

SYLLABUS OVERVIEW

The syllabus for the Diploma Programme chemistry course is divided into three parts: the core, the additional higher level (AHL) material and the options. A syllabus overview is provided below.

Core [80h]

Topics		Teaching hours
1	Stoichiometry	11
2	Atomic theory	4
3	Periodicity	6
4	Bonding	12
5	States of matter	5
6	Energetics	11
7	Kinetics	4
8	Equilibrium	5
9	Acids and bases	5
10	Oxidation and reduction	7
11	Organic chemistry	10

Additional Higher Level [55h]

Topics		Teaching hours
12	Atomic theory	4
13	Periodicity	4
14	Bonding	6
15	Energetics	4
16	Kinetics	6
17	Equilibrium	4
18	Acids and bases	11
19	Oxidation and reduction	7
20	Organic chemistry	9

Options

Options Standard Level

A Higher physical organic chemistry

Teaching
hours

15

Options Standard Level/Higher Level

B Medicines and drugs

15/22

C Human biochemistry

15/22

D Environmental chemistry

15/22

E Chemical industries

15/22

F Fuels and energy

15/22

Options Higher Level

G Modern analytical chemistry

22

H Further organic chemistry

22

Standard level candidates are required to study any **two** options from A–F.
The duration of each option is 15 hours.

Higher level candidates are required to study any **two** options from B–H.
The duration of each option is 22 hours.

SYLLABUS OUTLINE

Core [80h]		Teaching hours
Topic 1	Stoichiometry	[11]
	1.1 Mole concept and Avogadro's constant	2
	1.2 Formulas	3
	1.3 Chemical equations	1
	1.4 Mass and gaseous volume relationships in chemical reactions	3
	1.5 Solutions	2
Topic 2	Atomic theory	[4]
	2.1 The atom	2
	2.2 Electron arrangement	2
Topic 3	Periodicity	[6]
	3.1 The periodic table	1
	3.2 Physical properties	2
	3.3 Chemical properties	3
Topic 4	Bonding	[12]
	4.1 Ionic bond	2
	4.2 Covalent bond	6
	4.3 Intermolecular forces	1.5
	4.4 Metallic bond	0.5
	4.5 Physical properties	2
Topic 5	States of matter	[5]
	5.1 States of matter	5
Topic 6	Energetics	[11]
	6.1 Exothermic and endothermic reactions	2
	6.2 Calculation of enthalpy changes	3
	6.3 Hess's law	2
	6.4 Bond enthalpies	1
	6.5 Entropy	1
	6.6 Spontaneity	2
Topic 7	Kinetics	[4]
	7.1 Rates of reaction	2
	7.2 Collision theory	2

Core		Teaching hours
Topic 8	Equilibrium	[5]
8.1	Dynamic equilibrium	1
8.2	The position of equilibrium	4
Topic 9	Acids and bases	[5]
9.1	Properties of acids and bases	1
9.2	Strong and weak acids and bases	1
9.3	The pH scale	1
9.4	Buffer solutions	1
9.5	Acid–base titrations	1
Topic 10	Oxidation and reduction	[7]
10.1	Oxidation and reduction	2
10.2	Reactivity	2.5
10.3	Electrolysis	2.5
Topic 11	Organic chemistry	[10]
11.1	Homologous series	1
11.2	Hydrocarbons	2
11.3	Other functional groups	7
 Additional Higher Level [55h]		
Topic 12	Atomic theory	[4]
12.1	The mass spectrometer	1
12.2	Electron configuration of atoms	3
Topic 13	Periodicity	[4]
13.1	Periodic trends Na → Ar (the third period)	2
13.2	d-block elements (first row)	2
Topic 14	Bonding	[6]
14.1	Shapes of molecules and ions	1
14.2	Hybridization	2
14.3	Delocalization of electrons	2
14.4	Structures of allotropes of carbon	1
Topic 15	Energetics	[4]
15.1	Standard enthalpy changes of reaction	1
15.2	Lattice enthalpy	2
15.3	Spontaneity of a reaction	1

Additional Higher Level		Teaching hours
Topic 16	Kinetics	[6]
16.1	Rate expression	3
16.2	Reaction mechanism	1
16.3	Activation energy	2
Topic 17	Equilibrium	[4]
17.1	Phase equilibrium	2
17.2	The equilibrium law	2
Topic 18	Acids and bases	[11]
18.1	Brønsted–Lowry acids and bases	2
18.2	Lewis theory	1
18.3	Calculations involving acids and bases	5
18.4	Salt hydrolysis	1
18.5	Acid–base titrations	1
18.6	Indicators	1
Topic 19	Oxidation and reduction	[7]
19.1	Redox equations	2
19.2	Standard electrode potentials	3
19.3	Electrolysis	2
Topic 20	Organic chemistry	[9]
20.1	Determination of structure	4
20.2	Hydrocarbons	2
20.3	Nucleophilic substitution reactions	2
20.4	Alcohols	1

Option Standard Level	Teaching hours
Option A Higher physical organic chemistry	[15]
A.1 Determination of structure	5
A.2 Rate expression	3
A.3 Reaction mechanism	1
A.4 Nucleophilic substitution reactions	2
A.5 Acids, bases and buffers	4

Options Standard Level/Higher Level

Standard level students study the core of these options and higher level students study the whole option (ie the core and the extension material).

