Cambridge IGCSE · Chemistry 0620 · Notes

Relative masses and the mole

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Read first: **Formulae and equations**, **Elements, compounds and atomic structure**

What you need to know

By the end of this topic you should be able to:

1. **Describe relative atomic mass,** A_r: the average mass of the isotopes of an element, compared with $\frac{1}{12}$ of the mass of one atom of carbon-12.
2. **Define relative molecular mass,** M_r: the sum of the relative atomic masses of the atoms in the formula. For ionic compounds the same number is called the **relative formula mass** (still written M_r).
3. **Work out reacting masses by simple proportion** (for example, "if 2 g makes 5 g, then 4 g makes 10 g"), without using moles.
4. **State that concentration can be measured in g/dm³ or in mol/dm³.**
5. **Extended only:** state that the **mole** (mol) is the unit of **amount of substance**, and that one mole contains **6.02×10^{23} particles** (atoms, ions or molecules). This number is the **Avogadro constant**.
6. **Extended only:** use amount (mol) = mass (g) ÷ molar mass (g/mol) to find an amount, a mass, a molar mass, an A_r or M_r, and (with the Avogadro constant) a number of particles.
7. **Extended only:** use the **molar gas volume**, 24 dm³ at room temperature and pressure (r.t.p.), in calculations with gases.
8. **Extended only:** calculate reacting masses, **limiting reactants**, volumes of gases at r.t.p., and volumes and concentrations of solutions in g/dm³ and mol/dm³, converting between cm³ and dm³.
9. **Extended only:** use results from a **titration** to calculate the moles of solute, or the concentration or volume of a solution.
10. **Extended only:** calculate **empirical formulae** and **molecular formulae** from data.
11. **Extended only:** calculate **percentage yield**, **percentage composition by mass** and **percentage purity**.

Understand it

Relative atomic mass

Atoms are far too light to weigh in grams, so chemists compare them with a standard: an atom of **carbon-12**, the isotope with 6 protons and 6 neutrons. One twelfth of its mass is the unit. On this scale a hydrogen atom is about 1, a carbon-12 atom exactly 12 and a magnesium atom about 24.

Most elements are a mixture of isotopes, so the **relative atomic mass,** A_r, is an **average** of their masses, weighted by how common each one is, compared with $\frac{1}{12}$ of the mass of a carbon-12 atom. That is why chlorine's A_r is 35.5 and not a whole number: chlorine is about 75% chlorine-35 and 25% chlorine-37 (see the note on atomic structure for isotopes).

$$A_r(\mathrm{Cl}) = \frac{(35 \times 75) + (37 \times 25)}{100} = 35.5$$

A_r has **no units**, because it compares one mass with another. You never need to remember A_r values: they are printed on the Periodic Table in the exam.

Relative molecular mass and relative formula mass

Add up the A_r of **every atom** in the formula. For a molecule this is the **relative molecular mass**; for an ionic compound it is the **relative formula mass**. Both are written M_r and both have no units.

Substance	Working	M_r
water, H_2O	$(2 \times 1) + 16$	18
carbon dioxide, CO_2	$12 + (2 \times 16)$	44
calcium carbonate, $CaCO_3$	$40 + 12 + (3 \times 16)$	100
magnesium hydroxide, $Mg(OH)_2$	$24 + 2 \times (16 + 1)$	58
ammonium sulfate, $(NH_4)_2SO_4$	$2 \times (14 + 4 \times 1) + 32 + (4 \times 16)$	132

A number after a bracket multiplies **everything** inside it: $Mg(OH)_2$ has 2 O and 2 H.

Reacting masses by simple proportion

In a reaction, the masses that react are always in the same ratio. If you know one pair of masses, you can scale it. If 0.24 g of magnesium burns to make 0.40 g of magnesium oxide, then 1.20 g of magnesium (5 times as much) makes $5 \times 0.40 = 2.00$ g of magnesium oxide. This works for masses of gases, volumes of gases made, and any units (g, kg, tonnes), as long as both amounts are in the same units.

The mole

Extended only. Counting atoms one at a time is impossible, so chemists count them in huge batches. One batch is a **mole**: **6.02 × 10^{23} particles**. This number is the **Avogadro constant**. The mole is the unit of **amount of substance**, written **mol**.

