

Cambridge AS & A Level · Chemistry 9701 · Notes

Atoms, molecules and stoichiometry AS

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What you need to know

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

1. **Define the unified atomic mass unit:** one twelfth of the mass of one atom of carbon-12.
2. **Define relative atomic mass (A_r), relative isotopic mass, relative molecular mass (M_r) and relative formula mass**, each by comparing a mass with the unified atomic mass unit.
3. **Define the mole** using the Avogadro constant, and use it to count particles.
4. **Write the formula of an ionic compound** from the charges on its ions, including:
 - predicting an ion's charge from where its element sits in the Periodic Table;
 - using an oxidation number shown as a Roman numeral, as in iron(III);
 - remembering the names and formulas of nine ions: nitrate NO_3^- , carbonate CO_3^{2-} , sulfate SO_4^{2-} , hydroxide OH^- , ammonium NH_4^+ , zinc Zn^{2+} , silver Ag^+ , hydrogencarbonate HCO_3^- and phosphate PO_4^{3-} .
5. **Write balanced equations**, including ionic equations with no spectator ions, and add state symbols.
6. **Define and use empirical formula and molecular formula**, and calculate both from data.
7. **Use the words anhydrous, hydrated and water of crystallisation**.
8. **Do mole calculations** for reacting masses (with percentage yield), volumes of gases, volumes and concentrations of solutions, and limiting and excess reagents. Give answers to a sensible number of significant figures.
9. **Work out the ratio in which substances react** (the stoichiometry) from the results of such calculations.

Understand it

Relative masses

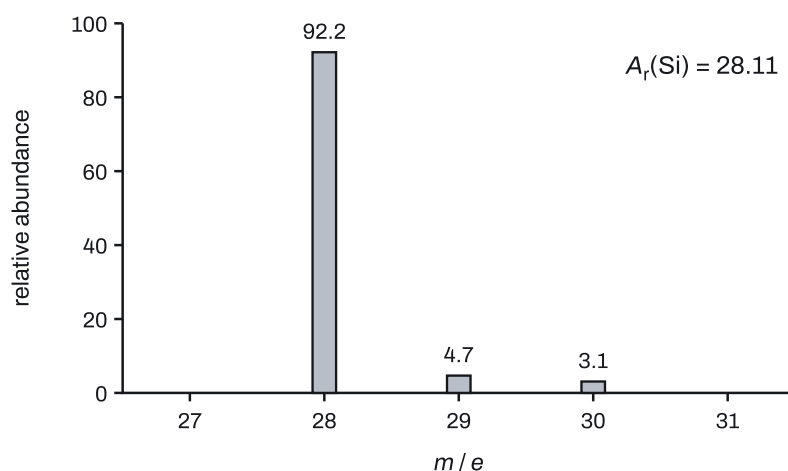
Atoms are far too light to weigh in grams, so chemists compare their masses with a standard: the **unified atomic mass unit**, one twelfth of the mass of a carbon-12 atom. On this scale a carbon-12 atom has a mass of exactly 12. Every "relative" mass below is a pure number with no units, because it is a mass divided by this unit.

- **Relative isotopic mass:** the mass of **one atom of a particular isotope**, compared with $\frac{1}{12}$ of the mass of a carbon-12 atom. It is very close to the isotope's mass number: potassium-41 has a relative isotopic mass of about 41.0.
- **Relative atomic mass, A_r :** the **weighted average** mass of the atoms of an element (taking all its isotopes in their natural amounts), compared with $\frac{1}{12}$ of the mass of a carbon-12 atom. Because it is an average, it belongs to the element, not to one atom: no single bromine atom has a mass of 79.9.
- **Relative molecular mass, M_r :** the weighted average mass of a **molecule**, on the same scale. Add up the A_r values of its atoms: $M_r(\text{N}_2) = 2 \times 14.0 = 28.0$.
- **Relative formula mass:** the same idea for a substance that is not made of molecules, such as an ionic compound. Add up the A_r values in its formula: $\text{MgCO}_3 = 24.3 + 12.0 + 3 \times 16.0 = 84.3$.

The A_r of each element is printed in the Periodic Table at the back of the paper, to one decimal place. Use those values.

Calculating A_r from isotopes

A mass spectrometer measures how much of each isotope a sample contains. Each isotope gives a peak at its mass; the height of the peak is its **relative abundance**.



The weighted average multiplies each mass by its abundance, adds them, and divides by the total abundance:

$$A_r(\text{Si}) = \frac{(28 \times 92.2) + (29 \times 4.7) + (30 \times 3.1)}{92.2 + 4.7 + 3.1} = \frac{2810.9}{100} = 28.11$$

The answer is nearer to 28 than to 30 because almost all silicon atoms are silicon-28. Peak heights don't have to add up to 100: always divide by their sum.

The mole and the Avogadro constant

Chemists count particles in **moles**, just as a baker counts eggs in dozens. 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** , and it is printed in the paper.

The mole is chosen so that the mass of one mole of a substance, in grams, equals its relative mass. One mole of carbon-12 atoms has a mass of exactly 12 g; one mole of water molecules ($M_r = 18.0$) has a mass of 18.0 g. This mass per mole is the **molar mass**, M , in g mol^{-1} . So

$$n = \frac{m}{M}$$

where n is the amount in moles and m is the mass in grams.

For 3.60 g of water: $n = \frac{3.60}{18.0} = 0.200 \text{ mol}$. That is $0.200 \times 6.02 \times 10^{23} = 1.20 \times 10^{23}$ molecules. Each molecule has three atoms, so there are 3.61×10^{23} atoms. Always ask **which particle** you are counting.