Option B Medicines and drugs	Teaching Hours
Core (SL + HL)	[15]
B.1 Pharmaceutical products	2
B.2 Antacids	1
B.3 Analgesics	3
B.4 Depressants	3
B.5 Stimulants	2.5
B.6 Antibacterials	2
B.7 Antivirals	1.5
Extension (HL only)	[7]
B.8 Stereochemistry in drug action and design	3
B.9 Anesthetics	2
B.10 Mind-altering drugs	2

Option C Human biochemistry	
Core (SL + HL)	[15]
C.1 Diet	2
C.2 Proteins	3
C.3 Carbohydrates	2.5
C.4 Fats	2.5
C.5 Vitamins	2.5
C.6 Hormones	2.5
Extension (HL only)	[7]
C.7 Enzymes	3
C.8 Nucleic acids	2
C.9 Metal ions in biological systems	2

Option D Environmental chemistry	Teaching Hours
Core (SL + HL)	[15]
D.1 Primary air pollution	3
D.2 Ozone depletion	2
D.3 Greenhouse effect and global warming	2
D.4 Acid rain	1.5
D.5 Water suitable for drinking	3
D.6 Dissolved oxygen in water	2
D.7 Waste water treatment	1.5
Extension (HL only)	[7]
D.8 Smog	2
D.9 Ozone depletion	2
D.10 Toxic substances in water	3
 Option E Chemical industries	
Core (SL + HL)	[15]
E.1 Initial overview	2
E.2 Principles of extraction and production of metals	2
E.3 Iron and aluminium	4
E.4 The oil industry	4
E.5 Polymers	3
Extension (HL only)	[7]
E.6 Silicon	1.5
E.7 Ellingham diagrams	1.5
E.8 Mechanisms in the organic chemicals industry	2
E.9 The chlor-alkali industry	2
 Option F Fuels and energy	
Core (SL + HL)	[15]
F.1 Energy sources	1
F.2 Fossil fuels	4
F.3 Nuclear energy	4
F.4 Solar energy	3
F.5 Electrochemical energy	3
Extension (HL only)	[7]
F.6 Storage of energy and limits of efficiency	1
F.7 Nuclear stability	2
F.8 Radioactive decay	2
F.9 Photovoltaics	2

Options Higher Level

Teaching
hours

Option G	Modern analytical chemistry	[22]
G.1	Analytical techniques	2
G.2	Principles of spectroscopy	2
G.3	Visible and ultraviolet spectroscopy	4
G.4	Infrared spectroscopy	3
G.5	Nuclear magnetic resonance (NMR) spectroscopy	4
G.6	Mass spectrometry	3
G.7	Chromatography	4
Option H	Further organic chemistry	[22]
H.1	Stereoisomerism	3
H.2	Free radical substitution reactions	3
H.3	Electrophilic addition reactions	4
H.4	Electrophilic substitution reactions	4
H.5	Nucleophilic addition reactions	1
H.6	Nucleophilic substitution reactions	2
H.7	Elimination reactions	2
H.8	Addition–elimination reactions	1
H.9	Acid–base reactions	2

SYLLABUS DETAILS

Topic 1: Stoichiometry

A.S.		Obj
	1.1 Mole Concept and Avogadro's Constant (2h)	
1.1.1	Describe the mole concept and apply it to substances. The mole concept applies to all kinds of particles: atoms, molecules, ions, formula units etc. The amount of substance is measured in units of moles. The approximate value of Avogadro's constant (L), $6.02 \times 10^{23} \text{ mol}^{-1}$, should be known.	2
1.1.2	Calculate the number of particles and the amount of substance (in moles). Convert between the amount of substance (in moles) and the number of atoms, molecules or formula units.	2
	1.2 Formulas (3h)	
1.2.1	Define the term <i>molar mass</i> (M) and calculate the mass of one mole of a species.	1, 2
1.2.2	Distinguish between <i>atomic mass</i> , <i>molecular mass</i> and <i>formula mass</i> . The term <i>molar mass</i> (in g mol^{-1}) can be used for all of these.	2
1.2.3	Define the terms <i>relative molecular mass</i> (M_r) and <i>relative atomic mass</i> (A_r). The terms have no units.	1
1.2.4	State the relationship between the amount of substance (in moles) and mass, and carry out calculations involving amount of substance, mass and molar mass.	1, 2
1.2.5	Define the terms <i>empirical formula</i> and <i>molecular formula</i> . The molecular formula is a multiple of the empirical formula.	1