The number was chosen so that **one mole of a substance has a mass in grams equal to its A_r or M_r**. This mass is the **molar mass**, in g/mol:

- 1 mol of carbon atoms has a mass of 12 g and contains 6.02 × 10^{23} atoms.
- 1 mol of water, H_2O, has a mass of 18 g and contains 6.02 × 10^{23} molecules.
- 1 mol of calcium carbonate has a mass of 100 g.

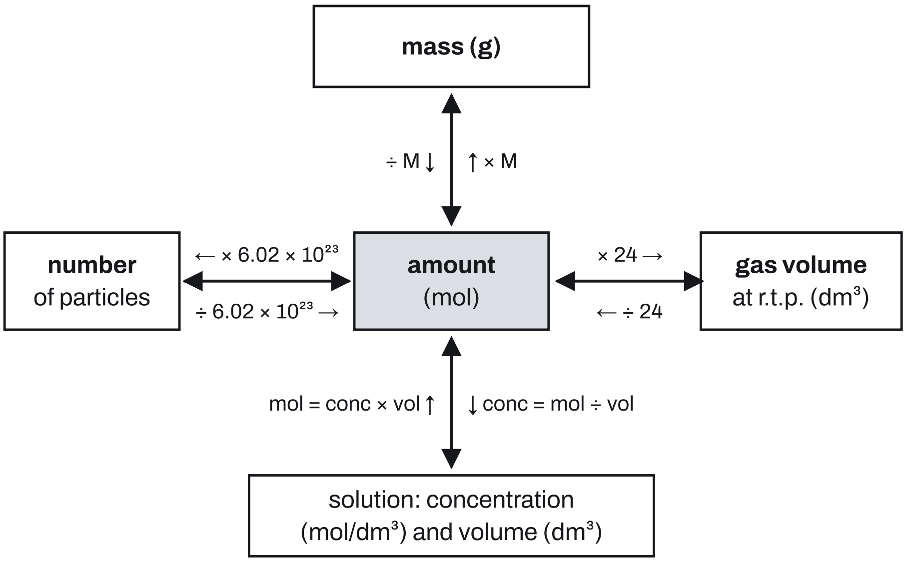
So to change a mass into moles, divide by the molar mass:

$$\text{amount (mol)} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

For example, 11.0 g of carbon dioxide (M_r = 44) is $11.0 \div 44 = 0.250$ mol. Rearranged, mass = amount × molar mass, and molar mass = mass ÷ amount. If 0.0500 mol of an element has a mass of 1.20 g, its molar mass is $1.20 \div 0.0500 = 24$ g/mol, so its A_r is 24: it is magnesium.

Moles are the centre of every calculation

Extended only. Almost every calculation in this topic goes **through moles**. Change what you are given into moles, use the equation to get the moles of the substance you want, then change those moles into whatever the question asks for.



11.0 g CO_2 ÷ 44 = 0.250 mol
= 6.00 dm^3 at r.t.p. = 1.51 × 10^{23} molecules

- **Particles:** number of particles = amount × 6.02 × 10^{23}. So 0.250 mol of CO_2 is $0.250 \times 6.02 \times 10^{23} = 1.51 \times 10^{23}$ molecules. Each molecule has 3 atoms, so that is 4.52×10^{23} atoms.
- **Gases:** at r.t.p. **one mole of any gas occupies 24 dm^3** (24 000 cm^3), whatever the gas. So 0.250 mol of CO_2 occupies $0.250 \times 24 = 6.00$ dm^3. Equal volumes of gases (at the same temperature and pressure) contain equal numbers of molecules.
- **Solutions:** concentration tells you how much solute is dissolved in each dm^3 of solution. In mol/dm^3: amount (mol) = concentration (mol/dm^3) × volume (dm^3).

Volumes must be in dm^3. 1 dm^3 = 1000 cm^3, so divide a volume in cm^3 by 1000: 25.0 cm^3 is 0.0250 dm^3.

Concentration

A concentration is either a **mass** per dm³ (**g/dm³**) or an **amount** per dm³ (**mol/dm³**). Dissolving 2.00 g of sodium hydroxide (M_r = 40) to make 250 cm³ of solution gives:

- in g/dm³: $2.00 \div 0.250 = 8.00$ g/dm³
- in mol/dm³ (**Extended only**): $2.00 \div 40 = 0.0500$ mol, and $0.0500 \div 0.250 = 0.200$ mol/dm³.

To change between them, multiply or divide by the molar mass: $0.200 \times 40 = 8.00$ g/dm³.

