Chemistry with Chloe

Expert A Level Chemistry Tuition

Edexcel A Level Chemistry

Paper 1 Revision Guide: Advanced Inorganic and Physical Chemistry

9CH0/01

TOPICS 1–5, 8, 10-15

Atomic Structure • Bonding & Structure • Redox I

Inorganic Chemistry • Formulae, Equations & Amounts

Energetics I • Equilibrium I • Equilibrium II • Acid-Base Equilibria

Energetics II • Redox II • Transition Metals

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How to Use This Guide

This revision guide covers every specification point for Topics 1–5, 8, 10, 11, 12, 13, 14 and 15 of the Edexcel A Level Chemistry Paper 1 exam (9CH0/01).

Each numbered box shows the official spec point in italics, followed by a concise, exam-ready explanation below. Key terms and phrases are highlighted in bold – these are the words that examiners are looking for in your answers.

Core Practicals are presented in orange boxes with step-by-step methods, sources of uncertainty, common errors, and suggested improvements – everything you need for the 6-mark practical questions.

Use this guide alongside past papers. Tick off spec points as you revise them, and focus your efforts on any areas you find challenging.

Good luck with your revision! – Chloe

Topic 1: Atomic Structure and the Periodic Table

1.1	*Know the structure of an atom in terms of electrons, protons and neutrons.*
	Atoms have a **small, dense, positively charged nucleus** containing protons and neutrons, surrounded by electrons in **quantum shells**. Most of the atom is empty space.

1.2	*Know the relative mass and relative charge of protons, neutrons and electrons.*
	Proton: relative mass = 1, relative charge = +1 **Neutron:** relative mass = 1, relative charge = 0 **Electron:** relative mass = 1/1836 (≈0), relative charge = −1

1.3	*Know what is meant by the terms 'atomic (proton) number' and 'mass number'.*
	Atomic (proton) number (Z): the number of protons in the nucleus of an atom. This defines the element. **Mass number (A):** the total number of protons and neutrons in the nucleus.

1.4	*Be able to determine the number of each type of sub-atomic particle in an atom, molecule or ion from the atomic (proton) number and mass number.*
	Protons = Z (atomic number) **Neutrons = A – Z** (mass number minus atomic number) **Electrons = Z** for a neutral atom. For ions: subtract the charge for cations, add the charge for anions.

1.5	*Understand the term 'isotopes'.*
	Isotopes are atoms of the **same element** (same number of protons) with **different numbers of neutrons**, and therefore different mass numbers. They have identical chemical properties but different physical properties (e.g. density, rate of diffusion).

1.6	*Be able to define the terms 'relative isotopic mass' and 'relative atomic mass', based on the ^{12}C scale.*
	Relative isotopic mass: the mass of one atom of an isotope compared to 1/12 of the mass of one atom of carbon-12. **Relative atomic mass (Ar):** the weighted mean mass of one atom of an element compared to 1/12 of the mass of one atom of carbon-12.

1.7	*Understand the terms 'relative molecular mass' and 'relative formula mass', including calculating these values from relative atomic masses.*
	Relative molecular mass (Mr): the sum of the relative atomic masses of all atoms in a molecular formula. Used for simple covalent molecules. **Relative formula mass:** the sum of the relative atomic masses of all atoms in a formula unit. Used for compounds with **giant structures** (e.g. ionic compounds like NaCl).

1.8	*Be able to analyse and interpret data from mass spectrometry to calculate relative atomic mass from relative abundance of isotopes and vice versa.*
	Ar = Σ(isotopic mass × relative abundance) / Σ(relative abundances) In a mass spectrum, the x-axis is m/z and the y-axis is relative abundance (%). Each peak represents an isotope. Multiply each mass by its abundance, sum, and divide by total abundance.

1.9	*Be able to predict the mass spectra, including relative peak heights, for diatomic molecules, including chlorine.*
	For Cl_2: chlorine has two isotopes, ^{35}Cl (75%) and ^{37}Cl (25%). Possible combinations: ^{35}Cl–^{35}Cl (m/z = 70), ^{35}Cl–^{37}Cl (m/z = 72), ^{37}Cl–^{37}Cl (m/z = 74). **Peak height ratio = 9 : 6 : 1** (using probability: $0.75^2 : 2\times0.75\times0.25 : 0.25^2$).

1.10	*Understand how mass spectrometry can be used to determine the relative molecular mass of a molecule.*
	The **molecular ion peak (M^+)** is the peak with the highest m/z value (ignoring the M+1 peak). This gives the **relative molecular mass** of the molecule directly. The M^+ peak is formed when the molecule loses one electron in the mass spectrometer.

1.11	*Be able to define the terms 'first ionisation energy' and 'successive ionisation energies'.*
	First ionisation energy: the energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of gaseous 1+ ions. • e.g. $Na(g) \rightarrow Na^+(g) + e^-$ **Successive ionisation energies:** the energies required to remove each subsequent electron from the increasingly positive ion.

1.12	*Understand how ionisation energies are influenced by the number of protons, the electron shielding and the electron sub-shell from which the electron is removed.*
	Ionisation energy **increases** with: • **More protons** in the nucleus (greater nuclear attraction) • **Less shielding** from inner electrons (outer electron feels greater effective nuclear charge) • **Electron removed from a sub-shell closer to the nucleus** (e.g. s < p in the same shell for removal difficulty) Electrons in the same sub-shell experience similar shielding.

1.13	*Understand reasons for the general increase in first ionisation energy across a period.*
	Across a period, **nuclear charge increases** (more protons) but **shielding remains approximately constant** (electrons are added to the same shell). This means the **effective nuclear charge increases**, so outer electrons are held more tightly, and **more energy is required** to remove them. Exceptions: drop from Group 2 to 3 (new sub-shell, p electron easier to remove) and Group 5 to 6 (paired electron in p-orbital experiences repulsion, easier to remove).

1.14	*Understand reasons for the decrease in first ionisation energy down a group.*
	Down a group: **more electron shells** means the **outer electron is further from the nucleus** and there is **increased shielding** from inner electrons. The **effective nuclear charge decreases**, so less energy is needed to remove the outer electron, despite the increase in nuclear charge.

1.15	*Understand how ideas about electronic configuration developed from: (i) atomic emission spectra provide evidence for quantum shells, (ii) successive ionisation energies provide evidence for quantum shells and the group, (iii) first ionisation energy of successive elements provides evidence for electron sub-shells.*
	(i) Atomic emission spectra show **discrete lines** (not continuous), proving electrons can only exist at **specific, quantised energy levels**. Each line corresponds to a transition between shells. **(ii)** A plot of successive IEs for an element shows **large jumps** where electrons are removed from a shell closer to the nucleus. The number of electrons removed before the first big jump = the group number. **(iii)** Plotting first IE across a period shows **sub-shell structure**: drops between Groups 2→3 (s→p) and 5→6 (half-full→paired p) indicate different sub-shell energies.

1.16	*Know the number of electrons that can fill the first four quantum shells.*
	Shell 1: 2 electrons ($1s^2$) **Shell 2:** 8 electrons ($2s^2\ 2p^6$) **Shell 3:** 18 electrons ($3s^2\ 3p^6\ 3d^{10}$) **Shell 4:** 32 electrons ($4s^2\ 3d^{10}\ 4p^6\ 4d^{10}\ 4f^{14}$) Maximum electrons per shell = **$2n^2$**.

1.17	*Know that an orbital is a region within an atom that can hold up to two electrons with opposite spins.*
	An orbital is a **region of space** around the nucleus where there is a **high probability** of finding an electron. Each orbital holds a **maximum of 2 electrons**, which must have **opposite (antiparallel) spins** (↑↓)

1.18	*Know the shape of an s-orbital and a p-orbital.*
	s-orbital: spherical shape, centred on the nucleus. **p-orbital:** dumbbell (figure-of-eight) shape, with a **node** at the nucleus. There are three p-orbitals per sub-shell (px, py, pz), oriented at 90° to each other.

1.19	*Know the number of electrons that occupy s, p and d-subshells.*
	s-subshell: 1 orbital → **2 electrons** **p-subshell:** 3 orbitals → **6 electrons** **d-subshell:** 5 orbitals → **10 electrons**

1.20	*Know that electrons fill subshells singly, before pairing up, and that two electrons in the same orbital must have opposite spins.*
	Electrons occupy orbitals **singly first** with **parallel spins** before pairing up. When electrons do pair, they must have **opposite spins.** This minimises electron-electron repulsion.

1.21	*Be able to predict the electronic configurations, using 1s notation and electrons-in-boxes notation, of: (i) atoms up to Z = 36, (ii) ions (s and p block only) up to Z = 36.*
	Fill in order: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$ **Key anomalies:** Cr = [Ar] $3d^5\ 4s^1$ and Cu = [Ar] $3d^{10}\ 4s^1$ (half-full/full d-subshell stability). **For ions:** remove electrons from the **outermost shell first** (i.e. 4s before 3d for d-block cations). Electrons-in-boxes: draw boxes for each orbital, fill with arrows (↑↓) **singly first** with **parallel spins** before pairing up.

1.22	*Know that elements can be classified as s, p and d-block elements.*
	Elements are classified by the **sub-shell into which the last electron is added**: • **s-block:** Groups 1 and 2 (including H and He) • **p-block:** Groups 3, 4, 5, 6, 7, 0 • **d-block:** transition metals

1.23	*Understand that electronic configuration determines the chemical properties of an element.*
	Elements with the **same outer shell configuration** show **similar chemical properties**. This is why elements in the same **group** of the Periodic Table react in similar ways. E.g. all Group 1 metals have one electron in their outer s-subshell and form 1+ ions.

1.24	*Understand periodicity in terms of a repeating pattern across different periods.*
	Periodicity is the **regular recurring variation** in properties of elements with increasing atomic number. As each new period begins, the pattern of properties repeats because the **outer shell electronic configuration repeats**.

1.25	*Understand reasons for the trends in: (i) ionisation energy across Periods 2 and 3, (ii) melting and boiling temperatures across Periods 2 and 3.*
	(i) Ionisation energy: general increase across the period due to increasing nuclear charge with constant shielding. Drops at Group 3 (p-electron easier to remove than s) and Group 6 (paired p-electron repulsion). **(ii) Melting/boiling points:** metals (Na, Mg, Al) increase due to increasing charge density and number of delocalised electrons → stronger metallic bonding. Si has very high m.p. (giant covalent). P_4, S_8, Cl_2, Ar are simple molecular – weak London forces, so low m.p. $S_8 > P_4 > Cl_2 > Ar$ (more electrons → stronger London forces).

1.26	*Be able to illustrate periodicity using data, including electronic configurations, atomic radii, melting and boiling temperatures and first ionisation energies.*
	You must be able to **interpret and sketch graphs** showing periodic trends. Look for repeating patterns across Periods 2 and 3. **Atomic radii decrease** across a period (increased effective nuclear charge pulls electrons closer). All properties show a **repeating pattern** that resets at the start of each new period.

Topic 2: Bonding and Structure

Topic 2A: Bonding

2.1	*Know that ionic bonding is the strong electrostatic attraction between oppositely charged ions.*
	Ionic bonding occurs between **metals and non-metals**. It is the **strong electrostatic attraction** between **oppositely charged ions** in all directions, forming a **giant ionic lattice**. This is an exam-mark phrase – always state "strong electrostatic attraction between oppositely charged ions".

2.2	*Understand the effects that ionic radius and ionic charge have on the strength of ionic bonding.*
	Ionic bonding is **stronger** when: • **Ionic charge is greater** (e.g. $Mg^{2+}O^{2-}$ stronger than $Na^{+}Cl^{-}$) – greater electrostatic attraction. • **Ionic radius is smaller** (ions are closer together) – greater charge density, stronger attraction. This explains trends in **melting points** and **lattice energies** of ionic compounds.

2.3	*Understand the formation of ions in terms of electron loss or gain.*
	Metals lose electrons to form **positive ions (cations)**, achieving a stable noble gas electronic configuration. **Non-metals gain electrons** to form **negative ions (anions)**, achieving a stable noble gas electronic configuration.

2.4	*Be able to draw electronic configuration diagrams of cations and anions using dot-and-cross diagrams.*
	Show the **transfer of electrons** from metal to non-metal. Use dots for one element and crosses for the other. Show the charge on each ion and square brackets around ions. Only outer shell electrons need to be shown (though full diagrams are accepted).

2.5	*Understand reasons for the trends in ionic radii down a group and for a set of isoelectronic ions.*
	Down a group: ionic radii **increase** because there are more electron shells, so the outermost electrons are further from the nucleus. **Isoelectronic ions** (e.g. N^{3-}, O^{2-}, F^{-}, Na^{+}, Mg^{2+}, Al^{3+}): all have the same number of electrons but **increasing nuclear charge** across the series. More protons pull the electron cloud in more tightly, so **ionic radius decreases** as nuclear charge increases.

2.6	*Understand that the physical properties of ionic compounds and the migration of ions provide evidence for the existence of ions.*
	Evidence for ions: • Ionic compounds conduct electricity **when molten or dissolved** (ions are free to move and carry charge) but **not when solid** (ions are fixed in the lattice). • **Electrolysis** causes ions to migrate to oppositely charged electrodes – cations to cathode, anions to anode. • These observations confirm the presence of **charged particles** (ions).

2.7	*Know that a covalent bond is the strong electrostatic attraction between two nuclei and the shared pair of electrons between them.*
	A covalent bond is the **strong electrostatic attraction** between the **two nuclei** and the **shared pair of electrons** between them. This is the exam-mark definition. Covalent bonds form between **non-metal atoms**.

2.8	*Be able to draw dot-and-cross diagrams to show electrons in covalent substances, including: (i) single, double and triple bonds, (ii) dative covalent (coordinate) bonding, including Al_2Cl_6 and ammonium ion.*
	Dot-and-cross diagrams show shared and lone pairs. Use dots for one atom and crosses for the other. • **Single bond:** one shared pair • **Double bond:** two shared pairs • **Triple bond:** three shared pairs (e.g. N_2) **Dative covalent bond:** both electrons in the shared pair come from **one atom** (the donor). Show with an arrow from donor to acceptor. E.g. NH_4^+ (N donates to H^+), Al_2Cl_6 (Cl donates lone pair to electron-deficient Al).

2.9	*Understand the relationship between bond lengths and bond strengths for covalent bonds.*
	Shorter bonds are stronger bonds. As bond order increases (single → double → triple), the bond gets **shorter** and **stronger** (higher bond enthalpy). More shared electrons means greater electrostatic attraction between nuclei and shared electrons.

2.10	*Understand that the shape of a simple molecule or ion is determined by the repulsion between the electron pairs that surround a central atom.*
	Electron pair repulsion theory (VSEPR): electron pairs around a central atom repel each other and arrange themselves to be **as far apart as possible** to minimise repulsion. **Lone pair–lone pair repulsion > lone pair–bonding pair > bonding pair–bonding pair.** Lone pairs are held closer to the nucleus and repel more strongly.