Formulas of ionic compounds

An ionic compound has no overall charge, so the positive and negative charges must cancel.

Charges from the Periodic Table. Metals lose their outer electrons and non-metals gain enough to fill their outer shell:

Group	1	2	13	15	16	17
Ion charge	1+	2+	3+ (Al)	3−	2−	1−

Roman numerals give the oxidation number, which for a simple metal ion is its charge: iron(II) is Fe^{2+} , iron(III) is Fe^{3+} , copper(I) is Cu^+ .

Ions to learn: NO_3^- , CO_3^{2-} , SO_4^{2-} , OH^- , NH_4^+ , Zn^{2+} , Ag^+ , HCO_3^- , PO_4^{3-} .

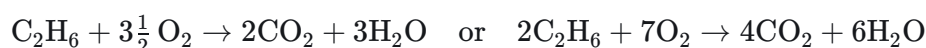
Balance the charges with the fewest ions, and put brackets round a group ion when you need more than one of it:

- Al^{3+} and SO_4^{2-} : two 3+ ions (6+) balance three 2− ions (6−), so $\text{Al}_2(\text{SO}_4)_3$.
- Ca^{2+} and NO_3^- : $\text{Ca}(\text{NO}_3)_2$.
- iron(III) oxide: Fe_2O_3 ; copper(I) oxide: Cu_2O ; ammonium carbonate: $(\text{NH}_4)_2\text{CO}_3$.

The same rule works for an unfamiliar ion given in the question. If you are told that the arsenate ion is AsO_4^{3-} , then calcium arsenate is $\text{Ca}_3(\text{AsO}_4)_2$: three 2+ ions balance two 3− ions.

Equations

An equation is **balanced** when every element has the same number of atoms on each side. Change only the numbers in front of formulas, never the formulas themselves. For combustion, balance C first, then H, then O last:



Fractions such as $3\frac{1}{2}$ are fine, and are handy when a question asks about "one mole" of fuel.

State symbols: (s) solid, (l) liquid, (g) gas, (aq) dissolved in water. Water made at room temperature is $\text{H}_2\text{O}(\text{l})$.

Ionic equations show only the particles that change. Ions that are the same before and after (**spectator ions**) are left out, and the charges must balance as well as the atoms.

- Silver nitrate with sodium chloride: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$. The Na^+ and NO_3^- are spectators.
- Any acid with any alkali: $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$.
- A carbonate with an acid: $\text{CO}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$.

A balanced equation is a **mole ratio**: $2\text{C}_2\text{H}_6 + 7\text{O}_2$ means 2 mol of ethane reacts with 7 mol of oxygen.

Empirical and molecular formulas

- The **empirical formula** is the **simplest whole-number ratio** of the atoms of each element in a compound.
- The **molecular formula** is the **actual number** of atoms of each element in one molecule. It is a whole-number multiple of the empirical formula.

Ethene, C_2H_4 , and cyclohexane, C_6H_{12} , both have the empirical formula CH_2 . Ethane-1,2-diol, $\text{C}_2\text{H}_6\text{O}_2$, has the empirical formula CH_3O .

To find an empirical formula, change masses (or percentages) into moles of atoms, then find the simplest ratio. A compound contains 55.8% carbon, 7.0% hydrogen and 37.2% oxygen by mass:

	C	H	O
mass in 100 g	55.8	7.0	37.2
$\div A_r$	4.65	7.0	2.325
$\div$ smallest	2.00	3.01	1

The empirical formula is $\text{C}_2\text{H}_3\text{O}$, with mass 43.0. If $M_r = 86.0$, the molecular formula has $86.0 \div 43.0 = 2$ lots of it: $\text{C}_4\text{H}_6\text{O}_2$.

Hydrated and anhydrous

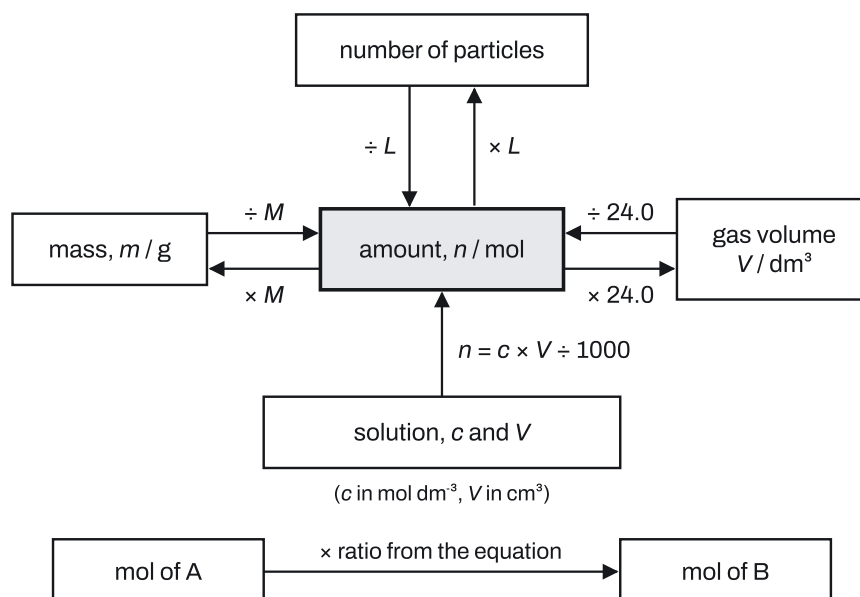
Many salts crystallise with water molecules built into the crystal. This is **water of crystallisation**, and the salt is **hydrated**: blue copper(II) sulfate crystals are $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. Heating drives the water off and leaves the **anhydrous** salt (no water), white CuSO_4 .