Core questions often give a concentration in g/dm³ (or mg/dm³) and ask for the mass in a smaller volume: scale by the volume. If river water contains 36 mg/dm³ of chloride ions, then 250 cm³ (0.250 dm³) contains $36 \times 0.250 = 9.0$ mg.

Using the equation: the mole ratio

Extended only. The numbers in a balanced equation give the ratio of **moles**, not masses. For calcium carbonate decomposing:

$$\mathrm{CaCO_3 \rightarrow CaO + CO_2}$$

1 mol $CaCO_3$ gives 1 mol CaO and 1 mol CO_2. So 25.0 g of $CaCO_3$ is $25.0 \div 100 = 0.250$ mol, which gives 0.250 mol CaO ($0.250 \times 56 = 14.0$ g) and 0.250 mol CO_2 ($0.250 \times 24 = 6.00$ dm³ at r.t.p.). Where the numbers are not 1, scale: in $\mathrm{2Mg + O_2 \rightarrow 2MgO}$, 0.10 mol of Mg needs only 0.050 mol of O_2.

Titrations

Extended only. In a titration you find the exact volume of one solution that reacts with a measured volume of another (the practical method is part of experimental techniques). One of the two solutions has a known concentration, so:

1. moles of the known solution = concentration × volume (dm³);
2. use the equation's ratio to get the moles of the other substance;
3. its concentration = moles ÷ its volume (dm³).

Limiting reactant

Extended only. When two reactants are mixed, one is usually used up first. That one is the **limiting reactant**: it decides how much product forms. The other is **in excess**, and some of it is left over. To find which is limiting, change both into moles and compare with the equation's ratio.

Empirical and molecular formulae from data

Extended only. The **empirical formula** is the simplest whole-number ratio of atoms. To find it from masses or percentages by mass:

1. Divide each mass (or percentage) by the element's A_r: this gives moles of atoms.
2. Divide every answer by the **smallest** one.
3. If a number ends in about .5, .33 or .67, multiply all of them by 2 or 3 to make whole numbers.

A compound contains 40.9% carbon, 4.5% hydrogen and 54.6% oxygen by mass:

	C	H	O
÷ A_r	$40.9 \div 12 = 3.41$	$4.5 \div 1 = 4.5$	$54.6 \div 16 = 3.41$
÷ smallest	1	1.32	1
× 3	3	4	3

The empirical formula is $C_3H_4O_3$, with $M_r = 36 + 4 + 48 = 88$. The **molecular formula** is a whole number of empirical units. If the compound's M_r is 176, then $176 \div 88 = 2$, so the molecular formula is $C_6H_8O_6$.

Percentages

Extended only. Three different percentages come up:

- **Percentage composition by mass** of an element in a compound $= \dfrac{\text{number of atoms} \times A_r}{M_r} \times 100$. In copper(II) sulfate, $CuSO_4$ (M_r = 160), copper is $\frac{64}{160} \times 100 = 40\%$.
- **Percentage yield** $= \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100$. The theoretical yield is the mass the equation says you should get; you usually get less, because some product is lost or the reaction is incomplete.
- **Percentage purity** $= \dfrac{\text{mass of the pure substance}}{\text{mass of the impure sample}} \times 100$. For example, a 5.00 g sample of impure calcium carbonate gives 0.0450 mol of CO_2 with excess acid. That is 0.0450 mol of $CaCO_3$, which is $0.0450 \times 100 = 4.50$ g, so the sample is $\frac{4.50}{5.00} \times 100 = 90.0\%$ pure.

Water of crystallisation

Extended only. Some salt crystals contain water molecules in a fixed ratio, the **water of crystallisation**, shown after a dot: $CuSO_4$•$5H_2O$. Heating drives the water off. Heat to **constant mass** (heat, weigh, and repeat until the mass stops changing) to be sure all the water has gone. To find x in $CuSO_4$•xH_2O, work out the moles of the anhydrous salt and the moles of water, and divide. If 2.50 g of crystals leave 1.60 g of $CuSO_4$ (M_r = 160): $CuSO_4$ is $1.60 \div 160 = 0.0100$ mol, water is $2.50 - 1.60 = 0.90$ g $= 0.90 \div 18 = 0.0500$ mol, and $x = 0.0500 \div 0.0100 = 5$.

Key facts and formulas

The exam paper prints the **Periodic Table**, with the A_r of every element, and under it the **molar gas volume**. Everything else is **Learn it**.