A.S.		Obj
1.2.6	Determine the empirical formula and/or the molecular formula of a given compound. Determine the: <ul style="list-style-type: none"> • empirical formula from the percentage composition or from other suitable experimental data • percentage composition from the formula of a compound • molecular formula when given both the empirical formula and the molar mass. 	3
1.3 Chemical Equations (1 h)		
1.3.1	Balance chemical equations when all reactants and products are given. Distinguish between coefficients and subscripts.	2
1.3.2	Identify the mole ratios of any two species in a balanced chemical equation. Use balanced chemical equations to obtain information about the amounts of reactants and products.	2
1.3.3	Apply the state symbols (s), (l), (g) and (aq). Encourage the use of state symbols in chemical equations.	2
1.4 Mass and Gaseous Volume Relationships in Chemical Reactions (3h)		
1.4.1	Calculate stoichiometric quantities and use these to determine experimental and theoretical yields. Mass is conserved in all chemical reactions. Given a chemical equation and the mass or amount (in moles) of one species, calculate the mass or amount of another species.	2, 3
1.4.2	Determine the limiting reactant and the reactant in excess when quantities of reacting substances are given. Given a chemical equation and the initial amounts of two or more reactants: <ul style="list-style-type: none"> • identify the limiting reactant • calculate the theoretical yield of a product • calculate the amount(s) of the reactant(s) in excess remaining after the reaction is complete. 	3
1.4.3	Apply Avogadro's law to calculate reacting volumes of gases.	2

A.S.		Obj
	1.5 Solutions (2h)	
1.5.1	Define the terms <i>solute</i> , <i>solvent</i> , <i>solution</i> and <i>concentration</i> (g dm^{-3} and mol dm^{-3}). Concentration in mol dm^{-3} is often represented by square brackets around the substance under consideration, eg $[\text{CH}_3\text{COOH}]$.	1
1.5.2	Carry out calculations involving concentration, amount of solute and volume of solution.	2
1.5.3	Solve solution stoichiometry problems. Given the quantity of one species in a chemical reaction in solution (in grams, moles or in terms of concentration), determine the quantity of another species.	3

Topic 2: Atomic Theory

A.S.

Obj

2.1 The Atom (2h)

- 2.1.1 State the relative mass and relative charge of protons, electrons and neutrons. 1

The accepted values are:

	Relative Mass	Charge
proton	1	+1
neutron	1	0
electron	$\frac{1}{1840}$	-1

- 2.1.2 State the position of protons, neutrons and electrons in the atom. 1

- 2.1.3 Define the terms *mass number (A)*, *atomic number (Z)* and *isotope*. 1

- 2.1.4 State the symbol for an isotope given its mass number and atomic number. 1

Use the notation ${}^A_Z\text{X}$, eg ${}^{12}_6\text{C}$.

- 2.1.5 Explain how the isotopes of an element differ. 3

Isotopes have the same chemical properties but different physical properties. Examples such as ${}^1_1\text{H}$, ${}^2_1\text{H}$, ${}^3_1\text{H}$; ${}^{12}_6\text{C}$, ${}^{14}_7\text{C}$; ${}^{35}_{17}\text{Cl}$ and ${}^{37}_{17}\text{Cl}$ should be considered.

- 2.1.6 Calculate and explain non-integer atomic masses from the relative abundance of isotopes. 2, 3

- 2.1.7 Calculate the number of protons, electrons and neutrons in atoms and ions from the mass number, atomic number and charge. 2

2.2 Electron Arrangement (2h)

- 2.2.1 Describe and explain the difference between a continuous spectrum and a line spectrum. 2, 3

- 2.2.2 Explain how the lines in the emission spectrum of hydrogen are related to the energy levels of electrons. 3

Students should be able to draw an energy-level diagram, show transitions between different energy levels and recognize that the lines in a line spectrum are directly related to these differences. An understanding of convergence is expected. Series should be considered in the ultraviolet, visible and infrared regions of the spectrum. Calculations, knowledge of quantum numbers and historical references are not required.

A.S.		Obj
2.2.3	Describe the electron arrangement of atoms in terms of main energy levels. Students should know the maximum number of electrons that can occupy a main energy level (up to $Z = 18$). No knowledge of sub-levels s, p, d and f is required. The term <i>valence electrons</i> is used to describe the electrons in the highest main energy level.	2
2.2.4	Determine the electron arrangement up to $Z = 20$. For example, 2.8.7 or 2,8,7 for $Z = 17$.	3