The number of water molecules per formula unit, x in $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$, is a mole ratio. If 4.99 g of the hydrated salt leaves 3.19 g of anhydrous salt, then 1.80 g of water was lost:

$$n(\text{CuSO}_4) = \frac{3.19}{159.6} = 0.0200 \text{ mol}, \quad n(\text{H}_2\text{O}) = \frac{1.80}{18.0} = 0.100 \text{ mol}, \quad x = \frac{0.100}{0.0200} = 5$$

Moles from gases and solutions

Every calculation in this topic passes through **moles**. Change what you are given into moles, use the equation's ratio, then change back into what you are asked for.



Gases. At the same temperature and pressure, one mole of **any** gas takes up the same volume, the **molar volume**. The paper gives it: $24.0 \text{ dm}^3 \text{ mol}^{-1}$ at room conditions and $22.4 \text{ dm}^3 \text{ mol}^{-1}$ at s.t.p. (273 K, 101 kPa). So 0.250 mol of any gas at room conditions fills $0.250 \times 24.0 = 6.00 \text{ dm}^3$, and 480 cm^3 is $\frac{480}{24\,000} = 0.0200 \text{ mol}$.

A useful result follows: for gases, **the ratio of volumes equals the ratio of moles**. From $2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2$, 10 cm^3 of carbon monoxide needs 5 cm^3 of oxygen and makes 10 cm^3 of carbon dioxide. You don't need moles at all when every substance is a gas at the same conditions. At room temperature, water is a **liquid** and takes up almost no volume.

Solutions. Concentration, c , is the amount of solute in each dm^3 (1000 cm^3) of solution, in mol dm^{-3} . With the volume V in cm^3 :

$$n = c \times \frac{V}{1000}$$

25.0 cm^3 of $0.200 \text{ mol dm}^{-3}$ acid contains $0.200 \times \frac{25.0}{1000} = 5.00 \times 10^{-3} \text{ mol}$. A concentration in g dm^{-3} is the mol dm^{-3} value times M : $0.200 \text{ mol dm}^{-3}$ sodium hydroxide ($M = 40.0$) is 8.00 g dm^{-3} .

Limiting reagent and percentage yield

Reactants are rarely mixed in exactly the equation's ratio. The reactant that runs out first is the **limiting reagent**: it decides how much product forms. The other is **in excess**, and some of it is left over. To find which is which, work out the moles of each and compare them with the ratio in the equation (Example 3 does this).

The mass of product calculated from the limiting reagent is the **theoretical yield**. Real experiments give less (some reactant doesn't react, some product is lost), so

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

Significant figures

Give an answer to the number of significant figures the question asks for. If it doesn't say, use the same number as the least precise data, usually 3 (sometimes 4 for titration data). Keep all the figures on your calculator until the final answer, and don't confuse significant figures with decimal places: 0.00552 is 3 s.f. but 5 d.p.

Key facts and formulas

"Given" means it is printed in the paper (the Periodic Table and the page of important values, constants and standards). "Learn it" means you must remember it.

Formula	Status
A_r of each element (one decimal place)	Given
Avogadro constant $L = 6.02 \times 10^{23} \text{ mol}^{-1}$	Given
Molar volume of gas: $24.0 \text{ dm}^3 \text{ mol}^{-1}$ at room conditions; $22.4 \text{ dm}^3 \text{ mol}^{-1}$ at s.t.p.	Given
$n = \frac{m}{M}$ (mass in g, molar mass in g mol^{-1})	Learn it
Number of particles = $n \times L$	Learn it
Gas: $n = \frac{V (\text{dm}^3)}{24.0} = \frac{V (\text{cm}^3)}{24\,000}$ at room conditions	Learn it
Solution: $n = c \times \frac{V (\text{cm}^3)}{1000}$; concentration in $\text{g dm}^{-3} = c \times M$	Learn it
$A_r = \frac{\sum(\text{relative isotopic mass} \times \text{abundance})}{\sum \text{abundance}}$	Learn it
Percentage by mass of an element = $\frac{\text{number of atoms} \times A_r}{M_r} \times 100$	Learn it
Percentage yield = $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$	Learn it
Molecular formula = (empirical formula) $\times \frac{M_r}{\text{empirical formula mass}}$	Learn it
x in salt $\cdot x\text{H}_2\text{O} = \frac{n(\text{H}_2\text{O})}{n(\text{anhydrous salt})}$	Learn it
Ions NO_3^- , CO_3^{2-} , SO_4^{2-} , OH^- , NH_4^+ , Zn^{2+} , Ag^+ , HCO_3^- , PO_4^{3-}	Learn it
Definitions: unified atomic mass unit, A_r , relative isotopic mass, M_r , mole, empirical formula	Learn it

How to do it

In the questions we have tagged for this topic (119 questions, 102 different, 2021–2025, Papers 1 and 2), the types came up this often. 37 of the questions mix more than one type, so the counts add up to 167.

Question type	Questions
Gas volumes from moles	24
Empirical and molecular formulas	20
Reacting masses and percentage by mass	17
A_r from isotopic abundances	16
Titration and solution calculations	16
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Limiting and excess reagents	12
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1. Gas volumes from moles (24 questions)

1. Write the balanced equation.
2. Change what you know into moles: $\text{mass} \div M$, or gas volume $\div 24.0 \text{ dm}^3$ ($\div 24\,000$ for cm^3).
3. Use the equation's ratio to get moles of the gas you want.
4. Multiply by $24.0 \text{ dm}^3 \text{ mol}^{-1}$ (or 22.4 if the question says s.t.p.) and give the units.

