Chemistry with Chloe

Expert A Level Chemistry Tuition

Edexcel A Level Chemistry

Paper 2: Advanced Organic and Physical Chemistry

9CH0/02

REVISION GUIDE

Topics 2, 3, 5, 6, 7, 9, 16, 17, 18 & 19

Bonding & Structure • Redox I • Formulae, Equations and Amounts of Substance

Organic Chemistry I & II & III • Analytical Techniques I & II • Kinetics I & II

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How to Use This Guide

This revision guide covers the specification points for Topics 2, 3, 5, 6, 7, 9, 16, 17, 18 and 19 of the Edexcel A Level Chemistry Paper 2 exam (9CH0/02). Topics 2, 3 and 5 also appear in Paper 1.

Quick notes for each topic at the start of the guide are then followed by complete (but concise) notes for each topic.

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

Core Practicals are presented in orange boxes where they fall in the specification, with step-by-step methods, sources of uncertainty, common errors, and improvements.

Use this guide alongside past papers. Tick off spec points as you revise them.

Good luck with your revision! – Chloe

Quick Reference: Key Organic Synthesis Routes

- ALKENE → ALCOHOL: steam + H_3PO_4 (300°C, 60 atm)
- ALKENE → HALOGENOALKANE: HBr (electrophilic addition)
- ALKENE → DIOL: cold, dilute, acidified $KMnO_4$

- ALCOHOL → ALDEHYDE: $K_2Cr_2O_7/H_2SO_4$, distil immediately
- ALCOHOL → CARBOXYLIC ACID: $K_2Cr_2O_7/H_2SO_4$, reflux
- ALCOHOL → ALKENE: conc. H_3PO_4, heat (elimination/dehydration)
- ALCOHOL → HALOGENOALKANE: PCl_5 (chloro) / KBr + 50% H_2SO_4 (bromo) / red Phosphorus + I_2 (iodo)

- HALOGENOALKANE → ALCOHOL: aq KOH, reflux (nucleophilic substitution)
- HALOGENOALKANE → NITRILE: KCN in ethanol, reflux (carbon chain +1)
- HALOGENOALKANE → AMINE: excess conc. NH_3, sealed tube, heat
- HALOGENOALKANE → ALKENE: ethanolic KOH, reflux (elimination)
- HALOGENOALKANE → GRIGNARD: Mg in dry ether → RMgX

- NITRILE → CARBOXYLIC ACID: acid hydrolysis (HCl/H_2O, reflux)
- NITRILE → AMINE: $LiAlH_4$ in dry ether (reduction)

- CARBOXYLIC ACID → ESTER: + alcohol, conc. H_2SO_4 catalyst, reflux
- CARBOXYLIC ACID → ACYL CHLORIDE: + PCl_5
- CARBOXYLIC ACID → ALCOHOL: $LiAlH_4$ in dry ether (reduction)

- GRIGNARD + CO_2 → CARBOXYLIC ACID (after acid hydrolysis)
- GRIGNARD + ALDEHYDE → SECONDARY ALCOHOL
- GRIGNARD + KETONE → TERTIARY ALCOHOL

- AROMATIC: nitrobenzene + Sn/conc. HCl, reflux → phenylamine

Topic 5: Formulae, Equations and Amounts of Substance Quick Notes

Formula / Equation	What It Means / When to Use It
n = m / M	Moles = mass (g) / molar mass (g mol^{-1})
n = c × V	Moles = concentration (mol dm^{-3}) × volume (dm^3)
n = V / 24,000	Moles of gas = volume (cm^3) / 24,000 at RTP
n = V / 24	Moles of gas = volume (dm^3) / 24 at RTP
pV = nRT	Ideal gas equation. p in Pa, V in m^3, T in K
Conc (g dm^{-3}) = c × M	Concentration (g dm^{-3}) = mol dm^{-3} × molar mass
No. of particles = n × L	Number of particles = moles × Avogadro constant

Yields and Atom Economy

Formula / Equation	What It Means / When to Use It
% yield = (actual / theoretical) × 100	Always < 100% (losses, side reactions)
Atom economy = (Mr desired / ΣMr all products) × 100	Higher = less waste. Addition = 100%

Uncertainties

Formula / Equation	What It Means / When to Use It
% uncertainty = (uncertainty / reading) × 100	Minimise by using larger measurements

- Burette: ±0.05 cm^3 per reading (±0.10 cm^3 per titre)
- 25 cm^3 pipette: ±0.06 cm^3 | 250 cm^3 volumetric flask: ±0.10 cm^3
- Balance: ±0.001 g (3 d.p.) or ±0.01 g (2 d.p.)

Topic 6: Organic Chemistry I Quick Notes

General Formulae

Formula / Reaction	What It Means / When to Use It
Alkanes: C_nH_{2n+2}	Saturated hydrocarbons. Cycloalkanes: C_nH_{2n}
Alkenes: C_nH_{2n}	Unsaturated (C=C). Cycloalkenes: C_nH_{2n-2}
Alcohols: $C_nH_{2n+1}OH$	Functional group: –OH

IUPAC Prefixes

- meth (1) eth (2) prop (3) but (4) pent (5) hex (6) hept (7) oct (8) non (9) dec (10)

Radical Substitution Mechanism (UV light)

- Initiation: $Cl_2 \rightarrow 2Cl\bullet$ (homolytic fission by UV)
- Propagation: $Cl\bullet + CH_4 \rightarrow CH_3\bullet + HCl$ then $CH_3\bullet + Cl_2 \rightarrow CH_3Cl + Cl\bullet$
- Termination: any two radicals combine, e.g. $Cl\bullet + Cl\bullet \rightarrow Cl_2$

Key Alkene Addition Reactions

Reaction	Reagents & Conditions	Product
Alkene + H_2	Ni catalyst, 150°C	Alkane (hydrogenation)
Alkene + Br_2	Room temp	Dihalogenoalkane
Alkene + HBr	Room temp, gas	Halogenoalkane (Markovnikov)
Alkene + H_2O	H_3PO_4 catalyst, 300°C, 60 atm	Alcohol (hydration)
Alkene + $KMnO_4$	Cold, dilute, acidified	Diol

- Markovnikov's rule: H adds to the C with MORE H atoms → more stable carbocation intermediate.
- Carbocation stability: tertiary > secondary > primary (inductive effect of alkyl groups).
- Test for C=C: bromine water is decolourised (orange → colourless).

Halogenoalkane Reactions

Reaction	Reagents & Conditions	Product
R–X + aq KOH	Reflux, aqueous	Alcohol (nucleophilic sub.)
R–X + KCN	Reflux, ethanolic	Nitrile (C chain +1)
R–X + $2NH_3$	Excess NH_3, sealed tube, heat	Primary amine
R–X + ethanolic KOH	Reflux, ethanol solvent	Alkene (elimination)

- Reactivity: iodoalkanes > bromoalkanes > chloroalkanes (C–I weakest bond → easiest to break).
- Hydrolysis rate: tertiary > secondary > primary.
- Primary halogenoalkanes → SN2 mechanism. Tertiary → SN1 mechanism.

Alcohol Reactions

Reaction	Reagents & Conditions	Product
1° alcohol + $K_2Cr_2O_7/H_2SO_4$	Distil immediately	Aldehyde
1° alcohol + $K_2Cr_2O_7/H_2SO_4$	Reflux	Carboxylic acid
2° alcohol + $K_2Cr_2O_7/H_2SO_4$	Reflux or distil	Ketone
3° alcohol + $K_2Cr_2O_7/H_2SO_4$	No reaction	—
Alcohol + conc. H_3PO_4	Heat	Alkene (elimination)
Alcohol + PCl_5	Room temp	Chloroalkane + HCl fumes

- Orange dichromate → green = oxidation has occurred. No colour change = tertiary.
- Aldehyde test: Fehling's (brick-red ppt) or Tollens' (silver mirror). Ketones give no reaction.

Topic 7: Modern Analytical Techniques I Quick Notes

Mass Spectrometry – Common Fragment Losses

Formula / Reaction	What It Means / When to Use It
M^+ peak	Molecular ion – gives relative molecular mass (Mr)
–15 = loss of CH_3	Methyl group lost
–17 = loss of OH	Hydroxyl group lost
–28 = loss of CO	Carbonyl group lost
–29 = loss of CHO or C_2H_5	Aldehyde group or ethyl group
–31 = loss of CH_2OH	From a primary alcohol
–45 = loss of OC_2H_5	Ethoxy group (from ethyl ester)

Infrared Spectroscopy – Key Absorptions

Formula / Reaction	What It Means / When to Use It
O–H (alcohol)	Broad, ~3230–3550 cm^{-1}
O–H (carboxylic acid)	Very broad, ~2500–3300 cm^{-1}
C=O (aldehyde, ketone, acid, ester)	Strong, sharp, ~1680–1750 cm^{-1}
C=C (alkene)	~1620–1680 cm^{-1}
C–H (alkane)	~2850–2960 cm^{-1}
N–H (amine)	~3300–3500 cm^{-1}

- Carboxylic acid: look for BOTH broad O–H AND strong C=O together.
- Fingerprint region (below ~1500 cm^{-1}) is unique to each compound.

Topic 9: Kinetics I Quick Notes

Formula / Reaction	What It Means / When to Use It
Rate = 1 / time	Simple rate from clock-type experiments
Rate = gradient of tangent	From concentration–time graph at time t

- Increasing temperature: Maxwell–Boltzmann curve shifts right and flattens. Much greater area beyond Ea.
- Catalyst: provides alternative pathway with lower Ea. MB curve unchanged, but Ea line moves left.
- Catalyst does NOT change ΔH, does NOT change equilibrium position, does NOT appear in the overall equation.

Topic 16: Kinetics II

Formula / Reaction	What It Means / When to Use It
Rate = $k[A]^m[B]^n$	Rate equation. m, n determined experimentally only.
ln k = ln A – Ea/RT	Arrhenius equation (y = mx + c form)
Plot: ln k vs 1/T	Straight line. Gradient = –Ea/R
Ea = –gradient × R	R = 8.314 J K^{-1} mol^{-1}. T in Kelvin.

Determining Order from Initial Rate Data

- [A] doubles, rate unchanged → zero order in A
- [A] doubles, rate doubles → first order in A
- [A] doubles, rate quadruples → second order in A

Recognising Order from Graphs

- Conc–time: zero order = straight line; first order = exponential decay (constant t½); second order = steeper curve
- Rate–conc: zero order = horizontal; first order = straight line through origin; second order = curve through origin

Units of Rate Constant k

- Zero order: mol dm^{-3} s^{-1} First order: s^{-1} Second order: dm^3 mol^{-1} s^{-1}

Rate-Determining Step and Mechanisms

- Species in the rate equation = species in (or before) the rate-determining step.
- SN1: rate = k[halogenoalkane] → only halogenoalkane in rds. Tertiary. Racemic mixture.
- SN2: rate = k[halogenoalkane][OH^-] → both in rds. Primary. Inversion of configuration.

Iodination of Propanone

- Rate = k[CH_3COCH_3][H^+] (zero order in I_2 – I_2 is NOT in the rds)

Topic 17: Organic Chemistry II Quick Notes

Carbonyl Compound Tests

Reaction	Reagents & Conditions	Product
2,4-DNPH + carbonyl	Room temp	Orange/yellow precipitate (tests for C=O)
Fehling's + aldehyde	Heat	Brick-red precipitate (Cu_2O)
Tollens' + aldehyde	Warm	Silver mirror (Ag deposited)
I_2/NaOH + CH_3CO–	Warm	Yellow ppt of CHI_3 (iodoform test)
$K_2Cr_2O_7/H^+$ + aldehyde	Warm	Orange → green (Cr^{3+})

- Ketones give NO reaction with Fehling's, Tollens', or acidified dichromate.
- Iodoform test is positive for ethanal AND methyl ketones (CH_3CO– group).
- 2,4-DNPH derivative: recrystallise and measure m.p. to identify the specific carbonyl compound.