Topic 3: Periodicity

A.S.		Obj
	3.1 The Periodic Table (1h)	
3.1.1	Describe the arrangement of elements in the periodic table in order of increasing atomic number. Names and symbols of the elements are given in the <i>Chemistry Data Booklet</i> . The history of the periodic table is not required.	2
3.1.2	Distinguish between the terms <i>group</i> and <i>period</i> . The numbering system for groups in the periodic table is shown in the data booklet. Students should also be aware of the position of the transition metals in the periodic table.	2
3.1.3	Deduce the relationship between the electron configuration of elements and their position in the periodic table. Explanations are only required for the first 20 elements, although general principles can extend to the whole of the periodic table. For example, students should know or be able to predict that K is in group 1 using $Z = 19$, but need only know that since Cs is in group 1, it has one electron in its outer shell.	3
	3.2 Physical Properties (2h)	
3.2.1	Describe and explain the periodic trends in atomic radii, ionic radii, ionization energies, electronegativity and melting points for the alkali metals (Li \rightarrow Cs), halogens (F \rightarrow I) and period 3 elements (Na \rightarrow Ar). Cross reference with topics 2, 4 and 5. Data for all these properties are listed in the data booklet. Explanations for the first four trends should be given in terms of the balance between the attraction of the nucleus for the electrons and the repulsion between electrons. Explanations based on effective nuclear charge are not required. Ionization energy is defined as the minimum energy required to remove one electron from an isolated gaseous atom.	2, 3
	3.3 Chemical Properties (3h)	
3.3.1	Discuss the similarities in chemical nature of elements in the same group. The following reactions should be covered: <ul style="list-style-type: none"> • alkali metals (Li, Na and K) with water and with halogens (Cl_2 and Br_2) • halogens (Cl_2, Br_2 and I_2) with halide ions (Cl^-, Br^- and I^-) • halide ions (Cl^-, Br^- and I^-) with silver ions. Reactions of the halogens with alkali and confirmation of the silver halide by reaction with ammonia solution are not required.	3

A.S.**Obj****3.3.2**

Discuss the change in nature, from metallic to non-metallic, of the elements across period 3.

3

Use the study of the period 3 oxides to illustrate, for example, the change from basic through amphoteric to acidic oxides and their reaction with water. Halides and hydrides are not required.

Topic 4: Bonding

A.S.		Obj
	4.1 Ionic Bond (2h)	
4.1.1	Describe the ionic bond as the result of electron transfer leading to attraction between oppositely charged ions.	2
4.1.2	Determine which ions will be formed when metals in groups 1, 2 and 3 lose electrons.	3
4.1.3	Determine which ions will be formed when elements in groups 6 and 7 gain electrons.	3
4.1.4	State that transition metals can form more than one ion. Restrict examples to simple ions eg Fe^{2+} and Fe^{3+} .	1
4.1.5	Predict whether a compound of two elements would be mainly ionic or mainly covalent from the position of the elements in the periodic table, or from their electronegativity values.	3
4.1.6	Deduce the formula and state the name of an ionic compound formed from a group 1, 2 or 3 metal and a group 5, 6 or 7 non-metal.	3, 1
	4.2 Covalent Bond (6h)	
4.2.1	Describe the covalent bond as the result of electron sharing. The electron pair is attracted by both nuclei leading to a bond which is directional in nature. Both single and multiple bonds should be considered. Dative covalent bonds are not required.	2
4.2.2	Draw the electron distribution of single and multiple bonds in molecules. Examples should include O_2 , N_2 , CO_2 , C_2H_4 (ethene) and C_2H_2 (ethyne).	1
4.2.3	State and explain the relationship between the number of bonds, bond length and bond strength. The comparison should include bond lengths and bond strengths of: <ul style="list-style-type: none"> • two carbon atoms joined by single, double and triple bonds • the carbon atom and the two oxygen atoms in the carboxyl group of a carboxylic acid. 	1, 3
4.2.4	Compare the relative electronegativity values of two or more elements based on their positions in the periodic table. Precise values of electronegativity are not required.	2

A.S.		Obj
4.2.5	Identify the relative polarity of bonds based on electronegativity values. In a covalent bond, electron distribution may not be symmetrical and the electron pair may not be equally shared.	2
4.2.6	Draw and deduce Lewis (electron dot) structures of molecules and ions for up to four electron pairs on each atom. A pair of electrons can be represented by dots, crosses, a combination of dots and crosses or by a line. For example, chlorine can be shown as:	1, 3
	$\begin{array}{c} \times \times \\ \times \text{Cl} \times \times \\ \times \times \end{array} \times \begin{array}{c} \times \times \\ \times \text{Cl} \times \times \\ \times \times \end{array} \quad \text{or} \quad \begin{array}{c} \cdot \cdot \\ \cdot \text{Cl} \cdot \cdot \\ \cdot \cdot \end{array} : \begin{array}{c} \cdot \cdot \\ \cdot \text{Cl} \cdot \cdot \\ \cdot \cdot \end{array} : \quad \text{or} \quad \overline{\text{Cl}} - \overline{\text{Cl}} $	
	Note: Cl – Cl is not a Lewis structure.	
4.2.7	Predict the shape and bond angles for molecules with four charge centres on the central atom. Use the valence shell electron pair repulsion (VSEPR) theory to predict the shapes and bond angles of molecules and ions having four pairs of electrons (charge centres) around the central atom. Suitable examples are NH ₃ , H ₂ O and alkanes (eg CH ₄).	3
4.2.8	Identify the shape and bond angles for species with two and three negative charge centres. Examples should include species with non-bonding as well as bonding electron pairs, eg CO ₂ , SO ₂ , C ₂ H ₂ , C ₂ H ₄ , CO ₃ ²⁻ and NO ₂ ⁻ .	2
4.2.9	Predict molecular polarity based on bond polarity and molecular shape. The polarity of a molecule depends on its shape and on the electronegativities of its atoms, eg CO ₂ , H ₂ O.	3