If every substance involved is a gas at the same conditions, skip the moles: **volumes react in the same ratio as moles**.

Variations you will meet:

- **Final volume after burning with oxygen.** 20 cm^3 of butane is burned in 200 cm^3 of oxygen, and the gases are cooled to room temperature. $\text{C}_4\text{H}_{10} + 6\frac{1}{2}\text{O}_2 \rightarrow 4\text{CO}_2 + 5\text{H}_2\text{O}$:
 1. oxygen used = $6\frac{1}{2} \times 20 = 130 \text{ cm}^3$, so $200 - 130 = 70 \text{ cm}^3$ is **left over**;
 2. carbon dioxide made = $4 \times 20 = 80 \text{ cm}^3$;
 3. water is a liquid at room temperature, so it doesn't count;
 4. final volume = $70 + 80 = 150 \text{ cm}^3$.
- **Passed through an alkali** (aqueous sodium hydroxide or potassium hydroxide): it absorbs the CO_2 (an acidic gas), so only the leftover oxygen remains, 70 cm^3 above. Other acidic gases, such as SO_2 , are also absorbed.
- **Water above 100°C** is a gas and does count in the volume.
- **A mixture of two fuels:** write the CO_2 and H_2O each fuel gives per molecule, and test each option's ratio against the measured volumes.

- **Gas volume to something else:** moles of gas $\rightarrow$ moles of reactant $\rightarrow$ concentration of a solution, the M_r of a carbonate (and so which metal it contains), or the percentage purity of a sample.
- **Molar volume at s.t.p.:** use 22.4 only when the question says s.t.p.; "room conditions" means 24.0.

2. Empirical and molecular formulas (20 questions)

1. Write the mass (or percentage) of each element. Find any missing one by difference: the percentage of oxygen is 100 minus the others.
2. Divide each by its A_r to get moles of atoms.
3. Divide every answer by the smallest.
4. If a ratio is close to a whole number (1.98, 3.02), round it. If it ends in about .5, .33 or .25, multiply everything by 2, 3 or 4.
5. Write the formula with the elements in the usual order (C, H, then others for organic compounds).
6. **Molecular formula:** divide M_r by the empirical formula mass and multiply the formula by that whole number.

Variations you will meet:

- **From a structure** (displayed or skeletal): count every C, H and O (skeletal formulas hide C and H atoms at every corner), then simplify the ratio: a molecule with 6 C, 10 H and 4 O atoms, $C_6H_{10}O_4$, gives $C_3H_5O_2$.
- **From burning a compound of C, H and O.** Mass of C = mass of $CO_2 \times \frac{12.0}{44.0}$; mass of H = mass of $H_2O \times \frac{2.0}{18.0}$; mass of O by difference. Then as above (Example 2).
- **From a titration:** moles of acid from moles of alkali, then $M_r = \frac{\text{mass}}{\text{moles}}$, then the formula whose mass fits. Check whether the options are empirical or molecular formulas: an empirical formula can't be simplified further, so a formula such as $C_4H_8O_4$ (all numbers divisible by 4) is not an empirical formula.
- **A hydrocarbon's M_r** (for example 70) with its empirical formula (CH_2 , 14) fixes the molecular formula (C_5H_{10}), but not the structure.
- **From a precipitate:** moles of a silver halide precipitate give moles of halogen atoms in the sample; combine them with other analyses of the same compound.

3. Reacting masses and percentage by mass (17 questions)

1. Moles of the substance you know: $n = \frac{m}{M}$.
2. Moles of the substance you want, using the equation's ratio.
3. Mass = $n \times M$. Convert units if needed (1 tonne = 10^6 g).

Percentage by mass of an element:

$$\% \text{ of element} = \frac{\text{number of its atoms} \times A_r}{M_r} \times 100$$

For iron in Fe_2O_3 : $\frac{2 \times 55.8}{159.6} \times 100 = 69.9\%$. To pick the compound with a given percentage, work out each option.

Variations you will meet:

- **Identify an element from masses.** Find the moles of the part you know, use the formula's ratio to get moles of the unknown element, then $A_r = \frac{\text{mass}}{\text{moles}}$ and look it up in the Periodic Table.
- **Deduce the ratio** (stoichiometry): if 0.0100 mol of X reacts with 0.0300 mol of Y, the equation has $1X : 3Y$. Divide both by the smaller.
- **Impure samples:** only the pure part reacts. For a sample with 80% of a carbonate, use $0.80 \times \text{mass}$.
- **Equal masses, which gives most product (or contains most of an element)?** Compare moles: work out the mass of each compound per mole of product (or per atom of the element); the smallest wins.
- **Mass lost on heating** (such as $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$): the lost mass is the gas that escaped.
- **Two reactions sharing one reagent:** set up the moles for each and add.

4. A_r from isotopic abundances (16 questions)

1. Multiply each relative isotopic mass by its abundance.
2. Add these and divide by the **sum** of the abundances (100 for percentages).
3. Give the answer to the decimal places or significant figures asked.

Abundances can be percentages, peak heights on a mass spectrum, or a count such as "3 atoms in every 4" (use 3 and 1, divide by 4). If you are given numbers of neutrons, add the proton number to get each mass number first.