2.11	*Understand reasons for the shapes of, and bond angles in, simple molecules and ions with up to six outer pairs of electrons.*
	2 electron pairs: linear, 180° (e.g. $BeCl_2$, CO_2) **3 electron pairs:** trigonal planar, 120° (e.g. BCl_3) **4 electron pairs:** tetrahedral, 109.5° (e.g. CH_4, NH_4^+); with 1 lone pair → pyramidal, 107° (e.g. NH_3); with 2 lone pairs → bent, 104.5° (e.g. H_2O) **5 electron pairs:** trigonal bipyramidal, 90°/120° (e.g. PCl_5) **6 electron pairs:** octahedral, 90° (e.g. SF_6) Lone pairs reduce bond angles by approximately 2.5° per lone pair.

2.12	*Be able to predict the shapes of, and bond angles in, simple molecules and ions analogous to those specified above using electron-pair repulsion theory.*
	Method: count the total number of electron pairs (bonding + lone) around the central atom. Determine the base shape from the total pairs, then adjust for lone pairs. Always state the **name of the shape** and the **bond angle**. Treat double/triple bonds as single "regions of electron density" for shape purposes.

2.13	*Know that electronegativity is the ability of an atom to attract the bonding electrons in a covalent bond.*
	Electronegativity is the **ability of an atom to attract the bonding electrons** in a covalent bond. Electronegativity **increases** across a period (more protons, same shielding) and **decreases** down a group (more shielding, greater atomic radius). **Fluorine is the most electronegative element.**

2.14	*Know that ionic and covalent bonding are the extremes of a continuum of bonding type and that electronegativity differences lead to bond polarity.*
	Bonding exists on a **spectrum**: purely covalent (equal sharing, e.g. Cl_2) → polar covalent (unequal sharing) → ionic (complete electron transfer). The **greater the electronegativity difference**, the more ionic the bond character. A small difference gives a polar covalent bond (δ+ and δ–).

2.15	*Understand that molecules with polar bonds may not be polar molecules and be able to predict whether or not a given molecule is likely to be polar.*
	A molecule is **polar** if it has an **overall dipole moment** (dipoles do not cancel). **Symmetric molecules** (e.g. CO_2, CCl_4, BF_3) have polar bonds but are **non-polar overall** because the dipoles cancel due to symmetry. **Asymmetric molecules** (e.g. H_2O, $CHCl_3$) are polar because the dipoles **do not cancel**.

2.16	*Understand the nature of intermolecular forces: (i) London forces, (ii) permanent dipoles, (iii) hydrogen bonds.*
	(i) London (dispersion) forces: instantaneous dipole–induced dipole attractions. Present in **all molecules**. Strength increases with number of electrons (larger Mr) and surface area of contact. **(ii) Permanent dipole–dipole:** attractions between molecules with permanent dipoles (δ+ attracted to δ–). Stronger than London forces alone for molecules of similar size. **(iii) Hydrogen bonds:** a special, strong type of permanent dipole interaction. Occurs when H is bonded to **N, O or F** (highly electronegative atoms with lone pairs). The lone pair on N/O/F is attracted to the δ+ H on an adjacent molecule.

2.17	*Understand the interactions in molecules, such as H_2O, liquid NH_3 and liquid HF, which give rise to hydrogen bonding.*
	Hydrogen bonding requires: • A **hydrogen atom bonded to N, O, or F** (δ+ H) • A **lone pair on N, O, or F** on an adjacent molecule In H_2O: each molecule can form up to **4 hydrogen bonds** (2 lone pairs + 2 H atoms). In NH_3: only 1 lone pair limits H-bonding. In HF: only 1 H and 3 lone pairs.

2.18	*Understand the anomalous properties of water resulting from hydrogen bonding: (i) relatively high melting and boiling temperatures, (ii) density of ice compared to water.*
	(i) Water has a **higher boiling point** than expected for its molecular mass because **extra energy is needed to break the hydrogen bonds** between molecules. **(ii)** Ice is **less dense than liquid water** because in ice, hydrogen bonds hold molecules in an **open, regular lattice structure** with large spaces. When ice melts, some H-bonds break, molecules move closer together, and density increases.

2.19	*Be able to predict the presence of hydrogen bonding in molecules analogous to those mentioned above.*
	Look for molecules containing **H bonded directly to N, O, or F**. If the molecule has this feature AND a lone pair on the N/O/F, hydrogen bonding can occur. Examples: alcohols (O–H), carboxylic acids (O–H), amines (N–H).

2.20	*Understand, in terms of intermolecular forces, physical properties shown by materials.*
	(i) Alkane boiling points increase with chain length: more electrons → stronger London forces → more energy needed. **(ii) Branching decreases boiling point:** less surface area of contact → weaker London forces. **(iii) Alcohols have higher boiling points than alkanes of similar Mr:** alcohols can form **hydrogen bonds** (extra energy needed). **(iv) HF has an anomalously high b.p.** compared to HCl due to hydrogen bonding. HCl < HBr < HI shows increasing London forces with more electrons. HF breaks the trend due to H-bonding.

2.21	*Understand factors that influence the choice of solvents.*
	(i) Water dissolves many ionic compounds: water molecules **hydrate the ions** (δ+ H attracted to anions, δ– O attracted to cations), overcoming the lattice energy. **(ii)** Water dissolves simple alcohols: the OH group forms **hydrogen bonds** with water. **(iii)** Water is a poor solvent for halogenoalkanes: they are polar but **cannot form hydrogen bonds** with water, and the energy released from interactions with water cannot overcome the existing intermolecular forces. **(iv)** Non-aqueous solvents dissolve non-polar substances: “like dissolves like” – similar intermolecular forces in solute and solvent.

2.22	*Know that metallic bonding is the strong electrostatic attraction between metal ions and the delocalised electrons.*
	Metallic bonding is the **strong electrostatic attraction** between **positive metal ions (cations)** and a **sea of delocalised electrons**. Metallic bond strength increases with **greater charge on the ion** and **smaller ionic radius** (e.g. Al > Mg > Na).

Topic 2B: Structure

2.23	*Know that giant lattices are present in: (i) ionic solids, (ii) covalently bonded solids such as diamond, graphite and silicon(IV) oxide, (iii) solid metals.*
	Giant ionic lattice: regular arrangement of alternating positive and negative ions (e.g. NaCl). High m.p. due to strong electrostatic attractions. **Giant covalent lattice:** atoms bonded by strong covalent bonds throughout (e.g. diamond, graphite, SiO_2). Very high m.p. **Giant metallic lattice:** regular arrangement of metal cations surrounded by delocalised electrons.

2.24	*Know that the structure of covalently bonded substances such as iodine and ice is simple molecular.*
	Simple molecular structures consist of small, discrete molecules held together by **weak intermolecular forces** (London forces, dipole–dipole, or hydrogen bonds). Within each molecule, atoms are joined by **strong covalent bonds**. Low melting and boiling points because only **weak intermolecular forces** need to be overcome (not covalent bonds).

2.25	*Know the different structures formed by carbon atoms, including graphite, diamond and graphene.*
	Diamond: each C bonded to 4 others tetrahedrally; very hard; does not conduct electricity (no free electrons/ions). **Graphite:** each C bonded to 3 others in flat hexagonal layers; layers held by weak London forces so can slide (lubricant); one delocalised electron per C → **conducts electricity**. **Graphene:** a **single layer of graphite**; strong, flexible, excellent conductor of electricity and heat.

2.26	*Be able to predict the type of structure and bonding present in a substance from numerical data and/or other information.*
	High m.p. + conducts when molten/dissolved → giant ionic **Very high m.p. + does not conduct (unless graphite) → giant covalent** **High m.p. + conducts when solid → giant metallic** **Low m.p. + does not conduct → simple molecular** Always link properties to the **type of bonding/forces that must be overcome**.

2.27	*Be able to predict the physical properties of a substance in terms of particles present, structure, type of bonding and intermolecular forces.*
	Always state: (1) the **type of structure** (giant ionic/covalent/metallic or simple molecular), (2) the **particles present**, (3) the **type of bonding or intermolecular forces**, (4) whether these are **strong or weak**, and therefore (5) whether **a lot or a little energy** is needed to overcome them. This determines melting/boiling points, electrical conductivity, and solubility.

Topic 3: Redox I

3.1	*Know what is meant by the term 'oxidation number'.*
	Oxidation number (oxidation state) is the **charge an atom would have if all bonds were ionic**. It indicates the degree of oxidation of an atom. **Rules:** uncombined elements = 0; monatomic ions = charge of ion; O = −2 (except peroxides = −1); H = +1 (except metal hydrides = −1); F = −1 always. Sum of oxidation numbers in a compound = 0; in an ion = charge of the ion.

3.2	*Be able to calculate the oxidation number of elements in compounds and ions.*
	Use the rules above algebraically. E.g. in SO_4^{2-}: let S = x, O = −2. x + 4(−2) = −2, so x = +6. Know that oxidation numbers in **peroxides** (O = −1, e.g. H_2O_2) and **metal hydrides** (H = −1, e.g. NaH) are exceptions to the usual rules.

3.3	*Understand oxidation and reduction in terms of electron transfer and changes in oxidation number, applied to reactions of s- and p-block elements.*
	Oxidation = loss of electrons = increase in oxidation number (OIL) **Reduction = gain of electrons = decrease in oxidation number** (RIG) Remember: **OIL RIG** (Oxidation is Loss, Reduction is Gain)

3.4	*Understand oxidation and reduction in terms of electron loss or electron gain.*
	In any redox reaction, **oxidation and reduction occur simultaneously**. The species that is oxidised **loses electrons**; these electrons are **gained by** the species that is reduced. Total electrons lost = total electrons gained.

3.5	*Know that oxidising agents gain electrons.*
	An **oxidising agent** (oxidant) causes another species to be oxidised. It does this by **gaining (accepting) electrons**, and is therefore **itself reduced** in the process.

3.6	*Know that reducing agents lose electrons.*
	A **reducing agent** (reductant) causes another species to be reduced. It does this by **losing (donating) electrons**, and is therefore **itself oxidised** in the process.

3.7	*Understand that a disproportionation reaction involves an element in a single species being simultaneously oxidised and reduced.*
	In **disproportionation**, the **same element** in **one species** is both oxidised and reduced simultaneously. E.g. $Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$: chlorine (0) is simultaneously reduced to −1 (in NaCl) and oxidised to +1 (in NaClO).

3.8	*Know that oxidation number is a useful concept in terms of the classification of reactions as redox and as disproportionation.*
	If oxidation numbers change during a reaction, it is a **redox reaction**. If the same element in one species shows both an increase and a decrease in oxidation number, it is **disproportionation**. Oxidation numbers allow us to identify and classify these reactions.

3.9	*Be able to indicate the oxidation number of an element in a compound or ion, using a Roman numeral.*
	The oxidation number is shown as a **Roman numeral in brackets** after the element name. E.g. iron(III) chloride = $FeCl_3$ (Fe is +3). Manganese(VII) = Mn is +7 (as in $KMnO_4$). This system is used for elements with **variable oxidation states**.

3.10	*Be able to write formulae given oxidation numbers.*
	Use the oxidation numbers to work out the ratio of atoms. E.g. iron(III) oxide: Fe = +3, O = −2. To balance charges: 2 × (+3) + 3 × (−2) = 0, so Fe_2O_3. Chromium(III) sulfate: Cr = +3, SO_4 = −2, so $Cr_2(SO_4)_3$.

3.11	*Understand that metals, in general, form positive ions by loss of electrons with an increase in oxidation number.*
	Metals are **electropositive** – they readily **lose electrons** to form **cations**. The oxidation number increases from 0 to a positive value. E.g. $Na \rightarrow Na^+ + e^-$ (oxidation number: 0 → +1).

3.12	*Understand that non-metals, in general, form negative ions by gain of electrons with a decrease in oxidation number.*
	Non-metals are **electronegative** – they readily **gain electrons** to form **anions**. The oxidation number decreases from 0 to a negative value. E.g. $Cl_2 + 2e^- \rightarrow 2Cl^-$ (oxidation number: 0 → −1).

3.13	*Be able to write ionic half-equations and use them to construct full ionic equations.*
	Half-equations show either the oxidation or reduction process separately. • **Balance atoms** first, then **balance charges** using electrons. • E.g. $Mg \rightarrow Mg^{2+} + 2e^-$ (oxidation); $Cl_2 + 2e^- \rightarrow 2Cl^-$ (reduction). • To form the **full equation**, balance electrons so they cancel, then combine. • $Mg + Cl_2 \rightarrow MgCl_2$ (electrons cancel: 2 lost = 2 gained).

Topic 4: Inorganic Chemistry and the Periodic Table

Topic 4A: The Elements of Groups 1 and 2

4A.1	*Understand reasons for the trend in ionisation energy down Group 2.*
	Ionisation energy **decreases** down Group 2 because: • **More electron shells** → outer electron is further from the nucleus • **Greater shielding** by inner electrons • These factors **outweigh** the increase in nuclear charge • **Less energy needed** to remove the outer electron

4A.2	*Understand reasons for the trend in reactivity of the Group 2 elements down the group.*
	Reactivity **increases** down Group 2 because **ionisation energy decreases** – it becomes easier to lose the two outer electrons. The outer electrons are further from the nucleus and more shielded, so less energy is required to remove them. Reactions occur more vigorously.

4A.3	*Know the reactions of the elements Mg to Ba in Group 2 with oxygen, chlorine and water.*
	With oxygen: $2M + O_2 \rightarrow 2MO$ (white solid oxides). Reactivity increases down the group. Burns with characteristic flame colours. **With chlorine:** $M + Cl_2 \rightarrow MCl_2$ (white solid chlorides). **With water:** $M + 2H_2O \rightarrow M(OH)_2 + H_2$. Mg reacts very slowly with cold water (forms $Mg(OH)_2$) but vigorously with steam ($Mg + H_2O \rightarrow MgO + H_2$). Ca to Ba react increasingly vigorously with cold water.

4A.4	*Know the reactions of the oxides of Group 2 elements with water and dilute acid, and their hydroxides with dilute acid.*
	Oxides + water → hydroxides: $MO + H_2O \rightarrow M(OH)_2$ (alkaline solutions, pH increases down group as hydroxides become more soluble). **Oxides + dilute acid → salt + water:** $MO + 2HCl \rightarrow MCl_2 + H_2O$ (neutralisation). **Hydroxides + dilute acid → salt + water:** $M(OH)_2 + 2HCl \rightarrow MCl_2 + 2H_2O$.

4A.5	*Know the trends in solubility of the hydroxides and sulfates of Group 2 elements.*
	Hydroxides: solubility **increases** down Group 2 ($Mg(OH)_2$ sparingly soluble → $Ba(OH)_2$ soluble). Solutions become more alkaline. **Sulfates:** solubility **decreases** down Group 2 ($MgSO_4$ soluble → $BaSO_4$ insoluble). $BaSO_4$ is used in the test for sulfate ions. Remember: Hydroxides Happy (increase), Sulfates Sad (decrease).