4.3 Intermolecular Forces (1.5h)

4.3.1	Describe the types of intermolecular force (hydrogen bond, dipole–dipole attraction and van der Waals' forces) and explain how they arise from the structural features of molecules. All these intermolecular forces are weaker than covalent bonds. For substances of similar molar mass, hydrogen bonds are stronger than dipole–dipole attractions which are stronger than van der Waals' forces. Van der Waals' forces arise from the electrostatic attraction between temporary induced dipoles in both polar and non-polar molecules.	2, 3
4.3.2	Describe and explain how intermolecular forces affect the boiling points of substances. The hydrogen bond can be illustrated by comparing physical properties of: <ul style="list-style-type: none"> • H₂O and H₂S • NH₃ and PH₃ • C₃H₈, CH₃CHO and C₂H₅OH. 	2, 3

A.S.		Obj
	4.4 Metallic Bond (0.5h)	
4.4.1	Describe metallic bond formation and explain the physical properties of metals. Metallic bonding is explained in terms of a lattice of positive ions surrounded by delocalized valence electrons. The delocalized electrons should be related to the high electrical conductivity, malleability and ductility of metals.	2, 3
	4.5 Physical Properties (2h)	
4.5.1	Compare and explain the following properties of substances resulting from different types of bonding: melting and boiling points, volatility, conductivity and solubility. Consider melting points, boiling points and volatility of similar substances, such as F ₂ , Cl ₂ , Br ₂ and I ₂ , and substances with different types of bonding and different intermolecular forces. Students should be aware of the effect of impurities on the melting point of a substance. The solubilities of compounds in non-polar and polar solvents should be compared and explained. Consider also the solubilities of alcohols in water as the length of the carbon chain increases.	2, 3
4.5.2	Predict the relative values of melting and boiling points, volatility, conductivity and solubility based on the different types of bonding in substances.	3

Topic 5: States of Matter

A.S.		Obj
	5.1 States of Matter (5h)	
5.1.1	Describe and compare solids, liquids and gases as the three states of matter. The movement of particles, the attractive forces between particles and interparticle spacing should be described. A molecular level description of what happens when evaporation, boiling, condensing, melting and freezing occur should be given. Students should understand what is meant by the term <i>diffusion</i> .	2
5.1.2	Describe kinetic theory in terms of the movement of particles whose average energy is proportional to absolute temperature. Kinetic theory should be interpreted in terms of ideal gases consisting of point masses in random motion whose energy is proportional to absolute temperature. Students should be able to describe what happens when the temperature is changed.	2
5.1.3	Describe the Maxwell–Boltzmann energy distribution curve.	2
5.1.4	Draw and explain qualitatively Maxwell–Boltzmann energy distribution curves for different temperatures.	1, 3
5.1.5	Describe qualitatively the effects of temperature, pressure and volume changes on a fixed mass of an ideal gas.	2
5.1.6	State the ideal gas equation, $PV = nRT$.	1
5.1.7	Apply the ideal gas equation in calculations. Use the relationship between P , V , n and T for gases. Students should be familiar with $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ and be able to calculate molar volume.	2

Topic 6: Energetics

A.S.		Obj
	6.1 Exothermic and Endothermic Reactions (2h)	
6.1.1	Define the terms <i>exothermic reaction</i> , <i>endothermic reaction</i> and <i>standard enthalpy change of reaction</i> (ΔH^\ominus). Standard enthalpy change is heat transferred under standard conditions—pressure 101.3 kPa, temperature 298 K. Only ΔH can be measured, not H for the initial or final state of a system.	1
6.1.2	State the relationship between temperature change, enthalpy change and whether a reaction is exothermic or endothermic. Combustion of organic compounds are good examples of exothermic reactions.	1
6.1.3	Deduce, from an enthalpy level diagram, the relative stabilities of reactants and products and the sign of the enthalpy change for the reaction. If the final state is more stable (lower on the enthalpy level diagram), this implies that $H_{\text{final}} < H_{\text{initial}}$ and ΔH must be negative. Energy must be released in going to a more stable state.	3
6.1.4	Describe and explain the changes which take place at the molecular level in chemical reactions. Relate bond formation to the release of energy and bond breaking to the absorption of energy.	2, 3
6.1.5	Suggest suitable experimental procedures for measuring enthalpy changes of reactions in aqueous solution. Explore different reactions operating at constant pressure (open containers). Use of the bomb calorimeter is not required.	3
	6.2 Calculation of Enthalpy Changes (3h)	
6.2.1	Calculate the heat change when the temperature of a pure substance is altered. Students should be able to calculate the heat change for a substance given the mass, specific heat and temperature change.	2
6.2.2	Explain that enthalpy changes of reaction relate to specific quantities of either reactants or products. Enthalpy changes are measured in joules (J) and are often quoted in kJ mol^{-1} of either a reactant or a product.	3
6.2.3	Analyse experimental data for enthalpy changes of reactions in aqueous solution.	3

A.S.		Obj
6.2.4	Calculate the enthalpy change for a reaction in aqueous solution using experimental data on temperature changes, quantities of reactants and mass of solution.	2

Enthalpy change of an acid–base reaction could be investigated.

