

Every AQA GCSE Chemistry calculation

25 calculations with units and worked examples. Unflagged calculations apply to Combined Science and Chemistry at both tiers.

Higher Higher tier only **Separate** Separate Chemistry only (not in Combined Science)

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CALCULATION	UNITS	WORKED EXAMPLE
Atomic structure		
Protons, neutrons and electrons in an atom number of neutrons = mass number – atomic number	mass number (protons + neutrons): no unit atomic number (protons): no unit	Sodium is written as $^{23}_{11}\text{Na}$. Work out its numbers of protons, neutrons and electrons. Atomic number (bottom) = 11, so there are 11 protons. Neutrons = mass number – atomic number = $23 - 11 = 12$. A sodium atom is neutral, so it has 11 electrons. In a neutral atom, electrons = protons.
Electrons in an ion electrons in an ion = protons – charge on the ion	protons (atomic number): no unit charge on the ion: e.g. +2 or –2	Group 2 metals lose 2 electrons, so they form +2 ions (e.g. Mg^{2+} , Ca^{2+}). Group 6 non-metals gain 2 electrons, so they form –2 ions (e.g. O^{2-} , the oxide ion - not 'oxygen ion'). Mg^{2+} (atomic number 12) has $12 - 2 = 10$ electrons; O^{2-} (atomic number 8) has $8 + 2 = 10$ electrons.
Relative atomic mass from isotopes $\text{Ar} = \text{sum of (isotope mass} \times \% \text{ abundance)} \div 100$	relative atomic mass, Ar: no unit abundance: per cent, %	A sample has 20% boron-10 and 80% boron-11: $\text{Ar} = (20 \times 10 + 80 \times 11) \div 100 = 10.8$, between the two isotope masses.
Quantitative chemistry		
Relative formula mass $\text{Mr} = \text{sum of the Ar of every atom in the formula}$	relative formula mass, Mr: no unit relative atomic mass, Ar: no unit	For $\text{Ca}(\text{OH})_2$ the Mr is $40 + 2 \times (16 + 1) = 40 + 34 = 74$ using Ca = 40, O = 16, H = 1. Relative formula mass has NO units - never add g or kg.
Percentage by mass of an element $\% \text{ by mass of an element} = (\text{Ar} \times \text{number of atoms of that element} \div \text{Mr}) \times 100$	percentage by mass: per cent, % Ar and Mr: no unit	% by mass of iron in Fe_2O_3 Mr = $(2 \times 56) + (3 \times 16) = 160$; iron = $2 \times 56 = 112$; $112 \div 160 \times 100 = 70\%$. Count both Fe atoms (not 35%).
Mean, range and uncertainty mean = total of the repeats ÷ number of repeats; range = highest – lowest	mean, range and uncertainty: same unit as the readings	Mean = $(24.8 + 25.0 + 25.2) \div 3 = 75.0 \div 3 = 25.0 \text{ cm}^3$. Range = highest – lowest = $25.2 - 24.8 = 0.4 \text{ cm}^3$. Estimate uncertainty as half the range: $0.4 \div 2 = 0.2 \text{ cm}^3$.
Moles from mass Higher $\text{moles} = \text{mass (g)} \div \text{Mr}$	amount: moles, mol mass: grams, g relative formula mass, Mr: no unit	For 80 g of calcium (Ar Ca = 40): moles = mass ÷ Mr = $80 \div 40 = 2 \text{ mol}$. Rearrange it to find mass from moles: mass (g) = moles × Mr.
Number of particles (Avogadro constant) Higher $\text{number of particles} = \text{moles} \times 6.02 \times 10^{23}$	number of particles: no unit amount: moles, mol	Two moles of a substance therefore contain $2 \times 6.02 \times 10^{23}$ particles. $2 \times 6.02 \times 10^{23} = 1.204 \times 10^{24}$ particles. That number of particles is the Avogadro constant, 6.02×10^{23} per mole.
Reacting masses from a balanced equation Higher $\text{moles} = \text{mass} \div \text{Mr}$, then use the equation's ratio, then mass = moles × Mr	mass: grams, g amount: moles, mol relative formula mass, Mr: no unit	$\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$. For 50 g CaCO_3 (Mr 100): moles = $50 \div 100 = 0.5$. Ratio 1 : 1, so 0.5 mol CaO. Mr CaO = 56, mass = $0.5 \times 56 = 28 \text{ g}$.
Balancing an equation from masses Higher $\text{moles of each substance} = \text{mass} \div \text{Mr}$; divide by the smallest; scale to the simplest whole-number ratio; then balance any other formulae from the atoms	mass: grams, g amount: moles, mol relative formula mass, Mr: no unit	Moles of Mg = $4.8 \div 24 = 0.2 \text{ mol}$. Moles of O_2 = $3.2 \div 32 = 0.1 \text{ mol}$. Divide by the smallest (0.1): Mg = $0.2 \div 0.1 = 2$; O_2 = $0.1 \div 0.1 = 1$. Ratio Mg : O_2 = 2 : 1; balancing the atoms then gives $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$. If a ratio comes out as 1 : 1.5, scale both up (× 2) to reach whole numbers such as 2 : 3 before writing the equation.

