Chemistry with Chloe

Expert A Level Chemistry Tuition

OCR A Level Chemistry A

Component 01: Periodic Table, Elements and Physical Chemistry

H432/01 • 2 hours 15 minutes • 100 marks • 37% of A-level

COMPLETE REVISION GUIDE

Modules 2, 3 and 5

Foundations • Periodic Table & Energy • Physical Chemistry & Transition Elements

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How to Use This Guide

This revision guide covers every specification point examined in OCR A Level Chemistry A Component 01 (H432/01), including Module 2 (Foundations in Chemistry), Module 3 (Periodic Table and Energy) and Module 5 (Physical Chemistry and Transition Elements).

Module 1 (Practical Skills) is assessed across all papers – PAG practicals are included inline where they appear in the specification.

Each numbered box shows the spec point in italics, followed by a concise, exam-ready explanation. Key terms are highlighted in bold.

Use this guide alongside past papers.

Good luck! – Chloe

MODULE 2: FOUNDATIONS IN CHEMISTRY

2.1 Atoms and Reactions

2.1.1 Atomic Structure and Isotopes

2.1.1a	*Isotopes as atoms of the same element with different numbers of neutrons and different masses. Atomic structure: protons, neutrons, electrons for atoms and ions.*
	Isotopes: atoms of the **same element** (same number of protons/atomic number) with **different numbers of neutrons** (different mass numbers). Same chemical properties, different physical properties. **Protons = Z** (atomic number). **Neutrons = A – Z. Electrons = Z** for neutral atoms. For ions: subtract charge for cations (+), add for anions (–).

2.1.1c	*Relative isotopic mass and relative atomic mass, based on ^{12}C scale. Use of mass spectrometry for determination of isotopic masses and abundances.*
	Relative isotopic mass: mass of one atom of an isotope compared to 1/12 the mass of ^{12}C. **Relative atomic mass (Ar):** weighted mean mass of an atom compared to 1/12 the mass of ^{12}C. **Ar = Σ(isotopic mass × abundance) / Σ(abundances)** Mass spectrometry measures isotopic masses and relative abundances (limited to singly charged ions for OCR).

2.1.1e	*Relative molecular mass (Mr) and relative formula mass. Calculation from relative atomic masses.*
	Relative molecular mass (Mr): sum of Ar values for all atoms in a molecular formula (simple molecules). **Relative formula mass:** sum of Ar values in a formula unit (giant structures, e.g. ionic compounds like NaCl).

2.1.2 Compounds, Formulae and Equations

2.1.2a	*Writing formulae of ionic compounds from ionic charges. Predict charge from periodic table position. Know NO_3^-, CO_3^{2-}, SO_4^{2-}, OH^-, NH_4^+, Zn^{2+}, Ag^+.*
	Predict ionic charge from group number: Group 1 = +1, Group 2 = +2, Group 6 = –2, Group 7 = –1. **Ions to know:** nitrate NO_3^-, carbonate CO_3^{2-}, sulfate SO_4^{2-}, hydroxide OH^-, ammonium NH_4^+, zinc Zn^{2+}, silver Ag^+. Balance charges to write ionic formulae: e.g. $Mg^{2+} + 2Cl^- \rightarrow MgCl_2$.

2.1.2b	*Balanced chemical equations (including ionic), with state symbols, for familiar and unfamiliar reactions.*
	State symbols: (s) solid, (l) liquid, (g) gas, (aq) aqueous. **Ionic equations:** show only the species that change. Remove **spectator ions** (unchanged on both sides). Check: atoms balance AND charges balance.

2.1.3 Amount of Substance

2.1.3a	*Amount of substance, the mole, Avogadro constant (6.02×10^{23} mol^{-1}), molar mass, molar gas volume.*
	Mole (mol): unit for amount of substance. One mole = **6.02×10^{23} particles** (Avogadro constant, NA – given on Data Sheet). **Molar mass (M):** mass per mole (g mol^{-1}). Numerically equal to Ar or Mr. **Molar gas volume:** volume per mole of gas at RTP. Given on the Data Sheet (~24.0 dm^3 mol^{-1} at 298 K, 100 kPa).

2.1.3b-d	*Empirical and molecular formulae. Calculate from composition by mass or %. Hydrated salts and water of crystallisation.*
	Empirical formula: simplest whole number ratio. **Method:** divide mass/% by Ar → divide by smallest → whole numbers. **Molecular formula:** actual number of atoms. Mr ÷ empirical formula mass = multiplier.

	Hydrated salts: contain water of crystallisation (e.g. $CuSO_4 \cdot 5H_2O$). Heat to remove water (anhydrous = without water). Calculate from mass changes.

2.1.3e-g	*Calculations using moles: mass, gas volume, solution volume and concentration. Ideal gas equation pV = nRT. Stoichiometry.*
	Key formulae: **n = m/M** (moles from mass) **n = c × V** (moles from concentration × volume in dm^3) **n = V/Vm** (moles from gas volume ÷ molar gas volume) **pV = nRT** (p in Pa, V in m^3, R = 8.314 J K^{-1} mol^{-1} – given, T in K) **Concentration** in mol dm^{-3} and g dm^{-3} (g dm^{-3} = c × M). **Stoichiometry:** use mole ratios from balanced equations to link known and unknown quantities.

2.1.3h-j	*Percentage yield and atom economy. Techniques for measuring mass, solution volumes, gas volumes. Sustainability benefits of high atom economy.*
	% yield = (actual yield / theoretical yield) × 100% (always < 100%). **Atom economy = (Mr desired product / ΣMr all products) × 100%** High atom economy = less waste = more sustainable. Addition reactions have 100% atom economy. **Sustainability:** developing processes with high atom economy reduces waste and use of resources.

⭐ PAG1: Moles Determination

Method:

1. Determine the formula of a hydrated salt by heating a known mass of hydrated salt (e.g. $CuSO_4 \cdot nH_2O$) in a crucible to constant mass.
2. Weigh crucible empty, then with hydrated salt, then after heating (repeat heating until mass is constant).
3. Calculate: mass of water lost = mass before – mass after. Mass of anhydrous salt = mass after – mass of crucible.
4. Calculate moles of water and moles of anhydrous salt. Find the simplest ratio to determine n.
5. Alternative: determine empirical formula of MgO by burning Mg ribbon in a crucible with a lid.