Formula	Status
A_r values of the elements (Periodic Table)	Given
Molar gas volume: 1 mol of any gas occupies 24 dm³ (24 000 cm³) at r.t.p.	Given
A_r: average mass of an element's isotopes compared with $\frac{1}{12}$ of the mass of a carbon-12 atom	Learn it
M_r = sum of the A_r of all the atoms in the formula	Learn it
Concentration units: g/dm³ and mol/dm³; 1 dm³ = 1000 cm³	Learn it
Extended only: 1 mol = 6.02 × 10^{23} particles (the Avogadro constant)	Learn it
Extended only: $\text{amount (mol)} = \dfrac{\text{mass (g)}}{\text{molar mass (g/mol)}}$	Learn it
Extended only: number of particles = amount × 6.02 × 10^{23}	Learn it
Extended only: volume of gas (dm³) = amount × 24	Learn it
Extended only: amount (mol) = concentration (mol/dm³) × volume (dm³)	Learn it
Extended only: concentration (g/dm³) = concentration (mol/dm³) × M_r	Learn it
Extended only: percentage composition $= \dfrac{\text{number of atoms} \times A_r}{M_r} \times 100$	Learn it
Extended only: percentage yield $= \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100$	Learn it
Extended only: percentage purity $= \dfrac{\text{mass of pure substance}}{\text{mass of sample}} \times 100$	Learn it

How to do it

In the questions we have tagged for this topic (182 questions, 2021–2025, on Papers 1–4), the types came up this often. 24 of the 182 questions mix types, so the counts add up to more than 182.

Question type	Questions
Moles and reacting masses	41
Concentrations and titrations	38
Relative formula mass	35
Gas volumes at r.t.p.	27
Empirical and molecular formulae	18
Number of particles	14
Percentage yield, composition and purity	12
Definitions and units	9
Water of crystallisation	8

Question type	Questions
Limiting reactant	5
Relative atomic mass from isotopes	3

1. Moles and reacting masses (41 questions)

Extended method, for "which mass of ... is formed?" or a structured "use the following steps" question:

1. **Moles of what you know** = mass ÷ M_r.
2. **Moles of what you want**, using the numbers in the equation (the mole ratio).
3. **Mass of what you want** = moles × M_r.

For example, magnesium burning, $2\mathrm{Mg} + \mathrm{O_2} \rightarrow 2\mathrm{MgO}$: 12 g of Mg is $12 \div 24 = 0.50$ mol, which makes 0.50 mol of MgO, which is $0.50 \times 40 = 20$ g.

Variations you will meet:

- **Core: simple proportion.** "When 0.24 g of Mg reacts, 0.40 g of MgO forms. Calculate the mass of MgO from 0.60 g." Find the scale factor ($0.60 \div 0.24 = 2.5$) and multiply: $2.5 \times 0.40 = 1.00$ g. No moles needed.
- **Masses in kg or tonnes.** The ratio is the same in any unit, so work in tonnes throughout. For $\mathrm{N_2} + 3\mathrm{H_2} \rightarrow 2\mathrm{NH_3}$, 34 tonnes of NH_3 (2 × 17) needs 28 tonnes of N_2.
- **"Which statement is correct?"** Each option makes a claim about masses. Turn each mass into moles and check it against the equation; only one survives.
- **A mass from a given number of moles**: mass = moles × M_r (0.250 mol of KNO_3, M_r = 101, is 25.3 g).
- **An unknown A_r from reacting masses**: find the moles of the part you know, use the ratio to get the moles of the unknown, then A_r = mass ÷ moles.

2. Concentrations and titrations (38 questions)

Extended method for a titration:

1. **Moles of the solution you know** = concentration × (volume in cm³ ÷ 1000).
2. **Moles of the other substance**, from the equation's ratio (watch for 2NaOH : 1 H_2SO_4).
3. **Concentration** = moles ÷ (its volume in cm³ ÷ 1000), in mol/dm³.
4. If asked for **g/dm³**, multiply the mol/dm³ by M_r.

Variations you will meet:

- **Core: mass of an ion in a sample.** A table gives concentrations in mg/dm³ or g/dm³; multiply by the volume in dm³.
- **The mass needed to make a solution**: moles = concentration × volume, then mass = moles × M_r. 250 cm³ of 0.40 mol/dm³ NaOH needs $0.40 \times 0.250 = 0.10$ mol, which is 4.0 g.