Carbonyl Reactions

Reaction	Reagents & Conditions	Product
Aldehyde + $LiAlH_4$	Dry ether	Primary alcohol
Ketone + $LiAlH_4$	Dry ether	Secondary alcohol
Carbonyl + HCN/KCN	Room temp (nucleophilic addition)	Hydroxynitrile (+ chiral centre)

- HCN mechanism: CN^- attacks δ+C of C=O (nucleophilic addition). Produces racemic mixture if new chiral centre.

Carboxylic Acid and Ester Reactions

Reaction	Reagents & Conditions	Product
RCOOH + NaOH	Room temp	RCOONa + H_2O (salt)
RCOOH + PCl_5	Room temp	RCOCl + $POCl_3$ + HCl
RCOOH + ROH	Conc. H_2SO_4 catalyst, reflux	Ester + H_2O (reversible)
RCOOH + $LiAlH_4$	Dry ether	Primary alcohol
Ester + dilute H_2SO_4	Reflux (acid hydrolysis)	RCOOH + ROH (reversible)
Ester + NaOH	Reflux (alkaline hydrolysis)	RCOONa + ROH (irreversible)

Acyl Chloride Reactions (vigorous, room temp, no catalyst)

Reaction	Reagents & Conditions	Product
RCOCl + H_2O		RCOOH + HCl
RCOCl + ROH		Ester + HCl
RCOCl + $2NH_3$		$RCONH_2$ + NH_4Cl (amide)
RCOCl + RNH_2		RCONHR + HCl (N-sub amide)

Chirality

- Chiral centre = carbon with 4 different groups → optical isomers (enantiomers).
- Racemic mixture = 50:50 mix of enantiomers → no net optical rotation.
- SN1 → racemic mixture (flat carbocation, attack from both sides). SN2 → inversion.

Topic 18: Organic Chemistry III Quick Notes

Benzene

- Delocalisation energy: benzene is 152 kJ mol^{-1} more stable than hypothetical cyclohexatriene.
- All C–C bonds equal length (140 pm) – between single (154 pm) and double (134 pm).
- Benzene undergoes electrophilic SUBSTITUTION (not addition) – preserves the stable delocalised ring.

Benzene Reactions and Electrophile Generation

Reaction	Reagents & Conditions	Product
Bromination	Br_2 + $AlBr_3$ catalyst	C_6H_5Br + HBr
Nitration	Conc. HNO_3 + conc. H_2SO_4, 50°C	$C_6H_5NO_2$ + H_2O
Friedel-Crafts (alkylation)	RCl + $AlCl_3$ catalyst	C_6H_5R + HCl
Friedel-Crafts (acylation)	RCOCl + $AlCl_3$ catalyst	C_6H_5COR + HCl

- Electrophile generation: $Br_2 + AlBr_3 \rightarrow Br^+ + AlBr_4^-$ | $HNO_3 + H_2SO_4 \rightarrow NO_2^+ + HSO_4^- + H_2O$
- Phenol + Br_2 water → 2,4,6-tribromophenol (white ppt). No catalyst needed (–OH activates ring).

Amines

Reaction	Reagents & Conditions	Product
RX + $2NH_3$	Excess NH_3, sealed tube, heat	RNH_2 + NH_4X (1° amine)
RCN + 4[H]	$LiAlH_4$ in dry ether (or H_2/Ni)	RCH_2NH_2 (1° amine, chain +1)
$ArNO_2$ + 6[H]	Sn + conc. HCl, reflux	$ArNH_2$ + $2H_2O$ (aromatic amine)

- Basicity order: aliphatic amine > NH_3 > aromatic amine.
- Aliphatic: alkyl group donates electrons to N (increases lone pair availability).
- Aromatic: N lone pair delocalised into ring (less available for protonation).

Condensation Polymers

- Polyester: dicarboxylic acid + diol → ester link (–COO–) + H_2O
- Polyamide: dicarboxylic acid + diamine → amide link (–CONH–) + H_2O
- Polypeptide: amino acids → peptide bond (–CONH–) + H_2O
- Condensation polymers are biodegradable (ester/amide bonds can be hydrolysed).

Amino Acids

- Zwitterion form (neutral pH): $^+NH_3$–CHR–COO^- (internal salt)
- In acid: $^+NH_3$–CHR–COOH (positive ion). In alkali: NH_2–CHR–COO^- (negative ion).
- All amino acids (except glycine) are optically active – chiral α-carbon.

Grignard Reagents (Carbon Chain Extension)

- Formation: RX + Mg → RMgX (in dry ether)
- RMgX + CO_2 → RCOOH (after acid hydrolysis) – carboxylic acid, chain extended
- RMgX + HCHO → primary alcohol | RMgX + RCHO → secondary alcohol | RMgX + RCOR → tertiary alcohol
- Must use DRY ether – Grignard reagents are destroyed by water.

Topic 19: Modern Analytical Techniques II Quick Notes

High-Resolution Mass Spectrometry

- Measures Mr to 4 decimal places – can distinguish compounds with same integer Mr.
- Accurate masses: ^{12}C = 12.0000, ^{1}H = 1.0078, ^{16}O = 15.9949, ^{14}N = 14.0031

^{13}C NMR

- Number of peaks = number of different carbon environments.
- Equivalent carbons (by symmetry) give ONE peak.
- Typical δ values: 10–40 (alkyl C), 50–80 (C–O or C–N), 110–150 (aromatic/C=C), 160–220 (C=O)

^{1}H NMR

Formula / Reaction	What It Means / When to Use It
Number of peaks	= number of different H environments
Peak area / integration	= ratio of H atoms in each environment
(n+1) rule for splitting	n = number of adjacent non-equivalent H atoms
D_2O shake	Peak disappears = O–H or N–H (exchangeable proton)

- Singlet (s) = 0 neighbours | Doublet (d) = 1 | Triplet (t) = 2 | Quartet (q) = 3
- Typical δ values: 0.9 ($R–CH_3$), 1.3 ($R–CH_2–R$), 3.3–3.9 (C–H next to O), 6.5–8.0 (ArH), 9–10 (CHO), 11–12 (COOH)

Chromatography

Formula / Reaction	What It Means / When to Use It
Rf = distance moved by substance / distance moved by solvent front	Always between 0 and 1. Compare with standards.

- TLC/paper: Rf values identify substances. GC/HPLC: retention times identify substances.
- GC-MS / HPLC-MS: separation + identification. Used in forensics and drug testing in sport.

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** or **LEO the lion says GER** (Loss of Electrons is Oxidation, Gain of Electrons is Reduction).

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 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 × 6.02 × 10^{23} = 3.01 × 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 (m^3), 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 cm^3 at RTP) or n = V/24 (dm^3) • n = c × V (concentration in mol dm^{-3} × volume in dm^3)

	• 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

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

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 6: Organic Chemistry I

Topic 6A: Introduction to Organic Chemistry

6.1	*Know that a hydrocarbon is a compound of hydrogen and carbon only.*
	A **hydrocarbon** contains **only carbon and hydrogen atoms**. Examples include alkanes (e.g. CH_4), alkenes (e.g. C_2H_4), and arenes (e.g. C_6H_6). They contain no other elements.

6.2	*Be able to represent organic molecules using empirical, molecular, general, structural, displayed and skeletal formulae.*
	Empirical formula: simplest whole number ratio (e.g. CH_2 for ethene). **Molecular formula:** actual number of atoms (e.g. C_2H_4). **General formula:** formula for a homologous series (e.g. C_nH_{2n+2} for alkanes). **Structural formula:** shows how atoms are grouped (e.g. CH_3CH_3). **Displayed formula:** shows every bond between every atom. **Skeletal formula:** shows the carbon backbone as zigzag lines; C and H atoms on the chain are not shown; functional groups and heteroatoms are shown.

6.3	*Know what is meant by the terms 'homologous series' and 'functional group'.*
	Homologous series: a family of compounds with the **same functional group** and **same general formula**, differing by CH_2 between successive members. They show a **gradual trend** in physical properties and **similar chemical properties**. **Functional group:** the atom or group of atoms responsible for the **characteristic chemical reactions** of a compound (e.g. C=C in alkenes, –OH in alcohols).

6.4	*Be able to name compounds using IUPAC nomenclature rules (prefixes up to C10).*
	IUPAC naming rules: (1) Find the **longest carbon chain** → parent name. (2) Number from the end nearest the functional group/branch. (3) Name branches as prefixes (methyl, ethyl, etc.). (4) Use di-, tri- for multiple identical groups. (5) Functional group determines the suffix. **Prefixes:** meth- (1), eth- (2), prop- (3), but- (4), pent- (5), hex- (6), hept- (7), oct- (8), non- (9), dec- (10).

6.5	*Be able to classify reactions as addition, elimination, substitution, oxidation, reduction, hydrolysis or polymerisation.*
	Addition: two molecules combine to form one product (no atoms lost). **Elimination:** a small molecule (e.g. H_2O, HBr) is removed from one molecule. **Substitution:** one atom/group is replaced by another. **Oxidation:** gain of oxygen / loss of hydrogen / loss of electrons. **Reduction:** loss of oxygen / gain of hydrogen / gain of electrons. **Hydrolysis:** a bond is broken by reaction with water (or aqueous OH^-/H^+). **Polymerisation:** many small molecules (monomers) join to form a long chain (polymer).

6.6	*Understand the term 'structural isomerism' and determine possible structural, displayed and skeletal formulae of an organic molecule.*
	Structural isomers have the **same molecular formula** but **different structural formulae** (different connectivity of atoms). Types include: **chain isomerism** (different carbon skeleton), **position isomerism** (same functional group in different positions), and **functional group isomerism** (different functional groups).

6.7	*Understand the term 'stereoisomerism', as illustrated by E/Z isomerism (including cis-trans isomerism).*
	Stereoisomers have the same structural formula but a **different arrangement of atoms in space**. **E/Z isomerism** occurs in alkenes where there is **restricted rotation** about the C=C and **two different groups on each carbon** of the double bond. **Z (cis):** higher priority groups on the **same side** of the double bond. **E (trans):** higher priority groups on **opposite sides**. Priority is assigned by **atomic number** (Cahn-Ingold-Prelog rules).

Topic 6B: Alkanes

6.8	*Know the general formula for alkanes.*
	C_nH_{2n+2} (where n = number of carbon atoms). E.g. methane CH_4, ethane C_2H_6, propane C_3H_8. Cycloalkanes have the general formula **C_nH_{2n}**.

6.9	*Know that alkanes and cycloalkanes are saturated hydrocarbons.*
	Saturated means the molecule contains **only single (sigma) bonds** between carbon atoms – no C=C double bonds. All bonds are C–C or C–H. This makes alkanes relatively **unreactive** compared to alkenes.

6.10	*Understand that alkane fuels are obtained from fractional distillation, cracking and reforming of crude oil.*
	Fractional distillation: crude oil is heated and separated into fractions based on **boiling point range** in a fractionating column. Shorter chains have lower b.p. and are collected near the top. **Cracking:** long-chain alkanes are broken into shorter, more useful molecules (shorter alkanes + alkenes). Thermal cracking (high T, high P) or catalytic cracking (zeolite catalyst, lower T). **Reforming:** straight-chain alkanes are converted into **branched-chain alkanes** and **cyclic hydrocarbons** for more efficient combustion (higher octane rating).