Limiting reactant and product mass Higher moles of each reactant = mass ÷ Mr; use the ratio to find which runs out; the product amount comes from the limiting reactant	mass: grams, g amount: moles, mol relative formula mass, Mr: no unit	For $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$, 0.10 mol Mg reacts with 0.30 mol HCl. Find the limiting reactant and the mass of H_2 made. (Ar: H = 1.) The ratio is Mg : HCl = 1 : 2, so 0.10 mol Mg needs $2 \times 0.10 = 0.20$ mol HCl. 0.30 mol HCl is available, which is more than 0.20 mol, so HCl is in excess. Ratio Mg : H_2 = 1 : 1, so 0.10 mol Mg makes 0.10 mol H_2 . Mass of H_2 = moles $\times$ Mr = $0.10 \times 2 = 0.20$ g. Always compare moles, not masses, because it is the mole ratio in the equation that decides which reactant runs out.
Concentration in g/dm³ concentration (g/dm³) = mass of solute (g) ÷ volume of solution (dm³)	concentration: grams per cubic decimetre, g/dm ³ mass of solute: grams, g volume of solution: cubic decimetres, dm ³ (cm ³ ÷ 1000)	Keep 20 g of solute but make the volume 0.5 dm ³ : concentration = $20 \div 0.5 = 40$ g/dm ³ . To convert a volume from cm ³ into dm ³ , divide by 1000: e.g. $250 \text{ cm}^3 = 0.25 \text{ dm}^3$, and $500 \text{ cm}^3 = 0.5 \text{ dm}^3$.
Percentage yield Separate % yield = (mass of product actually made ÷ maximum theoretical mass) × 100	percentage yield: per cent, % mass: grams, g	10.0 g of product is possible but only 9.0 g is collected because some sticks to the filter paper: yield = $(9.0 \div 10.0) \times 100 = 90\%$.
Theoretical mass of product (for % yield) Separate Higher moles of reactant = mass ÷ Mr; use the mole ratio; maximum theoretical mass of product = moles × Mr	mass: grams, g amount: moles, mol relative formula mass, Mr: no unit	Example: heating 25.0 g of CaCO_3 (Mr 100) gives 0.250 mol, so $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ makes 0.250 mol of CaO (Mr 56) - a theoretical mass of 14.0 g. Only then can you use % yield = actual ÷ theoretical $\times 100$ when just the reactant mass is given.
Atom economy Separate % atom economy = (Mr of the desired product ÷ sum of the Mr of all the reactants) × 100, each Mr multiplied by its coefficient in the balanced equation	atom economy: per cent, % relative formula mass, Mr: no unit	$\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$. Mr of desired product $\text{H}_2 = 2 \times 1 = 2$. Sum of Mr of all reactants = $65 + 98 = 163$, then % = $(2 \div 163) \times 100$. Atom economy = 1.23% (3 s.f.) - most reactant atoms end up in the ZnSO_4 by-product. Use the relative formula masses in the amounts shown by the balanced equation - multiply each Mr by its coefficient.
Concentration in mol/dm³ Separate Higher concentration (mol/dm³) = amount of solute (mol) ÷ volume of solution (dm³)	concentration: moles per cubic decimetre, mol/dm ³ amount of solute: moles, mol volume of solution: cubic decimetres, dm ³ (cm ³ ÷ 1000)	A NaOH solution has a concentration of 0.5 mol/dm ³ . Find the moles of NaOH in 2 dm ³ , then give the concentration in g/dm ³ . (Ar: Na = 23, O = 16, H = 1.) Moles = concentration $\times$ volume = $0.5 \times 2 = 1$ mol of NaOH. Mr of NaOH = $23 + 16 + 1 = 40$. Concentration in g/dm ³ = mol/dm ³ $\times$ Mr = $0.5 \times 40 = 20$ g/dm ³ . Rearrange it to find moles = concentration $\times$ volume, or volume = moles ÷ concentration.
Converting mol/dm³ to g/dm³ Separate Higher concentration (g/dm³) = concentration (mol/dm³) × Mr	concentration: g/dm ³ and mol/dm ³ relative formula mass, Mr: no unit	A solution is 0.1 mol/dm ³ of NaCl (Mr = 58.5). Concentration in g/dm ³ = $0.1 \times 58.5 = 5.85$ g/dm ³ , found by multiplying the molar concentration by Mr. Convert g/dm ³ back to mol/dm ³ by dividing the mass concentration by the solute's Mr.
Volume of a gas at room temperature and pressure Separate Higher volume of gas at rtp (dm³) = moles × 24	volume of gas: cubic decimetres, dm ³ amount: moles, mol	Example: what volume does 8 g of methane, CH_4 (Mr = 16), occupy at rtp? moles = $8 \div 16 = 0.5$ mol, so volume = $0.5 \times 24 = 12$ dm ³ at rtp. 1 dm ³ = 1000 cm ³ , so 24 dm ³ is the same as 24 000 cm ³ at rtp.
Gas volumes from a balanced equation Separate Higher volume ratio of gases = mole ratio in the balanced equation (same temperature and pressure)	gas volumes: same unit for each gas, e.g. dm ³	In $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, 48 dm ³ of H_2 needs 24 dm ³ of O_2 at rtp, because the 2:1 mole ratio is also the 2:1 volume ratio for gases.
Chemical changes		
pH and hydrogen ion concentration Higher each 1 unit decrease in pH = 10 times the H⁺ concentration	pH: no unit (0–14 scale) H ⁺ concentration: factor of 10 per pH unit	So a solution of pH 2 has ten times the H ⁺ concentration of one at pH 3. From pH 4 to pH 1 is 3 units, so the H ⁺ concentration is $10 \times 10 \times 10 = 1000$ times greater.