- **The volume of a solution** that contains a given number of moles: volume (dm^3) = moles ÷ concentration; multiply by 1000 for cm^3.
- **The units of concentration**: mol/dm^3 or g/dm^3.
- **The order of the steps**, or **what extra information you need**: you always need the balanced equation (the mole ratio) to go from one substance to the other.
- **A solution reacting with a solid or making a gas**: moles from concentration × volume, then carry on as in types 1 or 4.

3. Relative formula mass (35 questions)

1. Count every atom of each element (multiply everything inside a bracket by the number after it).
2. Multiply each count by the A_r and add.

Variations you will meet:

- **Complete the table** (Paper 3): fill in "number of atoms × A_r" for each element, then add the column. Show each product.
- **Multiple choice** with options that come from common slips: forgetting the bracket, or adding A_r values once each. $Ba(OH)_2$ is $137 + 2 \times 17 = 171$, not $137 + 17 = 154$.
- **Find an unknown element**: subtract the known atoms from the M_r and divide by the number of unknown atoms. If X_2O has M_r = 62, then $2 \times A_r(\mathrm{X}) = 62 - 16 = 46$, so $A_r = 23$: sodium.

4. Gas volumes at r.t.p. (27 questions)

Extended only.

1. Moles of the substance you know (mass ÷ M_r, or concentration × volume).
2. Moles of gas, from the equation's ratio.
3. Volume = moles × 24 dm^3, or moles × 24 000 cm^3. **Use the unit the question asks for.**

Variations you will meet:

- **From a volume to a mass**: moles of gas = volume ÷ 24 (dm^3) or ÷ 24 000 (cm^3), then use the ratio and × M_r.
- **Several gaseous products**: add their moles before multiplying by 24. For $\mathrm{NH_4HCO_3} \rightarrow \mathrm{NH_3} + \mathrm{CO_2} + \mathrm{H_2O(l)}$ at r.t.p., 1 mol gives 2 mol of gas, because the water is a liquid.
- **Gases reacting with gases**: volumes react in the same ratio as moles, so 50 cm^3 of methane burns in $2 \times 50 = 100$ cm^3 of oxygen.
- **An alkane from the CO_2 it makes**: 1 mol of alkane giving 96 dm^3 of CO_2 gives $96 \div 24 = 4$ mol of CO_2, so it has 4 carbon atoms: butane.
- **Core: proportion with measured volumes**: "0.10 g of $CaCO_3$ gives 24 cm^3 of CO_2; which mass gives 60 cm^3?" Scale: $0.10 \times 60 \div 24 = 0.25$ g.

5. Empirical and molecular formulae (18 questions)

Extended only.

1. Divide each percentage (or mass) by the A_r, not the atomic number.
2. Divide all by the smallest value.
3. Multiply by 2 or 3 if you get a .5, .33 or .67; never round 1.5 to 1 or 2.
4. For the molecular formula: $M_r \div M_r$ of the empirical formula = the multiplier.

Variations you will meet:

- **Only two percentages given** for a C, H, O compound: the oxygen is 100 minus the others (52.2% C and 13.0% H leaves 34.8% O).
- **Masses in grams instead of percentages**: exactly the same method.
- **Moles of atoms given** (such as X 0.25, Y 0.75, Z 0.50): skip step 1 and divide by the smallest: XY_3Z_2.
- **A hydrocarbon from its % carbon**: hydrogen is the rest. 80.0% C gives $6.67 : 20.0 = 1 : 3$, so CH_3.
- **Molecular formula from an empirical formula and M_r**, with no percentages: C_2H_5 has M_r 29; if the compound's M_r is 58, it is C_4H_{10}.

6. Number of particles (14 questions)

Extended only.

1. Moles of the substance (mass ÷ M_r, or volume ÷ 24).
2. Multiply by 6.02 × 10^{23} for **molecules** (or formula units).
3. For **atoms** or **ions**, multiply by how many there are in one formula.

Variations you will meet:

- **Atoms in a molecule**: 1 mol of NH_3 has 4 mol of atoms, so 2.00 mol of NH_3 has $8 \times 6.02 \times 10^{23}$ atoms.
- **Ions in an ionic compound**: Na_2O has 3 ions per formula (2 Na^+ and 1 O^{2-}), so 0.10 mol of Na_2O contains 0.30 mol of ions, 1.81×10^{23}.
- **"Which sample contains the most molecules?"** Change each into moles; the largest number of moles wins.
- **Give the answer in standard form**, such as 3.01×10^{22}, not 30 100 000 000 000 000 000 000.