6.11	*Know that pollutants are formed during combustion of alkane fuels, including CO, NOx, SOx, carbon particulates and unburned hydrocarbons.*
	Carbon monoxide (CO): from **incomplete combustion** (insufficient oxygen). **Nitrogen oxides (NO_x):** formed at high temperatures when N_2 and O_2 from the air react. **Sulfur dioxide (SO_2):** from combustion of sulfur impurities in the fuel. **Carbon particulates:** soot, from incomplete combustion. **Unburned hydrocarbons:** from incomplete combustion; contribute to smog.

6.12	*Understand problems from pollutants: toxicity of CO, acidity of NOx and SOx.*
	CO is toxic: it binds to haemoglobin more strongly than O_2, forming **carboxyhaemoglobin**, preventing oxygen transport in the blood. CO is colourless and odourless, making it especially dangerous. **NO_x and SO_x cause acid rain:** they dissolve in atmospheric water to form **nitric acid (HNO_3)** and **sulfuric acid (H_2SO_4)**, which damage buildings, ecosystems and aquatic life.

6.13	*Understand how a catalytic converter solves some pollutant problems.*
	A **catalytic converter** (Pt/Pd/Rh catalyst) converts toxic exhaust gases into less harmful products: **$2CO + 2NO \rightarrow 2CO_2 + N_2$** (removes both CO and NO simultaneously). CO is oxidised to CO_2; NO is reduced to N_2. The catalyst provides a surface for adsorption, weakening bonds and enabling reaction (see Topic 15).

6.14	*Understand the use of alternative fuels, including biodiesel and alcohols from renewable sources.*
	Biofuels (biodiesel, bioethanol) are made from **renewable** plant sources. **Advantages:** carbon-neutral in theory (CO_2 released = CO_2 absorbed during growth); reduce dependence on finite fossil fuels. **Disadvantages:** large land area needed (competition with food crops); energy input for farming/processing; not truly carbon-neutral when full lifecycle is considered.

6.15	*Know that a radical: (i) is a species with an unpaired electron, (ii) is formed by homolytic fission.*
	(i) A **radical** (free radical) has an **unpaired electron**, shown as a single dot (e.g. Cl•). Radicals are very **reactive** because they seek to pair their unpaired electron. **(ii) Homolytic fission:** a covalent bond breaks so that **each atom takes one electron** from the bond. This produces two radicals. Requires energy (UV light or heat). E.g. $Cl_2 \rightarrow 2Cl•$

6.16	*Understand reactions of alkanes with: (i) oxygen (combustion), (ii) halogens (radical substitution mechanism).*
	(i) Combustion: alkane + $O_2 \rightarrow CO_2 + H_2O$ (complete). Incomplete: produces CO, C, or H_2O. **(ii) Radical substitution** (e.g. $CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$ in UV light): **Initiation:** $Cl_2 \rightarrow 2Cl•$ (UV light breaks Cl–Cl bond homolytically) **Propagation:** $Cl• + CH_4 \rightarrow CH_3• + HCl$, then $CH_3• + Cl_2 \rightarrow CH_3Cl + Cl•$ **Termination:** two radicals combine, e.g. $Cl• + Cl• \rightarrow Cl_2$, or $CH_3• + Cl• \rightarrow CH_3Cl$ Curly half-arrows are NOT needed for this mechanism.

6.17	*Understand limitations of radical substitution in synthesis: further substitution and mixture of products.*
	Radical substitution is **not useful in synthesis** because: • **Further substitution** occurs: CH_3Cl can react again to form CH_2Cl_2, $CHCl_3$, CCl_4. • A **mixture of products** is obtained, making it difficult to isolate the desired product. • Different hydrogen atoms can be substituted, giving **positional isomers**.

Topic 6C: Alkenes

6.18	*Know the general formula for alkenes.*
	C_nH_{2n} (where n ≥ 2). E.g. ethene C_2H_4, propene C_3H_6. Cycloalkenes have the general formula **C_nH_{2n-2}**.

6.19	*Know that alkenes and cycloalkenes are unsaturated hydrocarbons.*
	Unsaturated means the molecule contains at least one **C=C double bond**. This makes alkenes more reactive than alkanes, as the **π-bond** is a region of high electron density that can be attacked by electrophiles.

6.20	*Understand the bonding in alkenes in terms of σ- and π-bonds.*
	The C=C double bond consists of: • A **σ-bond:** formed by the **head-on overlap** of two sp² hybrid orbitals. Strong, allows free rotation. • A **π-bond:** formed by the **sideways overlap** of two p-orbitals above and below the plane of the molecule. Weaker than the σ-bond; **prevents rotation** (causing E/Z isomerism). The π-bond has a region of high electron density above and below the bond, making it susceptible to **electrophilic attack**.

6.21	*Know what is meant by the term 'electrophile'.*
	An **electrophile** is an electron-pair **acceptor** – a species attracted to regions of **high electron density** (e.g. the π-bond in alkenes). Electrophiles are either **positively charged** (e.g. H^+, NO_2^+) or have a **partial positive charge** (e.g. δ+Br in Br_2).

6.22	*Understand the addition reactions of alkenes with: (i) H_2/Ni, (ii) halogens, (iii) hydrogen halides, (iv) steam/acid catalyst, (v) $KMnO_4$ in acid.*
	(i) H_2 + Ni catalyst: alkene → alkane (hydrogenation). Used in margarine manufacture (hardening unsaturated vegetable oils). **(ii) Halogens (e.g. Br_2):** alkene + Br_2 → dihalogenoalkane. Br_2 decolourises (test for C=C). **(iii) Hydrogen halides (e.g. HBr):** alkene + HBr → halogenoalkane. Markovnikov's rule: H adds to the C with more H atoms already (via the more stable carbocation). **(iv) Steam + H_3PO_4 catalyst (300°C, 60 atm):** alkene + H_2O → alcohol. Industrial production of ethanol. **(v) $KMnO_4$ (acidified, cold, dilute):** alkene → diol. Purple $KMnO_4$ decolourises.

6.23	*Understand that heterolytic bond fission results in the formation of ions.*
	Heterolytic fission: a covalent bond breaks so that **both electrons go to one atom**, forming a **cation** and an **anion**. E.g. HBr → H^+ + Br^-. This is the opposite of homolytic fission. It occurs in polar molecules and in ionic bond formation.

6.24	*Understand the mechanism of electrophilic addition between alkenes and: (i) halogens, (ii) hydrogen halides, (iii) other binary compounds.*
	Electrophilic addition mechanism: **(i) With Br_2:** the π-bond in the alkene **repels** the electrons in Br_2, inducing a dipole (δ+Br–Brδ–). A curly arrow from the C=C π-bond to the δ+Br shows bond formation. A second curly arrow

	from the Br–Br bond to the departing Br shows heterolytic fission. This forms a **carbocation intermediate** + Br^-. Then Br^- attacks the carbocation. **(ii) With HBr:** curly arrow from C=C to δ+H. The more stable **secondary or tertiary carbocation** forms preferentially (**Markovnikov's rule**). Br^- then attacks the carbocation. **Carbocation stability: tertiary > secondary > primary** (due to the inductive effect of alkyl groups stabilising the positive charge).

6.25	*Know the qualitative test for a C=C double bond using bromine or bromine water.*
	Add **bromine water** (orange/brown) to the substance. If a C=C double bond is present, the bromine water is **decolourised** (turns from orange to colourless). This is because Br_2 undergoes **electrophilic addition** across the double bond. Alkanes do NOT decolourise bromine water.

6.26	*Know that alkenes form polymers through addition polymerisation. Be able to identify repeat units.*
	In **addition polymerisation**, many alkene monomers join together to form a long-chain polymer with **no by-products** (atom economy = 100%). The C=C double bond in each monomer **opens up** and the monomers link by single C–C bonds. **Repeat unit:** draw the monomer with the double bond opened, and show the continuation bonds extending from each end. E.g. ethene → poly(ethene).

6.27	*Know that waste polymers can be: (i) recycled, (ii) incinerated for energy, (iii) used as feedstock for cracking.*
	(i) Recycling: polymers are sorted by type and reprocessed. Reduces waste and use of raw materials. **(ii) Incineration:** burning polymers releases energy that can generate electricity. Produces CO_2 and may release toxic gases (e.g. HCl from PVC). **(iii) Feedstock recycling/cracking:** polymers are broken down into monomers or other useful small molecules (e.g. for fuel or chemical synthesis).

6.28	*Understand, in terms of energy and resources over the life cycle of polymer products, that chemists can contribute to more sustainable use of materials.*
	Chemists consider the **entire life cycle** of a polymer: raw material extraction, manufacture, use, and disposal. Sustainable approaches include: using renewable feedstocks, reducing energy in production, designing polymers for easier recycling, and extending product lifetimes.

6.29	*Understand how chemists limit polymer disposal problems by: (i) developing biodegradable polymers, (ii) removing toxic waste gases from incineration.*
	(i) Biodegradable polymers (e.g. PLA from corn starch) can be broken down by microorganisms, reducing landfill waste. **(ii) Toxic gases** from incineration (e.g. HCl from PVC, dioxins) can be removed by **scrubbers** (e.g. using $Ca(OH)_2$ to neutralise acidic gases) or **filters** to trap particulates.

Topic 6D: Halogenoalkanes

6.30	*Know that halogenoalkanes can be classified as primary, secondary or tertiary.*
	Classification depends on the carbon bonded to the halogen: **Primary (1°):** the C bonded to the halogen is attached to **one** other carbon (or none, as in CH_3Br). **Secondary (2°):** the C bonded to the halogen is attached to **two** other carbons. **Tertiary (3°):** the C bonded to the halogen is attached to **three** other carbons.

6.31	*Understand what is meant by the term 'nucleophile'.*
	A **nucleophile** is an electron-pair **donor** – a species attracted to regions of **positive charge** (δ+ or +). Nucleophiles have a **lone pair of electrons** which they donate to form a new covalent bond. Examples: OH^-, CN^-, NH_3, H_2O.

6.32	*Understand the reactions of halogenoalkanes with: (i) aqueous KOH → alcohols, (ii) aqueous $AgNO_3$ in ethanol, (iii) KCN → nitriles, (iv) NH_3 → primary amines, (v) ethanolic KOH → alkenes.*
	(i) Aqueous KOH (nucleophilic substitution): $R–X + OH^- \rightarrow R–OH + X^-$. OH^- acts as a nucleophile. Produces an **alcohol**. Reflux with aqueous KOH. **(ii) Aqueous $AgNO_3$ in ethanol:** H_2O acts as the nucleophile. The halide ion released reacts with Ag^+ to form a silver halide precipitate (used to test for the halide and compare rates). **(iii) KCN (nucleophilic substitution):** $R–X + CN^- \rightarrow R–CN + X^-$. **Increases the carbon chain by one C.** Reflux with ethanolic KCN. **(iv) NH_3 (nucleophilic substitution):** $R–X + 2NH_3 \rightarrow R–NH_2 + NH_4X$. Produces a **primary amine**. Heat with excess concentrated ammonia in a sealed tube. **(v) Ethanolic KOH (elimination):** $R–X + KOH \rightarrow$ alkene $+ H_2O + KX$. OH^- acts as a **base** (removes H^+). Reflux with ethanolic (not aqueous) KOH. Produces an **alkene**.

6.33	*Understand that experimental observations can compare rates of hydrolysis of: (i) primary, secondary, tertiary halogenoalkanes, (ii) chloro-, bromo-, iodoalkanes, using aqueous $AgNO_3$ in ethanol.*
	Add aqueous $AgNO_3$ in ethanol to the halogenoalkane and time how long until a precipitate appears. **(i)** Tertiary fastest, primary slowest (tertiary undergoes SN1 – forms a stable carbocation). **(ii)** Iodoalkanes fastest, chloroalkanes slowest. The **C–I bond has the lowest bond enthalpy** (weakest, easiest to break), so reacts fastest. C–Cl is strongest, so slowest. The precipitate formed identifies the halide: AgCl (white), AgBr (cream), AgI (yellow).