4A.6	*Understand reasons for the trends in thermal stability of the nitrates and the carbonates of the elements in Groups 1 and 2.*
	Thermal stability **increases** down both groups. **Explanation:** the carbonate/nitrate anion is decomposed by the **polarising power** of the cation. **Smaller, more highly charged cations** (e.g. Mg^{2+}) have **greater charge density** and **distort (polarise)** the electron cloud of the anion more, weakening it and making it easier to decompose. Down the group, cations are larger with lower charge density, so less polarising → more thermally stable. Group 2 compounds are **less stable** than corresponding Group 1 compounds (2+ charge is more polarising than 1+).

4A.7	*Understand the formation of characteristic flame colours by Group 1 and 2 compounds in terms of electron transitions.*
	When heated, electrons are **promoted** to higher energy levels. When they **fall back** to lower levels, they emit light of **specific wavelengths** (colours). **Group 1:** Li = crimson/red, Na = yellow/orange, K = lilac **Group 2:** Ca = orange-red, Sr = red, Ba = green The energy gap between levels determines the **wavelength (colour)** of light emitted.

4A.8	*Understand experimental procedures for: (i) patterns in thermal decomposition of Group 1 and 2 nitrates and carbonates, (ii) flame colours.*
	(i) Thermal decomposition: heat samples in test tubes and test gases evolved. Carbonates produce CO_2 (turns limewater milky). Nitrates produce brown NO_2 gas and O_2 (relights a glowing splint). Time how long until gas is produced – shorter time or lower temperature = less stable. **(ii) Flame tests:** clean a nichrome wire loop in concentrated HCl, dip in the sample, hold in a blue Bunsen flame, and observe the colour.

Topic 4B: The Elements of Group 7 (Halogens)

4B.9	*Understand reasons for the trends in melting and boiling temperatures, physical state at room temperature, and electronegativity for Group 7 elements.*
	M.p./b.p. increase down Group 7: larger molecules have **more electrons → stronger London forces** between molecules → more energy needed to separate them. F_2 and Cl_2 are gases, Br_2 is a liquid, I_2 is a solid at room temperature. **Electronegativity decreases** down the group: atomic radius increases and shielding increases, so the nucleus attracts bonding electrons less strongly.

4B.10	*Understand reasons for the trend in reactivity of Group 7 elements down the group.*
	Reactivity **decreases** down Group 7. Halogens react by **gaining an electron**. Down the group, atomic radius increases and shielding increases, so the nucleus is **less able to attract an incoming electron**. It becomes harder to form a halide ion, so reactivity decreases.

4B.11	*Understand the trend in reactivity in terms of the redox reactions of Cl_2, Br_2 and I_2 with halide ions in aqueous solution, followed by the addition of an organic solvent.*
	A **more reactive halogen displaces a less reactive halide** from solution: • $Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$ (solution turns orange/brown) • $Cl_2 + 2I^- \rightarrow 2Cl^- + I_2$ (solution turns brown) • $Br_2 + 2I^- \rightarrow 2Br^- + I_2$ (solution turns brown) Adding a **non-polar organic solvent** (e.g. cyclohexane) helps identify products: Br_2 is **orange** in organic layer; I_2 is **purple/violet** in organic layer.

4B.12	*Understand, in terms of changes in oxidation number, the following reactions of the halogens: (i) oxidation with Group 1/2 metals, (ii) disproportionation of chlorine with water and use in water treatment, (iii) disproportionation with cold dilute NaOH (bleach), (iv) disproportionation with hot alkali, (v) analogous reactions.*
	(i) $2Na + Cl_2 \rightarrow 2NaCl$: Cl goes from 0 to −1 (reduced); Na goes from 0 to +1 (oxidised). **(ii)** $Cl_2 + H_2O \rightleftharpoons HOCl + HCl$: Cl disproportionates (0 → +1 in HOCl and 0 → −1 in HCl). HOCl kills bacteria – used in water treatment. **(iii) Cold dilute NaOH:** $Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$ (Cl: 0 → −1 and 0 → +1). NaClO is bleach. **(iv) Hot concentrated NaOH:** $3Cl_2 + 6NaOH \rightarrow 5NaCl + NaClO_3 + 3H_2O$ (Cl: 0 → −1 and 0 → +5).

4B.13	*Understand the following reactions: (i) solid Group 1 halides with concentrated sulfuric acid, (ii) precipitation reactions of Cl^-, Br^- and I^- with $AgNO_3(aq)$, (iii) hydrogen halides with ammonia and with water.*
	(i) $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$ (steamy white fumes). NaBr: also makes HBr, but H_2SO_4 oxidises HBr to Br_2 (orange fumes). NaI: HI is further oxidised to I_2 (purple fumes), H_2S (rotten eggs) and S (yellow solid). **Reducing ability: HI > HBr > HCl.** **(ii)** $Ag^+ + Cl^- \rightarrow AgCl$ (white, soluble in dilute NH_3); $Ag^+ + Br^- \rightarrow AgBr$ (cream, soluble in conc. NH_3); $Ag^+ + I^- \rightarrow AgI$ (yellow, insoluble in NH_3). **(iii)** $HCl + NH_3 \rightarrow NH_4Cl$ (white smoke). All hydrogen halides dissolve in water to form strong acids (HF is weak).

4B.14	*Be able to make predictions about fluorine and astatine and their compounds, in terms of knowledge of trends in halogen chemistry.*
	Fluorine: most reactive halogen, highest electronegativity, strongest oxidising agent, lowest b.p., HF is a weak acid (unlike other hydrogen halides), AgF is soluble. **Astatine:** least reactive halogen (of the common ones), lowest electronegativity, solid at room temperature, darkest colour, weakest oxidising agent. Predict: At_2 would be displaced by all other halogens; HAt would be the strongest reducing agent.

Topic 4C: Analysis of Inorganic Compounds

4C.15	*Know reactions for identifying: (i) CO_3^{2-} and HCO_3^- using acid, (ii) SO_4^{2-} using acidified $BaCl_2$, (iii) NH_4^+ using NaOH and warming.*
	(i) Carbonate/hydrogencarbonate test: add dilute acid (e.g. HCl). Both produce **CO_2 gas** which turns **limewater milky**. $CO_3^{2-} + 2H^+ \rightarrow H_2O + CO_2$. $HCO_3^- + H^+ \rightarrow H_2O + CO_2$. **(ii) Sulfate test:** add **dilute HCl** (to remove carbonate/sulfite ions), then add **$BaCl_2$(aq)**. A **white precipitate** of $BaSO_4$ confirms sulfate ions. $Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4$. **(iii) Ammonium test:** add **NaOH(aq)** and **warm gently**. **Ammonia gas** is produced, which turns **damp red litmus paper blue**. $NH_4^+ + OH^- \rightarrow NH_3 + H_2O$.

Topic 5: Formulae, Equations and Amounts of Substance

5.1	*Know that the mole (mol) is the unit for amount of a substance.*
	The **mole** is the SI unit for **amount of substance**. One mole contains exactly **6.02×10^{23} particles** (Avogadro's constant, L). The particles can be atoms, molecules, ions, electrons, or formula units – always specify the type.

5.2	*Be able to use the Avogadro constant, L, (6.02×10^{23} mol^{-1}) in calculations.*
	Number of particles = moles × L **Moles = number of particles / L** E.g. How many molecules in 0.5 mol of H_2O? $0.5 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$ molecules.

5.3	*Know that the molar mass of a substance is the mass per mole of the substance in g mol^{-1}.*
	Molar mass (M) has units of **g mol^{-1}** and is numerically equal to the relative atomic/molecular/formula mass. E.g. $M(H_2O) = 18.0$ g mol^{-1}. **Moles = mass / molar mass** (n = m/M).

5.4	*Know what is meant by the terms 'empirical formula' and 'molecular formula'.*
	Empirical formula: the simplest whole number ratio of atoms of each element in a compound (e.g. CH_2O). **Molecular formula:** the actual number of atoms of each element in one molecule (e.g. $C_6H_{12}O_6$ = glucose). Molecular formula is always a whole number multiple of the empirical formula.

5.5	*Be able to use experimental data to calculate: (i) empirical formulae, (ii) molecular formulae including the use of pV = nRT for gases and volatile liquids.*
	(i) Empirical formula method: divide each mass (or %) by the Ar to get moles; divide all by the smallest to get the ratio; scale up to whole numbers if needed. **(ii) Molecular formula:** find Mr (e.g. from mass spec or pV = nRT), then divide Mr by the empirical formula mass to find the multiplier. **Ideal gas equation:** pV = nRT where p = pressure (Pa), V = volume (m³), n = moles, R = 8.314 J K^{-1} mol^{-1}, T = temperature (K). Convert units carefully.

5.6	*Be able to write balanced full and ionic equations, including state symbols, for chemical reactions.*
	Full equations: show all reactants and products with correct formulae and state symbols: (s) solid, (l) liquid, (g) gas, (aq) aqueous. **Ionic equations:** only show the species that change during the reaction. Remove **spectator ions** (those unchanged on both sides). Always check charges balance as well as atoms.

5.7	*Be able to calculate amounts of substances (in mol) in reactions involving mass, volume of gas, volume of solution and concentration.*
	Key formulae: • n = m/M (mass and molar mass) • n = V/24,000 (gas volume in cm³ at RTP) or n = V/24 (dm³) • n = c × V (concentration in mol dm^{-3} × volume in dm³)

	• pV = nRT (ideal gas equation) Always convert units: 1 dm^3 = 1000 cm^3.

5.8	*Be able to calculate reacting masses from chemical equations, and vice versa.*
	Method: write a balanced equation → calculate moles of the known substance → use the **mole ratio** from the equation to find moles of the unknown → convert to mass using n = m/M. Always identify the **limiting reagent** if amounts of both reactants are given.

5.9	*Be able to calculate reacting volumes of gases from chemical equations using the concepts of amount of substance.*
	Use pV = nRT to calculate moles from gas volumes at non-standard conditions. Then use the mole ratio to find the required amount and convert back to volume. Remember to convert T to Kelvin and V to m^3 (or use consistent units).

5.10	*Be able to calculate reacting volumes of gases from chemical equations using the concepts of molar volume of gases.*
	At **room temperature and pressure (RTP)**, 1 mole of any gas occupies approximately **24 dm^3** (24,000 cm^3). Equal volumes of gases at the same T and P contain equal numbers of moles (**Avogadro's law**). Gas volume ratios = mole ratios from the balanced equation.

⭐ CORE PRACTICAL 1: Measure the Molar Volume of a Gas

Method:

1. Set up a gas syringe (or inverted measuring cylinder/burette over a water trough) connected to a conical flask via a delivery tube.
2. Weigh a known mass of a reactive metal (e.g. magnesium ribbon) or measure a known volume of a carbonate.
3. Add excess dilute acid (e.g. 1 mol dm^{-3} HCl) to the flask and immediately connect the gas syringe.
4. Record the volume of gas produced once the reaction is complete and no more gas is evolved.
5. Calculate moles of the solid reactant (n = m/M).
6. Use the mole ratio to find moles of gas produced.
7. Calculate molar volume: V_m = volume of gas / moles of gas. Compare to the theoretical value of 24,000 cm^3 at RTP.

Sources of Uncertainty:

- Gas syringe: ±0.5 cm^3 (or ±1 cm^3 depending on syringe)
- Balance: ±0.001 g
- Measuring cylinder for acid: ±0.5 cm^3
- Temperature: may fluctuate from assumed 25°C

Common Errors:

- Gas leaks from the apparatus (systematic error – volume recorded too low)
- Not all gas collected if reaction starts before apparatus fully connected
- Magnesium ribbon may have an oxide layer, reducing the actual mass of Mg reacting
- Temperature of the gas may not be at exactly 25°C (RTP), affecting the molar volume
- If collecting CO_2 over water, some gas may dissolve as soluble

Improvements:

- Ensure all connections are airtight; check for leaks using soapy water before starting
- Clean magnesium with sandpaper to remove the oxide layer before weighing
- Use a larger mass of reactant to reduce percentage uncertainty in the balance reading
- Record the room temperature and use pV = nRT for a more accurate molar volume calculation
- Repeat the experiment and take a mean value
- If collecting CO_2 over water, add a small pinch of carbonate to excess acid before starting to saturate the water with CO_2 and then no further CO_2 will dissolve in the water – more gas collected

5.11	*Be able to calculate solution concentrations, in mol dm $^{-3}$ and g dm $^{-3}$, including simple acid-base titrations.*
	Concentration (mol dm^{-3}): c = n/V (moles ÷ volume in dm^3) **Concentration (g dm^{-3}):** = c × M (mol dm^{-3} × molar mass) **Titration calculations:** use concordant titres (within 0.10 cm^3). Calculate moles of known → use mole ratio → calculate concentration of unknown. **Indicators:** phenolphthalein (colourless in acid, pink in alkali), methyl orange (red in acid, yellow in alkali).

⭐ CORE PRACTICAL 2: Prepare a Standard Solution from a Solid Acid and Use It to Find the Concentration of NaOH

Method:

1. Weigh accurately a known mass of solid acid (e.g. potassium hydrogen phthalate or oxalic acid) on a balance (±0.001 g) using a weighing bottle. Record mass before and after transfer (by difference).
2. Transfer the solid to a beaker. Rinse the weighing bottle with distilled water into the beaker to ensure all solid is transferred.
3. Add distilled water and stir with a glass rod to dissolve completely.
4. Transfer the solution to a 250 cm^3 volumetric flask using a funnel. Rinse the beaker and glass rod into the flask several times.
5. Make up to the 250 cm^3 graduation mark with distilled water, using a dropping pipette for the final additions. Stopper and invert to mix.
6. Use a 25.0 cm^3 pipette to transfer an aliquot of the NaOH solution into a conical flask.
7. Add a suitable indicator (e.g. phenolphthalein).
8. Fill a burette with the standard acid solution, recording the initial reading.
9. Titrate: add acid from the burette to the conical flask, swirling continuously, until the indicator changes colour. Record the final burette reading.
10. Repeat to obtain at least two concordant titres (within 0.10 cm^3).
11. Calculate the mean titre, moles of acid, use the mole ratio to find moles of NaOH, and calculate its concentration.

Sources of Uncertainty:

- Balance: ±0.001 g
- Volumetric flask (250 cm^3): ±0.1 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 the volumetric flask (systematic – concentration too low)
- Overshooting the end point (random error)
- Misreading the meniscus on the burette (parallax error)
- Not making up to the graduation mark accurately (systematic)
- Using a wet conical flask (no effect – this does NOT cause an error as it does not change moles of NaOH)

Improvements:

- Weigh by difference to account for any solid left in the weighing bottle
- Rinse all glassware (beaker, rod, funnel) into the volumetric flask thoroughly
- Use a dropping pipette for the final approach to the graduation mark
- Use a white tile under the conical flask to see the colour change more clearly
- Practice titrating to achieve concordant results within 0.10 cm^3

⭐ CORE PRACTICAL 3: Find the Concentration of a Solution of Hydrochloric Acid

Method:

1. A standard solution of known concentration (e.g. Na_2CO_3 or NaOH) is provided or prepared.
2. Use a 25.0 cm^3 pipette and pipette filler to transfer the standard solution into a clean conical flask. Touch the tip to the surface to drain.
3. Add a suitable indicator (methyl orange for Na_2CO_3/HCl; phenolphthalein for NaOH/HCl).
4. Fill a burette with the HCl solution, ensuring no air bubbles below the tap. Record the initial burette reading.
5. Titrate by adding HCl from the burette, swirling the conical flask, until the end point is reached. Record the final burette reading.
6. Repeat to obtain at least two concordant titres (within 0.10 cm^3).
7. Calculate the mean titre → moles of HCl → concentration of HCl.