7. Percentage yield, composition and purity (12 questions)

Extended only.

1. Decide which percentage the question wants.
2. For yield, first work out the **theoretical** mass by the reacting-mass method, then actual ÷ theoretical × 100.
3. For composition, number of atoms × $A_r \div M_r$ × 100.

Variations you will meet:

- **Mass of an element in a mass of compound**: in 100 g of $CaCO_3$ (M_r = 100) there is 12 g of carbon. In general, mass of element = mass of compound × (number of atoms × A_r) ÷ M_r.
- **Which compound has the greatest proportion** of an element: work out the percentage in each.
- **Percentage yield when a reactant is limiting**: base the theoretical yield on the limiting reactant.

8. Definitions and units (9 questions)

- A_r: the **average** mass of the **isotopes** of an element compared with $\frac{1}{12}$ **of the mass of a carbon-12 atom**. The standard is carbon-12, not hydrogen or oxygen.
- M_r: the **sum of the relative atomic masses** (not the atomic numbers) in the formula.
- **Mole**: the unit of amount of substance; 6.02×10^{23} particles. The mass of 6.02×10^{23} atoms of an isotope is its mass number in grams (for carbon-12, exactly 12 g).
- **Concentration**: in g/dm^3 or mol/dm^3.

9. Water of crystallisation (8 questions)

Extended only.

1. Moles of anhydrous salt = mass left ÷ its M_r.
2. Mass of water = mass before – mass after; moles of water = mass ÷ 18.
3. x = moles of water ÷ moles of anhydrous salt (a whole number).

Variations you will meet:

- **Why heat to constant mass?** To make sure all the water of crystallisation has been removed.
- **The moles of water are given**: skip step 2.
- **The name**: the water in the crystals is "water of crystallisation"; the crystals are **hydrated**, and after heating the salt is **anhydrous**.

10. Limiting reactant (5 questions)

Extended only.

1. Change the amount of each reactant into moles.
2. Divide each by its number in the equation. The smaller answer is the **limiting reactant**.
3. Use the limiting reactant (never the one in excess) to work out the product.

Variations you will meet:

- **Extra of one reactant makes no more product**: if 24 g of magnesium burns in 16 g of oxygen to make 40 g of magnesium oxide, using 20 g of oxygen still makes 40 g, because magnesium is the limiting reactant.
- **A solid added to a small amount of acid**: work out the moles of both; often the acid runs out.

- **Explain your answer**: give both mole numbers and compare them using the ratio.

11. Relative atomic mass from isotopes (3 questions)

Extended only. This builds on the note on atomic structure.

1. Multiply each isotope's mass number by its percentage abundance.
2. Add, and divide by 100.

Variation: find an abundance from the A_r. Gallium has isotopes of mass 69 and 71 and A_r = 69.8. If x% is gallium-69, $69x + 71(100 - x) = 6980$, so $x = 60\%$.

Worked examples

The bold codes show where marks are earned, as in Cambridge mark schemes. The number is how many marks.

- **M1** method mark: for a correct method, even if the numbers slip.
- **A1** accuracy mark: for a correct answer; it needs the M mark before it.
- **B1** mark for a correct result on its own, with no method needed.

Example 1

(a) Calculate the relative formula mass of potassium carbonate, K_2CO_3. **[1]**

(b) When 0.48 g of magnesium is heated in air, 0.80 g of magnesium oxide forms. Calculate the mass of magnesium oxide that forms from 1.44 g of magnesium. **[1]**

(c) A sample of lake water contains 48 mg/dm^3 of sulfate ions. Calculate the mass of sulfate ions in 250 cm^3 of the lake water. **[1]**

(a) $(2 \times 39) + 12 + (3 \times 16) = 78 + 12 + 48 =$ **138 B1**

(b) $1.44 \div 0.48 = 3$, so the mass is $3 \times 0.80 =$ **2.40 g B1**

(c) 250 cm^3 = 0.250 dm^3, so $48 \times 0.250 =$ **12 mg B1**

Example 2

Extended. Aluminium reacts with dilute hydrochloric acid.

$$2\mathrm{Al} + 6\mathrm{HCl} \rightarrow 2\mathrm{AlCl_3} + 3\mathrm{H_2}$$

1.35 g of aluminium reacts with excess dilute hydrochloric acid.