⭐ CORE PRACTICAL 4: Investigation of the Rates of Hydrolysis of Some Halogenoalkanes

Method:

1. Set up a series of labelled test tubes in a water bath at 50°C, each containing 1 cm^3 of ethanol and 1 cm^3 of 0.1 mol dm^{-3} aqueous silver nitrate solution.
2. In separate test tubes, place equal volumes (e.g. 2 drops) of each halogenoalkane: 1-chlorobutane, 1-bromobutane, and 1-iodobutane.
3. Add the halogenoalkane to the $AgNO_3$/ethanol mixture simultaneously and start the timer.
4. Observe the time taken for a precipitate to appear in each test tube. Record the colour of each precipitate.
5. Repeat to compare primary, secondary and tertiary halogenoalkanes (e.g. 1-bromobutane, 2-bromobutane, 2-bromo-2-methylpropane).

Sources of Uncertainty:
- Stopwatch: ±0.5 s (human reaction time)
- Thermometer: ±0.5°C
- Judgement of when precipitate first appears (subjective)

Common Errors:
- Halogenoalkanes not fully mixed with the aqueous $AgNO_3$ (they are immiscible – ethanol acts as a co-solvent)
- Temperature not maintained at a constant value during the experiment
- Using different amounts of each halogenoalkane (unfair test)
- Difficulty judging exactly when the precipitate first appears (parallax, cloudiness)

Improvements:
- Use a water bath to maintain a constant temperature
- Ensure equal volumes/concentrations of each halogenoalkane are used
- Use the same volume of $AgNO_3$ solution and ethanol for each test
- Shake or swirl the tubes to ensure mixing
- Repeat the experiment and take a mean time for each halogenoalkane
- Use a white card behind the tubes to see the precipitate more clearly

6.34	*Know the trend in reactivity of primary, secondary and tertiary halogenoalkanes.*
	Tertiary > secondary > primary for hydrolysis rate. Tertiary halogenoalkanes react via an **SN1 mechanism** (carbocation intermediate is stabilised by three alkyl groups). Primary halogenoalkanes react via an **SN2 mechanism** (direct attack by nucleophile).

6.35	*Understand, in terms of bond enthalpy, the trend in reactivity of chloro-, bromo- and iodoalkanes.*
	Reactivity: **iodoalkanes > bromoalkanes > chloroalkanes**. **Bond enthalpy: C–Cl > C–Br > C–I.** The C–I bond is the **weakest** (lowest bond enthalpy), so it is **easiest to break**, making iodoalkanes the most reactive. The C–Cl bond is the strongest, so chloroalkanes are the least reactive.

6.36	*Understand the mechanisms of nucleophilic substitution between primary halogenoalkanes and: (i) aqueous KOH, (ii) ammonia.*
	Both are **SN2 mechanisms** (one step, bimolecular). **(i) With OH^-:** curly arrow from the lone pair on OH^- to the δ+C. Simultaneously, curly arrow from the C–X bond to the halogen. The nucleophile attacks from the **opposite side** to the leaving group (backside attack). Transition state has 5 bonds to C. **(ii) With NH_3:** curly arrow from the lone pair on N to the δ+C. C–X bond breaks. An intermediate forms, then a second NH_3 molecule removes an H^+ to give the amine + NH_4^+. Show the lone pair on N, the δ+ on C, and the δ– on X.

Topic 6E: Alcohols

6.37	*Know that alcohols can be classified as primary, secondary or tertiary.*
	Classification depends on the carbon bonded to the –OH group: **Primary (1°):** the C bearing –OH has **one** (or zero) other C atoms attached. **Secondary (2°):** the C bearing –OH has **two** other C atoms. **Tertiary (3°):** the C bearing –OH has **three** other C atoms. This classification determines the **oxidation products**.

6.38	*Understand the reactions of alcohols with: (i) oxygen, (ii) halogenating agents, (iii) $K_2Cr_2O_7/H_2SO_4$, (iv) conc. H_3PO_4.*
	(i) Combustion: alcohol + O_2 → CO_2 + H_2O. Clean-burning fuels. **(ii) Halogenating agents:** PCl_5 + alcohol → chloroalkane + $POCl_3$ + HCl (steamy white fumes of HCl confirm –OH). KBr + 50% H_2SO_4 + alcohol → bromoalkane. Red phosphorus + I_2 + alcohol → iodoalkane. **(iii) Oxidation with acidified $K_2Cr_2O_7$: Primary alcohol** → aldehyde (distil immediately) → carboxylic acid (reflux). **Secondary alcohol** → ketone (reflux or distil). **Tertiary alcohol** → no reaction (cannot be oxidised). Orange dichromate turns **green** (Cr^{3+}) if oxidation occurs. Test aldehydes with **Fehling's/Benedict's** (blue solution changes to brick-red precipitate) or **Tollens' reagent** (colourless solution forms silver mirror). **(iv) Elimination with conc. H_3PO_4:** alcohol heated with conc. H_3PO_4 → alkene + H_2O. This is a **dehydration/elimination** reaction.

6.39	*Understand techniques used in preparation and purification of a liquid organic compound: (i) reflux, (ii) extraction, (iii) distillation, (iv) drying, (v) boiling point determination.*
	(i) Reflux: heating a reaction mixture in a flask fitted with a **vertical condenser**. Vapour condenses and returns to the flask, allowing prolonged heating without loss of volatile substances. **(ii) Solvent extraction (separating funnel):** shaking the product with an immiscible solvent to extract it from the reaction mixture. Drain the lower layer. **(iii) Distillation:** heating to separate the product by boiling point. Collect the fraction boiling at the correct temperature. **(iv) Drying:** add an **anhydrous drying agent** (e.g. anhydrous $MgSO_4$ or $CaCl_2$) to remove traces of water. Filter off the drying agent. **(v) Boiling point determination:** compare the observed boiling point with the literature value to confirm identity and assess purity.

⭐ CORE PRACTICAL 5: The Oxidation of Ethanol

Method:

1. Set up a pear-shaped flask with anti-bumping granules. Add approximately 10 cm^3 of ethanol.
2. Prepare a solution of acidified potassium dichromate(VI) by mixing $K_2Cr_2O_7$ solution with dilute sulfuric acid.
3. For making an aldehyde (ethanal): set up for immediate distillation (not reflux). Add the acidified dichromate dropwise to warm ethanol and distil the product as it forms. Collect the distillate (ethanal, b.p. 21°C) in an ice-cooled receiver.
4. For making a carboxylic acid (ethanoic acid): set up for reflux. Add excess acidified dichromate to the ethanol. Heat under reflux for 15–20 minutes. Then rearrange for distillation and distil the product.
5. Observe the colour change: orange ($Cr_2O_7^{2-}$) to green (Cr^{3+}).
6. Test the product: aldehyde gives a positive result with Fehling's (brick-red precipitate) or Tollens' (silver mirror). Carboxylic acid turns blue litmus red.

Sources of Uncertainty:

- Thermometer: ±0.5°C
- Volume of ethanol: ±0.5 cm^3
- Judgement of colour change end point

Common Errors:

- Not distilling the aldehyde quickly enough (it is further oxidised to the carboxylic acid in the flask)
- Not using excess dichromate when making the carboxylic acid (incomplete oxidation)
- Loss of volatile aldehyde product if apparatus is not properly connected
- Not heating sufficiently during reflux

Improvements:

- For aldehyde: distil immediately and collect in an ice-cooled receiver to prevent further oxidation and loss by evaporation
- For carboxylic acid: use excess dichromate and reflux for sufficient time
- Ensure all joints are secure to prevent loss of volatile products
- Test the product to confirm its identity

⭐ CORE PRACTICAL 6: Chlorination of 2-methylpropan-2-ol Using Concentrated Hydrochloric Acid

Method:

1. Place approximately 10 cm^3 of 2-methylpropan-2-ol in a separating funnel.
2. Add approximately 20 cm^3 of concentrated hydrochloric acid. Stopper the funnel.
3. Invert and shake the funnel gently, releasing pressure regularly by opening the tap (exothermic, gas produced).
4. Allow the layers to separate. The organic product (2-chloro-2-methylpropane) is the upper layer.
5. Run off and discard the lower aqueous layer.
6. Wash the organic layer with sodium hydrogencarbonate solution (to remove excess acid). Vent the funnel regularly.
7. Wash with water. Run off the aqueous layer.
8. Dry the organic product using anhydrous calcium chloride or $MgSO_4$. Filter off the drying agent.
9. Distil the product, collecting the fraction at the correct boiling point (51°C).

Sources of Uncertainty:

- Measuring cylinder: ±0.5 cm^3
- Thermometer (for distillation): ±0.5°C
- Balance (if weighing product): ±0.01 g

Common Errors:

- Not venting the separating funnel (pressure build-up from HCl gas)
- Incomplete separation of layers (product contaminated with aqueous layer)
- Not washing sufficiently (acid impurities remain in the product)
- Loss of product during multiple washing steps
- Product evaporating during transfer (low boiling point)

Improvements:

- Vent the separating funnel frequently during shaking
- Ensure complete layer separation before draining
- Wash with $NaHCO_3$ to neutralise acid, then with water
- Use anhydrous drying agent and allow sufficient time to dry
- Distil carefully and collect only the correct fraction
- Calculate percentage yield to assess efficiency

Topic 7: Modern Analytical Techniques I

Topic 7A: Mass Spectrometry

7.1	*Be able to use mass spectrometry data to: (i) determine Mr from the molecular ion peak, (ii) suggest structures from fragmentation patterns.*
	(i) Molecular ion peak (M^+): the peak with the **highest m/z value** (ignoring M+1) gives the **relative molecular mass** directly. **(ii) Fragmentation:** the molecular ion breaks into smaller pieces. Common losses: –15 (CH_3), –29 (CHO or C_2H_5), –31 (CH_2OH), –45 (OC_2H_5), –17 (OH), –28 (CO). Look at the difference between M^+ and fragment peaks to identify groups lost. Base peak = most abundant fragment (tallest peak).

Topic 7B: Infrared (IR) Spectroscopy

7.2	*Be able to use IR spectra to deduce functional groups and predict absorptions.*
	Key IR absorptions: • **O–H (alcohol):** broad absorption, ~3230–3550 cm^{-1} • **O–H (carboxylic acid):** very broad absorption, ~2500–3300 cm^{-1} (overlaps C–H) • **C=O (aldehyde/ketone):** strong, sharp, ~1680–1750 cm^{-1} • **C=O (carboxylic acid):** strong, ~1680–1750 cm^{-1} (plus broad O–H) • **C=C (alkene):** ~1620–1680 cm^{-1} • **C–H (alkane):** ~2850–2960 cm^{-1} • **N–H (amine):** ~3300–3500 cm^{-1} Use data provided in the exam. The **fingerprint region** (below ~1500 cm^{-1}) is unique to each compound.

⭐ CORE PRACTICAL 7: Analysis of Some Inorganic and Organic Unknowns

Method:

1. Carry out a series of chemical tests on unknown substances to identify the functional groups present.
2. For organic unknowns, use: bromine water (test for C=C), 2,4-DNPH (test for C=O), Tollens' reagent (distinguish aldehyde from ketone), Fehling's/Benedict's (aldehyde), PCl_5 (test for –OH), Na_2CO_3/$NaHCO_3$ (test for –COOH), iodoform test (CH_3CO– group).
3. For inorganic unknowns, use: flame tests (Group 1/2 cations), NaOH(aq) (metal hydroxide precipitates), $AgNO_3$ (halide ions), $BaCl_2$ (sulfate ions), acid + limewater (carbonate ions).
4. Record all observations carefully: colour changes, precipitate formation, gas evolution.
5. Use IR spectroscopy or mass spectrometry data if provided to support identifications.
6. Combine all test results to identify the unknown substance.