6.3 Hess's Law (2h)

6.3.1	Determine the enthalpy change of a reaction which is the sum of two or more reactions with known enthalpy changes.	3
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Use examples of simple two- and three-step processes. Students should be able to construct simple enthalpy cycles, but will not be required to state Hess's law.

6.4 Bond Enthalpies (1h)

6.4.1	Define the term <i>average bond enthalpy</i> .	1
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Bond enthalpies are quoted for the gaseous state and should be recognized as average values obtained from a number of similar compounds. Cross reference with 11.2.6.

6.4.2	Calculate the enthalpy change of a reaction using bond enthalpies.	2
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6.5 Entropy (1h)

6.5.1	State and explain the factors which increase the disorder (entropy) in a system.	1, 3
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An increase in disorder can result from the mixing of different types of particles, change of state (increased distance between particles), increased movement of particles or increased numbers of particles. An increase in the number of particles in the gaseous state usually has a greater influence than any other possible factor.

6.5.2	Predict whether the entropy change (ΔS) for a given reaction or process would be positive or negative.	3
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From a given equation, identify a single factor which affects the value of ΔS and predict the sign of ΔS .

6.6 Spontaneity (2h)

6.6.1	Define <i>standard free energy change of reaction</i> (ΔG^\ominus).	1
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6.6.2	State whether a reaction or process will be spontaneous by using the sign of ΔG^\ominus .	1
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6.6.3	State and predict the effect of a change in temperature on the spontaneity of a reaction, given standard entropy and enthalpy changes.	1, 3
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Use the equation $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$.

Topic 7: Kinetics

A.S.		Obj
	7.1 Rates of Reaction (2h)	
7.1.1	<p>Define the term <i>rate of reaction</i> and describe the measurement of reaction rates.</p> <p>Rate of reaction can be defined as the decrease in the concentration of reactants per unit time or the increase in the concentration of product per unit time.</p>	1, 2
7.1.2	<p>Analyse data from rate experiments.</p> <p>Graphs of changes in concentration, volume or mass against time should be interpreted qualitatively.</p>	3
	7.2 Collision Theory (2h)	
7.2.1	<p>Describe and explain the collision theory.</p> <p>Students should know that not all collisions lead to a reaction.</p>	2, 3
7.2.2	<p>Define <i>activation energy</i> (E_a) and explain that reactions occur when reacting species have $E \geq E_a$.</p> <p>Molecules must have a minimum energy and appropriate collision geometry in order to react. A simple treatment is all that is required. Cross reference with 5.1.3 and 5.1.4.</p>	1, 3
7.2.3	<p>Predict and explain, using collision theory, the qualitative effect of particle size, temperature, concentration and catalysts on the rate of a reaction.</p> <p>Increasing the temperature increases the frequency of collisions but, more importantly, the proportion of molecules with $E \geq E_a$ increases.</p>	3
7.2.4	<p>Explain that reactions can occur by more than one step and that one step can determine the rate of reaction.</p> <p>Few reactions involve just one step although one step in the reaction, the <i>rate determining step</i>, determines the reaction rate. Orders of reactions and rate laws are not required.</p>	3

Topic 8: Equilibrium

A.S.		Obj
	8.1 Dynamic Equilibrium (1h)	
8.1.1	Outline the characteristics of a system in a state of equilibrium. Many chemical reactions are reversible and never go to completion. Equilibrium can be approached from both directions. For a system in equilibrium the rate of the forward reaction equals the rate of the reverse reaction and the concentrations of all reactants and products remain constant. The system is closed and macroscopic properties remain constant. Use phase equilibrium as an example of dynamic equilibrium involving physical changes.	2
	8.2 The Position of Equilibrium (4h)	
8.2.1	State the equilibrium constant expression (K_c) for a homogeneous reaction. Consider equilibria involving one phase, gases or species in aqueous solution. The equilibrium constant is specific to a given system and varies with temperature. No calculations are required.	1
8.2.2	Deduce the extent of a reaction from the magnitude of the equilibrium constant. When $K_c \gg 1$, the reaction goes almost to completion. When $K_c \ll 1$, the reaction hardly proceeds.	3
8.2.3	Describe and predict the qualitative effects of changes of temperature, pressure and concentration on the position of equilibrium and the value of the equilibrium constant. Use Le Chatelier's principle to predict the effects of these changes on the position of equilibrium. The value of the equilibrium constant (K_c) is only affected by temperature. The position of equilibrium may change without the value of K_c changing.	2, 3
8.2.4	State and explain the effect of a catalyst on an equilibrium reaction.	1, 3
8.2.5	Describe and explain the application of equilibrium and kinetics concepts to the Haber process and the Contact process.	2, 3