(a) Calculate the volume of hydrogen made at r.t.p., in cm³. Use the following steps. **[3]**

- Calculate the number of moles of Al.
- Deduce the number of moles of H_2 made.
- Calculate the volume of H_2 at r.t.p.

(b) Calculate the number of hydrogen molecules made. Give your answer in standard form. **[1]**

(c) Calculate the mass of aluminium chloride made. **[2]**

(a) Moles of Al $= 1.35 \div 27 = 0.0500$ mol **M1**. The ratio Al : H_2 is 2 : 3, so moles of $H_2 = 0.0500 \times \frac{3}{2} = 0.0750$ mol **M1**. Volume $= 0.0750 \times 24\,000 =$ **1800 cm³ A1**

(b) $0.0750 \times 6.02 \times 10^{23} =$ **4.52 × 10^{22}** molecules **B1**

(c) Moles of $AlCl_3$ = moles of Al = 0.0500 mol. M_r of $AlCl_3 = 27 + (3 \times 35.5) = 133.5$ **M1**. Mass $= 0.0500 \times 133.5 =$ **6.68 g A1**

Example 3

Extended. 25.0 cm³ of aqueous sodium hydroxide, NaOH, is neutralised by 18.0 cm³ of 0.125 mol/dm³ dilute sulfuric acid.

$$2\mathrm{NaOH} + \mathrm{H_2SO_4} \rightarrow \mathrm{Na_2SO_4} + 2\mathrm{H_2O}$$

Calculate the concentration of the sodium hydroxide in mol/dm³ and in g/dm³. **[4]**

Moles of $H_2SO_4 = 0.125 \times \frac{18.0}{1000} = 0.00225$ mol **M1**

Moles of NaOH $= 2 \times 0.00225 = 0.00450$ mol **M1**

Concentration $= 0.00450 \div \frac{25.0}{1000} =$ **0.180 mol/dm³ A1**

In g/dm³: $0.180 \times 40 =$ **7.20 g/dm³ B1**

Example 4

Extended. 4.80 g of magnesium is added to 75.0 cm³ of 2.00 mol/dm³ hydrochloric acid.

$$\mathrm{Mg} + 2\mathrm{HCl} \rightarrow \mathrm{MgCl_2} + \mathrm{H_2}$$

(a) Show which reactant is limiting. **[2]**

(b) The student obtains 6.18 g of dry magnesium chloride, $MgCl_2$. Calculate the percentage yield. **[3]**

(a) Moles of Mg $= 4.80 \div 24 = 0.200$ mol; moles of HCl $= 2.00 \times 0.0750 = 0.150$ mol **M1**. 0.200 mol of Mg would need 0.400 mol of HCl, but there is only 0.150 mol, so **hydrochloric acid is the limiting reactant** (magnesium is in excess) **A1**.

(b) Moles of $MgCl_2 = 0.150 \div 2 = 0.0750$ mol **M1**. M_r of $MgCl_2 = 24 + 71 = 95$, so the theoretical yield $=$ $0.0750 \times 95 = 7.125$ g **M1**. Percentage yield $= \frac{6.18}{7.125} \times 100 =$ **86.7% A1**

Common mistakes

- **Adding each A_r once when a bracket multiplies it.** $Ca(NO_3)_2$ has 2 N and 6 O. Examiners see brackets cause the most trouble in M_r tables.
- **Using atomic numbers instead of A_r** in an M_r or empirical formula calculation. Use the relative atomic mass, the larger number on the Periodic Table.
- **Forgetting to convert cm^3 to dm^3** (or doing it twice), which leaves the answer 1000 times too big or too small. Write the unit next to each number.
- **Using 24 when the answer is wanted in cm^3 (or 24 000 for dm^3).** Check the unit asked for.
- **Ignoring the mole ratio**, or using it upside down. Examiners often see a 1 : 2 ratio doubled instead of halved. Say it in words: "2 mol NaOH react with 1 mol H_2SO_4, so moles of H_2SO_4 = moles of NaOH ÷ 2".
- **Empirical formulae: dividing percentages by the smallest percentage** instead of first dividing by A_r; then **rounding ratios such as 1 : 1.5 or 2.5 : 5 : 1** to whole numbers instead of multiplying them all by 2.
- **Rounding too early.** Keep at least three significant figures in every intermediate value; round only the final answer.
- **Stopping at the empirical formula** when the molecular formula is asked for, or giving a number (such as 2) instead of a formula.
- **Water of crystallisation: using the mass of the hydrated crystals** as the mass of anhydrous salt, or dividing moles of salt by moles of water (upside down).
- **Leaving an answer as a fraction or with no unit.** Give a decimal (or standard form) with the unit (g, cm^3, mol/dm^3).