Sources of Uncertainty:

- Pipette (25 cm^3): ±0.06 cm^3
- Burette: ±0.05 cm^3 per reading (±0.10 cm^3 per titre)
- Volumetric flask (if standard solution prepared): ±0.1 cm^3

Common Errors:

- Air bubbles in the burette below the tap (systematic – titre appears larger than true value)
- Overshooting the end point (random)
- Failing to rinse the burette with the acid solution before filling (dilutes the acid – systematic)
- Rinsing the conical flask with NaOH solution (increases moles of alkali – systematic)

Improvements:

- Rinse the burette with the solution it will contain before filling
- Rinse the pipette with the solution it will contain before use
- Do NOT rinse the conical flask with alkali – rinse with distilled water only (this is fine as it does not change moles)
- Add acid dropwise near the end point to avoid overshooting
- Use a white tile and read the burette at eye level to reduce parallax error
- Repeat and use concordant results to calculate a mean titre

5.12	*Be able to: (i) calculate measurement uncertainties and measurement errors, (ii) comment on sources of error.*
	(i) Percentage uncertainty = (uncertainty / measurement) × 100% Burette: ±0.05 cm^3 per reading (±0.10 cm^3 for a titre = 2 readings). Balance: ±0.001 g (or ±0.01 g). Volumetric flask: ±0.1 cm^3 (250 cm^3). Pipette: ±0.06 cm^3 (25 cm^3). **(ii) Sources of error:** systematic (affect accuracy, e.g. zero error on balance, indicator colour change not at true end point) and random (affect precision, e.g. reading meniscus, judging colour change).

5.13	*Understand how to minimise the percentage error and percentage uncertainty in experiments involving measurements.*
	Minimise percentage uncertainty by making measurements **as large as possible**. E.g. use a larger mass on the balance, use a burette rather than a measuring cylinder, ensure titre volumes are not too small. **Repeat measurements** to reduce random error and improve precision. Use **more precise equipment** (e.g. volumetric pipette instead of measuring cylinder).

5.14	*Be able to calculate percentage yields and percentage atom economies using chemical equations and experimental results.*
	Percentage yield = (actual yield / theoretical yield) × 100% Yield < 100% due to: incomplete reaction, side reactions, losses during transfers/purification. **Atom economy = (molar mass of desired product / sum of molar masses of ALL products) × 100%** High atom economy = less waste. Addition reactions have 100% atom economy. Substitution/elimination reactions have lower atom economy.

5.15	*Be able to relate ionic and full equations to observations from simple test tube reactions.*
	You must be able to write equations AND describe what you would **see**: **(i) Displacement:** e.g. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$. Observation: grey zinc dissolves, blue solution fades, brown copper deposited. **(ii) Acids:** e.g. $Mg + 2HCl \rightarrow MgCl_2 + H_2$. Observation: effervescence, metal dissolves. **(iii) Precipitation:** e.g. $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$. Observation: yellow precipitate forms.

5.16	*Understand risks and hazards in practical procedures and suggest appropriate precautions.*
	Hazard = something that can cause harm. **Risk** = the likelihood of harm occurring. Precautions include: wearing **safety goggles** (corrosive/irritant substances), working in a **fume cupboard** (toxic/volatile gases), using **small quantities**, keeping away from naked flames (flammable substances), washing hands after handling chemicals. Always link the precaution to the **specific hazard**.

Topic 8: Energetics I

8.1	*Know that standard conditions are 100 kPa and a specified temperature, usually 298 K.*
	Standard conditions: pressure = **100 kPa**, temperature = **298 K (25°C)**. Standard state is the physical state of a substance under these conditions. The standard state symbol is ⊖ (or °). All standard enthalpy changes are measured under these conditions.

8.2	*Know that the enthalpy change is the heat energy change measured at constant pressure.*
	Enthalpy change (ΔH) is the heat energy transferred in a reaction at **constant pressure**. If heat is given out to the surroundings, the reaction is **exothermic** (ΔH is negative). If heat is absorbed from the surroundings, the reaction is **endothermic** (ΔH is positive).

8.3	*Be able to construct and interpret enthalpy level diagrams showing an enthalpy change, including appropriate signs for exothermic and endothermic reactions.*
	Enthalpy level diagrams show reactants and products on an enthalpy axis. **Exothermic:** products are **lower** than reactants; the arrow points **downward**; ΔH is **negative**. **Endothermic:** products are **higher** than reactants; the arrow points **upward**; ΔH is **positive**. Note: activation energy is NOT shown on enthalpy level diagrams (it is shown on reaction profile diagrams).

8.4	*Be able to define standard enthalpy change of: (i) reaction, (ii) formation, (iii) combustion, (iv) neutralisation.*
	(i) Standard enthalpy of reaction (ΔH⊖_r): the enthalpy change when the amounts shown in the equation react under standard conditions. **(ii) Standard enthalpy of formation (ΔH⊖_f):** the enthalpy change when **one mole** of a compound is formed from its **elements in their standard states** under standard conditions. By definition, ΔH⊖_f of an element in its standard state = 0. **(iii) Standard enthalpy of combustion (ΔH⊖_c):** the enthalpy change when **one mole** of a substance is **completely burned in excess oxygen** under standard conditions. Always exothermic (negative). **(iv) Standard enthalpy of neutralisation (ΔH⊖_neut):** the enthalpy change when an acid and base react to form **one mole of water** under standard conditions.

8.5	*Understand experiments to measure enthalpy changes in terms of: (i) processing results using Q = mcΔT, (ii) evaluating sources of error and assumptions made.*
	Q = mcΔT where Q = energy transferred (J), m = mass of solution (g), c = specific heat capacity (4.18 J g^{-1} K^{-1} for water), ΔT = temperature change (K or °C). **Then:** ΔH = –Q / n (in kJ mol^{-1}), where n = moles of the limiting reagent. The negative sign accounts for the direction of energy transfer. **Key assumptions:** all heat is transferred to the solution; the solution has the density and specific heat capacity of water; no heat is lost to the surroundings. **Types of experiment:** mixing in an insulated container (e.g. polystyrene cup), spirit burner combustion, or indirect measurement via Hess's Law.

8.6	*Be able to calculate enthalpy changes in kJ mol^{-1} from given experimental results.*
	Method: (1) Calculate Q = mcΔT in joules. (2) Convert to kJ by dividing by 1000. (3) Calculate moles of the substance of interest. (4) Divide Q (kJ) by moles to get kJ mol^{-1}. (5) Assign the correct **sign** (– for exothermic, + for endothermic). Always give both a **sign and units** in the final answer.

8.7	*Be able to construct enthalpy cycles using Hess's Law.*
	Hess's Law: the total enthalpy change for a reaction is **independent of the route taken**, provided the initial and final conditions are the same. This allows calculation of enthalpy changes that cannot be measured directly. Draw a **cycle** with two routes from reactants to products: **Route 1:** direct (the unknown ΔH) **Route 2:** indirect, via intermediate compounds (using known ΔH values) By Hess's Law: ΔH (Route 1) = ΔH (Route 2). Common cycles use **enthalpies of formation** or **enthalpies of combustion** as the intermediate step.

8.8	*Be able to calculate enthalpy changes from data using Hess's Law.*
	Using enthalpies of formation: ΔHr = ΣΔHf (products) – ΣΔHf (reactants) **Using enthalpies of combustion:** ΔHr = ΣΔHc (reactants) – ΣΔHc (products) Remember: multiply ΔH values by the stoichiometric coefficients. ΔHf of elements in their standard states = 0.

⭐ CORE PRACTICAL 8: To Determine the Enthalpy Change of a Reaction Using Hess's Law

Method:

1. This practical typically involves measuring two enthalpy changes that cannot be combined directly, then using Hess's Law to calculate a third. A common example is finding ΔH for: $NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$.

2. Reaction 1: Weigh a known mass of $NaHCO_3$ (e.g. ~3 g) using a balance (±0.001 g). Record the mass precisely.

3. Measure 25.0 cm^3 of approximately 2 mol dm^{-3} HCl in a measuring cylinder and pour into a polystyrene cup. Record the initial temperature of the acid.

4. Add the $NaHCO_3$ to the acid, stir with a thermometer, and record the lowest temperature reached (endothermic reaction). Calculate ΔT_1.

5. Reaction 2: Weigh a known mass of Na_2CO_3 (equivalent moles, e.g. ~2.65 g). Repeat the above procedure with a fresh 25.0 cm^3 of HCl. Record the maximum/minimum temperature change. Calculate ΔT_2.

6. Calculate Q = mcΔT for each reaction (use total mass of solution, c = 4.18 J g^{-1} K^{-1}). Then calculate ΔH_1 and ΔH_2 in kJ mol^{-1}.

7. Construct a Hess's Law cycle and calculate the enthalpy change for the target reaction: $\Delta H = \Delta H_1 - \Delta H_2$ (or as appropriate from the cycle).

Sources of Uncertainty:

- Balance: ±0.001 g
- Measuring cylinder (25 cm^3): ±0.5 cm^3
- Thermometer: ±0.5°C (standard) or ±0.1°C (digital)

Common Errors:

- Heat loss to the surroundings (systematic – ΔT measured is less than the true value, giving a less exothermic/more endothermic result)
- Not all solid dissolving before the maximum/minimum temperature is reached
- Assuming the specific heat capacity and density of the solution are the same as water
- Heat absorbed by the polystyrene cup, thermometer, and stirrer (not accounted for)
- Mass of solid may include adsorbed moisture

Improvements:

- Use a lid on the polystyrene cup to reduce heat loss to the surroundings
- Use a more precise thermometer (e.g. digital to ±0.1°C) to reduce uncertainty in ΔT
- Use a larger mass of solid to increase ΔT, reducing percentage uncertainty in the temperature reading

- Stir continuously to ensure even heat distribution
- Use a plotting method (temperature vs time graph with extrapolation back to the time of mixing) to correct for heat loss
- Repeat the experiment and take a mean of the calculated ΔH values

8.9	*Know what is meant by the terms 'bond enthalpy' and 'mean bond enthalpy'.*
	Bond enthalpy: the energy required to break **one mole** of a specific covalent bond in a gaseous molecule, with all species in the **gaseous state**. It is always **endothermic** (positive). **Mean bond enthalpy:** the **average** energy required to break one mole of a particular type of bond across a range of different gaseous molecules. It is a mean value because the exact bond enthalpy varies depending on the molecular environment.

8.10	*Be able to calculate an enthalpy change of reaction using mean bond enthalpies and explain the limitations of this method.*
	ΔH = Σ(bonds broken) – Σ(bonds formed) Bonds broken = energy **in** (endothermic, positive). Bonds formed = energy **out** (exothermic, negative). **Limitations:** • Mean bond enthalpies are **averages** – the actual bond enthalpy in a specific molecule may differ, so the calculated ΔH is only an **estimate**. • Only applies to reactions where **all reactants and products are gaseous** (otherwise you'd need to account for enthalpy changes of vaporisation/fusion). • Does not account for intermolecular forces.

8.11	*Be able to calculate mean bond enthalpies from enthalpy changes of reaction.*
	Rearrange ΔH = Σ(bonds broken) – Σ(bonds formed) to solve for the unknown bond enthalpy. Substitute all known bond enthalpies and the overall ΔH, then solve algebraically. You may need to account for multiple bonds of the same type (e.g. 4 × C–H in CH_4).

Topic 10: Equilibrium I

10.1	*Know that many reactions are readily reversible and that they can reach a state of dynamic equilibrium in which: (i) the rate of the forward reaction is equal to the rate of the backward reaction, (ii) the concentrations of reactants and products remain constant.*
	Dynamic equilibrium occurs in a **closed system** when: • The **rate of the forward reaction equals the rate of the reverse reaction**. • The **concentrations of reactants and products remain constant** (but are NOT necessarily equal). • Both forward and reverse reactions are **still occurring** – it is dynamic, not static. Equilibrium can only be achieved in a **closed system** (no substances can enter or leave).

10.2	*Be able to predict and justify the qualitative effect of a change in temperature, concentration or pressure on a homogeneous system in equilibrium.*
	Le Chatelier's Principle: if a system at equilibrium is disturbed, it will shift to **oppose the change** and restore equilibrium. **Temperature:** increasing T favours the **endothermic direction**; decreasing T favours the **exothermic direction**. This is the only change that alters the value of K. **Concentration:** increasing [reactant] shifts equilibrium to the **right** (towards products); increasing [product] shifts it to the **left**. **Pressure (gases only):** increasing pressure shifts equilibrium towards the side with **fewer moles of gas**; decreasing pressure shifts it towards **more moles of gas**. **Catalyst:** does NOT affect the position of equilibrium or K – it speeds up both forward and reverse reactions equally.

10.3	*Evaluate data to explain the necessity, for many industrial processes, to reach a compromise between yield and rate of reaction.*
	Industrial compromise: conditions are chosen to balance **yield** and **rate** (and cost). E.g. **Haber process** ($N_2 + 3H_2 \rightleftharpoons 2NH_3$, exothermic): • **Low temperature** favours yield (forward is exothermic) but gives a **slow rate**. • **High pressure** favours yield (fewer moles on right) but is **expensive** and has safety risks. • **Compromise:** moderate temperature (~450°C) for a reasonable rate; high pressure (~200 atm) for good yield; **iron catalyst** to increase rate without affecting equilibrium position.

10.4	*Be able to deduce an expression for Kc, for homogeneous and heterogeneous systems, in terms of equilibrium concentrations.*
	For the reaction: $aA + bB \rightleftharpoons cC + dD$ **$K_c = [C]^c[D]^d / [A]^a[B]^b$** (products over reactants, raised to stoichiometric powers) **Homogeneous:** all species in the same phase – all appear in the Kc expression. **Heterogeneous:** species in different phases – **pure solids and pure liquids are omitted** from the Kc expression (their concentrations are constant). E.g. $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$: $K_c = [CO_2]$ only.

Topic 11: Equilibrium II

11.1	*Be able to deduce an expression for Kp, for homogeneous and heterogeneous systems, in terms of equilibrium partial pressures in atm.*
	For the reaction: aA(g) + bB(g) ⇌ cC(g) + dD(g) **$K_p = p(C)^c \times p(D)^d / p(A)^a \times p(B)^b$** (partial pressures in atm) **Partial pressure:** p(A) = mole fraction of A × total pressure. **Mole fraction** = moles of A / total moles of gas. **Heterogeneous systems:** only **gaseous species** are included in the Kp expression. Solids and liquids are omitted.