Sources of Uncertainty:

- Visual judgement of colour changes and precipitate formation
- Volume of reagents used (consistency between tests)

Common Errors:

- Not using controls or reference samples for comparison
- Contaminating samples with previous reagents (cross-contamination between tests)
- Misidentifying precipitate colours (e.g. cream AgBr vs white AgCl)
- Not carrying out tests in a logical order

Improvements:

- Use clean test tubes and droppers for each test
- Carry out tests systematically, recording all observations immediately
- Use small, consistent quantities of reagents

- Compare results with known standards where possible
- Use a white tile or background to observe colours accurately

Topic 9: Kinetics I

9.1	*Understand, in terms of collision theory, the effect of changes in concentration, temperature, pressure and surface area on rate.*
	For a reaction to occur, particles must **collide** with sufficient energy (≥ Ea) and the correct **orientation**. **Increasing concentration/pressure (gases):** more particles per unit volume → more frequent collisions → faster rate. **Increasing temperature:** particles move faster → more frequent collisions AND a **greater proportion** of collisions exceed Ea → much faster rate. **Increasing surface area:** more exposed particles available for collision → faster rate.

9.2	*Understand that reactions only take place when collisions have sufficient energy, known as activation energy.*
	Activation energy (Ea) is the **minimum energy** required for a collision to result in a reaction. Only collisions with energy ≥ Ea lead to product formation. This is why not every collision is successful.

9.3	*Be able to calculate rate of reaction from: (i) time data, (ii) gradient of a graph.*
	(i) Rate from time: rate = 1/time (or change in quantity / time). E.g. if a reaction takes 50 s, rate ≈ 1/50 = 0.02 s^{-1}. **(ii) Rate from a graph:** draw a **tangent** to the curve at the required point and calculate its **gradient**. For initial rate: draw tangent at t = 0. Units depend on the axes (e.g. mol dm^{-3} s^{-1}).

9.4	*Understand qualitatively, using the Maxwell-Boltzmann distribution, how temperature affects rate.*
	The **Maxwell-Boltzmann distribution** shows the range of energies of molecules at a given temperature. At higher temperature: the **peak shifts right and is lower** (same total area, but flatter and broader). The **area under the curve to the right of Ea increases significantly** – meaning a much greater proportion of molecules have energy ≥ Ea. This is the main reason rate increases with temperature.

9.5	*Understand the role of catalysts in providing alternative reaction routes of lower activation energy.*
	A **catalyst** increases the rate of reaction by providing an **alternative reaction pathway** with a **lower activation energy**. This means a **greater proportion** of collisions have energy ≥ Ea (at the same temperature), so more successful collisions occur per unit time. The catalyst is **not consumed** in the overall reaction.

9.6	*Be able to draw reaction profiles for uncatalysed and catalysed reactions.*
	A **reaction profile** shows enthalpy on the y-axis and reaction progress on the x-axis. It shows: reactants, products, and the **activation energy (Ea)** as a hump between them. The **catalysed profile** has a **lower Ea** (smaller hump) but the same ΔH (same start and end points). The overall enthalpy change is **not affected** by the catalyst.

9.7	*Be able to interpret catalyst action using the Maxwell-Boltzmann distribution.*
	On a Maxwell-Boltzmann diagram, a catalyst is shown by **moving the Ea line to the left** (lower Ea), while the distribution curve stays the same. The area under the curve to the right of the new (lower) Ea is **larger**, showing that **more molecules** now have sufficient energy to react.

9.8	*Understand the use of a solid (heterogeneous) catalyst for industrial gas-phase reactions.*
	A **heterogeneous catalyst** is in a different phase from the reactants (usually a solid catalyst with gaseous reactants). It provides a **surface** for the reaction: reactant molecules **adsorb**, bonds weaken, reaction occurs, and products **desorb**. E.g. iron in the Haber process, V_2O_5 in the Contact process.

9.9	*Understand the economic benefits of catalysts in industrial reactions.*
	Catalysts provide **economic benefits** by: • Allowing reactions at **lower temperatures and pressures** (reducing energy costs) • Increasing the **rate** of reaction (more product per unit time) • Being **reusable** (not consumed) • Reducing the need for extreme conditions (improving safety) • Reducing waste and improving atom economy in some cases

Topic 16: Kinetics II

16.1	*Understand the terms: rate of reaction, rate equation, order, overall order, rate constant, half-life, rate-determining step, activation energy, catalysts.*
	Rate of reaction: change in concentration per unit time (mol dm^{-3} s^{-1}). **Rate equation:** rate = $k[A]^m[B]^n$ – determined **experimentally only** (cannot be deduced from the balanced equation). **Order** with respect to A: the power m to which [A] is raised. **Overall order** = m + n. **Rate constant (k):** proportionality constant; depends on **temperature** (increases with T). **Half-life ($t_{½}$):** time for the concentration of a reactant to halve. For **first-order** reactions, $t_{½}$ is **constant**. **Rate-determining step (rds):** the **slowest step** in a multi-step mechanism. **Activation energy (Ea):** minimum energy for a successful collision. **Heterogeneous catalyst:** different phase. **Homogeneous catalyst:** same phase.

16.2	*Be able to determine and use rate equations of the form rate = $k[A]^m[B]^n$, where m and n are 0, 1 or 2.*
	Determining order from initial rate data: if doubling [A] has no effect on rate → **zero order** in A. If doubling [A] doubles the rate → **first order**. If doubling [A] quadruples the rate → **second order**. **Using the rate equation:** substitute known values of rate and concentrations to calculate k. **Units of k** depend on the overall order: zero = mol dm^{-3} s^{-1}, first = s^{-1}, second = dm^3 mol^{-1} s^{-1}.

16.3	*Be able to select and justify suitable experimental techniques to obtain rate data: (i) titration, (ii) colorimetry, (iii) mass change, (iv) volume of gas, (v) others.*
	(i) Titration: remove samples (aliquots) at timed intervals, quench the reaction (e.g. add to cold water or $NaHCO_3$), and titrate to find the concentration of a reactant/product. **(ii) Colorimetry:** measure absorbance/transmission of light if a coloured species changes concentration (e.g. I_2 is brown, MnO_4^- is purple). **(iii) Mass change:** place the reaction on a balance; suitable if a gas is produced (mass decreases over time). **(iv) Gas volume:** collect gas in a syringe or over water and record volume at timed intervals. **(v) Others:** conductivity (if ionic concentration changes), pH measurement (if H^+ produced/consumed), clock reactions (time to a visible change).

16.4	*Understand experiments to investigate rates by: (i) initial-rate method, (ii) continuous monitoring.*
	(i) Initial-rate method: carry out multiple experiments, each with a **different initial concentration** of one reagent (keeping others constant). Measure the initial rate for each. Compare rates to deduce orders. **(ii) Continuous monitoring:** follow the reaction over time (e.g. measuring concentration at intervals). Plot **concentration vs time** or **volume vs time** graphs. A clock reaction measures the time for a fixed small amount of reaction to occur – 1/time ≈ initial rate.

16.5	*Be able to calculate rate and half-life of a first-order reaction from a concentration-time or volume-time graph.*
	Rate at time t: draw a **tangent** to the concentration-time curve at time t. Gradient of tangent = rate. **Half-life:** read off the time for [A] to halve (e.g. from 1.0 to 0.5 mol dm^{-3}, then from 0.5 to 0.25, etc.). If the half-life is **constant**, the reaction is **first order**.

16.6	*Be able to deduce order (0, 1 or 2) from: (i) concentration-time graph, (ii) rate-concentration graph.*
	(i) Concentration-time graphs: • **Zero order:** straight line (constant rate of decrease). • **First order:** exponential decay (constant half-life). • **Second order:** steep initial decrease that levels off more gradually. **(ii) Rate-concentration graphs:** • **Zero order:** horizontal line (rate is independent of concentration). • **First order:** straight line through the origin (rate ∝ [A]). • **Second order:** upward curve through the origin (rate ∝ $[A]^2$).

16.7	*Be able to deduce order from initial-rate data.*
	Compare experiments where only one concentration changes: **[A] doubles, rate unchanged → zero order** in A. **[A] doubles, rate doubles → first order** in A. **[A] doubles, rate quadruples → second order** in A. **[A] triples, rate ×9 → second order** in A. Repeat for each reactant. Then write the full rate equation.

16.8	*Understand how to: (i) obtain data for the acid-catalysed iodination of propanone, (ii) use data to predict species in the rds, (iii) deduce a mechanism.*
	The reaction: $CH_3COCH_3 + I_2 \rightarrow CH_3COCH_2I + HI$ (catalysed by H^+). Experimentally: rate = $k[CH_3COCH_3][H^+]$. The rate is **zero order** in $[I_2]$ – iodine does NOT appear in the rate equation. This means the rds does **not involve I_2**. The rds involves propanone and H^+ only. The mechanism involves: (1) protonation of propanone by H^+ (slow/rds), (2) enolisation, (3) reaction with I_2 (fast). Since I_2 is involved only in a fast step after the rds, it does not affect the rate.

16.9	*Be able to deduce a rate-determining step from a rate equation and vice versa.*
	The **rate-determining step** contains the species that appear in the rate equation, in the same quantities as their orders. E.g. if rate = k[A][B], the rds involves **one molecule of A and one molecule of B**. If rate = $k[A]^2$, the rds involves **two molecules of A** (or A appears in steps up to and including the rds). Conversely, given a mechanism with a slow step, write the rate equation from the species in the rds.

16.10	*Be able to deduce a reaction mechanism from the rate equation and stoichiometric equation.*
	Method: (1) Write the rate equation from experimental data. (2) Identify which species are in the rds (those in the rate equation). (3) The rds + subsequent fast steps must give the **overall stoichiometric equation**. (4) Propose a mechanism with the rds first, followed by fast step(s) that account for the remaining reactants/products. Check that all species balance.

16.11	*Understand that rate equations for halogenoalkane hydrolysis provide evidence for SN1 or SN2 mechanisms.*
	Primary halogenoalkanes: rate = $k[\text{halogenoalkane}][OH^-]$ → **SN2** (both species in rds, one-step bimolecular). Backside attack by OH^-.

	Tertiary halogenoalkanes: rate = k[halogenoalkane] → **SN1** (only the halogenoalkane in the rds, two-step unimolecular). Step 1: slow formation of carbocation (rds). Step 2: fast attack by nucleophile. SN1 produces a **racemic mixture** (nucleophile can attack from either side of the planar carbocation). SN2 produces **inversion of configuration**.

16.12	*Be able to use graphical methods to find activation energy from experimental data.*
	The **Arrhenius equation:** $k = Ae^{-Ea/RT}$ Taking natural logs: **ln k = ln A – Ea/RT** This is in the form y = c + mx, where: y = ln k, x = 1/T, gradient = **–Ea/R**, intercept = ln A. Plot **ln k vs 1/T** → straight line with **negative gradient**. **Ea = –gradient × R** (where R = 8.314 J K^{-1} mol^{-1}). T must be in **Kelvin**. The Arrhenius equation will be given in the exam if needed.