Variations you will meet:

- **A missing abundance.** With two isotopes, call one x and the other $100 - x$. With "two isotopes of equal abundance", call both y and the third $100 - 2y$. Then solve the A_r equation (Example 1(c)).
- **The mass of a third isotope.** Its abundance is 100 minus the other two; put everything into the A_r equation and solve for the unknown mass.
- **Two isotopes whose mass numbers differ by 2:** call them m and $m + 2$, and use the fact that A_r is nearer the more common one.
- **Sense check:** A_r always lies between the lightest and heaviest isotopic masses, and nearest the most abundant.

5. Titration and solution calculations (16 questions)

1. Moles of the solution you know exactly: $n = c \times \frac{V}{1000}$.
2. Use the equation's ratio to get moles of the other substance. It is often **not** 1 : 1: $\text{H}_2\text{SO}_4 + 2\text{KOH}$ is 1 : 2; $\text{I}_2 + 2\text{S}_2\text{O}_3^{2-}$ is 1 : 2; $\text{Ca}(\text{OH})_2 + 2\text{HCl}$ is 1 : 2.
3. **Scale up** from the portion you titrated (the aliquot) to the whole solution: 25.0 cm^3 taken from 250.0 cm^3 means $\times 10$; 25.0 cm^3 to 1.00 dm^3 means $\times 40$.
4. Change to what is asked: concentration ($c = n \div \frac{V}{1000}$), mass ($n \times M$), M_r ($\text{mass} \div n$) or percentage by mass ($\frac{\text{mass found}}{\text{mass of sample}} \times 100$).

For example, 25.0 cm^3 of sulfuric acid is neutralised by 20.0 cm^3 of $0.150 \text{ mol dm}^{-3}$ potassium hydroxide.

$$n(\text{KOH}) = 3.00 \times 10^{-3}, \text{ so } n(\text{H}_2\text{SO}_4) = 1.50 \times 10^{-3} \text{ and } c = \frac{1.50 \times 10^{-3}}{0.0250} = 0.0600 \text{ mol dm}^{-3}.$$

Variations you will meet:

- **Dilution:** the moles don't change when you add water, so $c_1 V_1 = c_2 V_2$. 5.00 cm³ of 4.00 mol dm⁻³ acid made up to 250 cm³ is 0.0800 mol dm⁻³.
- **Purity:** find the mass of the substance that reacted, divide by the mass of the impure sample, $\times 100$.
- **An acid with several acidic H atoms:** count the -COOH groups. Each carbonate ion takes two H⁺, so 1 mol of an acid with n acidic H atoms reacts with $\frac{n}{2}$ mol of Na₂CO₃.
- **Paper 2 chains:** each step is a mark, and if the paper says "if you were unable to ..., use the value ...", use it and carry on. Your later steps still earn marks.

6. Define the basic terms (14 questions)

Learn these in your own words, with both parts:

- **Unified atomic mass unit:** one twelfth of the mass of one atom of carbon-12.
- **Relative atomic mass:** the weighted average mass of the atoms of an element (**1**), compared with one twelfth of the mass of a carbon-12 atom (**1**).
- **Relative isotopic mass:** the mass of one atom of an isotope compared with one twelfth of the mass of a carbon-12 atom.
- **Relative molecular mass:** the weighted average mass of a molecule compared with one twelfth of the mass of a carbon-12 atom.
- **Mole:** the amount of substance containing 6.02×10^{23} (the Avogadro constant) particles.
- **Empirical formula:** the simplest whole-number ratio of the atoms of each element in a compound.

Multiple-choice traps. Test each statement against the definitions:

- A statement that gives **one atom** an A_r is wrong: A_r is an average for the element.
- A relative isotopic mass that is not close to a whole number is really an A_r .
- An M_r for a diatomic element must count **both** atoms ($\text{Cl}_2 = 71.0$, not 35.5).
- The Avogadro constant is a number of particles per mole (units mol⁻¹), not a mass.
- The molecular formula is the actual numbers of atoms; "simplest ratio" describes the empirical formula.

7. Count particles with the Avogadro constant (14 questions)

1. Moles of the substance ($\frac{m}{M}$ or $\frac{V}{24.0}$).
2. Multiply by the number of the named particle in one formula: O atoms in CO₂ is 2; atoms in S₈ is 8; ions in CaCl₂ is 3.
3. Multiply by $L = 6.02 \times 10^{23}$.

For "which sample contains the most", you can compare **moles of the named particle** and skip step 3. For example, 3.01×10^{23} nitrogen atoms is $\frac{3.01 \times 10^{23}}{6.02 \times 10^{23}} = 0.500$ mol of N atoms: that is 0.250 mol of N₂ or 0.500 mol of NH₃.

8. Limiting and excess reagents (12 questions)

1. Moles of each reactant.
2. Pick one reactant and use the equation's ratio to work out how much of the other it needs.
3. If there isn't enough of the other, the other one is limiting; if there is more than enough, the one you picked is limiting.
4. Work out the product from the **limiting** reagent only.
5. **Excess left over** = moles present – moles used; change to mass or volume if asked.

For example, 5.60 g of nitrogen (0.200 mol) and 1.50 g of hydrogen (0.750 mol) with $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$: the nitrogen needs 0.600 mol of hydrogen, and there is 0.750 mol, so **nitrogen is limiting** and 0.150 mol (0.300 g) of hydrogen is left over. Ammonia = $0.200 \times 2 = 0.400 \text{ mol} = 9.60 \text{ dm}^3$ at room conditions.

Variations you will meet:

- **From a graph:** if one reactant's concentration stops falling while it is still above zero, the reaction stopped because the **other** reactant ran out. Say which reactant is limiting and give that reason.
- **Neutralisation with one reagent limiting:** the amount of water (or salt) formed comes from the limiting one.
- **Gas volumes:** leftover oxygen is part of the final volume (type 1).

