
- (a) $M(\text{NaOH}) = 23.0 + 16.0 + 1.0 = 40.0 \text{ g mol}^{-1}$; $n = 8.00 / 40.0 = 0.200 \text{ mol}$. (b) $n = 12.6 / 55.8 = 0.226 \text{ mol}$. (c) $M(\text{CaCO}_3) = 100.1 \text{ g mol}^{-1}$; $n = 2.50 / 100.1 = 0.0250 \text{ mol}$.
- (a) $M(\text{MgCl}_2) = 24.3 + 71.0 = 95.3 \text{ g mol}^{-1}$; $m = 0.350 \times 95.3 = 33.4 \text{ g}$. (b) $M(\text{KNO}_3) = 39.1 + 14.0 + 48.0 = 101.1 \text{ g mol}^{-1}$; $m = 2.50 \times 10^{-3} \times 101.1 = 0.253 \text{ g}$. (c) $n = 1.20 \times 10^{22} / 6.02 \times 10^{23} = 0.0199 \text{ mol}$; $m = 0.0199 \times 44.0 = 0.877 \text{ g}$.
- (a) $0.400 \times 6.02 \times 10^{23} = 2.41 \times 10^{23}$ molecules. (b) Each NH_3 has 3 H atoms: $3 \times 2.408 \times 10^{23} = 7.22 \times 10^{23}$ atoms. (c) $M(\text{NaCl}) = 58.5 \text{ g mol}^{-1}$; $n = 5.85 / 58.5 = 0.100 \text{ mol}$ of NaCl, which is 0.100 mol of Na^+ and 0.100 mol of Cl^- , 0.200 mol of ions in total; $0.200 \times 6.02 \times 10^{23} = 1.20 \times 10^{23}$ ions.
- (a) $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$. (b) $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$. (c) $2\text{KOH} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$. (d) $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2$.
- (a) $\text{Mg}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Cu}(\text{s})$; sulfate ions are spectators. (b) $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$; K^+ and Cl^- are spectators. (c) $\text{Ca}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s})$; Na^+ and Cl^- are spectators.
- $(\text{NH}_4)_2\text{SO}_4$: $M_r = (2 \times 14.0) + (8 \times 1.0) + 32.1 + (4 \times 16.0) = 132.1$; mass of N = 28.0; % N =

$28.0 / 132.1 \times 100 = 21.2\%$. NH_4NO_3 : $M_r = 28.0 + 4.0 + 48.0 = 80.0$; % N = $28.0 / 80.0 \times 100 = 35.0\%$. Ammonium nitrate supplies more nitrogen per gram.

8. $M(\text{Fe}_2\text{O}_3) = (2 \times 55.8) + (3 \times 16.0) = 159.6 \text{ g mol}^{-1}$. $n(\text{Fe}_2\text{O}_3) = 1000 / 159.6 = 6.266 \text{ mol}$. From the equation $n(\text{Fe}) = 2 \times 6.266 = 12.53 \text{ mol}$. Mass of Fe = $12.53 \times 55.8 = 699 \text{ g}$.

9. (a) In 100 g: Mg 72.2 g, N 27.8 g. Mg: $72.2 / 24.3 = 2.97 \text{ mol}$; N: $27.8 / 14.0 = 1.99 \text{ mol}$. Divide by 1.99: Mg 1.49, N 1.00. Multiply by 2: Mg 3, N 2. Empirical formula Mg_3N_2 . (b) K: $31.9 / 39.1 = 0.816$; Cl: $29.0 / 35.5 = 0.817$; O: $39.1 / 16.0 = 2.44$. Divide by 0.816: K 1.00, Cl 1.00, O 2.99. Empirical formula KClO_3 .

10. Hydrogen = $100 - 85.7 = 14.3\%$. C: $85.7 / 12.0 = 7.14$; H: $14.3 / 1.0 = 14.3$. Divide by 7.14: C 1.00, H 2.00. Empirical formula CH_2 , empirical formula mass = $12.0 + 2.0 = 14.0$. $56.0 / 14.0 = 4$, so the molecular formula is C_4H_8 .

11. $M(\text{CuSO}_4) = 63.5 + 32.1 + 64.0 = 159.6 \text{ g mol}^{-1}$. $n(\text{CuSO}_4) = 3.19 / 159.6 = 0.0200 \text{ mol}$. Mass of water lost = $4.99 - 3.19 = 1.80 \text{ g}$; $n(\text{H}_2\text{O}) = 1.80 / 18.0 = 0.100 \text{ mol}$. $x = 0.100 / 0.0200 = 5.00$, so $x = 5$ and the formula is $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.