Sources of Uncertainty:

- Balance: ±0.001 g
- Crucible may absorb heat unevenly
- Judgement of 'constant mass'

Common Errors:

- Not heating to constant mass (water not fully removed – mass of water too low)
- Salt spitting out of crucible (mass of salt too low)
- Not replacing lid on crucible during heating (MgO experiment: Mg smoke escapes)
- Oxidation of metal on the balance pan

Improvements:

- Heat, cool, weigh, reheat until two consecutive masses agree (constant mass)
- Use a lid on the crucible (lift occasionally to allow air in for MgO experiment)
- Use a desiccator to cool the crucible (prevents reabsorption of moisture)
- Handle crucible with tongs, not fingers

2.1.4 Acids

2.1.4a-c	*Common acids and alkalis. Strong and weak acids (qualitative). Neutralisation reactions: $H^+ + OH^- \rightarrow H_2O$. Acids with bases, carbonates, metal oxides, alkalis.*
	Common acids: HCl, H_2SO_4, HNO_3, CH_3COOH. **Common alkalis:** NaOH, KOH, NH_3. Acids release **H^+** in aqueous solution. Alkalis release **OH^-**. **Strong acid:** fully dissociated (e.g. $HCl \rightarrow H^+ + Cl^-$). **Weak acid:** partially dissociated (e.g. $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$). **Neutralisation:** $H^+ + OH^- \rightarrow H_2O$. Acid + base → salt + water. Acid + carbonate → salt + water + CO_2.

2.1.4d-e	*Preparing standard solutions and carrying out acid-base titrations. Structured and non-structured titration calculations.*
	Standard solution: known concentration, made by dissolving an accurately weighed mass in a volumetric flask.

	Titration method: pipette known volume of one solution → add indicator → titrate from burette → concordant results (within 0.10 cm^3) → mean titre. **Calculations:** mean titre → moles of known → mole ratio → moles of unknown → concentration.

⭐ PAG2: Acid–Base Titration

Method:

1. Prepare a standard solution: weigh solid acid accurately (±0.001 g) by difference. Dissolve in distilled water. Transfer quantitatively to a volumetric flask (250 cm^3). Make up to the mark with a dropping pipette. Stopper and invert.
2. Pipette 25.0 cm^3 of unknown solution into a conical flask. Add indicator.
3. Fill burette with standard solution. Record initial reading.
4. Titrate: add from burette, swirling, until end point. Record final reading.
5. Repeat for concordant titres (within 0.10 cm^3). Calculate mean titre.
6. Calculate: moles of known → mole ratio → moles of unknown → concentration.

Sources of Uncertainty:

- Balance: ±0.001 g
- Volumetric flask (250 cm^3): ±0.10 cm^3
- Pipette (25 cm^3): ±0.06 cm^3
- Burette: ±0.05 cm^3 per reading (±0.10 cm^3 per titre)

Common Errors:

- Not all solid transferred to flask (concentration too low)
- Overshooting end point
- Parallax error reading burette
- Rinsing conical flask with the solution it will contain (changes moles)

Improvements:

- Weigh by difference
- Rinse all glassware into volumetric flask
- Use dropping pipette for final approach to graduation mark
- Rinse burette with its solution; rinse pipette with its solution
- White tile under conical flask
- Add dropwise near end point
- Concordant titres within 0.10 cm^3

2.1.5 Redox

2.1.5a-c	*Oxidation number rules and calculations. Writing formulae using oxidation numbers. Roman numeral notation.*
	Oxidation number rules: uncombined elements = 0; monatomic ions = charge; O = −2 (peroxides = −1); H = +1 (metal hydrides = −1); F = −1 always. Sum in compound = 0; in ion = charge. **Roman numerals:** e.g. iron(III) chloride = $FeCl_3$ (Fe = +3). Can write formulae from names like chlorate(I) = ClO^-, chlorate(V) = ClO_3^-.

2.1.5d-f	*Oxidation and reduction: electron transfer and oxidation number changes. Redox reactions of metals with acids. Interpretation of unfamiliar redox reactions.*
	Oxidation = loss of electrons = increase in oxidation number (OIL RIG). **Reduction = gain of electrons = decrease in oxidation number.** **Metals + acids:** M + acid → salt + H_2 (limited to those producing salt + hydrogen). Interpret unfamiliar redox reactions by tracking oxidation number changes and electron transfer.

2.2 Electrons, Bonding and Structure

2.2.1 Electron Structure

2.2.1a-d	*Electron shells, sub-shells and orbitals. Shapes of s- and p-orbitals. Filling order. Electron configurations of atoms (up to Z=36) and s/p-block ions.*
	Shells contain sub-shells: s (1 orbital, $2e^-$), p (3 orbitals, $6e^-$), d (5 orbitals, $10e^-$). **s-orbital:** spherical. **p-orbital:** dumbbell/figure-of-eight. **Fill order:** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$. Orbitals fill singly before pairing (Hund's rule). Paired electrons have opposite spins. **Electron configurations:** use sub-shell notation (e.g. O = $1s^22s^22p^4$). For ions: remove/add electrons from the outermost shell.

2.2.2 Bonding and Structure

2.2.2a-c	*Ionic bonding as electrostatic attraction between ions. Dot-and-cross diagrams. Giant ionic lattice structure. Properties of ionic compounds.*
	Ionic bonding: strong electrostatic attraction between **oppositely charged ions**. Formed by electron transfer (metal → non-metal). **Giant ionic lattice:** ions strongly attracted in all directions (e.g. NaCl). **Properties:** high m.p./b.p. (strong ionic bonds); conduct when molten/dissolved (ions free to move); do not conduct when solid (ions fixed).