9. Construct and balance equations (11 questions)

1. Write the correct formula of every reactant and product.
2. Balance metals first, then non-metals, then H and O last.
3. Check every element, and the total charge in an ionic equation.
4. Add state symbols if asked.

Unfamiliar reactions: use the products you are told about and balance atoms to deduce any missing one. For magnesium nitride, Mg_3N_2 , with water giving magnesium hydroxide and one other product, the 2 N atoms must end up in 2NH_3 : $\text{Mg}_3\text{N}_2 + 6\text{H}_2\text{O} \rightarrow 3\text{Mg}(\text{OH})_2 + 2\text{NH}_3$.

Ionic equations: write the full equation, split soluble ionic compounds into ions, cross out the ions that appear unchanged on both sides, and keep solids, water and gases whole (Example 3(a)).

10. Find x in a hydrate (9 questions)

1. Mass of anhydrous salt (after heating) and mass of water (before – after).
2. Moles of each: divide by M of the **anhydrous** salt and by 18.0.
3. $x = \frac{n(\text{water})}{n(\text{anhydrous salt})}$, rounded to a whole number unless the question says it isn't one.

Variations you will meet:

- **From M_r of the hydrate:** $x = \frac{M_r(\text{hydrate}) - M_r(\text{anhydrous})}{18.0}$ (Example 4).

- **From a percentage of water by mass:** treat 100 g of hydrate as that many grams of water and the rest as anhydrous salt.
- **An expression with letters.** For $\text{MgSO}_4 \cdot x\text{H}_2\text{O}$ ($M = 120.4$ for the anhydrous salt): empty dish a , dish + hydrate b , dish + residue c . Water = $b - c$ and anhydrous salt = $c - a$, so $x = \frac{(b - c)/18.0}{(c - a)/120.4} = \frac{120.4(b - c)}{18.0(c - a)}$.
- **From the percentage of one ion:** the mass of that ion gives the moles of salt, and the rest of the mass is water.

11. Combustion ratios (9 questions)

Complete combustion turns every C into CO_2 and every H into H_2O .

Moles of oxygen for one mole of fuel $\text{C}_x\text{H}_y\text{O}_z$: it gives x mol CO_2 and $\frac{y}{2}$ mol H_2O , which contain $2x + \frac{y}{2}$ mol of O atoms; subtract the fuel's own z and halve. For butanone, $\text{C}_4\text{H}_8\text{O}$: $\frac{8+4-1}{2} = 5\frac{1}{2}$ mol of O_2 . For a hydrocarbon this is simply $x + \frac{y}{4}$.

Variations you will meet:

- **Moles of products:** add x and $\frac{y}{2}$ for one mole of fuel.
- **Incomplete combustion:** if the products are carbon (or CO) and water, balance with those instead; less oxygen is needed.
- **A mixture of fuels:** work out the oxygen for each and add.
- **"Which compound needs n mol of oxygen?":** use the formula above on each option. An O atom in the fuel lowers the oxygen needed by $\frac{1}{2}$ mol.

12. Formulas of ionic compounds (5 questions)

1. Write each ion with its charge (from its group, its Roman numeral, the list of ions to learn, or an acid given in the question: an acid with two $-\text{COOH}$ groups gives a $2-$ ion).
2. Choose the fewest ions whose charges cancel; bracket a group ion when there are two or more.

Common checks: silver is always Ag^+ , so silver sulfate is Ag_2SO_4 ; phosphate is $3-$, so calcium phosphate is $\text{Ca}_3(\text{PO}_4)_2$; hydrogencarbonate is $1-$, so magnesium hydrogencarbonate is $\text{Mg}(\text{HCO}_3)_2$.

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) Define relative isotopic mass. **[1]**

(b) A sample of lithium contains 7.6% of an isotope with relative isotopic mass 6.02 and 92.4% of an isotope with relative isotopic mass 7.02. Calculate the relative atomic mass of this lithium. Give your answer to two decimal places. **[2]**

(c) Thallium has two isotopes, thallium-203 and thallium-205 (relative isotopic masses 203.0 and 205.0). Its relative atomic mass is 204.4. Calculate the percentage abundance of thallium-203. **[2]**

(d) Calculate the number of atoms in 2.50 g of thallium. **[2]**

(a) The mass of one atom of an isotope, compared with one twelfth of the mass of an atom of carbon-12. **B1**

(b)

$$A_r = \frac{(6.02 \times 7.6) + (7.02 \times 92.4)}{100} = \frac{694.4}{100} = 6.944 = 6.94$$

M1 for the expression, **A1** for 6.94 (two decimal places, as asked).

(c) Let thallium-203 be $x\%$, so thallium-205 is $(100 - x)\%$. **M1**

$$\frac{203.0x + 205.0(100 - x)}{100} = 204.4 \Rightarrow 20\,500 - 2x = 20\,440 \Rightarrow x = 30$$

Thallium-203 is 30%. **A1**

(d) $n = \frac{2.50}{204.4} = 0.01223 \text{ mol}$. **M1** Atoms = $0.01223 \times 6.02 \times 10^{23} = 7.36 \times 10^{21}$. **A1**

Example 2

Compound Y contains carbon, hydrogen and oxygen only. When 0.900 g of Y is burned completely, it forms 1.76 g of carbon dioxide and 0.900 g of water.