Titration calculations Separate Higher moles of known solution = concentration × volume (dm³); use the equation's ratio; unknown concentration = moles ÷ volume (dm³)	concentration: mol/dm ³ (or g/dm ³ after × Mr) amount: moles, mol volume of solution: cubic decimetres, dm ³ (cm ³ ÷ 1000)	25.0 cm ³ of sodium hydroxide is exactly neutralised by 20.0 cm ³ of 0.100 mol/dm ³ hydrochloric acid. <chem>HCl + NaOH -> NaCl + H2O</chem> shows the acid and alkali react in a 1 : 1 ratio. Moles of HCl = concentration × volume in dm ³ = 0.100 × 20.0/1000 = 0.00200 mol. 1 : 1 ratio → moles of NaOH = 0.00200 mol. Concentration of NaOH = moles ÷ volume in dm ³ = 0.00200 ÷ 0.0250 = 0.0800 mol/dm ³ . Always convert cm ³ to dm ³ (÷1000) before using concentration = moles ÷ volume.
Energy changes		
Energy change from bond energies Higher overall energy change = energy needed to break bonds – energy released making bonds	energy: kilojoules per mole, kJ/mol	Breaking bonds takes 500 kJ/mol; forming bonds releases 650 kJ/mol. Overall change = 500 – 650 = –150 kJ/mol: exothermic. Break = takes IN (endo). Make = gives OUT (exo).
Rates of reaction		
Mean rate of reaction mean rate of reaction = quantity of reactant used (or product formed) ÷ time taken	rate: g/s or cm ³ /s (mol/s at Higher tier) quantity: grams, g, or cubic centimetres, cm ³ (moles, mol, at Higher tier) time: seconds, s	24 cm ³ gas collected in 40 s gives a mean rate of 24 ÷ 40 = 0.60 cm ³ /s. Mean rate = quantity used or formed ÷ time. Watch the units: g/s or cm ³ /s.
Rate from the gradient of a tangent Higher rate at one moment = gradient of the tangent = change in quantity ÷ change in time , both changes read along the tangent	rate: vertical-axis unit ÷ horizontal-axis unit, e.g. cm ³ /s	On a reactant-remaining curve the gradient is negative - the rate is the magnitude of that negative gradient.
Chemical analysis		
Rf value (chromatography) Rf = distance moved by the substance ÷ distance moved by the solvent	Rf value: no unit (ratio, never more than 1) distances: same unit for both, e.g. mm or cm	A spot travels 54 mm and the solvent front 90 mm: Rf = 54 ÷ 90 = 0.60. Measure from the same baseline; compare under the same conditions. Measure from the baseline to the centre of each spot, and from the baseline to the solvent front.