Topic 9: Acids and Bases

A.S.		Obj
	9.1 Properties of Acids and Bases (1h)	
9.1.1	Outline the characteristic properties of acids and bases in aqueous solution. The properties that must be considered are: effects on indicators and reactions of acids with bases, metals and carbonates. Bases which are not hydroxides, such as ammonia, soluble carbonates and hydrogencarbonates, should be included. Alkalis are bases that dissolve in water.	2
	9.2 Strong and Weak Acids and Bases (1h)	
	Note: Brønsted–Lowry definitions of acids and bases are not required for this sub-topic.	
9.2.1	Describe and explain the differences between strong and weak acids and bases in terms of the extent of dissociation, reaction with water and conductivity. The term <i>ionization</i> can be used instead of <i>dissociation</i> . Solutions of equal concentration can be compared by pH and/or conductivity.	2, 3
9.2.2	State whether a given acid or base is strong or weak. Specified strong acids are hydrochloric acid, nitric acid and sulfuric acid. Specified weak acids are ethanoic acid and carbonic acid (aqueous carbon dioxide). Specified strong bases are all group 1 hydroxides and barium hydroxide. Specified weak bases are ammonia and ethylamine.	1
9.2.3	Describe and explain data from experiments to distinguish between strong and weak acids and bases, and to determine the relative acidities and basicities of substances.	2, 3
	9.3 The pH Scale (1h)	
9.3.1	Distinguish between aqueous solutions that are acidic, neutral or basic using the pH scale.	2
9.3.2	Identify which of two or more aqueous solutions is more acidic or basic, using pH values. Measure pH using a pH meter or pH paper. Students should know that pH paper contains a mixture of indicators. The theory of pH meters is not required.	2

A.S.		Obj
9.3.3	State that each change of one pH unit represents a tenfold change in the hydrogen ion concentration $[H^+(aq)]$. Relate integral values of pH to $[H^+(aq)]$ expressed as powers of ten. Calculation of pH from $[H^+(aq)]$ is not required.	1
9.3.4	Deduce changes in $[H^+(aq)]$ when the pH of a solution changes by more than one pH unit.	3
9.4 Buffer Solutions (1h)		
9.4.1	Describe a buffer solution in terms of its composition and behaviour. A buffer resists change in pH when a small amount of a strong acid or base is added. Suitable examples include ammonium chloride/ammonia solution and ethanoic acid/sodium ethanoate. Blood is an example of a buffer solution.	2
9.4.2	Describe ways of preparing buffer solutions.	2
9.5 Acid–base Titrations (1h)		
9.5.1	Draw and explain a graph showing pH against volume of titrant for titrations involving strong acids and bases.	1, 3

Topic 10: Oxidation and Reduction

A.S.		Obj
	10.1 Oxidation and Reduction (2h)	
10.1.1	Define <i>oxidation</i> and <i>reduction</i> in terms of electron loss and gain. Introduce the concept of the half-equation.	1
10.1.2	Calculate the oxidation number of an element in a compound. Oxidation numbers should be shown by a sign (+ or -) and a number, eg +7 for Mn in KMnO_4 .	2
10.1.3	State and explain the relationship between oxidation numbers and the names of compounds. Oxidation numbers in names of compounds are represented by Roman numerals, eg iron(II) oxide, iron(III) oxide.	1, 3
10.1.4	Identify whether an element is oxidized or reduced in simple redox reactions, using oxidation numbers. Appropriate reactions to illustrate this can be found in topics 3 and 11. Possible examples include: iron(II) and (III), manganese(II) and (VII), chromium(III) and (VI), copper(I) and (II), oxides of sulfur and oxyacids, halogens and halide ions.	2
10.1.5	Define the terms <i>oxidizing agent</i> and <i>reducing agent</i> .	1
	10.2 Reactivity (2.5h)	
10.2.1	Deduce a reactivity series based upon the chemical behaviour of a group of oxidizing and reducing agents. Displacement reactions of metals and halogens (see 3.3.1) provide a good experimental illustration of reactivity. Standard electrode potentials or reduction potentials are not required.	3
10.2.2	Deduce the feasibility of a redox reaction from a given reactivity series.	3
10.2.3	Describe and explain how a redox reaction is used to produce electricity in a voltaic cell. Students should be able to draw a diagram of a simple half-cell, and show how two half-cells can be connected by a salt bridge to form a whole cell. Suitable examples of half-cells are Mg, Zn, Fe and Cu in solutions of their ions.	2, 3

A.S.		Obj
	10.3 Electrolysis (2.5h)	
10.3.1	Draw a diagram showing the essential components of an electrolytic cell. An electrolytic cell converts electrical energy to chemical energy. The diagram should include the source of electric current and conductors, positive and negative electrodes and the electrolyte.	1
10.3.2	Describe how current is conducted in an electrolytic cell.	2
10.3.3	Deduce the products for the electrolysis of a molten salt. Equations showing the formation of products at each electrode should be given.	3
10.3.4	Distinguish between the use of a spontaneous redox reaction to produce electricity in a voltaic cell and the use of electricity to carry out a non-spontaneous redox reaction in an electrolytic cell. Some teachers may wish to describe reactions at the electrodes in a cell in terms of reduction at the cathode and oxidation at the anode, but this is not required.	2
10.3.5	Describe and explain the use of electrolysis in electroplating. Restrict this to copper plating.	2, 3

