GCSE CHEMISTRY - AQA AND PEARSON EDEXCEL

GCSE Chemistry revision handbook

A connected review of particles, equations, calculations, reactions and evidence, followed by original exam-style practice.

Use the confirmed course

This handbook covers core ideas shared by AQA and Pearson Edexcel. Topic order, required practicals, tier content and equation sheets can differ. Always use the learner's current specification and teacher guidance as the final authority.

Inside this handbook

Each chapter gives the essential knowledge, a worked example and original exam-style practice. Complete the questions before using the answer section.

CHAPTER 1

Atomic structure and the periodic table

Atomic structure explains element identity, isotope mass and repeating chemical behaviour across the periodic table.

Coverage map

Subatomic particles; isotopes; electronic structure; development of the periodic table; Group 1; Group 7; Group 0; transition metals.

Essential knowledge

- Atomic number is proton number; mass number is protons plus neutrons.
- Isotopes have the same protons but different neutrons.
- Electrons occupy shells or energy levels; outer electrons drive chemical behaviour.
- Group number links to outer-electron pattern for main-group elements.
- Group 1 reactivity increases down the group; Group 7 reactivity decreases.

Worked example	Chlorine-35 has atomic number 17: 17 protons, 18 neutrons and 17 electrons in a neutral atom.
Exam focus	Use proton number to identify the element, then use electron arrangement to explain ion formation and group trends.
Common mistake to avoid	Relative atomic mass is a weighted mean of isotopes, not simply the mass number of the most abundant isotope.

CHAPTER 1 PRACTICE

Exam-style practice

1. An ion has 12 protons, 10 electrons and mass number 24. State its charge and number of neutrons.

2. Explain why Group 1 metals become more reactive down the group.

3. Chlorine is 75% chlorine-35 and 25% chlorine-37. Calculate its relative atomic mass.

CHAPTER 2

Bonding, structure and properties

Properties depend on bonding, particles and structure rather than on a substance merely containing a particular element.

Coverage map

Ionic bonding; covalent bonding; metallic bonding; simple molecules; polymers; giant ionic, metallic and covalent structures; graphene, fullerenes and nanoparticles.

Essential knowledge

- Ionic bonding is electrostatic attraction between oppositely charged ions.
- Covalent bonds are shared pairs of electrons; simple molecules and giant covalent structures behave differently.
- Metallic bonding is attraction between positive ions and delocalised electrons.
- Melting and boiling involve overcoming forces or bonds; be precise about which.
- Electrical conduction requires mobile charged particles.

Worked example	Solid sodium chloride does not conduct because ions are fixed in a lattice. Molten sodium chloride conducts because ions can move and carry charge.
Exam focus	Name the particles and forces present, then explain the property by saying what can move or what must be overcome.
Common mistake to avoid	Melting a simple molecular substance overcomes intermolecular forces, not the covalent bonds inside each molecule.

CHAPTER 2 PRACTICE

Exam-style practice

1. Explain why magnesium oxide has a high melting point.

2. Compare electrical conductivity in graphite and diamond.

3. Explain why metals are malleable without losing metallic bonding.

CHAPTER 3

Quantitative chemistry

Chemical calculations connect balanced equations with measurable mass, concentration, gas volume and yield.

Coverage map

Relative masses; balancing equations; moles; reacting masses; limiting reactants; concentration; gas volumes; percentage yield; atom economy; titration calculations.

Essential knowledge

- Amount in moles = mass ÷ relative formula mass.
- Balanced-equation coefficients give mole ratios, not mass ratios.
- Concentration in mol/dm^3 = moles ÷ volume in dm^3.
- Percentage yield = actual yield ÷ theoretical yield × 100.
- Atom economy compares desired product mass with total product mass from the equation.

Worked example	CaCO3 has Mr 100. A 25 g sample contains $^{25}/_{100}$ = 0.25 mol.
Exam focus	Write the balanced equation, convert to moles, apply the mole ratio and only then convert to the requested quantity.
Common mistake to avoid	Equation coefficients are mole ratios, not mass ratios; convert volume from cm^3 to dm^3 before using molar concentration.