(a) Show that the empirical formula of Y is $\text{C}_2\text{H}_5\text{O}$. **[3]**

(b) The relative molecular mass of Y is 90.0. Deduce its molecular formula. **[1]**

(c) Write a balanced equation for the complete combustion of Y. **[1]**

(d) Calculate the volume of oxygen, measured at room conditions, needed to burn 0.900 g of Y. **[2]**

(a) Mass of C = $1.76 \times \frac{12.0}{44.0} = 0.480$ g; mass of H = $0.900 \times \frac{2.0}{18.0} = 0.100$ g. **M1**

Mass of O = $0.900 - 0.480 - 0.100 = 0.320$ g. **M1**

	C	H	O
moles	$\frac{0.480}{12.0} = 0.0400$	$\frac{0.100}{1.0} = 0.100$	$\frac{0.320}{16.0} = 0.0200$
ratio	2	5	1

So the empirical formula is C₂H₅O. **A1**

(b) Empirical formula mass = 45.0; $90.0 \div 45.0 = 2$, so C₄H₁₀O₂. **B1**

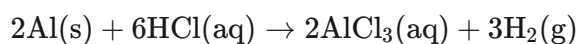
(c) C₄H₁₀O₂ + 5½O₂ → 4CO₂ + 5H₂O **B1**

(d) $n(Y) = \frac{0.900}{90.0} = 0.0100$ mol, so $n(O_2) = 5\frac{1}{2} \times 0.0100 = 0.0550$ mol. **M1**

Volume = $0.0550 \times 24.0 = 1.32$ dm³. **A1**

Example 3

A student adds 0.540 g of aluminium powder to 45.0 cm³ of 1.00 mol dm⁻³ hydrochloric acid.



(a) Write the ionic equation for this reaction, with state symbols. **[1]**

(b) Show which reactant is in excess. **[2]**

(c) Calculate the maximum volume of hydrogen, in cm³, at room conditions. **[1]**

(d) Calculate the mass of aluminium left over at the end. **[1]**

(e) The student collects 495 cm³ of hydrogen. Calculate the percentage yield. **[1]**

(a) $2\text{Al(s)} + 6\text{H}^+\text{(aq)} \rightarrow 2\text{Al}^{3+}\text{(aq)} + 3\text{H}_2\text{(g)}$ **B1** (the Cl⁻ ions are spectators)

(b) $n(\text{Al}) = \frac{0.540}{27.0} = 0.0200$ mol; $n(\text{HCl}) = 1.00 \times \frac{45.0}{1000} = 0.0450$ mol. **M1**

0.0200 mol of Al would need $3 \times 0.0200 = 0.0600$ mol of HCl, but there is only 0.0450 mol. So HCl is limiting and **aluminium is in excess**. **A1**

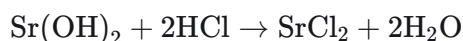
(c) $n(\text{H}_2) = 0.0450 \times \frac{3}{6} = 0.0225$ mol; volume = $0.0225 \times 24\,000 = 540$ cm³. **B1**

(d) Al used = $0.0450 \times \frac{2}{6} = 0.0150$ mol = $0.0150 \times 27.0 = 0.405$ g, so $0.540 - 0.405 = 0.135$ g is left. **B1**

(e) $\frac{495}{540} \times 100 = 91.7\%$ **B1**

Example 4

A student dissolves 2.66 g of hydrated strontium hydroxide, $\text{Sr}(\text{OH})_2 \cdot x\text{H}_2\text{O}$, in water and makes the solution up to 250.0 cm^3 . A 25.0 cm^3 portion of this solution needs 20.00 cm^3 of $0.100 \text{ mol dm}^{-3}$ hydrochloric acid to react completely.



- (a) Calculate the amount, in mol, of hydrochloric acid used. **[1]**
- (b) Calculate the amount, in mol, of strontium hydroxide in the 250.0 cm^3 of solution. **[2]**
- (c) Calculate the relative formula mass of the hydrated strontium hydroxide, and hence find x . **[3]**

$$\text{(a)} \quad n(\text{HCl}) = 0.100 \times \frac{20.00}{1000} = 2.00 \times 10^{-3} \text{ mol } \mathbf{B1}$$

$$\text{(b)} \quad \text{In } 25.0 \text{ cm}^3: \frac{2.00 \times 10^{-3}}{2} = 1.00 \times 10^{-3} \text{ mol } \mathbf{M1} \text{ (the 1 : 2 ratio)}$$

$$\text{In } 250.0 \text{ cm}^3: 1.00 \times 10^{-3} \times 10 = 0.0100 \text{ mol } \mathbf{A1} \text{ (scaling up)}$$

$$\text{(c)} \quad M_r = \frac{2.66}{0.0100} = 266 \mathbf{M1}$$

$$M_r(\text{Sr}(\text{OH})_2) = 87.6 + 2(16.0 + 1.0) = 121.6, \text{ so water} = 266.0 - 121.6 = 144.4 \mathbf{M1}$$

$$x = \frac{144.4}{18.0} = 8.02, \text{ so } x = 8. \mathbf{A1} \text{ (a whole number: the formula is } \text{Sr}(\text{OH})_2 \cdot 8\text{H}_2\text{O)}$$

Common mistakes

- **Rounding too early, or to too few figures.** Examiners regularly see a middle value such as 0.0123 cut to 0.012, and final answers to one significant figure. Keep full calculator values and round only at the end, usually to 3 s.f.
- **Mixing up significant figures and decimal places.** "Three significant figures" for 0.0055213 is 0.00552; "three decimal places" is 0.006. Read which one is asked.
- **Leaving out the equation's ratio.** In a thiosulfate or sulfuric acid titration the ratio is 1 : 2, not 1 : 1. Write the ratio down before you use it.
- **Forgetting to scale up.** A titration uses one portion of the solution; multiply by 10 (from 25.0 of 250 cm^3) or 40 (from 25.0 cm^3 to 1 dm^3) to get the whole.
- **Using the wrong molar volume.** "Room conditions" is $24.0 \text{ dm}^3 \text{ mol}^{-1}$; 22.4 is only for s.t.p. Using the wrong one gives a wrong answer that often appears as a distractor.
- **Counting water as a gas at room temperature.** After cooling to room temperature, water is a liquid and adds nothing to the gas volume.
- **Forgetting leftover reactant in a final gas volume.** Excess oxygen is still there at the end.
- **Weak definitions.** Answers that don't mention carbon-12, or talk about "the mass of a compound", lose marks. Learn both parts of each definition.
- **Giving the molecular formula when asked for the empirical formula,** or the other way round.