11.2	*Be able to calculate a value, with units where appropriate, for the equilibrium constant (Kc and Kp) for homogeneous and heterogeneous reactions, from experimental data.*
	Method for Kc: (1) Write the balanced equation. (2) Use an **ICE table** (Initial, Change, Equilibrium) to find equilibrium concentrations. (3) Substitute into the Kc expression. (4) Calculate the **units** by substituting (mol dm^{-3}) raised to the appropriate powers. **Method for Kp:** (1) Calculate moles at equilibrium. (2) Find mole fractions. (3) Multiply by total pressure for partial pressures. (4) Substitute into Kp expression. (5) Calculate **units** by substituting (atm) raised to appropriate powers. If Σ(product powers) = Σ(reactant powers), K has **no units**.

11.3	*Know the effect of changing temperature on the equilibrium constant (Kc and Kp), for both exothermic and endothermic reactions.*
	Exothermic reaction (forward ΔH negative): • Increasing temperature **decreases K** (equilibrium shifts left, less products at equilibrium). • Decreasing temperature **increases K** (equilibrium shifts right, more products). **Endothermic reaction (forward ΔH positive):** • Increasing temperature **increases K** (equilibrium shifts right, more products). • Decreasing temperature **decreases K** (equilibrium shifts left, less products).

11.4	*Understand that the effect of temperature on the position of equilibrium is explained using a change in the value of the equilibrium constant.*
	When temperature changes, **K changes**, and this is what causes the position of equilibrium to shift. A new equilibrium is established with different concentrations/partial pressures that satisfy the new value of K. The shift in position is a **consequence** of K changing, not the other way around.

11.5	*Understand that the value of the equilibrium constant is not affected by changes in concentration or pressure or by the addition of a catalyst.*
	Only temperature changes K. Changing concentration or pressure causes the position of equilibrium to shift, but the system adjusts so that the **ratio of concentrations/pressures still equals K**. A catalyst speeds up both forward and reverse reactions equally, so equilibrium is reached **faster** but K and the equilibrium position are **unchanged**.

Topic 12: Acid-Base Equilibria

12.1	*Know that a Brønsted–Lowry acid is a proton donor and a Brønsted–Lowry base is a proton acceptor.*
	Brønsted–Lowry acid: a species that **donates a proton (H^+)** to another species. **Brønsted–Lowry base:** a species that **accepts a proton (H^+)** from another species. This definition is broader than the Arrhenius definition and includes species like NH_3 (base, accepts H^+) that do not contain OH^-.

12.2	*Know that acid-base reactions involve the transfer of protons.*
	In every Brønsted–Lowry acid-base reaction, a **proton (H^+) is transferred** from the acid to the base. E.g. $HCl + NH_3 \rightarrow NH_4^+ + Cl^-$: HCl donates H^+ (acid), NH_3 accepts H^+ (base).

12.3	*Be able to identify Brønsted–Lowry conjugate acid-base pairs.*
	A **conjugate pair** differs by **one proton**. The **conjugate base** of an acid is formed when the acid loses H^+ (e.g. HCl/Cl^-). The **conjugate acid** of a base is formed when the base gains H^+ (e.g. NH_3/NH_4^+). In $HA + B \rightleftharpoons A^- + BH^+$: conjugate pairs are HA/A^- and B/BH^+.

12.4	*Be able to define the term 'pH'.*
	$pH = -\log_{10}[H^+]$ pH is a logarithmic scale that measures the **concentration of hydrogen ions** in solution. A decrease of 1 pH unit = a **10-fold increase** in $[H^+]$. Lower pH = more acidic; higher pH = more alkaline.

12.5	*Be able to calculate pH from hydrogen ion concentration.*
	$pH = -\log_{10}[H^+]$ E.g. If $[H^+] = 0.01$ mol dm^{-3}, $pH = -\log(0.01) = 2.00$. Always give pH to **2 decimal places** in calculations.

12.6	*Be able to calculate the concentration of hydrogen ions from pH, using $[H^+] = 10^{-pH}$.*
	$[H^+] = 10^{-pH}$ E.g. If pH = 3.40, then $[H^+] = 10^{-3.40} = 3.98 \times 10^{-4}$ mol dm^{-3}. This is the inverse of the pH calculation.

12.7	*Understand the difference between a strong acid and a weak acid in terms of degree of dissociation.*
	A **strong acid** is **fully (completely) dissociated** in aqueous solution – every molecule releases H^+. E.g. $HCl \rightarrow H^+ + Cl^-$ (use a single arrow →). A **weak acid** is **partially dissociated** in aqueous solution – an equilibrium is established. E.g. $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$ (use an equilibrium arrow ⇌). The degree of dissociation is small.

12.8	*Be able to calculate the pH of a strong acid.*

	For a **strong monobasic acid** (e.g. HCl): $[H^+]$ = concentration of the acid (fully dissociated). **$pH = -\log_{10}[H^+]$** For a **strong dibasic acid** (e.g. H_2SO_4): $[H^+]$ = 2 × concentration of the acid (each molecule releases $2H^+$).

12.9	*Be able to deduce the expression for the acid dissociation constant, Ka, for a weak acid and carry out relevant calculations.*
	For the weak acid $HA \rightleftharpoons H^+ + A^-$: **$K_a = [H^+][A^-] / [HA]$** Ka is the **acid dissociation constant** – it measures the strength of a weak acid. A **larger Ka** means a **stronger weak acid** (more dissociated). Units are usually **mol dm^{-3}**.

12.10	*Be able to calculate the pH of a weak acid making relevant assumptions.*
	Assumptions for weak acid pH calculation: • $[H^+] = [A^-]$ (each HA that dissociates produces one of each). • [HA] at equilibrium ≈ initial [HA] (dissociation is so small that the change is negligible). **Method:** $K_a = [H^+]^2 / [HA]$, so **$[H^+] = \sqrt{(K_a \times [HA])}$**, then $pH = -\log[H^+]$. State both assumptions in exam answers for full marks.

12.11	*Be able to define the ionic product of water, Kw.*
	Water undergoes autoionisation: $H_2O \rightleftharpoons H^+ + OH^-$ **$K_w = [H^+][OH^-]$** At 25°C, Kw = **1.00×10^{-14} mol^2 dm^{-6}**. Kw increases with temperature because the forward reaction is **endothermic**.

12.12	*Be able to calculate the pH of a strong base from its concentration, using Kw.*
	For a strong base like NaOH: $[OH^-]$ = concentration of base (fully dissociated). **$[H^+] = K_w / [OH^-]$** Then: **$pH = -\log_{10}[H^+]$** E.g. 0.10 mol dm^{-3} NaOH: $[H^+] = 1.00 \times 10^{-14} / 0.10 = 1.00 \times 10^{-13}$, pH = 13.00.

12.13	*Be able to define the terms 'pKa' and 'pKw'.*
	$pK_a = -\log_{10}(K_a)$ and **$K_a = 10^{-pK_a}$** **$pK_w = -\log_{10}(K_w)$** and at 25°C, pKw = 14.00. A **lower pKa** means a **stronger weak acid** (larger Ka). The p-notation is a convenient way to handle very small numbers.

12.14	*Be able to analyse data from experiments: (i) measuring pH of strong/weak acids, bases, and salts, (ii) comparing pH of strong and weak acids after dilution.*
	(i) A strong acid (e.g. 0.1 M HCl) has a **lower pH** than a weak acid of the same concentration (e.g. 0.1 M CH_3COOH) because it is fully dissociated. Strong bases have higher pH than weak bases of the same concentration. **(ii)** Diluting a strong acid by a factor of 10 increases pH by exactly **1 unit** (e.g. pH 1 → pH 2). Diluting a weak acid by a factor of 10 increases pH by approximately **0.5 units** because the equilibrium shifts to produce more H^+, partially compensating for the dilution.

12.15	*Be able to calculate Ka for a weak acid from experimental data given the pH of a solution containing a known mass of acid.*
	Method: (1) Calculate $[H^+]$ from pH: $[H^+] = 10^{-pH}$. (2) $[A^-] = [H^+]$ (from dissociation). (3) Calculate initial [HA] from mass, Mr, and volume. (4) [HA] at equilibrium = initial [HA] – $[H^+]$ (or assume ≈ initial [HA] if $[H^+]$ is very small compared to [HA]). (5) Substitute into $K_a = [H^+][A^-] / [HA]$.

12.16	*Be able to draw and interpret titration curves using all combinations of strong and weak monobasic acids and bases.*
	Key features of titration curves: • **Strong acid + strong base:** sharp vertical section at equivalence point; equivalence pH = 7. • **Weak acid + strong base:** starts at higher pH; gradual rise then sharp at equivalence; equivalence pH > 7 (due to hydrolysis of conjugate base); buffer region before equivalence point. • **Strong acid + weak base:** equivalence pH < 7. • **Weak acid + weak base:** no sharp change at equivalence; very difficult to determine end point. The equivalence point is where stoichiometric amounts have reacted.

12.17	*Be able to select a suitable indicator, using a titration curve and appropriate data.*
	A suitable indicator must have its **colour change (pKa range)** within the **vertical (steep) section** of the titration curve. • **Strong acid/strong base:** phenolphthalein or methyl orange (sharp change covers both ranges). • **Weak acid/strong base:** phenolphthalein only (equivalence pH > 7; methyl orange changes too early). • **Strong acid/weak base:** methyl orange only (equivalence pH < 7; phenolphthalein changes too late). • **Weak/weak:** no suitable indicator (no sharp change).

12.18	*Know what is meant by the term 'buffer solution'.*
	A **buffer solution** is one that **resists changes in pH** when small amounts of acid or base are added, or when the solution is diluted. Buffer solutions are made from: • A **weak acid and its conjugate base** (e.g. CH_3COOH + CH_3COONa) – acidic buffer. • A **weak base and its conjugate acid** (e.g. NH_3 + NH_4Cl) – basic buffer.

12.19	*Understand the action of a buffer solution.*
	Acidic buffer (e.g. CH_3COOH / CH_3COO^-): **When H^+ is added:** the conjugate base (CH_3COO^-) **reacts with the added H^+** to form the weak acid: $CH_3COO^- + H^+ \rightarrow CH_3COOH$. This **removes** the added H^+, so pH barely changes. **When OH^- is added:** the weak acid (CH_3COOH) **reacts with the added OH^-**: $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$. This **removes** the added OH^-. The large **reservoir** of weak acid and conjugate base allows the buffer to absorb small additions of either acid or base.

12.20	*Be able to calculate the pH of a buffer solution given appropriate data.*
	Method: Use $K_a = [H^+][A^-] / [HA]$. Rearrange: **$[H^+] = K_a \times [HA] / [A^-]$**

	Then: pH = –log[H^+]. [HA] = concentration of the **weak acid** in the buffer. [A^-] = concentration of the **conjugate base** (salt) in the buffer. Since both are in the same solution, you can use moles instead of concentrations (the volumes cancel). Alternatively: **pH = pKa + log([A^-] / [HA])** (Henderson–Hasselbalch equation).

12.21	*Be able to calculate the concentrations of solutions required to prepare a buffer solution of a given pH.*
	Method: Rearrange [H^+] = Ka × [HA] / [A^-] to find the required **ratio** of [HA] to [A^-]. (1) Calculate [H^+] from the target pH. (2) Ratio: [HA]/[A^-] = [H^+]/Ka. (3) Choose convenient concentrations/volumes that give this ratio. You only need the **ratio** to set the pH; the actual concentrations determine the **capacity** of the buffer.

12.22	*Understand how to use a weak acid–strong base titration curve to: (i) demonstrate buffer action, (ii) determine Ka from the pH at the half-equivalence point.*
	(i) Buffer region: the relatively **flat section** of the curve before the equivalence point shows buffer action – pH changes only gradually as base is added because the solution contains both the weak acid and its conjugate base. **(ii) At the half-equivalence point** (half the acid neutralised): [HA] = [A^-], so Ka = [H^+], meaning **pH = pKa**. Read the pH from the curve at exactly half the equivalence volume to determine Ka.

12.23	*Understand why there is a difference in enthalpy changes of neutralisation values for strong and weak acids.*
	Strong acid + strong base: $\Delta H_{neut} \approx -57.1$ kJ mol^{-1} (the standard value for $H^+ + OH^- \rightarrow H_2O$). **Weak acid + strong base:** ΔH_{neut} is **less exothermic** (less negative) because some energy is used to **dissociate the weak acid** further (the dissociation of the weak acid is endothermic). This **endothermic dissociation** partially offsets the exothermic neutralisation.

12.24	*Understand the roles of carbonic acid molecules and hydrogencarbonate ions in controlling the pH of blood.*
	Blood pH is maintained at approximately **7.4** by the **carbonate buffer system:** $CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$ **If blood pH falls** (too much H^+): HCO_3^- ions **react with excess H^+** → H_2CO_3 → CO_2 (exhaled). This removes H^+ and raises pH. **If blood pH rises** (too much OH^- or too little H^+): H_2CO_3 **dissociates** to release H^+, lowering the pH. The equilibrium between CO_2, H_2CO_3 and HCO_3^- acts as the buffer. Breathing rate also helps regulate CO_2 levels.

⭐ CORE PRACTICAL 9: Finding the Ka Value for a Weak Acid

Method:

1. Prepare a solution of the weak acid (e.g. ethanoic acid) of known concentration by dissolving a known mass in distilled water and making up to a known volume in a volumetric flask.
2. Calibrate a pH meter using buffer solutions of known pH (e.g. pH 4.00 and pH 7.00). Rinse the electrode with distilled water between calibrations.
3. Measure the pH of the weak acid solution using the calibrated pH meter. Record the pH to 2 decimal places.
4. Calculate $[H^+]$ from the pH: $[H^+] = 10^{-pH}$.
5. Since the weak acid $HA \rightleftharpoons H^+ + A^-$: $[A^-] = [H^+]$ (from the dissociation).
6. Calculate [HA] at equilibrium: $[HA]_{eq}$ = initial [HA] – $[H^+]$ (or assume $[HA]_{eq} \approx$ initial [HA] if $[H^+]$ is very small).
7. Calculate $K_a = [H^+][A^-] / [HA]$. Include units (mol dm^{-3}).

Sources of Uncertainty:

- pH meter: typically ±0.01 pH units (after calibration)
- Balance: ±0.001 g (if preparing solution from solid acid)
- Volumetric flask: ±0.1 cm^3 (250 cm^3 flask)

Common Errors:

- pH meter not properly calibrated (systematic – pH reading is inaccurate, leading to incorrect $[H^+]$ and Ka)
- Temperature not controlled – Ka varies with temperature, so the value is only valid at the temperature of the experiment
- Assuming [HA] ≈ initial [HA] when the acid is not sufficiently weak (introduces error if degree of dissociation is significant)
- CO_2 from the air dissolving in the solution and lowering the pH (systematic)

Improvements:

- Calibrate the pH meter with at least two buffer solutions immediately before use
- Record the temperature and quote Ka at that temperature
- Repeat measurements at different concentrations of the weak acid and take a mean Ka value
- Use a freshly prepared solution and measure pH promptly to minimise CO_2 absorption
- Check the assumption: if $[H^+]$ is more than ~5% of initial [HA], use the full quadratic expression rather than the approximation

Topic 13: Energetics II

Topic 13A: Lattice Energy

13.1	*Be able to define lattice energy as the energy change when one mole of an ionic solid is formed from its gaseous ions.*
	Lattice energy (ΔLE): the enthalpy change when **one mole of an ionic solid is formed from its gaseous ions** under standard conditions. E.g. $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$ Lattice energy is always **exothermic (negative)** because energy is released when the ionic lattice forms due to the strong electrostatic attraction between oppositely charged ions. Note: some textbooks define it as the energy to break apart the lattice (endothermic) – Edexcel uses the **formation definition** (exothermic).