⭐ CORE PRACTICAL 13a: Rates of Reaction – Iodine-Propanone Reaction (Titrimetric Method)

Method:

1. Mix known volumes of propanone, sulfuric acid (catalyst) and iodine solutions, starting the timer on mixing.
2. At regular timed intervals (e.g. every 2 minutes), pipette a fixed volume (aliquot) of the reaction mixture into a flask containing excess $NaHCO_3$ solution (to quench the reaction by neutralising the acid catalyst).
3. Titrate the quenched sample with standard sodium thiosulfate solution ($Na_2S_2O_3$) using starch indicator to find the concentration of remaining I_2.
4. Plot $[I_2]$ vs time. The graph should be a straight line (zero order in I_2), confirming rate is independent of $[I_2]$.
5. Repeat with different initial concentrations of propanone and H^+ to determine their orders.

Sources of Uncertainty:

- Pipette: ±0.06 cm^3
- Burette: ±0.05 cm^3 per reading
- Stopwatch: ±0.5 s (reaction time for quenching)

Common Errors:

- Not quenching quickly enough (reaction continues during titration, giving inaccurate $[I_2]$)
- Inconsistent timing of aliquot removal
- Not mixing reagents thoroughly at the start
- Using incorrect indicator (must use starch for iodine-thiosulfate titration)

Improvements:

- Use excess $NaHCO_3$ to ensure complete quenching
- Take aliquots at precise, regular intervals
- Use a thermostatted water bath to maintain constant temperature
- Ensure thorough mixing of reagents at the start
- Repeat titrations to obtain concordant results

⭐ CORE PRACTICAL 13b: Investigating a Clock Reaction (Harcourt-Esson / Iodine Clock)

Method:

1. Prepare a series of experiments with different concentrations of one reactant (e.g. H_2O_2 or KIO_3), keeping all others constant. Make up the volume with water to keep total volume the same.
2. Add a fixed small amount of sodium thiosulfate and starch indicator to the reaction mixture.
3. Add the final reagent (e.g. acid), start the timer immediately, and swirl to mix.

4. Record the time for the solution to turn blue/black (indicating all the thiosulfate has been used up and free I_2 is now present to react with starch).
5. Calculate rate ≈ 1/time for each experiment.
6. Plot rate vs concentration to determine the order with respect to the varied reagent.

Sources of Uncertainty:

- Measuring cylinders: ±0.5 cm^3
- Stopwatch: ±0.5 s (human reaction time for observing colour change)
- Temperature control: ±0.5°C

Common Errors:

- Inconsistent swirling/mixing (affects rate of reaction and time to colour change)
- Temperature not controlled (rate depends on temperature)
- Difficulty judging the exact moment of colour change
- Not keeping total volume constant when varying concentration

Improvements:

- Use a water bath to maintain constant temperature
- Ensure thorough and consistent mixing for all experiments
- Use a white tile beneath the flask to observe the colour change more clearly
- Keep total volume constant by adding distilled water to compensate
- Repeat each concentration at least twice and take a mean time

⭐ CORE PRACTICAL 14: Finding the Activation Energy of a Reaction

Method:

1. Choose a reaction with a measurable rate at different temperatures (e.g. reaction of sodium thiosulfate with HCl, or iodine clock at different temperatures).
2. Set up a water bath at the first temperature (e.g. 20°C). Allow the reagents to equilibrate to this temperature.
3. Mix the reagents, start the timer, and record the time for the reaction (e.g. time for the cross to disappear, or time for colour change).
4. Repeat at several different temperatures (e.g. 25, 30, 35, 40, 45°C), keeping all concentrations and volumes the same.
5. For each temperature, calculate rate = 1/time.
6. Calculate ln(rate) and 1/T (T in Kelvin).
7. Plot ln(rate) vs 1/T. This should give a straight line with gradient = –Ea/R.
8. Calculate Ea = –gradient × R (R = 8.314 J K^{-1} mol^{-1}).

Sources of Uncertainty:

- Thermometer: ±0.5°C
- Stopwatch: ±0.5 s
- Measuring cylinders: ±0.5 cm^3
- Temperature may not be perfectly constant during the reaction

Common Errors:

- Temperature changing during the reaction (not equilibrating reagents properly before mixing)
- Inconsistent timing of the end point at different temperatures
- Not converting temperature to Kelvin before calculating 1/T
- Using too few temperatures (need at least 5 data points for a reliable line of best fit)

Improvements:

- Allow reagents to reach the water bath temperature before mixing (at least 5 minutes)
- Use at least 5 different temperatures across a reasonable range (e.g. 20–50°C)
- Use the same volumes and concentrations for each experiment
- Plot the graph carefully and draw a line of best fit
- Calculate the gradient using two widely separated points on the line
- Convert Ea to kJ mol^{-1} by dividing by 1000

Topic 17: Organic Chemistry II

Topic 17A: Chirality

17.1	*Know that optical isomerism results from chirality in molecules with a single chiral centre.*
	Optical isomerism is a type of stereoisomerism that occurs when a molecule has a **chiral centre** – a carbon atom bonded to **four different groups**. This produces two non-superimposable mirror image forms called **enantiomers**.

17.2	*Understand that optical isomers are object and non-superimposable mirror images from asymmetric carbon atoms.*
	An **asymmetric (chiral) carbon** has **four different substituents**. The two enantiomers are mirror images that **cannot be superimposed** on each other (like left and right hands). They have **identical physical properties** (m.p., b.p., solubility) but rotate plane-polarised light in **opposite directions**.

17.3	*Know that optical activity is the ability of a single optical isomer to rotate plane-polarised monochromatic light.*
	A single enantiomer rotates the plane of **plane-polarised monochromatic light** either clockwise (+, dextrorotatory) or anticlockwise (–, laevorotatory) by the same angle but in opposite directions.

17.4	*Understand the nature of a racemic mixture.*
	A **racemic mixture** (racemate) contains **equal amounts** of both enantiomers. It shows **no overall optical activity** because the rotations cancel out. Racemic mixtures are formed in reactions where a planar intermediate (e.g. carbocation in SN1) allows attack from **both sides equally**.

17.5	*Be able to use data on optical activity as evidence for SN1 and SN2 mechanisms.*
	SN1: produces a **racemic mixture** (no net optical activity) because the planar carbocation intermediate is attacked equally from both sides. **SN2:** produces **inversion of configuration** (one specific enantiomer) because the nucleophile attacks from the **opposite side** to the leaving group. If a single enantiomer undergoes SN2, the product is a single enantiomer with inverted configuration.

Topic 17B: Carbonyl Compounds

17.6	*Be able to identify the aldehyde and ketone functional groups.*
	Aldehyde: contains a **–CHO** group (C=O with H on the carbonyl carbon, always at the end of the chain). E.g. ethanal, CH_3CHO. **Ketone:** contains a **C=O** group where the carbonyl carbon is bonded to **two other carbon atoms** (in the middle of the chain). E.g. propanone, CH_3COCH_3.

17.7	*Understand that: (i) aldehydes and ketones cannot form intermolecular H-bonds with each other, (ii) they can H-bond with water.*
	(i) Aldehydes and ketones have a C=O group (δ– on O, δ+ on C) but **no O–H or N–H bond**, so they cannot form hydrogen bonds **with each other**. Their boiling points are lower than corresponding alcohols. **(ii)** The lone pairs on the C=O oxygen **can** form hydrogen bonds **with water**, so small aldehydes/ketones are **soluble in water**.

17.8	*Understand reactions of carbonyl compounds with: (i) Fehling's/Benedict's, Tollens', acidified dichromate, (ii) $LiAlH_4$, (iii) HCN/KCN, (iv) 2,4-DNPH, (v) I_2/NaOH.*
	(i) Distinguishing aldehydes from ketones: Aldehydes are easily oxidised; ketones are not. • **Fehling's/Benedict's:** aldehyde gives a **brick-red precipitate** (Cu_2O). Ketone: no change. • **Tollens' (ammoniacal $AgNO_3$):** aldehyde gives a **silver mirror**. Ketone: no change. • **Acidified $K_2Cr_2O_7$:** aldehyde turns it from **orange to green**. Ketone: no change. **(ii) Reduction ($LiAlH_4$ in dry ether):** aldehyde → primary alcohol; ketone → secondary alcohol. [H] is used in equations. **(iii) Nucleophilic addition with HCN/KCN:** C=O + HCN → hydroxynitrile. **Mechanism:** CN^- (nucleophile) attacks the δ+C of C=O (curly arrow from lone pair on C of CN^- to δ+C). The π-bond breaks, electrons go to O. Then H^+ from HCN protonates the O^-. If the carbonyl is not at a chiral centre, the product is a **racemic mixture** (attack from either side of the planar C=O). **(iv) 2,4-DNPH:** produces an **orange/yellow precipitate** (a 2,4-dinitrophenylhydrazone derivative). This is a **test for the presence of a C=O group** (both aldehydes and ketones). The precipitate can be recrystallised and its **melting point measured** to identify the specific carbonyl compound. **(v) Iodoform reaction (I_2/NaOH):** compounds with a **$CH_3CO–$** group (methyl ketones and ethanal) give a **yellow precipitate of CHI_3** (triiodomethane/iodoform) with a characteristic antiseptic smell.

Topic 17C: Carboxylic Acids, Esters and Acyl Chlorides

17.9	*Be able to identify the carboxylic acid functional group.*
	Carboxylic acid: contains a **–COOH** group (C=O and O–H on the same carbon, always at the end of the chain). E.g. ethanoic acid, CH_3COOH. Named with the suffix **-oic acid**.

17.10	*Understand that hydrogen bonding affects physical properties of carboxylic acids.*
	Carboxylic acids can form **hydrogen bonds** via both the O–H and C=O groups. They often form **dimers** (pairs of molecules linked by two H-bonds). This gives them **higher boiling points** than alcohols or aldehydes of similar Mr. Small carboxylic acids are **soluble in water** (H-bonding with water).

17.11	*Understand that carboxylic acids can be prepared by: oxidation of alcohols/aldehydes, hydrolysis of nitriles.*
	From primary alcohols/aldehydes: reflux with acidified $K_2Cr_2O_7$ (orange → green). The alcohol is oxidised to the aldehyde, then further to the carboxylic acid. **From nitriles:** R–CN + $2H_2O$ + HCl → R–COOH + NH_4Cl (acid hydrolysis). This produces a carboxylic acid with **one more carbon** than the original halogenoalkane used to make the nitrile.

17.12	*Understand reactions of carboxylic acids with: (i) $LiAlH_4$, (ii) bases, (iii) PCl_5, (iv) alcohols/acid catalyst.*
	(i) Reduction ($LiAlH_4$ in dry ether): carboxylic acid → primary alcohol. **(ii) With bases:** RCOOH + NaOH → RCOONa + H_2O (neutralisation to form a salt). **(iii) With PCl_5:** RCOOH + PCl_5 → RCOCl + $POCl_3$ + HCl. Steamy white fumes of HCl (test for –OH or –COOH). Produces an **acyl chloride**. **(iv) With alcohols (acid catalyst, e.g. conc. H_2SO_4):** RCOOH + R′OH ⇌ RCOOR′ + H_2O. This is **esterification** (reversible, reflux). Esters have **fruity smells**.

17.13	*Be able to identify acyl chloride and ester functional groups.*
	Acyl chloride: contains **–COCl** (C=O bonded to Cl). Very reactive. Named with suffix **-oyl chloride** (e.g. ethanoyl chloride, CH_3COCl). **Ester:** contains **–COO–** (C=O bonded to O bonded to another C). Named as alkyl alkanoate (e.g. ethyl ethanoate, $CH_3COOC_2H_5$).

17.14	*Understand reactions of acyl chlorides with: (i) water, (ii) alcohols, (iii) conc. ammonia, (iv) amines.*
	Acyl chlorides are **very reactive** (vigorous reactions at room temperature, no catalyst needed): **(i) + Water:** RCOCl + H_2O → RCOOH + HCl (carboxylic acid + steamy white fumes). **(ii) + Alcohol:** RCOCl + R′OH → RCOOR′ + HCl (ester + HCl fumes). Better for ester synthesis than acid + alcohol (faster, goes to completion). **(iii) + Conc. ammonia:** RCOCl + $2NH_3$ → $RCONH_2$ + NH_4Cl (amide + ammonium salt). **(iv) + Amine:** RCOCl + R′NH_2 → RCONHR′ + HCl (N-substituted amide + HCl).