2.2.2d-f	*Covalent bonding: strong electrostatic attraction between shared pair and nuclei. Dot-and-cross diagrams (single, multiple, dative). Average bond enthalpy as a measure of bond strength.*
	Covalent bond: strong electrostatic attraction between the **shared pair of electrons** and the **nuclei** of the bonded atoms. **Multiple bonds:** double = 2 shared pairs; triple = 3. **Dative (coordinate):** both electrons from one atom (arrow from donor). **Average bond enthalpy:** larger value = stronger bond. Note: actual values vary between molecules.

2.2.2g-h	*Shapes of molecules and ions (up to 6 electron pairs) by electron pair repulsion. Bond angles. Lone pair repulsion greater than bonded pair.*
	VSEPR: electron pairs repel and arrange to be as far apart as possible. **Lone pair–lone pair > lone pair–bond pair > bond pair–bond pair.** **Linear:** 180° (2 pairs). **Trigonal planar:** 120° (3). **Tetrahedral:** 109.5° (4). **Pyramidal:** 107° (4, 1 lp). **Bent:** 104.5° (4, 2 lp). **Octahedral:** 90° (6). Lone pairs reduce bond angles by ~2.5° each.

2.2.2i-j	*Electronegativity (Pauling values). Polar bonds, permanent dipoles. Polar vs non-polar molecules based on dipoles and shape.*
	Electronegativity: ability of an atom to attract bonding electrons. Increases towards F in the periodic table. **Polar bond:** between atoms with different electronegativities (δ+ and δ–). **Polar molecule:** has an overall dipole (dipoles don't cancel). Symmetric molecules (e.g. CO_2, CCl_4) have polar bonds but are **non-polar overall**.

2.2.2k-o	*Intermolecular forces: permanent dipole–dipole, induced dipole–dipole (London forces). Hydrogen bonding (N, O, F). Anomalous water properties. Simple molecular structures and properties.*
	London forces (induced dipole–dipole): in ALL molecules. Increase with more electrons/higher Mr. **Permanent dipole–dipole:** between polar molecules. **Hydrogen bonding:** H bonded to **N, O, or F** + lone pair on adjacent N/O/F. Strongest intermolecular force. OCR note: both permanent dipole–dipole and London forces can be called **van der Waals' forces**. **Water anomalies:** high b.p. (H-bonds); ice less dense than water (open H-bonded lattice). **Simple molecular:** low m.p./b.p. (weak IMFs overcome); do not conduct (no free ions/electrons).

MODULE 3: PERIODIC TABLE AND ENERGY

3.1 The Periodic Table

3.1.1 Periodicity

3.1.1a	*Periodic table: arrangement by increasing atomic number. Periods show repeating trends (periodicity). Groups have similar properties.*
	Elements arranged by **increasing atomic (proton) number**. **Periods:** horizontal rows showing repeating patterns in properties. **Groups:** vertical columns with similar chemical properties (same outer shell electron configuration).

3.1.1b-c	*Periodic trend in electron configuration across Periods 2 and 3. s-, p-, d-block classification. First and successive ionisation energies. Trends across periods and down groups.*
	s-block: Groups 1–2. **p-block:** Groups 13–18. **d-block:** transition metals. **First IE definition:** energy to remove 1 mol of electrons from 1 mol of gaseous atoms. **Across Periods 2/3:** general increase (more protons, same shielding). Small drops at: **Group 13** (p vs s electron, easier to remove) and **Group 16** (paired p-electron repulsion, easier to remove). **Down a group:** decrease (more shells, greater shielding, electron further from nucleus). **Successive IEs:** large jump = electron removed from new inner shell → evidence for shell structure.

3.1.1d-g	*Metallic bonding and giant metallic lattice. Giant covalent structures (diamond, graphite, graphene, silicon). Properties of giant structures. Variation in melting points across Periods 2 and 3.*
	Metallic bonding: strong electrostatic attraction between cations and delocalised electrons. Giant metallic lattice. **Giant covalent:** diamond (4 bonds, very hard, doesn't conduct), graphite (3 bonds + delocalised e^-, conducts, layers slide), graphene (single layer, excellent conductor), silicon. **Melting points across Period 3:** Na < Mg < Al (metallic, strengthening). Si very high (giant covalent). P_4, S_8 = simple molecular (London forces). Cl_2, Ar very low. Trend: giant metallic → giant covalent → simple molecular.

3.1.2 Group 2

3.1.2a-c	*Group 2: outer shell s^2 configuration. Reactivity of Mg–Ba with oxygen, water, dilute acids. Trend in reactivity explained by ionisation energies.*
	Group 2 metals lose 2 outer s-electrons to form **2+ ions**. **Reactions:** with O_2 → MO; with H_2O → $M(OH)_2$ + H_2; with dilute acid → salt + H_2. **Reactivity increases** down the group: IE decreases (more shells, greater shielding) → easier to lose 2 electrons. Mg reacts slowly with cold water (vigorous with steam). Ca–Ba increasingly vigorous with cold water.

3.1.2d-e	*Action of water on Group 2 oxides and pH of solutions. Uses of Group 2 compounds as bases.*
	Oxides + water → hydroxides: MO + H_2O → $M(OH)_2$. Solutions are **alkaline** (pH increases down the group as hydroxides become more soluble). **Uses:** $Ca(OH)_2$ to neutralise acid soils in agriculture. $Mg(OH)_2$ and $CaCO_3$ as antacids for indigestion.

3.1.3 The Halogens

3.1.3a	*Halogens as diatomic molecules. Trend in boiling points of Cl_2, Br_2, I_2 in terms of London forces.*
	Halogens exist as **diatomic molecules** (Cl_2, Br_2, I_2). **Boiling point increases** down the group: more electrons → stronger **London forces** → more energy to overcome. Cl_2 = gas, Br_2 = liquid, I_2 = solid at RTP.

3.1.3b-d	*Halogen outer shell s^2p^5 configuration. Trend in reactivity (displacement reactions). Explanation in terms of atomic radius, electron shielding, attraction.*

	Halogens gain 1 electron to form **1– ions**. **Reactivity decreases** down the group: larger atomic radius, more shielding → nucleus attracts incoming electron less strongly. **Displacement:** a more reactive halogen displaces a less reactive halide. $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$. Observe colour changes in aqueous and organic layers.