13.2	*Be able to define the terms: (i) enthalpy change of atomisation, ΔatH, (ii) electron affinity.*
	(i) Enthalpy change of atomisation (ΔatH⊖): the enthalpy change when **one mole of gaseous atoms** is formed from the element in its **standard state**. Always **endothermic** (positive). E.g. $½Cl_2(g) \rightarrow Cl(g)$ **(ii) First electron affinity (ΔEA1):** the enthalpy change when **one mole of gaseous atoms** gains **one mole of electrons** to form **one mole of gaseous 1– ions**. E.g. $Cl(g) + e^- \rightarrow Cl^-(g)$ — this is **exothermic** (negative) for most elements. **Second electron affinity (ΔEA2):** adding a second electron to an already negative ion is **endothermic** (positive) due to repulsion. E.g. $O^-(g) + e^- \rightarrow O^{2-}(g)$.

13.3	*Be able to construct Born-Haber cycles and carry out related calculations.*
	A **Born-Haber cycle** is an application of Hess's Law to calculate lattice energy (or any other unknown step). The cycle links: **ΔfH⊖** (enthalpy of formation of the ionic compound) = ΔatH (metal) + ΔatH (non-metal) + IE (ionisation energies of metal) + EA (electron affinities of non-metal) + ΔLE (lattice energy) Rearrange to find the unknown. Always be careful with **signs** – some steps are endothermic (+) and some exothermic (–). Draw the cycle as an energy level diagram with gaseous atoms/ions at the top and the ionic solid at the bottom.

13.4	*Know that lattice energy provides a measure of ionic bond strength.*
	A **more exothermic (more negative) lattice energy** indicates **stronger ionic bonding**. This means more energy is released when the lattice forms, and more energy would be required to break it apart. Lattice energy depends on **ionic charge** and **ionic radius**.

13.5	*Understand that a comparison of experimental lattice energy (from Born-Haber cycle) with theoretical value (from electrostatic theory) indicates the degree of covalent bonding.*
	Theoretical lattice energy assumes the compound is **perfectly ionic** (spherical ions with no distortion). **Experimental lattice energy** (from a Born-Haber cycle) reflects the **actual bonding**. If the experimental value is **more exothermic** than the theoretical value, the compound has **some covalent character** – the cation has polarised the anion, increasing electron density between the ions. E.g. AgI has significant covalent character (large discrepancy); NaCl has very little (values are close).

13.6	*Understand the meaning of polarisation as applied to ions.*
	Polarisation is the **distortion of the electron cloud** of an anion by a nearby cation. The cation attracts the electron density of the anion towards itself, pulling the electron cloud out of shape. This introduces **covalent character** into an ionic bond.

13.7	*Know that the polarising power of a cation depends on its radius and charge.*
	A cation has **high polarising power** if it has: • A **small ionic radius** (charge is concentrated in a small volume) • A **high charge** (greater attraction for the anion's electrons) This gives a **high charge density**. E.g. Al^{3+} is highly polarising (small, 3+ charge); Na^{+} is weakly polarising (larger, 1+ charge).

13.8	*Know that the polarisability of an anion depends on its radius and charge.*
	An anion is **easily polarised** if it has: • A **large ionic radius** (outer electrons are far from the nucleus, loosely held) • A **high negative charge** (more electrons to be distorted) E.g. I^{-} is much more polarisable than F^{-} (larger radius). This is why AgI has more covalent character than AgF.

13.9	*Be able to define the terms 'enthalpy change of solution, ΔsolH', and 'enthalpy change of hydration, ΔhydH'.*
	Enthalpy of solution (ΔsolH⊖): the enthalpy change when **one mole of a solute** dissolves in sufficient water to form an **infinitely dilute solution**. Can be exothermic or endothermic. **Enthalpy of hydration (ΔhydH⊖):** the enthalpy change when **one mole of gaseous ions** is surrounded by water molecules to form **aqueous ions**. Always **exothermic** (negative) because ion-dipole bonds form between the ion and water molecules.

13.10	*Be able to use energy cycles and energy level diagrams to carry out calculations involving enthalpy change of solution, enthalpy change of hydration and lattice energy.*
	The key relationship: **ΔsolH = –ΔLE + ΔhydH(cation) + ΔhydH(anion)** Draw a Hess's Law cycle: • Top route: ionic solid → gaseous ions (reverse of lattice energy, so +ΔLE) → aqueous ions (ΔhydH for each ion) • Bottom route: ionic solid → aqueous ions (ΔsolH, directly) If ΣΔhydH > |ΔLE|, dissolving is exothermic. If ΣΔhydH < |ΔLE|, dissolving is endothermic.

13.11	*Understand the effect of ionic charge and ionic radius on the values of: (i) lattice energy, (ii) enthalpy change of hydration.*
	(i) Lattice energy becomes more exothermic with **smaller ions** (ions are closer, stronger attraction) and **higher ionic charges** (greater electrostatic attraction). E.g. MgO has a much more exothermic lattice energy than NaCl. **(ii) Enthalpy of hydration becomes more exothermic** with **smaller ions** (higher charge density attracts water molecules more strongly) and **higher charges**. E.g. ΔhydH of Mg^{2+} is more exothermic than Na^{+}.

Topic 13B: Entropy

13.12	*Understand that, since some endothermic reactions can occur at room temperature, enthalpy changes alone do not control whether reactions occur.*
	Some endothermic reactions occur **spontaneously** at room temperature (e.g. dissolving NH_4NO_3 in water – the solution gets cold but it still dissolves). This shows that ΔH alone does not determine feasibility. There must be **another factor** – **entropy** – that also drives reactions.

13.13	*Know that entropy is a measure of the disorder of a system and that the natural direction of change is increasing total entropy.*
	Entropy (S) is a measure of the **disorder** (number of possible arrangements) of a system. Units: **J K^{-1} mol^{-1}**. The **Second Law of Thermodynamics:** the total entropy of the universe always **increases** for a spontaneous process. A reaction is feasible if **ΔStotal > 0** (positive).

13.14	*Understand why entropy changes occur during: (i) changes of state, (ii) dissolving of a solid ionic lattice, (iii) reactions with a change in moles of gas.*
	(i) Changes of state: solid → liquid → gas involves **increasing disorder**, so **ΔS is positive**. Reverse: ΔS is negative. **(ii) Dissolving:** when an ionic solid dissolves, the ordered lattice breaks apart and ions become dispersed in solution → **ΔS is usually positive**. **(iii) Change in moles of gas:** if moles of gas **increase**, ΔS is positive. If they decrease, ΔS is negative. Gases have much higher entropy than solids or liquids. **Examples:** dissolving NH_4NO_3 (+ΔS, endothermic but occurs); burning Mg (–ΔS system but large –ΔH drives it); $Ba(OH)_2 \cdot 8H_2O$ + NH_4Cl (endothermic, large +ΔS from solid → liquid/gas).

13.15	*Understand that the total entropy change is: ΔStotal = ΔSsystem + ΔSsurroundings.*
	For a reaction to be **feasible (spontaneous): ΔStotal must be positive**. ΔStotal = ΔSsystem + ΔSsurroundings ΔSsystem comes from the change in disorder of the reactants/products. ΔSsurroundings comes from the heat transferred to/from the surroundings.

13.16	*Be able to calculate the entropy change for the system, ΔSsystem, given the entropies of reactants and products.*
	ΔSsystem = ΣS(products) – ΣS(reactants) Multiply each entropy value by its stoichiometric coefficient. Units: **J K^{-1} mol^{-1}**. A positive value means the products are more disordered than the reactants.

13.17	*Be able to calculate ΔSsurroundings, and hence ΔStotal, using: ΔSsurroundings = –ΔH / T.*
	ΔSsurroundings = –ΔH / T Where ΔH is in **J mol^{-1}** (not kJ!) and T is in **K**. For an **exothermic** reaction (ΔH negative): ΔSsurroundings is **positive** (heat released increases disorder of surroundings). For an **endothermic** reaction (ΔH positive): ΔSsurroundings is **negative**. Then: ΔStotal = ΔSsystem + ΔSsurroundings. If positive, the reaction is feasible.

13.18	*Know that the balance between entropy and enthalpy determines feasibility, represented by ΔG = ΔH – TΔSsystem.*
	Gibbs free energy: ΔG = ΔH – TΔSsystem A reaction is **thermodynamically feasible** when **ΔG < 0** (negative). ΔG < 0 can occur when: • ΔH is negative and TΔS is positive (always feasible) • ΔH is negative and |ΔS| is small (feasible at all temperatures) • ΔH is positive but TΔS is large enough to overcome it (feasible at **high temperatures**) • ΔH is negative but ΔS is negative (feasible at **low temperatures**) Note: ΔH must be in **kJ** and ΔS in **kJ K⁻¹** (or convert consistently).

13.19	*Be able to use ΔG = ΔH – TΔSsystem to: (i) predict feasibility, (ii) determine the temperature at which a reaction becomes feasible.*
	(i) Feasibility: calculate ΔG. If ΔG < 0, the reaction is **feasible**. If ΔG > 0, it is **not feasible** under those conditions. **(ii) Temperature for feasibility:** set ΔG = 0 and rearrange: **T = ΔH / ΔSsystem** Above this temperature (if ΔH > 0 and ΔS > 0) the reaction becomes feasible. Below it (if ΔH < 0 and ΔS < 0) the reaction is feasible. Ensure ΔH and ΔS have **consistent units** (both in kJ or both in J).

13.20	*Be able to use ΔG = –RT ln K to show that feasible reactions have large K values and vice versa.*
	ΔG = –RT ln K where R = 8.314 J K⁻¹ mol⁻¹, T in K. Rearranging: **ln K = –ΔG / RT** If ΔG is **negative** (feasible): ln K is **positive**, so K > 1 (products favoured). If ΔG is **positive** (not feasible): ln K is **negative**, so K < 1 (reactants favoured). If ΔG = 0: K = 1. This links thermodynamics to equilibrium quantitatively.

13.21	*Understand why a reaction for which ΔG is negative may not occur in practice.*
	ΔG < 0 means a reaction is **thermodynamically feasible** but it may still **not occur** because: • The **activation energy** may be too high (kinetic barrier) – the reaction is too slow to observe. • The reaction may need a **catalyst** or an **energy input** to overcome the activation energy. **Thermodynamic feasibility ≠ kinetic feasibility.** E.g. diamond → graphite is thermodynamically feasible but kinetically inert at room temperature.

13.22	*Know that reactions that are thermodynamically feasible may be inhibited by kinetic factors.*
	Kinetic stability means a reaction does not proceed at a measurable rate despite being thermodynamically favourable. The **rate of reaction** depends on the activation energy, temperature, and presence of catalysts. A high activation energy acts as a **kinetic barrier**, preventing an otherwise feasible reaction from occurring.

Topic 14: Redox II

14.1	*Understand the terms 'oxidation' and 'reduction' in terms of electron transfer, applied to s-, p- and d-block elements.*
	Oxidation = loss of electrons. Reduction = gain of electrons. (OIL RIG) This applies to all blocks of the Periodic Table. d-block elements show **variable oxidation states** because they can lose different numbers of electrons from both the 4s and 3d sub-shells.

14.2	*Understand the terms 'oxidation' and 'reduction' in terms of changes in oxidation number, applied to s-, p- and d-block elements.*
	Oxidation = increase in oxidation number. Reduction = decrease in oxidation number. E.g. $Fe^{2+} \rightarrow Fe^{3+}$ is oxidation (+2 → +3). $MnO_4^- \rightarrow Mn^{2+}$ is reduction (+7 → +2).

14.3	*Know what is meant by the term 'standard electrode potential', E⊖.*
	Standard electrode potential (E⊖) is the **emf (voltage)** of a half-cell measured under **standard conditions**, relative to the **standard hydrogen electrode (SHE)**, which is assigned a value of **0.00 V**. It measures the tendency of a species to be **reduced**. A more positive E⊖ means a greater tendency to be reduced (stronger oxidising agent).

14.4	*Know that E⊖ refers to conditions of: (i) 298 K, (ii) 100 kPa pressure of gases, (iii) 1.00 mol dm^{-3} concentration of ions.*
	Standard conditions for electrode potentials: • Temperature: **298 K** (25°C) • Gas pressure: **100 kPa** • Ion concentrations: **1.00 mol dm^{-3}** All measurements are taken relative to the standard hydrogen electrode under these conditions.

14.5	*Know the features of the standard hydrogen electrode and understand why a reference electrode is necessary.*
	Standard hydrogen electrode (SHE): • **Pt electrode** (inert, provides a surface for the equilibrium) • **H_2 gas at 100 kPa** bubbled over the Pt electrode • **1.00 mol dm^{-3} $H^+(aq)$** solution • Temperature: **298 K** The half-equation: $2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$, E⊖ = 0.00 V. A **reference electrode is necessary** because you **cannot measure the absolute potential** of a single half-cell – you can only measure the **potential difference** between two half-cells.

14.6	*Understand that different methods are used to measure E⊖ of: (i) metals/non-metals in contact with their ions, (ii) ions of the same element with different oxidation numbers.*
	(i) Metal/non-metal in contact with its ions: the metal electrode is dipped into a solution of its ions (e.g. Zn in $ZnSO_4$). Connected to SHE via a salt bridge and high-resistance voltmeter. **(ii) Ions with different oxidation states:** use a **Pt electrode** dipped into a solution containing **both** oxidation states (e.g. Fe^{2+} and Fe^{3+}). The Pt acts as an inert electrode. Connected to SHE.

⭐ CORE PRACTICAL 10: Investigating Some Electrochemical Cells

Method:

1. Set up a series of electrochemical cells by connecting pairs of half-cells (e.g. Zn/Zn^{2+}, Cu/Cu^{2+}, Fe/Fe^{2+}, Ag/Ag^{+}).
2. Each half-cell: place the metal electrode into a beaker containing 1.00 mol dm^{-3} solution of its metal ion.
3. Connect the two half-cells with a salt bridge (e.g. filter paper soaked in saturated KNO_3 solution).
4. Connect the electrodes to a high-resistance voltmeter using crocodile clips and wires.
5. Record the voltage (emf) displayed on the voltmeter. Note the polarity (+/–) to identify the cathode and anode.
6. Repeat with different combinations of half-cells.
7. Compare experimental emf values with calculated E⊖cell values from data book E⊖ values.