- **Counting particles wrongly.** If the question asks for atoms (or carbon atoms, or ions), multiply the moles of molecules by the number in one formula before using L .
- **Mixing up hydrated and anhydrous.** Moles of salt come from the **anhydrous** mass and M ; and when the question says x is an integer, give a whole number.
- **Spectator ions and missing state symbols** in ionic equations. Leave out ions that don't change, and give every species a state symbol.
- **Wrong formulas** such as "CaCO₂" for calcium carbonate. Build formulas from the ion charges.

How to get the marks

- **Show every step in calculations.** Paper 2 gives a mark for each step (moles, then ratio, then scaling, then mass), so a slip early on costs one mark, not all of them.
- **Error carried forward.** If you got an earlier part wrong, use your answer (or the value the paper offers) in the next part. Correct method on a wrong number still scores.
- **Significant figures.** Match the question, or the data (usually 3 s.f.; 3 or 4 for titration results). Never round to 1 s.f. Use A_r values from the Periodic Table as printed.
- **Units.** Give them when the answer line doesn't: mol, g, dm³ or cm³, mol dm⁻³, g dm⁻³.
- **"Define":** learn the key phrases: "weighted average", "one twelfth of the mass of a carbon-12 atom", "simplest whole-number ratio".
- **"Show your working"** for an empirical formula: the mark scheme looks for the line of moles (mass $\div A_r$) and the ratio, then the formula.
- **"Deduce":** reason from the information given, such as counting atoms in a structure or picking the limiting reagent from a graph. Say why.
- **Equations:** balanced, correct formulas, state symbols when asked, and an ionic equation with no spectator ions.
- **"Show that":** the answer is given, so every step must be visible and your working must reach it.
- **Multiple choice:** write the equation and the mole numbers in the margin. Many wrong options are the answers to a common slip (wrong ratio, water counted as a gas, molecular instead of empirical formula).

Quick check

1. Calculate the total number of ions in 5.55 g of calcium chloride, CaCl₂.
2. Write the formulas of ammonium sulfate, chromium(III) nitrate, silver phosphate, zinc hydroxide and ammonium hydrogencarbonate.
3. 20 cm³ of ethyne, C₂H₂, is burned in 70 cm³ of oxygen. The gases are cooled to room temperature. What is the final volume of gas? What volume remains after the gas is passed through aqueous sodium hydroxide?
4. Calculate the percentage by mass of magnesium in magnesium nitride, Mg₃N₂.
5. 25.0 cm³ of limewater (aqueous calcium hydroxide) is neutralised by 17.20 cm³ of 0.0500 mol dm⁻³ hydrochloric acid. Calculate the concentration of the calcium hydroxide in mol dm⁻³ and in g dm⁻³.

Answers

1. $M_r(\text{CaCl}_2) = 111.1$, so $n = \frac{5.55}{111.1} = 0.04995$ mol. Each formula unit has 3 ions: $0.04995 \times 3 \times 6.02 \times 10^{23} = 9.02 \times 10^{22}$ ions.
2. $(\text{NH}_4)_2\text{SO}_4$, $\text{Cr}(\text{NO}_3)_3$, Ag_3PO_4 , $\text{Zn}(\text{OH})_2$, NH_4HCO_3 .
3. $\text{C}_2\text{H}_2 + 2\frac{1}{2}\text{O}_2 \rightarrow 2\text{CO}_2 + \text{H}_2\text{O}$. Oxygen used = 50 cm^3 , leaving 20 cm^3 ; $\text{CO}_2 = 40 \text{ cm}^3$; water is liquid. Final volume = 60 cm^3 . Sodium hydroxide absorbs the CO_2 , leaving 20 cm^3 of oxygen.
4. $M_r = 3(24.3) + 2(14.0) = 100.9$; $\frac{72.9}{100.9} \times 100 = 72.2\%$.
5. $n(\text{HCl}) = 0.0500 \times \frac{17.20}{1000} = 8.60 \times 10^{-4}$ mol; $n(\text{Ca}(\text{OH})_2) = 4.30 \times 10^{-4}$ mol; $c = \frac{4.30 \times 10^{-4}}{0.0250} = 0.0172$ mol dm^{-3} . $M_r(\text{Ca}(\text{OH})_2) = 74.1$, so $0.0172 \times 74.1 = 1.27 \text{ g dm}^{-3}$.

Past questions to try

- **9701/22/M/J/25 Q2**: define relative atomic mass, then find the mass of a third isotope.
- **9701/24/M/J/25 Q6**: define empirical formula and calculate one from a percentage by mass.
- **9701/21/M/J/25 Q4**: a titration with a 1 : 1 ratio, scaled up to the mass in a tablet.
- **9701/23/O/N/24 Q1**: part (d)(i): find x in a hydrate from its percentage of water by mass.