3.1.3e-f	*Disproportionation of chlorine with water and with cold dilute NaOH. Benefits and risks of chlorine in water treatment.*
	$Cl_2 + H_2O \rightleftharpoons HCl + HOCl$ (disproportionation: Cl 0 → −1 and +1). HOCl kills bacteria. **$Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$** (bleach). **Benefits:** kills bacteria in water, prevents disease. **Risks:** Cl_2 is toxic, possible formation of chlorinated hydrocarbons. Society judges benefits outweigh risks.

3.1.3g	*Precipitation reactions of Cl^-, Br^-, I^- with aqueous silver ions, followed by aqueous ammonia, as a test for halide ions.*
	Add **$AgNO_3$(aq)** to the solution: **AgCl = white** precipitate (soluble in **dilute** NH_3). **AgBr = cream** precipitate (soluble in **concentrated** NH_3). **AgI = yellow** precipitate (**insoluble** in NH_3). Note for OCR: complexes with ammonia are not required beyond observations.

3.1.4 Qualitative Analysis

3.1.4a	*Qualitative analysis: tests for CO_3^{2-} (acid → CO_2), SO_4^{2-} (Ba^{2+} precipitation), Cl^-/Br^-/I^- ($AgNO_3$), NH_4^+ (warm NaOH → NH_3).*
	CO_3^{2-}: add dilute acid → effervescence, CO_2 turns limewater milky. **SO_4^{2-}:** add Ba^{2+}(aq) (from $BaCl_2$) → white precipitate of $BaSO_4$. **Halides:** $AgNO_3$(aq) → AgCl white, AgBr cream, AgI yellow. Confirm with NH_3. **NH_4^+:** warm with NaOH(aq) → NH_3 gas turns damp red litmus paper blue. **Test sequence:** carbonate first, then sulfate, then halide ($BaCO_3$ and Ag_2SO_4 are both insoluble).

⭐ PAG4: Qualitative Analysis of Ions

Method:

1. Test for CO_3^{2-}: add dilute acid. Effervescence → bubble gas through limewater. Turns milky = CO_2 = carbonate confirmed.
2. Test for SO_4^{2-}: add dilute HCl (to remove CO_3^{2-} which would interfere), then add $BaCl_2$(aq). White precipitate of $BaSO_4$ = sulfate confirmed.
3. Test for halides: add dilute HNO_3, then $AgNO_3$(aq). White = Cl^-, cream = Br^-, yellow = I^-. Add NH_3: AgCl dissolves in dilute NH_3, AgBr in conc. NH_3, AgI insoluble.
4. Test for NH_4^+: add NaOH(aq), warm gently. Test gas with damp red litmus → turns blue = NH_3 = ammonium confirmed.
5. OCR test sequence: carbonate FIRST, then sulfate, then halide ($BaCO_3$ and Ag_2SO_4 are both insoluble).

Sources of Uncertainty:

- Visual judgement of precipitate colours
- Subjectivity of colour matching

Common Errors:

- Not testing in correct sequence (false positives from interfering ions)
- Not acidifying before $BaCl_2$ test ($BaCO_3$ gives false positive for sulfate)
- Using HCl instead of HNO_3 before $AgNO_3$ test (adds Cl^- ions)
- Confusing cream AgBr with white AgCl

Improvements:

- Follow the OCR test sequence: carbonate → sulfate → halide
- Use dilute HCl before $BaCl_2$; use dilute HNO_3 before $AgNO_3$
- Use a white tile for colour comparison
- Clean test tubes between tests to avoid cross-contamination

3.2 Physical Chemistry

3.2.1 Enthalpy Changes

3.2.1a-d	*Exothermic and endothermic reactions. Enthalpy profile diagrams. Activation energy. Standard conditions (100 kPa, 298 K). Definitions of ΔrH, ΔfH, ΔcH, ΔneutH.*
	Exothermic: ΔH negative (products lower). **Endothermic:** ΔH positive (products higher). **Activation energy (Ea):** minimum energy for reaction to occur. Shown on enthalpy profile diagrams. **Standard conditions:** 100 kPa, 298 K. Standard states = physical state under these conditions. **ΔrH:** enthalpy change for a stated equation. **ΔfH:** formation of 1 mol from elements in standard states (= 0 for elements). **ΔcH:** complete combustion of 1 mol. **ΔneutH:** formation of 1 mol of water from neutralisation. OCR requires definitions of ΔfH, ΔcH, ΔneutH.

3.2.1e	*Determination of enthalpy changes from experimental results using q = mcΔT.*
	q = mcΔT where q = energy (J), m = mass of solution (g), c = specific heat capacity (given), ΔT = temperature change. **ΔH = –q/n** (kJ mol^{-1}). Negative sign for direction. Assumptions: no heat loss, solution has density/c of water.

3.2.1f	*Average bond enthalpy. Exothermic/endothermic in terms of bond breaking and making. Calculations using average bond enthalpies.*
	Average bond enthalpy: energy to break 1 mol of a specific bond in gaseous molecules, averaged across different molecules. **ΔH = Σ(bonds broken) – Σ(bonds formed)** Bonds broken = endothermic (+). Bonds formed = exothermic (–). Actual bond enthalpies may differ from average values (limitations).

3.2.1g-h	*Hess' law: enthalpy cycles using ΔcH and ΔfH. Direct and indirect determination of enthalpy changes.*
	Hess' law: total ΔH is independent of the route. **Using ΔfH:** ΔrH = ΣΔfH(products) – ΣΔfH(reactants) **Using ΔcH:** ΔrH = ΣΔcH(reactants) – ΣΔcH(products) Can also use unfamiliar enthalpy cycles (provided in exam).