Sources of Uncertainty:

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

Common Errors:

- Poor electrical contact between the salt bridge and solutions (reduces measured voltage)
- Metal electrodes not clean/polished (oxide layer increases resistance, lowers emf)
- Concentrations not exactly 1.00 mol dm^{-3} (non-standard conditions, E differs from E⊖)
- Salt bridge drying out during measurements
- Temperature not controlled at 298 K

Improvements:

- Sand/polish metal electrodes immediately before use to remove oxide layers
- Use freshly prepared 1.00 mol dm^{-3} solutions made from accurate weighings
- Ensure the salt bridge makes good contact with both solutions
- Use a high-resistance voltmeter to minimise current draw (prevents the cell from discharging)
- Record the temperature and note any deviation from 298 K
- Repeat each measurement and take a mean

14.7	*Be able to calculate a standard emf, E⊖cell, by combining two standard electrode potentials.*
	Ecell = E(more positive half-cell) – E(less positive half-cell) Or equivalently: **Ecell = E(cathode) – E(anode)** The species with the more positive E is **reduced** (cathode). The species with the less positive E is **oxidised** (anode). A positive Ecell means the reaction is **thermodynamically feasible**.

14.8	*Be able to write cell diagrams using the conventional representation of half-cells.*
	Convention: anode (oxidation) on the **left**, cathode (reduction) on the **right**. **Format:** $Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$ • **|** = phase boundary (between different states) • **||** = salt bridge • Oxidised form is next to the salt bridge on each side • For ion/ion half-cells: $Pt \mid Fe^{2+}(aq), Fe^{3+}(aq) \parallel$ (comma separates species in the same phase)

14.9	*Understand the importance of the conditions when measuring electrode potential, E.*
	Changing conditions from standard alters the electrode potential: • Changing **concentration** shifts the half-cell equilibrium (Le Chatelier's), changing E. • Changing **temperature** changes E and may change which reaction is feasible. • Changing **pressure** of gases affects E.

	This is why **standard conditions** are specified – to allow meaningful comparisons between different half-cells.

14.10	*Be able to predict the thermodynamic feasibility of a reaction using standard electrode potentials.*
	If **Ecell > 0** (positive), the reaction is **thermodynamically feasible** under standard conditions. The species with the **more positive E is reduced** (it is the better oxidising agent). The species with the **less positive E is oxidised**. Write out both half-equations, reverse the one being oxidised, and combine to give the overall equation.

14.11	*Understand that E⊖cell is directly proportional to total entropy change and to ln K.*
	Ecell ∝ ΔStotal and **Ecell ∝ ln K** The relationship: ΔG = –nFEcell (where n = moles of electrons, F = Faraday constant = 96,485 C mol^{-1}). Since ΔG = –RT ln K, it follows that **nFEcell = RT ln K**. A larger positive Ecell means a more negative ΔG, a larger K, and a greater ΔStotal – the reaction goes further to completion.

14.12	*Understand the limitations of predictions made using E⊖ values, in terms of kinetic inhibition and departure from standard conditions.*
	Limitations of E predictions: • E predicts **thermodynamic feasibility** only – it says nothing about **rate**. A feasible reaction may be too slow (high activation energy / kinetic barrier). • E values apply only under **standard conditions**. Non-standard concentrations, temperatures, or pressures may change E and alter feasibility. • A reaction predicted to be feasible may not occur if conditions differ significantly from standard.

14.13	*Know that standard electrode potentials can be listed as an electrochemical series.*
	The **electrochemical series** lists half-cells in order of **increasing E** (most negative at the top, most positive at the bottom). Species at the **bottom** are the **strongest oxidising agents** (most easily reduced). Species at the **top** are the **strongest reducing agents** (most easily oxidised). Any species on the right of a half-equation can oxidise any species on the left of a half-equation above it in the series.

14.14	*Understand how disproportionation reactions relate to standard electrode potentials.*
	Disproportionation occurs when a species is simultaneously oxidised and reduced. Using E values, disproportionation is feasible if the **Ecell is positive** when the relevant half-cells are combined. E.g. Cu^+ disproportionates to Cu and Cu^{2+} because the E for Cu^+/Cu is more positive than E for Cu^{2+}/Cu^+, giving a positive Ecell.

14.15	*Understand the application of electrode potentials to storage cells.*
	Storage cells (rechargeable batteries) use **reversible** redox reactions. During discharge, the spontaneous redox reaction generates a voltage (Ecell > 0). During charging, an external voltage is applied to **reverse** the reaction, regenerating the reactants. The voltage of the cell = E(cathode) – E(anode). The greater the difference, the higher the voltage.

14.16	*Understand that fuel cells utilise the energy from fuel + oxygen reactions to generate a voltage.*
	A **fuel cell** converts the chemical energy of a fuel (e.g. H_2, methanol, other hydrogen-rich fuels) directly into **electrical energy** by reacting it with oxygen. Unlike combustion, the energy is not converted to heat first, so fuel cells are **more efficient**. They produce electricity continuously as long as fuel and oxygen are supplied.

14.17	*Know the electrode reactions that occur in a hydrogen-oxygen fuel cell.*
	Acidic electrolyte (e.g. PEM fuel cell): • Anode: $H_2 \rightarrow 2H^+ + 2e^-$ (oxidation) • Cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$ (reduction) **Alkaline electrolyte:** • Anode: $H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$ (oxidation) • Cathode: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ (reduction) **Overall in both cases:** $2H_2 + O_2 \rightarrow 2H_2O$. The only product is water – fuel cells are environmentally clean.

14.18	*Be able to carry out both structured and non-structured titration calculations including Fe^{2+}/MnO_4^-, and $I_2/S_2O_3^{2-}$.*
	$KMnO_4$ titrations: $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$. Mole ratio: 1 MnO_4^- : 5 Fe^{2+}. No indicator needed – MnO_4^- is purple; end point is colourless to the first permanent pink colour. **Iodine-thiosulfate titrations:** $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$. Mole ratio: 1 I_2 : 2 $S_2O_3^{2-}$. Add starch indicator near the end point – turns blue/black with I_2. End point: blue/black → colourless. **Non-structured calculations:** you must identify the relevant equations and mole ratios yourself. Often involves multiple steps.

14.19	*Understand the methods used in redox titrations.*
	$KMnO_4$ titration method: pipette the Fe^{2+} solution into a conical flask. Add excess dilute H_2SO_4 (to provide H^+ ions – never use HCl as Cl^- is oxidised by MnO_4^-). Titrate with $KMnO_4$ from the burette. End point: first **permanent pink colour** that persists for 30 seconds. **Iodine-thiosulfate method:** react the oxidising agent with excess KI(aq) to liberate I_2. Titrate with $Na_2S_2O_3$ from a burette. Add **starch indicator** when the solution is pale straw-yellow. End point: **blue/black → colourless**.

⭐ CORE PRACTICAL 11: Redox Titration

Method:

1. Typically a $KMnO_4/Fe^{2+}$ titration or an iodine/thiosulfate titration.
2. FOR $KMnO_4$ TITRATION: pipette 25.0 cm^3 of the Fe^{2+} solution (e.g. iron tablet dissolved in excess dilute H_2SO_4) into a conical flask using a pipette and filler.
3. Add approximately 20 cm^3 of dilute sulfuric acid (to acidify – never use HCl with $KMnO_4$).
4. Fill a burette with standardised $KMnO_4$ solution. Record the initial burette reading.
5. Titrate by adding $KMnO_4$ dropwise to the conical flask, swirling continuously.
6. The end point is the first permanent pink/purple colour that persists for at least 30 seconds (no indicator needed – $KMnO_4$ is self-indicating).
7. Record the final burette reading. Repeat to obtain concordant titres (within 0.10 cm^3).
8. Calculate mean titre → moles of MnO_4^- → moles of Fe^{2+} (1:5 ratio) → concentration or mass of iron.

Sources of Uncertainty:

- Pipette (25 cm^3): ±0.06 cm^3
- Burette: ±0.05 cm^3 per reading (±0.10 cm^3 per titre)
- Balance (if dissolving tablets): ±0.001 g

Common Errors:

- Using HCl instead of H_2SO_4 (Cl^- is oxidised by MnO_4^-, giving falsely high titres)
- Not adding enough acid (MnO_2 precipitates instead of forming Mn^{2+}, brown colour)
- Overshooting the end point (adding too much $KMnO_4$)
- Fe^{2+} solution being partially oxidised to Fe^{3+} before titration (lower titre)
- Parallax error when reading the burette ($KMnO_4$ is dark – read from top of meniscus)

Improvements:

- Always use dilute H_2SO_4 to acidify, never HCl or HNO_3
- Prepare Fe^{2+} solutions freshly to minimise oxidation to Fe^{3+}
- Add $KMnO_4$ dropwise near the end point and swirl after each drop
- For $KMnO_4$, read the burette at the top of the meniscus (solution is too dark for the bottom)
- Warm the solution to ~60°C if the reaction is slow ($MnO_4^-/C_2O_4^{2-}$ titrations require warming)
- Repeat and use concordant titres for calculation

Topic 15: Transition Metals

Topic 15A: Principles of Transition Metal Chemistry

15.1	*Be able to deduce the electronic configurations of atoms and ions of the d-block elements of Period 4 (Sc–Zn), given the atomic number and charge.*
	Atoms: fill 4s before 3d. E.g. Fe = [Ar] $3d^6$ $4s^2$. **Anomalies:** Cr = [Ar] $3d^5$ $4s^1$ and Cu = [Ar] $3d^{10}$ $4s^1$ (stability of half-full/full d-subshell). **Ions:** remove electrons from **4s first** (it is the higher energy sub-shell in ions). E.g. Fe^{2+} = [Ar] $3d^6$ (lost both 4s electrons). Fe^{3+} = [Ar] $3d^5$.

15.2	*Know that transition metals are d-block elements that form one or more stable ions with incompletely-filled d-orbitals.*
	A **transition metal** is a d-block element that forms at least one stable ion with an **incomplete d-subshell**. **Sc is NOT** a transition metal (Sc^{3+} = [Ar] $3d^0$ – empty d-orbitals). **Zn is NOT** a transition metal (Zn^{2+} = [Ar] $3d^{10}$ – full d-orbitals). All other Period 4 d-block elements (Ti to Cu) are transition metals.

15.3	*Understand why transition metals show variable oxidation number.*
	Transition metals have electrons in both the **3d and 4s sub-shells**, which have **similar energies**. Different numbers of electrons can be lost, giving **multiple stable oxidation states**. E.g. Fe can be +2 (lose $4s^2$) or +3 (lose $4s^2$ and one 3d).

15.4	*Know what is meant by the term 'ligand'.*
	A **ligand** is a molecule or ion that has a **lone pair of electrons** which it can **donate** to a central metal ion to form a **dative (coordinate) covalent bond**. Examples: H_2O, NH_3, Cl^-, OH^-, CN^-.

15.5	*Understand that dative (coordinate) bonding is involved in the formation of complex ions.*
	A **complex ion** is formed when ligands donate **lone pairs** to the central metal ion via **dative covalent (coordinate) bonds**. The metal ion acts as an electron pair **acceptor** (Lewis acid) and the ligand acts as an electron pair **donor** (Lewis base).

15.6	*Know that a complex ion is a central metal ion surrounded by ligands.*
	A **complex ion** consists of a **central transition metal ion** bonded to a number of **ligands** by dative covalent bonds. The complex ion carries an overall charge. E.g. $[Cu(H_2O)_6]^{2+}$, $[CuCl_4]^{2-}$, $[Fe(CN)_6]^{3-}$.

15.7	*Know that transition metals form coloured ions in solution.*
	Transition metal ions in solution are **coloured** because of **partially filled d-orbitals**. Common colours: $[Cu(H_2O)_6]^{2+}$ = blue, $[Fe(H_2O)_6]^{2+}$ = pale green, $[Fe(H_2O)_6]^{3+}$ = yellow/brown, $[Cr(H_2O)_6]^{3+}$ = green/violet, $[Co(H_2O)_6]^{2+}$ = pink.

15.8	*Understand that colour results from the splitting of d-orbital energy levels by ligands.*
	In a complex, the **d-orbitals are split** into two groups of different energy by the ligands. Electrons can be **promoted** from the lower to the higher energy d-orbitals by absorbing **visible light** of a specific frequency. The colour we see is the **complementary colour** of the light absorbed. The energy gap (ΔE) between the split d-orbitals determines the frequency: ΔE = hv.

15.9	*Understand why there is a lack of colour in some aqueous ions and other complex ions.*
	Ions with **d^0** (e.g. Sc^{3+}) or **d^{10}** (e.g. Zn^{2+}, Cu^+) configurations are **colourless** because there are either **no d-electrons** to promote or the d-orbitals are **full** (no empty orbital to accept a promoted electron). Without a d-d transition, no visible light is absorbed.

15.10	*Understand that colour changes may arise from changes in: (i) oxidation number, (ii) ligand, (iii) coordination number.*
	Any change that alters the **d-orbital splitting (ΔE)** changes the colour: **(i) Oxidation number:** different charge on the metal ion changes the interaction with ligands. E.g. Fe^{2+} (pale green) vs Fe^{3+} (yellow). **(ii) Ligand:** different ligands split d-orbitals by different amounts. E.g. $[Cu(H_2O)_6]^{2+}$ (blue) vs $[Cu(NH_3)_4(H_2O)_2]^{2+}$ (deep blue). **(iii) Coordination number:** changing the number of ligands changes the splitting. E.g. $[Co(H_2O)_6]^{2+}$ (pink, octahedral) vs $[CoCl_4]^{2-}$ (blue, tetrahedral).

15.11	*Understand the meaning of the term 'coordination number'.*
	The **coordination number** is the number of **dative covalent bonds** from ligands to the central metal ion. E.g. $[Cu(H_2O)_6]^{2+}$ has coordination number 6. $[CuCl_4]^{2-}$ has coordination number 4.

15.12	*Understand why H_2O, OH^- and NH_3 act as monodentate ligands.*
	Monodentate ligands form **one dative bond** per ligand. H_2O has two lone pairs on O but only uses **one** to bond. NH_3 has one lone pair on N. OH^- has lone pairs on O but uses **one** to bond. Each donates **one lone pair** to the metal ion.

15.13	*Understand why complexes with six-fold coordination have an octahedral shape.*
	When **six monodentate ligands** surround a central metal ion, they arrange themselves at **90°** to each other to minimise repulsion, giving an **octahedral shape**. Bond angles: **90°**. E.g. $[Fe(H_2O)_6]^{2+}$, $[Cr(NH_3)_6]^{3+}$.

15.14	*Know that transition metal ions may form tetrahedral complexes with relatively large ligands such as Cl^-.*
	Large ligands like Cl^- cannot fit six around a central metal ion due to **steric effects**. Only **four** Cl^- ligands can coordinate, giving a **tetrahedral** shape with bond angles of **109.5°**. E.g. $[CuCl_4]^{2-}$, $[CoCl_4]^{2-}$. The coordination number changes from 6 to 4.