17.15	*Understand the hydrolysis of esters in acidic and alkaline solution.*
	Acid hydrolysis: RCOOR′ + H_2O ⇌ RCOOH + R′OH (reversible, reflux with dilute H_2SO_4). Produces carboxylic acid + alcohol. **Alkaline hydrolysis (saponification):** RCOOR′ + NaOH → RCOONa + R′OH (**irreversible**, reflux with NaOH). Produces the sodium salt of the carboxylic acid + alcohol. This is the basis of soap-making.

17.16	*Understand how polyesters are formed by condensation polymerisation.*
	Condensation polymerisation: monomers join by **losing a small molecule** (usually H_2O) at each link. **Polyesters** form from a **dicarboxylic acid** + a **diol** (or a hydroxy acid). An ester bond (–COO–) forms at each link, with loss of H_2O. E.g. Terylene (PET): benzene-1,4-dicarboxylic acid + ethane-1,2-diol. Unlike addition polymers, condensation polymers are **biodegradable** (ester bonds can be hydrolysed).

Topic 18: Organic Chemistry III

Topic 18A: Arenes – Benzene

18.1	*Understand that benzene bonding has been represented using Kekulé and delocalised models.*
	Kekulé model: benzene as cyclohexatriene with alternating C=C and C–C bonds. This model is **incorrect** because it fails to explain why benzene does not react like alkenes. **Delocalised model:** each carbon is sp^2 hybridised with one p-orbital perpendicular to the ring. The six p-orbitals overlap **sideways** to form a **delocalised π-system** above and below the plane of the ring. The electrons are shared equally around the ring, giving benzene extra stability.

18.2	*Understand that evidence for delocalisation comes from enthalpy of hydrogenation data and bond length data.*
	Enthalpy of hydrogenation: theoretical cyclohexatriene (Kekulé) would have ΔH = 3 × (ΔH for cyclohexene). But benzene's actual ΔH of hydrogenation is **152 kJ mol^{-1} less exothermic** than predicted. This difference is the **delocalisation energy** (benzene is more stable than expected). **Bond lengths:** all C–C bonds in benzene are **equal length** (140 pm) – intermediate between C–C (154 pm) and C=C (134 pm). The Kekulé model would predict alternating short and long bonds.

18.3	*Understand why benzene is resistant to bromination compared to alkenes.*
	In alkenes, the π-bond is **localised** between two carbon atoms, creating a region of high electron density that attracts electrophiles → electrophilic **addition** occurs readily. In benzene, the π-electrons are **delocalised** over all six carbons, so the electron density is **spread out** and less available to electrophiles. Also, addition would destroy the stable delocalised ring, which is thermodynamically unfavourable. Therefore benzene undergoes electrophilic **substitution** (preserving the ring) rather than addition, and requires a **catalyst**.

18.4	*Understand reactions of benzene with: (i) O_2, (ii) Br_2/catalyst, (iii) conc. HNO_3/H_2SO_4, (iv) halogenoalkanes and acyl chlorides with $AlCl_3$.*
	(i) Combustion: burns with a **smoky/sooty flame** (high C:H ratio). **(ii) Bromination:** $C_6H_6 + Br_2 \rightarrow C_6H_5Br + HBr$ (catalyst: $AlBr_3$ or $FeBr_3$). Electrophilic substitution. **(iii) Nitration:** $C_6H_6 + HNO_3 \rightarrow C_6H_5NO_2 + H_2O$ (catalyst: conc. H_2SO_4, temperature ~50°C). Electrophile = NO_2^+ (nitronium ion). **(iv) Friedel-Crafts alkylation:** $C_6H_6 + RCl \rightarrow C_6H_5R + HCl$ (catalyst: $AlCl_3$). Friedel-Crafts acylation: $C_6H_6 + RCOCl \rightarrow C_6H_5COR + HCl$ (catalyst: $AlCl_3$).

18.5	*Understand the mechanism of electrophilic substitution of benzene (halogenation, nitration, Friedel-Crafts).*
	General mechanism: (1) Generate the electrophile. (2) Electrophile attacks the delocalised π-system (curly arrow from ring to E^+). (3) Intermediate forms (H still attached). (4) H^+ is lost to restore the stable ring (curly arrow from C–H bond back into the ring). **Electrophile generation:** • Halogenation: $Br_2 + AlBr_3 \rightarrow Br^+ + AlBr_4^-$ • Nitration: $HNO_3 + H_2SO_4 \rightarrow NO_2^+ + HSO_4^- + H_2O$ • Friedel-Crafts: $RCl + AlCl_3 \rightarrow R^+ + AlCl_4^-$ (or RCO^+ for acylation)

18.6	*Understand the reaction of phenol with bromine water.*
	Phenol reacts with **bromine water** (no catalyst needed) to form **2,4,6-tribromophenol** (a white precipitate) + 3HBr. The bromine water is **decolourised**. This is much more readily than benzene, which requires a catalyst.

18.7	*Understand reasons for the relative ease of bromination of phenol compared to benzene.*
	The **–OH group** on phenol donates a lone pair of electrons from the oxygen **into the benzene ring** (via the p-orbital overlap), **increasing the electron density** of the ring. This makes the ring more attractive to electrophiles and lowers the activation energy for attack. Therefore phenol undergoes bromination much more easily than benzene – no catalyst is required.

Topic 18B: Amines, Amides, Amino Acids and Proteins

18.8	*Be able to identify: (i) amine and amide functional groups, (ii) amino acids.*
	Amine: contains **$-NH_2$** (primary), **–NHR** (secondary), or **$-NR_2$** (tertiary). Named with suffix **-amine** or prefix **amino-**. **Amide:** contains **$-CONH_2$** (or –CONHR, $-CONR_2$). A C=O bonded to N. **Amino acid:** contains both an **$-NH_2$ group** and a **–COOH group** on the same molecule. The α-amino acids have both groups on the carbon adjacent to the COOH (α-carbon).

18.9	*Understand reactions of primary aliphatic amines (e.g. butylamine) with: (i) water, (ii) acids, (iii) ethanoyl chloride, (iv) halogenoalkanes, (v) Cu^{2+} ions.*
	(i) + Water: $RNH_2 + H_2O \rightleftharpoons RNH_3^+ + OH^-$ (weakly alkaline; the lone pair on N accepts H^+ from water). **(ii) + Acids:** $RNH_2 + HCl \rightarrow RNH_3^+Cl^-$ (salt formation; amine acts as a base). **(iii) + Ethanoyl chloride:** $RNH_2 + CH_3COCl \rightarrow CH_3CONHR + HCl$ (N-substituted amide). **(iv) + Halogenoalkanes:** $RNH_2 + R'X \rightarrow R'NHR + HX$ (secondary amine; further substitution can give tertiary amines and quaternary ammonium salts). **(v) + Cu^{2+} ions:** the amine acts as a **ligand** (donates lone pair on N to Cu^{2+}), forming a **deep blue complex ion**. E.g. $[Cu(NH_2CH_2CH_2CH_2CH_3)_4]^{2+}$.

18.10	*Understand reasons for the difference in basicity of ammonia, primary aliphatic and primary aromatic amines.*
	Basicity order: **aliphatic amine > ammonia > aromatic amine**. **Aliphatic amines** are stronger bases than ammonia because the **alkyl group is electron-donating** (positive inductive effect), which increases the electron density on the N lone pair, making it a better proton acceptor. **Aromatic amines** (e.g. phenylamine) are weaker bases because the N lone pair is **delocalised into the benzene ring**, making it less available to accept a proton.

18.11	*Understand the preparation of primary aliphatic amines: (i) from halogenoalkanes, (ii) by reduction of nitriles.*
	(i) From halogenoalkanes: $RX + 2NH_3 \rightarrow RNH_2 + NH_4X$. Heat with **excess concentrated ammonia** in a sealed tube. Excess NH_3 is needed to minimise further substitution. **(ii) Reduction of nitriles:** $RCN + 4[H] \rightarrow RCH_2NH_2$. Reduce using **$LiAlH_4$ in dry ether** or **H_2/Ni catalyst**. This gives a primary amine with **one more carbon** than the original nitrile.

18.12	*Know that aromatic nitro-compounds can be reduced to form amines using tin and concentrated HCl.*
	$ArNO_2 + 6[H] \rightarrow ArNH_2 + 2H_2O$ **Reagents:** tin (Sn) and **concentrated hydrochloric acid** (HCl). The mixture is **heated under reflux**. The reducing agent is Sn/HCl. This produces the aromatic amine (e.g. nitrobenzene → phenylamine).

18.13	*Understand that amides can be prepared from acyl chlorides.*
	$RCOCl + NH_3 \rightarrow RCONH_2 + HCl$ (primary amide). $RCOCl + RNH_2 \rightarrow RCONHR + HCl$ (secondary/N-substituted amide). Acyl chlorides react vigorously with ammonia and amines at **room temperature**.

18.14	*Know that polyamide formation is a condensation polymerisation reaction.*
	A **polyamide** (e.g. nylon) is formed by condensation polymerisation between a **dicarboxylic acid** (or diacyl chloride) and a **diamine**. An **amide bond (–CONH–)** forms at each link, with loss of **H_2O** (or HCl if using acyl chloride). Proteins are natural polyamides.

18.15	*Be able to draw repeat units of condensation polymers formed by: (i) dicarboxylic acids + diols, (ii) dicarboxylic acids + diamines, (iii) amino acids.*
	(i) Polyester: show the ester link **–COO–** in the repeat unit, with continuation bonds. Loss of H_2O at each link. **(ii) Polyamide:** show the amide link **–CONH–** in the repeat unit. Loss of H_2O at each link. **(iii) From amino acids:** each amino acid has $–NH_2$ and –COOH, so it condenses with itself to form a polyamide (polypeptide). The repeat unit contains **one amino acid residue** with a peptide bond. Always show the bonds extending from the repeat unit and indicate loss of nH_2O.

18.16	*Understand properties of 2-amino acids: (i) zwitterion formation, (ii) effect on plane-polarised light.*
	(i) Zwitterions: in neutral solution, the $–NH_2$ group is protonated ($–NH_3^+$) and the –COOH is deprotonated ($–COO^-$), forming a **zwitterion** (internal salt with both + and – charges). This explains why amino acids have **high melting points** (ionic interactions) and are soluble in water. In acid: NH_3^+ and COOH (positive ion). In alkali: NH_2 and COO^- (negative ion). Amino acids are **amphoteric**. **(ii)** All amino acids (except glycine) have a chiral centre (α-carbon has 4 different groups) and are therefore **optically active** – they rotate plane-polarised light.

18.17	*Understand that the peptide bond in proteins: (i) is formed by condensation, (ii) can be hydrolysed and products separated by chromatography.*
	(i) Amino acids join by **condensation** – the –COOH of one reacts with the $–NH_2$ of another, forming a **peptide bond (–CONH–)** and releasing H_2O. Proteins are **polypeptides** (long chains of amino acids). **(ii)** Proteins can be **hydrolysed** (broken down) by heating with 6 mol dm^{-3} HCl under reflux, breaking the peptide bonds to release the individual amino acids. The amino acids can then be separated and identified by **chromatography** (e.g. thin-layer or paper chromatography, comparing Rf values with known standards).