⭐ PAG3: Enthalpy Determination

Method:

1. CALORIMETRY: measure a known volume of solution into a polystyrene cup. Record initial temperature (every minute for 2–3 min baseline). Add the second reagent at a known time. Stir, record temperature every minute for 5–6 min.
2. Plot temperature vs time. Extrapolate both lines back to time of mixing to find true ΔT.
3. Calculate q = mcΔT, then ΔH = –q/n (kJ mol^{-1}).
4. HESS' LAW: measure ΔH for two reactions that share a common starting or finishing point (e.g. $NaHCO_3$ + HCl and Na_2CO_3 + HCl). Use the cycle to calculate ΔH for the target reaction indirectly.

Sources of Uncertainty:

- Thermometer: ±0.5°C (or ±0.1°C digital)
- Measuring cylinder: ±0.5 cm^3
- Balance: ±0.001 g

Common Errors:

- Heat loss to surroundings (systematic – ΔT too small)
- Assuming c and density of solution = water
- Incomplete reaction
- Not using extrapolation method (simple max/min T underestimates ΔT)

Improvements:

- Use polystyrene cup with lid to reduce heat loss
- Use plotting/extrapolation method to correct for heat loss
- Use digital thermometer for precision
- Stir continuously
- Use excess of solid reagent to ensure complete reaction of solution
- Repeat and take mean

3.2.2 Reaction Rates

3.2.2a-b	*Effect of concentration/pressure on rate (collision frequency). Calculation of reaction rate from gradients of graphs.*
	Increasing concentration/pressure: more particles per unit volume → more frequent collisions → faster rate. **Rate from graph:** draw a **tangent** to the concentration–time curve and calculate its **gradient**. Units depend on axes (e.g. mol dm^{-3} s^{-1}).

3.2.2c-e	*Catalysts: increase rate via lower Ea. Homogeneous vs heterogeneous. Economic importance and sustainability. Techniques for investigating rates.*
	Catalyst: increases rate by providing an **alternative pathway with lower Ea**. Not consumed. **Homogeneous:** same phase as reactants. **Heterogeneous:** different phase (usually solid catalyst, gaseous/liquid reactants). **Economic benefits:** lower temperatures/pressures → reduced energy costs and CO_2 emissions. Improved sustainability.

⭐ PAG9: Rates of Reaction

Method:

1. Choose a suitable method to follow the reaction: mass loss (balance), gas volume (gas syringe), colour change (colorimeter), time to form precipitate (disappearing cross).
2. Set up the reaction with measured quantities. Start the timer on mixing.
3. Record the measured quantity at regular time intervals.
4. Plot a graph of the quantity vs time. Calculate rate from the gradient (tangent).
5. Investigate the effect of a variable (e.g. concentration, temperature, catalyst) by repeating with one change at a time, keeping all other variables constant.

Sources of Uncertainty:

- Balance: ±0.01 g
- Gas syringe: ±0.5 cm^3
- Stopwatch: ±0.5 s
- Thermometer: ±0.5°C

Common Errors:

- Temperature not controlled between experiments
- Gas escaping before apparatus connected
- Not keeping total volume constant when varying concentration
- Inconsistent mixing

Improvements:

- Use a water bath for temperature control
- Connect apparatus quickly; check for gas leaks
- Keep total volume constant by adding distilled water
- Use consistent mixing technique
- Repeat and take means

3.2.2f-g	*Boltzmann distribution and its relationship with Ea. Effect of temperature changes and catalysts on the distribution and rate.*
	Boltzmann distribution: shows range of molecular energies. Peak = most probable energy. Area under curve = total molecules (constant). **Higher temperature:** peak shifts right, becomes lower and broader. Area beyond Ea **increases significantly** → much faster rate. Main factor: greater proportion with E ≥ Ea. **Catalyst:** distribution unchanged; Ea moves **left** (lower) → greater proportion exceeds new Ea → faster rate.

3.2.3 Chemical Equilibrium

3.2.3a-e	*Dynamic equilibrium (closed system). Le Chatelier's principle. Catalyst unchanged position. Compromise conditions for industry. Techniques for investigating equilibrium.*
	Dynamic equilibrium (closed system): forward rate = reverse rate; concentrations constant but NOT equal. **Le Chatelier's:** system shifts to **oppose** the change. Increase T → favours endothermic. Increase P → favours fewer moles of gas. Increase [reactant] → shifts right. **Catalyst:** does NOT change position of equilibrium. Speeds up both forward and reverse equally. Equilibrium reached **faster**. **Industrial compromise:** e.g. Haber process: moderate T (~450°C) for rate, high P (~200 atm) for yield, Fe catalyst.

3.2.3f-g	*Kc expressions for homogeneous reactions. Calculate Kc from equilibrium concentrations. Estimate position from magnitude of Kc.*
	$Kc = [C]^c[D]^d / [A]^a[B]^b$ (products over reactants, raised to stoichiometric powers). OCR note: learners will **not** need to determine units for Kc at this stage. **Large Kc** ($\gg 1$) = equilibrium lies to the right (products favoured). **Small Kc** ($\ll 1$) = lies to the left (reactants favoured). Qualitative estimation only.

MODULE 5: PHYSICAL CHEMISTRY AND TRANSITION ELEMENTS

5.1 Rates, Equilibrium and pH

5.1.1 How Fast? (Rate Equations)

5.1.1a-c	*Rate of reaction, order, overall order, rate constant, half-life, rate-determining step. Deduce orders from data. Rate equation rate = $k[A]^m[B]^n$. Calculate k and units.*
	Rate equation: rate = $k[A]^m[B]^n$. Determined **experimentally only**. **Order** = power to which concentration is raised (0, 1 or 2). **Overall order** = m + n. **Rate constant (k):** increases with temperature. **Units of k:** zero order = mol dm^{-3} s^{-1}; first = s^{-1}; second = dm^3 mol^{-1} s^{-1}. **Half-life (t½):** time for concentration to halve. Constant for **first order** reactions. **Determining order:** double [A]: no effect = 0; rate doubles = 1; rate quadruples = 2.