12. (a) $n(\text{CO}_2) = 0.880 / 44.0 = 0.0200 \text{ mol}$, so $n(\text{C}) = 0.0200 \text{ mol}$ and mass of C = $0.0200 \times 12.0 = 0.240 \text{ g}$. $n(\text{H}_2\text{O}) = 0.540 / 18.0 = 0.0300 \text{ mol}$, so $n(\text{H}) = 0.0600 \text{ mol}$ and mass of H = 0.0600 g . Mass of O = $0.460 - 0.240 - 0.060 = 0.160 \text{ g}$. (b) $n(\text{O}) = 0.160 / 16.0 = 0.0100 \text{ mol}$. Ratio C : H : O = $0.0200 : 0.0600 : 0.0100 = 2 : 6 : 1$. Empirical formula $\text{C}_2\text{H}_6\text{O}$. (c) Empirical formula mass = $24.0 + 6.0 + 16.0 = 46.0$, equal to M_r , so the molecular formula is also $\text{C}_2\text{H}_6\text{O}$ (ethanol).

13. (a) $M(\text{Na}_2\text{CO}_3) = 46.0 + 12.0 + 48.0 = 106.0 \text{ g mol}^{-1}$; $n = 5.30 / 106.0 = 0.0500 \text{ mol}$; $c = 0.0500 / 0.250 = 0.200 \text{ mol dm}^{-3}$. (b) $n(\text{KOH}) = 0.150 \times 0.500 = 0.0750 \text{ mol}$; $M(\text{KOH}) = 39.1 + 16.0 + 1.0 = 56.1 \text{ g mol}^{-1}$; mass = $0.0750 \times 56.1 = 4.21 \text{ g}$. (c) $0.200 \times 106.0 = 21.2 \text{ g dm}^{-3}$.

14. $n(\text{NaOH}) = 0.100 \times 25.0 / 1000 = 2.50 \times 10^{-3} \text{ mol}$. $2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$, so $n(\text{H}_2\text{SO}_4) = 2.50 \times 10^{-3} / 2 = 1.25 \times 10^{-3} \text{ mol}$. $c = 1.25 \times 10^{-3} / 0.0200 = 0.0625 \text{ mol dm}^{-3}$.

15. $\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\text{O} + \text{CO}_2$. $n(\text{HCl})$ in the titre = $0.105 \times 24.60 / 1000 = 2.583 \times 10^{-3} \text{ mol}$. $n(\text{Na}_2\text{CO}_3)$ in $25.0 \text{ cm}^3 = 2.583 \times 10^{-3} / 2 = 1.292 \times 10^{-3} \text{ mol}$. In the whole 250 cm^3 : $1.292 \times 10^{-3} \times 10 = 0.01292 \text{ mol}$. Mass of pure $\text{Na}_2\text{CO}_3 = 0.01292 \times 106.0 = 1.369 \text{ g}$. Purity = $1.369 / 1.50 \times 100 = 91.3\%$.

16. (a) $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$. (b) $n(\text{Mg}) = 5.00 / 24.3 = 0.206 \text{ mol}$. $n(\text{HCl}) = 2.00 \times 100 / 1000 = 0.200 \text{ mol}$. 0.206 mol of Mg would need 0.412 mol of HCl, but only 0.200 mol is present, so hydrochloric acid is the limiting reagent and magnesium is in excess. (c) $n(\text{H}_2) = n(\text{HCl}) / 2 = 0.100 \text{ mol}$; $V = 0.100 \times 24.0 = 2.40 \text{ dm}^3$. (d) $n(\text{MgCl}_2) = 0.100 \text{ mol}$; mass = $0.100 \times 95.3 = 9.53 \text{ g}$.

17. (a) $n(\text{Fe}) = 10.0 / 55.8 = 0.179 \text{ mol}$; $n(\text{S}) = 10.0 / 32.1 = 0.312 \text{ mol}$. The ratio is 1 : 1, so iron is the limiting reagent. (b) $n(\text{FeS}) = 0.179 \text{ mol}$; $M(\text{FeS}) = 55.8 + 32.1 = 87.9 \text{ g mol}^{-1}$; mass = $0.1792 \times 87.9 = 15.8 \text{ g}$. (c) Sulfur left = $0.3115 - 0.1792 = 0.1323 \text{ mol}$; mass = $0.1323 \times 32.1 = 4.25 \text{ g}$.

18. $\text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$. $n(\text{CaCO}_3) = 10.0 / 100.1 = 0.0999 \text{ mol} = n(\text{CO}_2)$. $V = 0.0999 \times 24.0 = 2.40 \text{ dm}^3$.