⭐ CORE PRACTICAL 15: Analysis of Some Inorganic and Organic Unknowns

Method:

1. This practical extends CP7 to A2-level functional groups and compounds.
2. Carry out systematic chemical tests on unknown organic substances: 2,4-DNPH (C=O), Tollens' (aldehyde vs ketone), iodoform test ($CH_3CO–$), $FeCl_3$ (phenol → purple), bromine water (phenol or alkene), $NaHCO_3$ (carboxylic acid → fizzing), PCl_5 (–OH or –COOH → steamy HCl fumes).
3. For inorganic unknowns: NaOH(aq) for transition metal hydroxide precipitates, flame tests, halide tests with $AgNO_3$, sulfate test with $BaCl_2$.
4. Use IR spectra, mass spectra and NMR data if provided to confirm the identity.
5. Record all observations and link them to the functional groups identified.

Sources of Uncertainty:

- Visual judgement of colour changes
- Consistency of reagent volumes

Common Errors:

- Cross-contamination between tests
- Misidentifying similar colours (e.g. green $Cr(OH)_3$ vs green $Fe(OH)_2$)

- Not using appropriate controls
- Forgetting to test for multiple functional groups in the same compound

Improvements:

- Use clean apparatus for each test
- Work systematically through a logical testing sequence
- Record observations immediately and compare with data tables
- Use small, consistent quantities

Topic 18C: Organic Synthesis

18.18	*Be able to deduce empirical, molecular and structural formulae from combustion analysis, elemental composition, functional group tests, IR, mass spec and NMR data.*
	Combine analytical techniques: • **Mass spectrum:** Mr from M^+ peak; fragmentation suggests groups present. • **IR spectrum:** identifies functional groups (O–H, C=O, N–H, etc.). • **Combustion analysis/elemental %:** calculate empirical formula, then molecular formula using Mr. • **NMR:** determines the carbon/hydrogen environments and connectivity. • **Chemical tests:** confirms functional groups (e.g. 2,4-DNPH for C=O, bromine water for C=C). Use all available data to build up the structure systematically.

18.19	*Be able to plan reaction schemes, of up to four steps, to form familiar and unfamiliar compounds.*
	You must know the **reagents, conditions, and products** for all reactions in the specification and be able to **chain them together** in multi-step syntheses. **Key synthetic routes to know:** Alkene → alcohol (steam/H_3PO_4) → halogenoalkane (HBr or PCl_5) → nitrile (KCN) → carboxylic acid (acid hydrolysis) or amine ($LiAlH_4$). Alcohol → aldehyde ($K_2Cr_2O_7$, distil) → carboxylic acid ($K_2Cr_2O_7$, reflux). Halogenoalkane → amine (NH_3) or alcohol (aq KOH) or alkene (ethanolic KOH). Always state reagents, conditions, and show intermediate products.

18.20	*Understand methods of increasing carbon chain length using Grignard reagents and their reactions with CO_2 and carbonyl compounds.*
	Grignard reagent: formed by reacting a halogenoalkane with **Mg in dry ether**: RX + Mg → RMgX. **Reactions:** • **With CO_2:** RMgX + CO_2 → $RCOO^-MgX^+$, then acid hydrolysis → **RCOOH** (carboxylic acid with one extra C). • **With aldehydes (RCHO):** RMgX + R′CHO → secondary alcohol (after acid hydrolysis). • **With ketones (RCOR):** RMgX + R′COR″ → tertiary alcohol (after acid hydrolysis). • **With methanal (HCHO):** RMgX + HCHO → primary alcohol. Grignard reagents are destroyed by water, so **dry ether** solvent is essential.

18.21	*Be able to select and justify practical procedures for reactions, including identifying control measures to reduce risk.*
	You must be able to choose appropriate techniques (reflux, distillation, etc.) for a given synthesis, and justify why. Always link **hazards** to specific **risks** and suggest appropriate **precautions**: • Flammable solvents → no naked flames, use water/electric heating baths. • Toxic/corrosive chemicals → fume cupboard, gloves, safety goggles. • Volatile products → work in well-ventilated area or fume cupboard. • Exothermic reactions → add reagents slowly, cool if necessary.

18.22	*Understand techniques for preparation and purification: reflux, washing, extraction, recrystallisation, drying, distillation (including steam), melting and boiling point determination.*
	(i) Reflux: prolonged heating without loss of volatile substances.

	(ii) Washing: removing water-soluble impurities using a separating funnel. **(iii) Solvent extraction:** dissolving the product in an organic solvent, separating layers. **(iv) Recrystallisation:** dissolve impure solid in minimum hot solvent, filter hot to remove insoluble impurities, cool to crystallise pure product, filter and dry. **(v) Drying:** anhydrous drying agent (e.g. $MgSO_4$) for liquids; filter paper/desiccator for solids. **(vi) Distillation/steam distillation:** separate by boiling point; steam distillation for heat-sensitive or immiscible products. **(vii) Melting/boiling point:** compare with literature values to confirm identity and assess purity. Pure substances have sharp melting points.

⭐ CORE PRACTICAL 16: The Preparation of Aspirin

Method:

1. Weigh approximately 2 g of 2-hydroxybenzoic acid (salicylic acid) into a pear-shaped or round-bottomed flask. Record the exact mass.
2. Add approximately 4 cm^3 of ethanoic anhydride (the acylating agent) and 5 drops of concentrated phosphoric acid (catalyst).
3. Heat the mixture under reflux in a water bath at 80–90°C for approximately 15 minutes. Swirl occasionally.
4. Allow to cool. Carefully add approximately 20 cm^3 of cold distilled water to decompose excess ethanoic anhydride.
5. Cool the mixture further in an ice bath until crystallisation is complete. Stir/scratch the inside of the flask with a glass rod to encourage crystallisation.
6. Filter the crystals under reduced pressure using a Buchner funnel. Wash with a small amount of cold distilled water.
7. Recrystallise the crude product: dissolve in minimum hot ethanol, add warm water, cool slowly, then ice-cool. Filter again.
8. Dry the purified crystals between sheets of filter paper or in a desiccator.
9. Weigh the dry product and calculate percentage yield.
10. Measure the melting point and compare with the literature value (135°C) to assess purity.

Sources of Uncertainty:

- Balance: ±0.001 g
- Thermometer: ±0.5°C
- Melting point apparatus: ±1°C

Common Errors:

- Not heating for long enough under reflux (incomplete reaction, low yield)
- Adding water too quickly to hot reaction mixture (can cause splattering)
- Not cooling sufficiently before filtering (product remains dissolved, lower yield)
- Impure product due to incomplete recrystallisation (melting point below 135°C or broad range)
- Losses during transfer, filtration and recrystallisation

Improvements:

- Heat under reflux for the full 15 minutes to ensure complete reaction
- Add cold water slowly and carefully to the cooled mixture
- Cool thoroughly in an ice bath and scratch the flask to maximise crystallisation
- Recrystallise carefully using minimum hot solvent
- Dry completely before weighing and measuring melting point
- A sharp melting point close to 135°C indicates high purity
- Calculate percentage yield: actual mass / theoretical mass × 100%

Topic 19: Modern Analytical Techniques II

Topic 19A: Mass Spectrometry

19.1	*Be able to use high-resolution mass spec data to: (i) suggest structures from accurate Mr (4 d.p.), (ii) calculate accurate Mr and identify compounds.*
	High-resolution mass spectrometry measures Mr to **4 decimal places**, allowing you to distinguish between compounds with the same integer Mr. E.g. CO (Mr = 27.9949) vs C_2H_4 (Mr = 28.0313) – same nominal mass (28) but different accurate masses. **Method:** calculate the accurate Mr by summing the accurate relative atomic masses of each atom (^{12}C = 12.0000, ^{1}H = 1.0078, ^{16}O = 15.9949, ^{14}N = 14.0031). Compare with the measured value.

Topic 19B: Nuclear Magnetic Resonance (NMR)

19.2	*Understand that ^{13}C NMR provides information about positions of ^{13}C atoms in a molecule.*
	^{13}C NMR detects the different **chemical environments** of carbon atoms. Each peak represents a carbon atom (or group of equivalent carbons) in a **different environment**. The chemical shift (δ) indicates the type of environment (e.g. C in C=O, C bonded to O, C in alkyl chain).

19.3	*Be able to use ^{13}C NMR data to: (i) predict carbon environments from δ values, (ii) justify the number of peaks.*
	(i) Use a **^{13}C chemical shift table** (provided in exams): e.g. δ ~ 10–40 = C in alkyl groups; δ ~ 50–80 = C bonded to O or N; δ ~ 110–150 = aromatic C or C=C; δ ~ 160–220 = C=O. **(ii)** The number of peaks = number of **different carbon environments**. Equivalent carbons (related by symmetry) give **one peak**. E.g. propanone (CH_3COCH_3) has **2 peaks** (the two CH_3 carbons are equivalent; the C=O is different).

19.4	*Understand that high-resolution ^{1}H NMR provides information about positions of ^{1}H atoms.*
	^{1}H (proton) NMR detects the different **chemical environments** of hydrogen atoms. It provides: the **number of different H environments** (number of peaks), the **ratio of H atoms** in each environment (peak areas/integration), and **splitting patterns** (from adjacent non-equivalent H atoms).

19.5	*Be able to use ^{1}H NMR data to: (i) predict H environments from δ, (ii) use peak areas for ratios, (iii) use the (n+1) rule for splitting, (iv) predict shifts and splitting.*
	(i) Chemical shift (δ): use a table (provided): e.g. δ ~ 0.9 = R–CH_3; δ ~ 3.3–3.9 = C–H next to O or halogen; δ ~ 9.0–10.0 = CHO; δ ~ 11–12 = COOH. **(ii) Peak areas/integration:** the ratio of the areas under each peak = the ratio of H atoms in each environment. E.g. a 3:2:1 ratio in ethanol (CH_3CH_2OH). **(iii) (n+1) rule:** a peak is split into **(n+1) lines** where n = the number of **adjacent non-equivalent H atoms**. Singlet (s) = 0 neighbours; doublet (d) = 1; triplet (t) = 2; quartet (q) = 3. **(iv) Prediction:** given a structure, identify each set of equivalent H atoms, assign the δ value from the table, count adjacent non-equivalent H's, and predict the splitting pattern. **D_2O shake:** adding D_2O removes the –OH or –NH peak (exchanges H for D, which is NMR-inactive). If a peak disappears with D_2O, it's an –OH or –NH.

Topic 19C: Chromatography

19.6	*Know that chromatography separates components between a mobile phase and a stationary phase.*
	Chromatography separates mixtures based on the **differential distribution** of components between a **stationary phase** (does not move; e.g. silica on a plate, liquid on paper) and a **mobile phase** (moves; e.g. solvent, gas). Components with **greater affinity** for the stationary phase move **more slowly**.

19.7	*Be able to calculate Rf values from one-way chromatograms.*
	Rf = distance moved by substance / distance moved by solvent front Rf values are always between 0 and 1. They are characteristic of a substance under specific conditions (same solvent, temperature, stationary phase). Comparing Rf values with known standards allows **identification** of unknown substances.

19.8	*Know that HPLC and GC: (i) are types of column chromatography, (ii) separate by different retention times, (iii) may be used with mass spectrometry (GC-MS) in forensics and drug testing.*
	HPLC (High Performance Liquid Chromatography): liquid mobile phase is forced through a packed column at high pressure. Separates non-volatile and thermally unstable compounds. **GC (Gas Chromatography):** gaseous mobile phase carries the sample through a heated column. Separates volatile compounds. Both produce chromatograms with **peaks at different retention times**. The area under each peak indicates the **amount** of each component. **GC-MS / HPLC-MS:** the output from the chromatography column feeds into a mass spectrometer, giving both separation AND identification. Used in **forensics** (crime scene analysis), **drug testing in sport**, and pharmaceutical quality control.

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