5.1.1d-g	*Concentration–time graphs: deduce order (0 or 1). Rate from gradients. Half-life for first order (constant). k = ln 2/t½. Rate–concentration graphs: deduce order (0, 1, 2). k for first order from gradient.*
	Conc–time graphs: zero order = straight line; first order = exponential decay (constant t½). **Rate from graph:** gradient of tangent at time t. **First order:** t½ is constant. **$k = \ln 2 / t_{½}$** (OCR: learners won't derive this but must use it). **Rate–conc graphs:** zero = horizontal; first = straight through origin; second = curve through origin. For first order, gradient of rate–conc graph = k.

5.1.1h-i	*Techniques for initial rates and continuous monitoring (including colorimetry). Rate-determining step: predict rate equation from mechanism and vice versa.*
	Initial rates: separate experiments, different initial concentrations. Clock reactions approximate this. **Continuous monitoring:** follow concentration over time (mass loss, gas volume, colorimetry). **Rate-determining step (rds):** the slowest step. Species in the rate equation appear in or before the rds. If rate = k[A][B] → rds involves one A and one B. Proposed mechanism must sum to overall equation.

⭐ PAG10: Initial Rates and Continuous Monitoring

Method:

1. INITIAL RATES (e.g. iodine clock): prepare flasks with different concentrations of one reactant. Keep total volume and other concentrations constant.
2. Add $Na_2S_2O_3$ + starch. Add final reagent, start timer, swirl. Record time for blue/black colour. Rate ≈ 1/time.
3. Plot rate vs [varied reactant] to determine order.
4. CONTINUOUS MONITORING (e.g. colorimetry): mix reagents, start timer. Use a colorimeter to measure absorbance at regular intervals.
5. Plot absorbance (or concentration) vs time. Draw tangents to find rate at different times. Determine order from graph shape.

Sources of Uncertainty:

- Stopwatch: ±0.5 s
- Measuring cylinders: ±0.5 cm^3
- Colorimeter calibration

Common Errors:

- Temperature not controlled
- Not keeping total volume constant
- Inconsistent timing/mixing
- Colorimeter not calibrated properly

Improvements:

- Thermostatted water bath
- Constant total volume (add water to compensate)
- Consistent swirling/mixing
- Calibrate colorimeter with known standards
- Repeat and take means

5.1.1j-k	*Qualitative effect of temperature on rate constant. Arrhenius equation: $k = Ae^{-Ea/RT}$. Graphical determination of Ea using ln k = –Ea/RT + ln A.*
	Increasing temperature increases k (rate constant): more molecules exceed Ea. **Arrhenius equation:** $k = Ae^{-Ea/RT}$ (A = pre-exponential factor, Ea = activation energy, R = 8.314, T in K). Rearranged: **ln k = –Ea/RT + ln A** (y = mx + c). Plot **ln k vs 1/T** → straight line, gradient = **–Ea/R**. So **Ea = –gradient × R**. Equations given on Data Sheet.

5.1.2 How Far? (Equilibrium Kc and Kp)

5.1.2a-e	*Mole fraction and partial pressure. Kc and Kp for homogeneous and heterogeneous equilibria (including units). Calculations.*
	Mole fraction = moles of A / total moles. **Partial pressure** = mole fraction × total pressure. **$Kp = p(C)^c \times p(D)^d / p(A)^a \times p(B)^b$** (gases only). **Heterogeneous equilibria:** omit pure solids and liquids from Kc and Kp expressions. OCR: learners **will** need to determine units for Kc and Kp at A level. Learners will not solve quadratics.

5.1.2f-h	*Effect of temperature on Kc and Kp. Constancy with concentration/pressure/catalyst changes. How K controls position of equilibrium.*
	Only temperature changes K. Exothermic forward: increase T → K decreases. Endothermic forward: increase T → K increases. Concentration, pressure and catalysts do **not** change K. The system adjusts to restore the ratio. **K controls position:** if conditions change, concentrations shift until the Kc/Kp expression is satisfied again.

5.1.3 Acids, Bases and Buffers

5.1.3a-c	*Brønsted–Lowry acids and bases. Conjugate pairs. Monobasic, dibasic, tribasic. Ka and pKa.*
	Acid: proton donor. **Base:** proton acceptor. **Conjugate pair:** differs by one H^+. **Monobasic:** donates 1 H^+ (e.g. HCl). **Dibasic:** 2 H^+ (e.g. H_2SO_4). **Tribasic:** 3 H^+ (e.g. H_3PO_4). **Ka** = acid dissociation constant (larger Ka = stronger weak acid). **pKa = –log Ka** (lower pKa = stronger).

5.1.3d-h	*pH = –log[H^+]. Kw. pH of strong acids, strong bases, weak acids (with approximations). Limitations of approximations.*
	pH = –log[H^+] and **[H^+] = 10^{-pH}**. **Kw = [H^+][OH^-] = 1.00×10^{-14}** (at 25°C). Kw increases with T (endothermic). **Strong acid pH:** [H^+] = concentration (fully dissociated). **Strong base pH:** [H^+] = Kw / [OH^-]. **Weak acid pH:** approximations: [H^+] ≈ [A^-] and [HA] ≈ initial concentration. **[H^+] = $\sqrt{(Ka \times [HA])}$**. **Limitations:** approximations break down for "stronger" weak acids where [H^+] is not negligible compared to [HA].

5.1.3i-m	*Buffer solutions: definition, formation, action. Buffer pH calculations. Blood carbonate buffer system.*
	Buffer: minimises pH change on addition of small amounts of acid or base. **Formation:** weak acid + its salt (e.g. CH_3COOH + CH_3COONa), OR excess weak acid + strong alkali. **Action:** added H^+ reacts with A^-; added OH^- reacts with HA. Large reservoir of both. **pH calculation:** [H^+] = Ka × [HA]/[A^-], then pH = –log[H^+]. **Blood buffer:** H_2CO_3/HCO_3^- system maintains blood pH 7.35–7.45.