CHAPTER 3 PRACTICE

Exam-style practice

1. Calculate the moles in 9.8 g of sulfuric acid, H_2SO_4. Use H = 1, S = 32, O = 16.

2. $2H_2 + O_2 \rightarrow 2H_2O$. What mass of water forms from 4 mol of hydrogen? $M_r(H_2O) = 18$.

3. 25.0 cm^3 of solution contains 0.005 mol solute. Calculate concentration in mol/dm^3.

CHAPTER 4

Chemical changes and energy

Reactivity, acids, electrolysis and energy profiles explain how and why chemical changes occur.

Coverage map

Reactivity series; oxidation and reduction; extraction; acids, alkalis and salts; electrolysis; exothermic and endothermic reactions; reaction profiles; cells and fuel cells.

Essential knowledge

- A more reactive metal displaces a less reactive metal from its compound.
- Oxidation is loss of electrons; reduction is gain of electrons.
- Acids form H+ ions in water; alkalis form OH- ions.
- Electrolysis uses electrical energy to decompose an ionic substance; reduction occurs at the cathode.
- Exothermic reactions transfer energy to surroundings; endothermic reactions take energy in.

Worked example	For Zn + CuSO4 → ZnSO4 + Cu, zinc is oxidised to Zn2+ and Cu2+ is reduced to copper. Zinc displaces copper because zinc is more reactive.
Exam focus	Track electron transfer or ion movement explicitly and use the signs of energy changes consistently.
Common mistake to avoid	At the cathode reduction occurs; at the anode oxidation occurs, regardless of whether the cell is molten or aqueous.

CHAPTER 4 PRACTICE

Exam-style practice

1. Write the ionic equation for neutralisation.

2. Predict the products of electrolysis of molten lead bromide and identify each electrode.

3. Bond breaking requires 700 kJ/mol and bond making releases 910 kJ/mol. Calculate the overall energy change.

CHAPTER 5

Rates, equilibrium and organic chemistry

Collision theory explains rate, while reversible reactions and carbon structures explain industrial and organic processes.

Coverage map

Measuring rate; collision theory; temperature, concentration, pressure and surface area; catalysts; reversible reactions; equilibrium; crude oil; alkenes, alcohols, carboxylic acids and polymers.

Essential knowledge

- Reaction rate increases when successful collisions occur more often.
- Higher temperature increases collision frequency and the fraction exceeding activation energy.
- A catalyst provides an alternative pathway with lower activation energy.
- At dynamic equilibrium, forward and reverse rates are equal in a closed system.
- Crude oil is separated by fractional distillation; cracking produces shorter alkanes and alkenes.

Worked example	Powdered calcium carbonate reacts faster than large chips because it has a larger surface area, so acid particles collide with carbonate more frequently.
Exam focus	Use collision theory for rate questions and identify the endothermic or exothermic direction before predicting an equilibrium shift.
Common mistake to avoid	A catalyst speeds up both directions and does not change the equilibrium position or overall energy change.

CHAPTER 5 PRACTICE

Exam-style practice

1. Explain how a catalyst changes a reaction profile without changing the reaction's overall energy change.

2. For an exothermic reversible reaction, predict the effect of increasing temperature on equilibrium yield.

3. Write the displayed change represented by cracking C10H22 → C8H18 + ? and identify the second product type.

CHAPTER 6

Chemical analysis, atmosphere and resources

Analytical tests identify substances, while environmental chemistry evaluates evidence, treatment and resource use.

Coverage map

Purity and formulations; chromatography; gas tests; ion tests; instrumental analysis; atmospheric change; pollutants; potable water; life-cycle assessment; recycling and the Haber process.