15.15	*Know that square planar complexes are formed by transition metal ions and that cisplatin is an example.*
	Square planar complexes have coordination number **4** with bond angles of **90°**, with all four ligands in the same plane. E.g. **cisplatin** = $[PtCl_2(NH_3)_2]$ – used in cancer treatment. The Pt^{2+} ion is surrounded by 2 Cl^- and 2 NH_3 in a square planar arrangement.

15.16	*Understand why cisplatin is supplied as a single isomer and not mixed with the trans form.*
	Cisplatin (cis isomer) has both Cl ligands on the **same side**. The **trans isomer** (transplatin) has Cl on opposite sides. Only the **cis isomer is active** against cancer – its geometry allows it to bind to DNA in a specific way (cross-linking adjacent bases on the same strand), preventing replication and triggering cell death. The trans isomer binds differently and is **not therapeutically active**.

15.17	*Be able to identify bidentate ligands (e.g. $NH_2CH_2CH_2NH_2$) and multidentate ligands (e.g. $EDTA^{4-}$).*
	Bidentate ligands have **two lone pairs** that can both form dative bonds simultaneously. E.g. ethane-1,2-diamine (en = $NH_2CH_2CH_2NH_2$) – each N donates a lone pair. Ethanedioate ($C_2O_4^{2-}$) is also bidentate. **Multidentate** ligands form **more than two** dative bonds. E.g. $EDTA^{4-}$ forms **6 dative bonds** (it is hexadentate).

15.18	*Know that haemoglobin is an iron(II) complex containing a multidentate ligand.*
	Haemoglobin contains an **Fe^{2+}** ion at the centre of a **haem group**, which is a multidentate ligand. The Fe^{2+} forms coordinate bonds to the nitrogen atoms of the porphyrin ring. One coordination site is available for **O_2 binding**, allowing haemoglobin to transport oxygen in the blood.

15.19	*Know that a ligand exchange reaction occurs when an O_2 molecule bound to haemoglobin is replaced by a CO molecule.*
	Carbon monoxide (CO) is a stronger ligand than O_2 – it binds **more strongly** to Fe^{2+} in haemoglobin. When CO is inhaled, it **displaces O_2** from haemoglobin in a ligand exchange reaction, forming **carboxyhaemoglobin**. This is effectively **irreversible** under normal conditions, preventing oxygen transport. This is why CO is **toxic**.

Topic 15B: Reactions of Transition Metal Elements

15.20	*Know the colours of the oxidation states of vanadium (+5, +4, +3 and +2) in its compounds.*
	V^{5+} (VO_2^+): yellow **V^{4+} (VO^{2+}):** blue **V^{3+}:** green **V^{2+}:** purple/violet The stepwise reduction is typically carried out using zinc in acidic conditions: yellow → blue → green → purple/violet. (Remember: It's very nearly 5 o'clock – You Better Get Prepared)

15.21	*Understand redox reactions for the interconversion of vanadium oxidation states, in terms of the relevant E⊖ values.*
	Vanadium can be stepwise reduced from V^{5+} to V^{2+} using zinc (E = −0.76 V) in acid: $VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$ (E = +1.00 V) $VO^{2+} + 2H^+ + e^- \rightarrow V^{3+} + H_2O$ (E = +0.34 V) $V^{3+} + e^- \rightarrow V^{2+}$ (E = −0.26 V) Zinc is a strong enough reducing agent to reduce V through all states to V^{2+}. Each step is feasible because Ecell is positive when combined with the Zn^{2+}/Zn half-cell.

15.22	*Understand, in terms of E⊖ values, that $Cr_2O_7^{2-}$: (i) can be reduced to Cr^{3+} and Cr^{2+} using zinc in acid, (ii) can be produced by oxidation of Cr^{3+} using H_2O_2 in alkaline conditions.*
	(i) $Cr_2O_7^{2-}$ (orange, Cr^{6+}) is reduced to Cr^{3+} (green) and then Cr^{2+} (blue) using zinc in acidic conditions. Both steps are feasible because Ecell is positive. **(ii)** Cr^{3+} (green) can be oxidised back to CrO_4^{2-} (yellow, Cr^{6+}) using H_2O_2 in **alkaline** conditions, then acidified to give $Cr_2O_7^{2-}$ (orange).

15.23	*Know that $Cr_2O_7^{2-}$ can be converted into CrO_4^{2-} via the equilibrium: $2CrO_4^{2-} + 2H^+ \rightleftharpoons Cr_2O_7^{2-} + H_2O$.*
	This is an **acid-base equilibrium** (not redox – Cr stays as +6 throughout): **$2CrO_4^{2-}(aq) + 2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + H_2O(l)$** • In **acidic conditions:** equilibrium shifts **right** → orange $Cr_2O_7^{2-}$ predominates. • In **alkaline conditions:** equilibrium shifts **left** → yellow CrO_4^{2-} predominates. Adding H^+ (acid) shifts right; adding OH^- (alkali) removes H^+ and shifts left.

15.24	*Be able to record observations and write equations for reactions of Cr^{3+}, Fe^{2+}, Fe^{3+}, Co^{2+} and Cu^{2+} with aqueous NaOH and aqueous ammonia, including in excess.*
	With NaOH(aq) – small amount: all form coloured hydroxide precipitates: • $Cr^{3+} \rightarrow Cr(OH)_3$ (green); $Fe^{2+} \rightarrow Fe(OH)_2$ (green); $Fe^{3+} \rightarrow Fe(OH)_3$ (brown/rust); $Co^{2+} \rightarrow Co(OH)_2$ (blue, turns pink on standing); $Cu^{2+} \rightarrow Cu(OH)_2$ (blue). **With excess NaOH:** $Cr(OH)_3$ **dissolves** (amphoteric) to form green $[Cr(OH)_6]^{3-}$. Others: **no change**. **With NH_3(aq) – small amount:** same precipitates as NaOH. **With excess NH_3:** $Cr(OH)_3$ – no change. $Cu(OH)_2$ dissolves to form deep blue $[Cu(NH_3)_4(H_2O)_2]^{2+}$. $Co(OH)_2$ dissolves to form straw/brown $[Co(NH_3)_6]^{2+}$. $Fe(OH)_2$ and $Fe(OH)_3$ – **no change**.

15.25	*Be able to write ionic equations to show the difference between ligand exchange and amphoteric behaviour.*
	Ligand exchange (with excess NH_3): NH_3 molecules **replace H_2O** ligands. E.g. $[Cu(H_2O)_6]^{2+} + 4NH_3 \rightarrow [Cu(NH_3)_4(H_2O)_2]^{2+} + 4H_2O$ **Amphoteric behaviour** (with excess OH^-): the hydroxide precipitate **dissolves** in excess OH^-. E.g. $Cr(OH)_3 + 3OH^- \rightarrow [Cr(OH)_6]^{3-}$ Key difference: ligand exchange changes the **ligands**; amphoteric behaviour involves the hydroxide **reacting with OH^-** to dissolve.

15.26	*Understand ligand exchange and colour changes in the formation of: (i) $[Cu(NH_3)_4(H_2O)_2]^{2+}$ from $[Cu(H_2O)_6]^{2+}$; (ii) $[CuCl_4]^{2-}$ from $[Cu(H_2O)_6]^{2+}$; (iii) $[CoCl_4]^{2-}$ from $[Co(H_2O)_6]^{2+}$.*
	(i) $[Cu(H_2O)_6]^{2+}$ (**pale blue**) + $4NH_3 \rightarrow [Cu(NH_3)_4(H_2O)_2]^{2+}$ (**deep blue**) + $4H_2O$. Partial exchange: 4 of 6 water ligands replaced. Coordination number stays 6. **(ii)** $[Cu(H_2O)_6]^{2+}$ (**pale blue**) + $4Cl^- \rightarrow [CuCl_4]^{2-}$ (**yellow/green**) + $6H_2O$. Coordination number 6 → 4 (octahedral → tetrahedral). **(iii)** $[Co(H_2O)_6]^{2+}$ (**pink**) + $4Cl^- \rightarrow [CoCl_4]^{2-}$ (**blue**) + $6H_2O$. Octahedral → tetrahedral. Reversible.

15.27	*Understand that substitution of small, uncharged ligands by larger, charged ligands can lead to a change in coordination number.*
	When **small neutral ligands** (e.g. H_2O) are replaced by **larger charged ligands** (e.g. Cl^-), the **coordination number decreases** from 6 to 4. The larger Cl^- ligands experience greater **steric repulsion** and cannot fit 6 around the metal ion. The shape changes from **octahedral to tetrahedral**.

15.28	*Understand, in terms of a large positive ΔSsystem, that substitution of monodentate by bidentate or multidentate ligands leads to a more stable complex (chelate effect).*
	The **chelate effect:** replacing monodentate ligands with bidentate/multidentate ligands is **thermodynamically favourable** because: • The number of **free particles increases** (e.g. 1 $EDTA^{4-}$ replaces 6 H_2O → net increase of 5 free particles). • This gives a **large positive ΔSsystem** (increased disorder). • ΔG becomes more negative, so the chelated complex is more **stable** (larger K). This is why EDTA is used in water treatment, medicine, and analytical chemistry.

15.29	*Know that transition metals and their compounds can act as heterogeneous and homogeneous catalysts.*
	Transition metals are effective catalysts because they can: • **Change oxidation state** easily (accept and donate electrons) • **Provide a surface** for reactants to adsorb onto (heterogeneous) • Form **intermediates** with reactants (homogeneous) Examples: Fe in the Haber process, V_2O_5 in the Contact process (heterogeneous), Fe^{2+} catalysing $I^-/S_2O_8^{2-}$ (homogeneous).

15.30	*Know that a heterogeneous catalyst is in a different phase from the reactants and that the reaction occurs at the surface.*
	A **heterogeneous catalyst** is in a **different phase** (usually solid catalyst, gaseous/liquid reactants). The reaction occurs at the **surface** of the catalyst. Reactant molecules **adsorb** onto the surface, bonds are weakened, reaction occurs, and products **desorb**. Surface area matters – finely divided catalysts are more effective.

15.31	*Understand, in terms of oxidation number, how V_2O_5 acts as a catalyst in the Contact process.*
	In the Contact process: $2SO_2 + O_2 \rightarrow 2SO_3$ **Step 1:** V_2O_5 oxidises SO_2 to SO_3. Vanadium is **reduced** from +5 to +4: $SO_2 + V_2O_5 \rightarrow SO_3 + V_2O_4$ **Step 2:** V_2O_4 is **re-oxidised** back to V_2O_5 by oxygen: $V_2O_4 + ½O_2 \rightarrow V_2O_5$ The V_2O_5 is **regenerated** – not consumed overall. The catalyst provides an **alternative pathway** with lower activation energy, using its variable oxidation state.

15.32	*Understand how a catalytic converter decreases CO and NO emissions by: (i) adsorption, (ii) weakening of bonds and reaction, (iii) desorption of products.*
	(i) Adsorption: CO and NO molecules **adsorb** onto the surface of the catalyst (Pt/Pd/Rh). This holds them in place and in the correct orientation. **(ii) Bonds are weakened:** adsorption weakens the bonds within CO and NO. The molecules react on the surface: $2CO + 2NO \rightarrow 2CO_2 + N_2$. **(iii) Desorption:** the products (CO_2 and N_2) are **less strongly adsorbed** and leave (desorb from) the surface, freeing it for more reactant molecules.

15.33	*Know that a homogeneous catalyst is in the same phase as the reactants and that the reaction proceeds via an intermediate.*
	A **homogeneous catalyst** is in the **same phase** as the reactants (typically all in aqueous solution). It works by forming an **intermediate species** in one step, which then reacts in a second step to regenerate the catalyst. The catalyst provides an **alternative pathway** with a lower activation energy.

15.34	*Understand the role of Fe^{2+} ions in catalysing the reaction between I^- and $S_2O_8^{2-}$ ions.*
	The direct reaction between I^- and $S_2O_8^{2-}$ is slow (both are negatively charged – repulsion). **Step 1:** $2Fe^{2+} + S_2O_8^{2-} \rightarrow 2Fe^{3+} + 2SO_4^{2-}$ (Fe^{2+} is oxidised to Fe^{3+}) **Step 2:** $2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$ (Fe^{3+} is reduced back to Fe^{2+}) **Overall:** $S_2O_8^{2-} + 2I^- \rightarrow 2SO_4^{2-} + I_2$. Fe^{2+} is regenerated. Each step involves **oppositely charged ions** (less repulsion, lower Ea).

15.35	*Know the role of Mn^{2+} ions in autocatalysing the reaction between MnO_4^- and $C_2O_4^{2-}$ ions.*
	Autocatalysis: the reaction starts **slowly** because there is no catalyst initially. As the reaction proceeds, **Mn^{2+} ions are produced** as a product: $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$ The Mn^{2+} produced **catalyses** the further reaction. So the reaction **speeds up** as it progresses (the product is the catalyst). The purple $KMnO_4$ is decolourised slowly at first, then increasingly rapidly.

⭐ CORE PRACTICAL 12: The Preparation of a Transition Metal Complex

Method:

1. Typically the preparation of $[Cu(NH_3)_4(H_2O)_2]SO_4$ (tetraamminecopper(II) sulfate).
2. Dissolve copper(II) sulfate (e.g. ~6 g $CuSO_4 \cdot 5H_2O$) in the minimum volume of warm distilled water (~10 cm^3) in a beaker.
3. Add concentrated ammonia solution dropwise with stirring. A pale blue precipitate of $Cu(OH)_2$ forms initially.
4. Continue adding ammonia until the precipitate dissolves completely to give a deep blue/violet solution of $[Cu(NH_3)_4(H_2O)_2]^{2+}$.
5. Add ethanol (approximately equal volume) to reduce the solubility of the complex and cause crystallisation. Stir gently.
6. Filter the crystals under reduced pressure using a Buchner funnel and flask. Wash the crystals with a small amount of ethanol, then a small amount of diethyl ether.
7. Allow the crystals to dry on filter paper (do NOT use a hot oven as the complex may decompose and lose NH_3 ligands).
8. Weigh the dry product and calculate the percentage yield.

Sources of Uncertainty:

- Balance: ±0.001 g (for mass of reactant and product)
- Measuring cylinder for volumes: ±0.5 cm^3
- Judgement of when precipitate has fully dissolved

Common Errors:

- Not adding enough ammonia ($Cu(OH)_2$ precipitate remains, reducing yield and purity)
- Adding too much water when dissolving $CuSO_4$ (crystals are slow to form or too small)
- Not washing crystals sufficiently (impurities remain, reducing purity)
- Drying at too high a temperature (complex decomposes, losing ammonia ligands)
- Losses during filtration and transfer (reduces yield)

Improvements:

- Add ammonia dropwise with continuous stirring to ensure the precipitate fully dissolves before adding excess
- Use the minimum volume of water to dissolve $CuSO_4$ (supersaturated solutions crystallise more readily)
- Wash crystals with a small volume of cold ethanol then diethyl ether (ether evaporates quickly, aiding drying)
- Dry at room temperature or in a desiccator, not in an oven
- Weigh the product accurately and calculate percentage yield to assess the success of the preparation
- Confirm identity by recording the colour and, if available, measuring the melting point

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