5.1.3n-o	*pH titration curves for all combinations. Indicator selection (pH range within steep section). Indicator equilibrium (HA/A^- colour change). pH meter techniques.*
	Strong/strong: sharp vertical, equivalence pH = 7. **Weak acid/strong base:** equivalence pH > 7, buffer region. **Strong acid/weak base:** equivalence pH < 7. **Weak/weak:** no sharp change, no suitable indicator. **Indicator:** a weak acid, HA, with different colours for HA and A^- forms. Colour changes when equilibrium shifts.

	Selection: pKa of indicator must fall within the steep section. Phenolphthalein (~8–10); methyl orange (~3–5).

⭐ PAG11: pH Measurement and Titration Curves

Method:

1. Calibrate a pH meter using buffer solutions (e.g. pH 4.00 and pH 7.00). Rinse electrode with distilled water between calibrations.
2. Pipette 25.0 cm^3 of acid into a beaker. Record initial pH.
3. Fill burette with base. Add in small increments (1 cm^3), recording pH after each addition. Swirl/stir after each addition.
4. Near equivalence point, use very small increments (0.1–0.2 cm^3).
5. Continue well past equivalence. Plot pH vs volume of base added.
6. Identify: equivalence point (steepest section), buffer region (flat section before equivalence for weak acid/strong base), pH at half-equivalence = pKa.

Sources of Uncertainty:

- pH meter: ±0.01 pH
- Burette: ±0.05 cm^3 per reading
- Pipette: ±0.06 cm^3

Common Errors:

- pH meter not calibrated
- Large increments near equivalence (miss steep section)
- Not mixing after each addition
- Temperature changes affecting Kw and pH

Improvements:

- Calibrate with two buffer solutions immediately before use
- Use 0.1–0.2 cm^3 increments near equivalence point
- Stir/swirl thoroughly and wait for reading to stabilise
- Record temperature
- Draw a smooth curve, not straight lines between points

5.2 Energy

5.2.1 Lattice Enthalpy and Born–Haber Cycles

5.2.1a-b	*Lattice enthalpy (formation of 1 mol of ionic lattice from gaseous ions). Born–Haber cycles and calculations.*
	Lattice enthalpy (ΔLEH): enthalpy change when 1 mol of ionic lattice is formed from gaseous ions. Always **exothermic** (OCR uses formation definition). **Born–Haber cycle:** uses Hess' law. Links: ΔfH, ΔatH (atomisation), IE (ionisation energy), EA (electron affinity), ΔLEH. Rearrange to find the unknown step. Watch signs carefully.

5.2.1c-e	*Enthalpy of solution (ΔsolH) and hydration (ΔhydH). Enthalpy cycles for dissolution. Effect of ionic charge and radius on lattice enthalpy and hydration enthalpy.*
	ΔsolH: dissolving 1 mol of solute (can be exo- or endothermic). **ΔhydH:** dissolving 1 mol of gaseous ions in water (always exothermic – ion-dipole bonds form). **Cycle:** ΔsolH = –ΔLEH + ΣΔhydH (or using lattice dissociation: ΔsolH = +ΔLEH(dissociation) + ΣΔhydH). **Effect of charge/radius:** higher charge and smaller radius → more exothermic lattice enthalpy and hydration enthalpy (stronger attraction).

5.2.2 Enthalpy and Entropy

5.2.2a-c	*Entropy as dispersal of energy/disorder. Entropy differences: solids < liquids < gases. Change with number of gaseous molecules. Calculate ΔS from entropies.*
	Entropy (S): measure of disorder/dispersal of energy. Units: J K^{-1} mol^{-1}. S increases: solid → liquid → gas. S increases when more moles of gas are produced. **$\Delta S_{system} = \Sigma S^{\ominus}(products) - \Sigma S^{\ominus}(reactants)$**

5.2.2d-f	*Feasibility depends on ΔH and TΔS. Gibbs free energy: ΔG = ΔH – TΔS. Feasible when ΔG < 0. Limitations (kinetics).*
	ΔG = ΔH – TΔS (Gibbs' equation). Feasible when **ΔG < 0** (negative). **T for feasibility:** set ΔG = 0, T = ΔH/ΔS. **UNITS:** ΔH in kJ, ΔS in J K^{-1} → must convert ΔS to kJ K^{-1} (÷ 1000). **Limitations:** ΔG < 0 does NOT guarantee reaction occurs – kinetic barriers (high Ea) may prevent it.

5.2.3 Redox and Electrode Potentials

5.2.3a-c	*Oxidising/reducing agents. Construct redox equations from half-equations and oxidation numbers. Interpret and predict electron transfer reactions.*
	Oxidising agent: gains electrons (is reduced). **Reducing agent:** loses electrons (is oxidised). Construct redox equations: write half-equations, balance electrons, combine. Use oxidation numbers to track changes.

5.2.3d-e	*Redox titrations: Fe^{2+}/MnO_4^- and I_2/$S_2O_3^{2-}$. Structured and non-structured calculations.*
	MnO_4^-/Fe^{2+}: $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$. Ratio 1:5. Self-indicating (purple → pink end point). Use H_2SO_4. **I_2/$S_2O_3^{2-}$:** $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$. Ratio 1:2. Starch indicator (blue → colourless).

5.2.3f-i	*Standard electrode potential (E⊖). Measurement with hydrogen electrode. Cell potential by combining E⊖ values. Feasibility predictions and limitations.*
	E⊖: measured under standard conditions (298 K, 100 kPa, 1.00 mol dm^{-3}) relative to **standard hydrogen electrode (SHE) = 0.00 V**. Measure using Pt electrode for ions in solution, or metal in contact with its ions. Salt bridge connects half-cells. **EMF = E⊖(cathode) – E⊖(anode)** = E⊖(more positive) – E⊖(less positive). Positive EMF → feasible. **Limitations:** predicts thermodynamics only (not rate); non-standard conditions change E.

⭐ PAG8: Electrochemical Cells

Method:

1. Set up half-cells: e.g. Zn in $ZnSO_4$(aq) and Cu in $CuSO_4$(aq), each 1.00 mol dm^{-3}.
2. Connect with a salt bridge (filter paper soaked in KNO_3).
3. Connect electrodes to a high-resistance voltmeter.
4. Record voltage and polarity (+/–). Identify cathode and anode.
5. Repeat with different half-cell combinations.
6. For ions in different oxidation states: use a Pt electrode in a solution containing both ions.
7. Compare experimental values with calculated E⊖cell from data.