Essential knowledge

- A pure substance has a sharp melting or boiling point under fixed conditions.
- Chromatography separates substances by different attractions to stationary and mobile phases.
- Gas tests: hydrogen squeaky pop, oxygen relights a glowing splint, carbon dioxide turns limewater cloudy.
- Greenhouse gases absorb outgoing infrared radiation and affect Earth's energy balance.
- Life-cycle assessment compares environmental impacts across extraction, manufacture, use and disposal.

Worked example	Rf = distance moved by substance ÷ distance moved by solvent front. If a spot moves 3.2 cm and solvent moves 8.0 cm, Rf = 0.40.
Exam focus	Separate observation from inference and state enough procedural detail for a test to be repeatable.
Common mistake to avoid	An Rf value has no unit and is meaningful only when experimental conditions are comparable.

CHAPTER 6 PRACTICE

Exam-style practice

1. A chromatogram spot moves 4.5 cm while the solvent front moves 7.5 cm. Calculate Rf.

2. Describe how to produce potable water from fresh water containing suspended solids and microorganisms.

3. Explain why a life-cycle assessment may not give one unambiguous answer.

ANSWERS AND WORKED METHODS

Check the reasoning as well as the final answer

Award credit only for a method that answers the question asked. Where a response is unclear, return to the matching chapter and correct it before trying a fresh question.

Chapter 1: Atomic structure and the periodic table

1. There are two more protons than electrons, so the charge is 2+. Neutrons = 24 - 12 = 12.

2. The outer electron is further from the nucleus and more shielded by inner shells, so attraction to the nucleus is weaker and the electron is lost more easily.

3. Relative atomic mass = (0.75 × 35) + (0.25 × 37) = 26.25 + 9.25 = 35.5.

Chapter 2: Bonding, structure and properties

1. It has a giant ionic lattice with strong electrostatic attractions between Mg2+ and O2- ions in all directions. A large amount of energy is needed to overcome these attractions.

2. Graphite has one delocalised electron per carbon atom that can move through layers, so it conducts. In diamond all four outer electrons form covalent bonds, so there are no mobile charged particles.

3. Layers of positive metal ions can slide past each other while attraction to the sea of delocalised electrons remains, so the structure changes shape without shattering.

Chapter 3: Quantitative chemistry

1. Mr(H_2SO_4) = 2 + 32 + 64 = 98. Moles = 9.8 ÷ 98 = 0.10 mol.

2. The ratio H_2:H_2O is 2:2, so 4 mol H_2 forms 4 mol H_2O. Mass = 4 × 18 = 72 g.

3. 25.0 cm^3 = 0.0250 dm^3. Concentration = 0.005 ÷ 0.0250 = 0.20 mol/dm^3.

Chapter 4: Chemical changes and energy

1. $H^+ + OH^- \rightarrow H_2O$.

2. Pb^{2+} gains electrons at the negative cathode to form lead: $Pb^{2+} + 2e^- \rightarrow Pb$. Br^- loses electrons at the positive anode to form bromine: $2Br^- \rightarrow Br_2 + 2e^-$.

3. Energy change = energy to break bonds - energy released making bonds = 700 - 910 = -210 kJ/mol. The negative value means exothermic.

Chapter 5: Rates, equilibrium and organic chemistry

1. The catalyst lowers activation energy by providing another pathway, so both forward and reverse reactions can proceed faster. The energy difference between reactants and products is unchanged.

2. The endothermic reverse direction is favoured, so the equilibrium shifts towards reactants and the product yield decreases.

3. Balance atoms: $C_{10}H_{22} \rightarrow C_8H_{18} + C_2H_4$. C_2H_4 is ethene, an alkene with a carbon-carbon double bond.

Chapter 6: Chemical analysis, atmosphere and resources

1. Rf = 4.5 ÷ 7.5 = 0.60. It has no unit.

2. Allow suspended solids to settle or filter them out, then sterilise using chlorine, ozone or ultraviolet radiation. The exact process depends on the water source.

3. Some impacts are difficult to quantify, data may be incomplete, and weighting carbon emissions, water use, toxicity and habitat damage involves value judgements.