How to get the marks

- **"Use the following steps"** questions give a mark for each step, and later steps get credit for correct working from an earlier wrong answer (error carried forward). So **show every step**, labelled, even if you are unsure of the first one.
- **"Calculate"** means show the working. A correct answer with no working can earn full marks, but a wrong answer with no working earns nothing.
- **Units and accuracy.** Give the unit if the answer line doesn't. Three significant figures is usually right; follow the data (if the question gives 2.00 g, don't answer to one figure). Standard form when asked: 4.52×10^{22}.
- **"Deduce"** (for example the moles of product) means use the equation's ratio and the number you just worked out.
- **"Determine the value of x"** expects a whole number, from moles of water ÷ moles of salt.
- **"Explain" a limiting reactant**: state both mole numbers and compare them with the ratio in the equation.

- **Definitions** are marked on key words. A_r: "average mass of the isotopes" + "compared to $\frac{1}{12}$ of the mass of a carbon-12 atom". M_r: "sum of the relative atomic masses". Mole: "6.02×10^{23} particles".
- **Multiple-choice options** are built from the usual slips: the wrong ratio, cm^3 against dm^3, and a bracket missed. Work the answer out before looking at the options.

Quick check

1. (a) Calculate the relative formula mass of aluminium hydroxide, $Al(OH)_3$. (b) When 0.56 g of iron is heated with sulfur, 0.88 g of iron(II) sulfide forms. Calculate the mass of iron needed to make 2.20 g of iron(II) sulfide.
2. **Extended.** How many moles of ammonia, NH_3, are in 8.50 g? How many atoms does this contain in total?
3. **Extended.** Potassium chlorate decomposes on heating: $2KClO_3 \rightarrow 2KCl + 3O_2$. Calculate the volume of oxygen, in dm^3 at r.t.p., made from 12.25 g of $KClO_3$ (M_r = 122.5).
4. **Extended.** 22.5 cm^3 of 0.150 mol/dm^3 nitric acid, HNO_3, neutralises 25.0 cm^3 of aqueous potassium hydroxide, KOH. $HNO_3 + KOH \rightarrow KNO_3 + H_2O$. Calculate the concentration of the potassium hydroxide in mol/dm^3.
5. **Extended.** 14.3 g of hydrated sodium carbonate, $Na_2CO_3 \cdot xH_2O$, is heated to constant mass. 5.30 g of anhydrous Na_2CO_3 (M_r = 106) remains. Determine the value of x.

Answers

1. (a) $27 + 3 \times (16 + 1) =$ **78**. (b) $2.20 \div 0.88 = 2.5$, so $2.5 \times 0.56 =$ **1.40 g** of iron.
2. M_r of $NH_3 = 14 + 3 = 17$, so $8.50 \div 17 =$ **0.500 mol**. Each molecule has 4 atoms: $0.500 \times 4 \times 6.02 \times 10^{23} =$ **1.20×10^{24} atoms**.
3. Moles of $KClO_3 = 12.25 \div 122.5 = 0.100$ mol. The ratio $KClO_3 : O_2$ is 2 : 3, so $O_2 = 0.150$ mol. Volume $= 0.150 \times 24 =$ **3.60 dm³**.
4. Moles of $HNO_3 = 0.150 \times 0.0225 = 0.003375$ mol $=$ moles of KOH (1 : 1). Concentration $= 0.003375 \div 0.0250 =$ **0.135 mol/dm³**.
5. Na_2CO_3: $5.30 \div 106 = 0.0500$ mol. Water: $14.3 - 5.30 = 9.00$ g, so $9.00 \div 18 = 0.500$ mol. $x = 0.500 \div 0.0500 =$ **10**.

Past questions to try

- **0620/32/M/J/22 Q8**: reacting masses and gas volumes by simple proportion (Core).
- **0620/42/F/M/24 Q3**: a titration calculation, ending with a concentration in g/dm³.
- **0620/43/M/J/21 Q7**: an empirical formula from percentages, then the molecular formula.
- **0620/41/O/N/22 Q4**: finding the water of crystallisation in hydrated calcium nitrate.