Sources of Uncertainty:

- Voltmeter: ±0.01 V
- Concentrations may not be exactly 1.00 mol dm^{-3}
- Temperature may differ from 298 K

Common Errors:

- Oxide layer on metal electrodes (lower voltage)
- Poor salt bridge contact
- Solutions not at standard concentration
- Drawing current (internal resistance)

Improvements:

- Polish/sand electrodes before use
- Ensure good salt bridge contact with both solutions
- Use high-resistance voltmeter
- Use freshly prepared 1.00 mol dm^{-3} solutions
- Record temperature
- Repeat and take means

5.2.3j-k	*Electrode potentials applied to storage cells. Fuel cells: energy from fuel + O_2 to generate voltage.*
	Storage cells: reversible redox reactions. Discharge = spontaneous. Charge = external voltage reverses reaction. **Fuel cells:** convert chemical energy of fuel (e.g. H_2) + O_2 directly to electrical energy. More efficient than combustion. Continuous supply needed. OCR: details and equations for storage cells and fuel cells will be **provided** in the exam.

5.3 Transition Elements

5.3.1 Properties, Ligands, Complexes and Reactions

5.3.1a-c	*d-block electron configurations (Sc–Zn). Transition elements = Ti–Cu (incomplete d-subshell in at least one ion). Variable oxidation states, coloured ions, catalytic activity.*
	Electron configs: fill 4s before 3d. E.g. Fe = [Ar] $3d^6$ $4s^2$. For ions: remove **4s first**. **Transition metal:** forms at least one ion with an **incomplete d-subshell**. Sc (d^0 in Sc^{3+}) and Zn (d^{10} in Zn^{2+}) are **NOT** transition metals. **Properties:** variable oxidation states, coloured compounds, catalytic activity, complex ion formation.

5.3.1d-g	*Ligands (coordinate bonding). Complex ions and coordination number. Shapes: octahedral (6), tetrahedral (4), square planar (4). Stereoisomerism: cis-trans and optical. Cisplatin.*
	Ligand: molecule/ion with a lone pair that donates to a metal ion (dative bond). **Monodentate:** H_2O, Cl^-, NH_3. **Bidentate:** en ($NH_2CH_2CH_2NH_2$). **Complex ion:** central metal ion + ligands. **Coordination number** = number of dative bonds. **Octahedral** (6 ligands, 90°): e.g. $[Cu(H_2O)_6]^{2+}$. **Tetrahedral** (4 large, 109.5°): e.g. $[CuCl_4]^{2-}$. **Square planar** (4, 90°): e.g. cisplatin $[PtCl_2(NH_3)_2]$. **Cis-trans isomerism** in square planar and octahedral complexes. **Optical isomerism** with bidentate ligands. **Cisplatin:** cis isomer is anticancer drug (binds to DNA, prevents cell division).

5.3.1h-i	*Ligand substitution with colour changes: $[Cu(NH_3)_4(H_2O)_2]^{2+}$ and $[CuCl_4]^{2-}$ from $[Cu(H_2O)_6]^{2+}$. $[Cr(NH_3)_6]^{3+}$ from $[Cr(H_2O)_6]^{3+}$. Haemoglobin: O_2 and CO substitution.*
	$[Cu(H_2O)_6]^{2+}$ (pale blue) + $4NH_3$ → **$[Cu(NH_3)_4(H_2O)_2]^{2+}$** (deep blue) + $4H_2O$. **$[Cu(H_2O)_6]^{2+}$** + $4Cl^-$ → **$[CuCl_4]^{2-}$** (yellow/green) + $6H_2O$ (octahedral → tetrahedral). **$[Cr(H_2O)_6]^{3+}$** (green/violet) + $6NH_3$ → **$[Cr(NH_3)_6]^{3+}$** (purple) + $6H_2O$. **Haemoglobin:** Fe^{2+} complex. O_2 binds reversibly. **CO binds more strongly** → carboxyhaemoglobin (toxic, prevents O_2 transport).

5.3.1j	*Precipitation reactions of Cu^{2+}, Fe^{2+}, Fe^{3+}, Mn^{2+}, Cr^{3+} with NaOH(aq) and NH_3(aq), including excess.*
	With NaOH(aq): Cu^{2+} → $Cu(OH)_2$ blue (insoluble in excess). Fe^{2+} → $Fe(OH)_2$ green (insoluble). Fe^{3+} → $Fe(OH)_3$ brown (insoluble). Mn^{2+} → $Mn(OH)_2$ pale brown/cream (insoluble). Cr^{3+} → $Cr(OH)_3$ green (**dissolves in excess NaOH** → $[Cr(OH)_6]^{3-}$). **With excess NH_3:** $Cu(OH)_2$ dissolves → deep blue $[Cu(NH_3)_4(H_2O)_2]^{2+}$. $Cr(OH)_3$ dissolves → purple $[Cr(NH_3)_6]^{3+}$. Others insoluble.

5.3.1k-l	*Redox reactions and colour changes: Fe^{2+}/Fe^{3+} interconversions, $Cr^{3+}/Cr_2O_7^{2-}$ interconversions, $Cu^{2+}/Cu^+/Cu$. Prediction of unfamiliar reactions.*
	Fe^{2+} → Fe^{3+}: oxidised by H^+/MnO_4^- (pale green → yellow). **Fe^{3+} → Fe^{2+}:** reduced by I^-. **Cr^{3+} → $Cr_2O_7^{2-}$:** oxidised by H_2O_2/OH^- (green → orange). **$Cr_2O_7^{2-}$ → Cr^{3+}:** reduced by Zn/H^+ (orange → green). **Cu^{2+} → Cu^+:** reduced by I^-. **Cu^+ disproportionates:** $2Cu^+$ → Cu^{2+} + Cu. OCR: learners won't recall equations but must construct them from half-equations and oxidation numbers.

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