19. From the equation, 1 volume of propane needs 5 volumes of oxygen and gives 3 volumes of carbon dioxide. Oxygen needed = $5 \times 500 = 2500 \text{ cm}^3$; carbon dioxide formed = $3 \times 500 = 1500 \text{ cm}^3$. The molar gas volume is not needed because equal volumes of gases at the same temperature and pressure contain equal numbers of moles, so the mole ratio is also the volume ratio.

20. $n = 0.480 / 24.0 = 0.0200$ mol. $M_r = 0.560 / 0.0200 = 28.0$. Possible gases with M_r 28.0: nitrogen, N_2 ; ethene, C_2H_4 ; carbon monoxide, CO. Any two are acceptable.

21. $M(CuO) = 63.5 + 16.0 = 79.5$ g mol⁻¹; $n(CuO) = 4.00 / 79.5 = 0.0503$ mol. The ratio CuO : CuSO₄·5H₂O is 1 : 1, so theoretical $n(\text{crystals}) = 0.0503$ mol; theoretical mass = $0.05031 \times 249.6 = 12.56$ g. Percentage yield = $10.2 / 12.56 \times 100 = 81.2\%$.

22. (a) Hydration: ethanol is the only product, so atom economy = $46.0 / 46.0 \times 100 = 100\%$. Fermentation: $M_r(\text{glucose}) = 72.0 + 12.0 + 96.0 = 180.0$; mass of desired product = $2 \times 46.0 = 92.0$; atom economy = $92.0 / 180.0 \times 100 = 51.1\%$. (b) Any two of: fermentation uses a renewable feedstock (sugar from crops) rather than ethene from crude oil; it runs at low temperature and pressure, so it needs less energy and cheaper equipment; the carbon dioxide by-product can be collected and sold.

23. (a) 100%, because ammonia is the only product, so all the atoms of the reactants end up in the desired product. (b) $n(N_2) = 56.0 / 28.0 = 2.00$ mol. From the equation, theoretical $n(NH_3) = 2 \times 2.00 = 4.00$ mol; $M(NH_3) = 17.0$ g mol⁻¹; theoretical mass = $4.00 \times 17.0 = 68.0$ g. Percentage yield = $20.4 / 68.0 \times 100 = 30.0\%$. (c) The reaction is reversible and the equilibrium lies well short of complete conversion under the conditions used; in practice the gases leave the reactor before equilibrium is reached.

24. (a) $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$; $HCl + NaOH \rightarrow NaCl + H_2O$. (b) $n(HCl)$ added = $1.00 \times 50.0 / 1000 = 0.0500$ mol. (c) $n(NaOH)$ used = $1.00 \times 28.40 / 1000 = 0.0284$ mol; the ratio is 1 : 1, so $n(HCl)$ unreacted = 0.0284 mol. (d) $n(HCl)$ that reacted with $CaCO_3 = 0.0500 - 0.0284 = 0.0216$ mol. From the equation, $n(CaCO_3) = 0.0216 / 2 = 0.0108$ mol. (e) Mass of $CaCO_3 = 0.0108 \times 100.1 = 1.081$ g. Percentage by mass = $1.081 / 1.25 \times 100 = 86.5\%$. (f) $n(CO_2) = n(CaCO_3) = 0.0108$ mol; $V = 0.0108 \times 24.0 = 0.259$ dm³ = 259 cm³.

25. (a) Mass of anhydrous $MgSO_4 = 2.46 - 1.26 = 1.20$ g. $M(MgSO_4) = 24.3 + 32.1 + 64.0 = 120.4$ g mol⁻¹; $n(MgSO_4) = 1.20 / 120.4 = 9.97 \times 10^{-3}$ mol. $n(H_2O) = 1.26 / 18.0 = 0.0700$ mol. $x = 0.0700 / 9.97 \times 10^{-3} = 7.02$, so $x = 7$ and the salt is $MgSO_4 \cdot 7H_2O$. (b) $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$. $n(Mg(OH)_2) = n(Mg^{2+}) = 9.97 \times 10^{-3}$ mol; $M(Mg(OH)_2) = 24.3 + 2 \times 17.0 = 58.3$ g mol⁻¹; maximum mass = $9.967 \times 10^{-3} \times 58.3 = 0.581$ g. (c) $n(MgO) = 9.97 \times 10^{-3}$ mol; $M(MgO) = 24.3 + 16.0 = 40.3$ g mol⁻¹; maximum mass = $9.967 \times 10^{-3} \times 40.3 = 0.402$ g. (d) Percentage yield = $0.350 / 0.4017 \times 100 = 87.1\%$.