Topic 11: Organic Chemistry

A.S.		Obj
	11.1 Homologous Series (1h)	
11.1.1	Describe the features of a homologous series. Features include a general formula and neighbouring members differing by CH_2 , with similar chemical properties and with a gradation in physical properties.	2
11.1.2	Predict and explain the trends in boiling points of members of a homologous series. In a homologous series there is a gradual increase in boiling point as the number of carbon atoms increases. Cross reference with 4.3.	3
	11.2 Hydrocarbons (2h)	
11.2.1	Draw structural formulas for the isomers of the non-cyclic alkanes up to C_6 . Structural formulas should indicate clearly the bonding between atoms. For example, for pentane:	1
	$ \begin{array}{ccccccccc} & \text{H} & & \text{H} & & \text{H} & & \text{H} & & \text{H} & & \\ & & & & & & & & & & & \\ \text{H} & -\text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{C} & - & \text{H} \\ & & & & & & & & & & & \\ & \text{H} & & \text{H} & & \text{H} & & \text{H} & & \text{H} & & \end{array} $ or $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$	
	Or $\text{CH}_3\text{---}(\text{CH}_2)_3\text{---CH}_3$, but not C_5H_{12} .	
11.2.2	State the names of alkanes up to C_6 . Name these using IUPAC rules. Consider both straight and branch-chained alkanes.	1
11.2.3	Explain the relative inertness of alkanes. Refer to bond enthalpies. See 6.4.	3
11.2.4	Draw structural formulas and state the names for straight-chain alkenes (C_nH_{2n} , where n is between 2 and 5). Geometric (<i>cis-trans</i>) isomers are not required.	1
11.2.5	Describe complete and incomplete combustion of hydrocarbons. The formation of CO and C during incomplete combustion should be related to environmental impacts and oxidation–reduction.	2
11.2.6	State that the combustion of hydrocarbons is an exothermic process. See 6.3 and 6.4.	1

A.S.

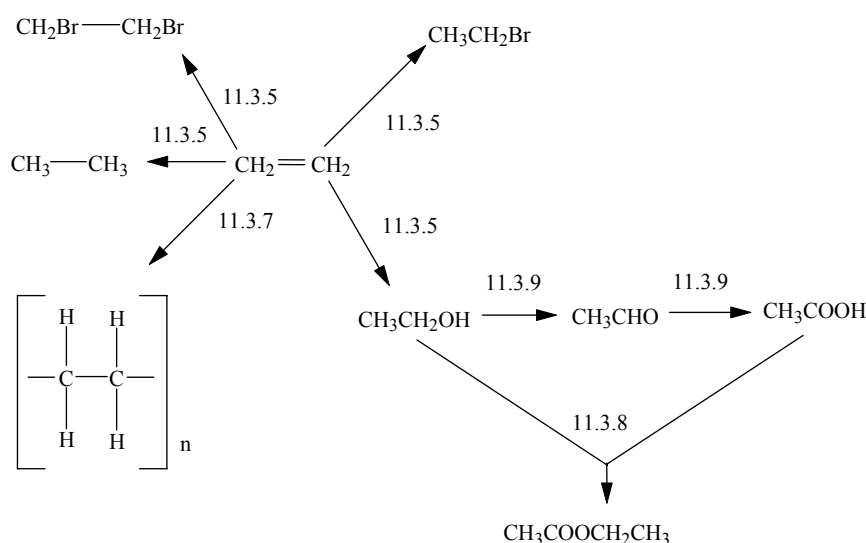
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11.3 Other Functional Groups (7h)

Along with alkanes and alkenes, compounds containing one or more functional groups have been chosen to introduce students to:

- interrelationships involving significant functional groups
- important reaction types such as addition, substitution, oxidation, condensation, esterification and polymerization.

This is expressed in the following scheme:

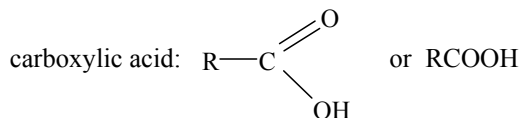
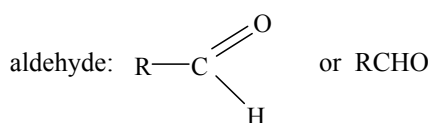


11.3.1

Draw and state the names of compounds containing up to five carbon atoms with one of the following functional groups: aldehyde, ketone, carboxylic acid, alcohol, amide, amine, ester and halogenoalkane.

1

Functional groups in full and condensed forms are required, eg:



11.3.2

Explain that functional groups can exist as isomers.

3

Examples include:

- ethanoic acid (CH_3COOH) and methyl methanoate (HCOOCH_3)
- propanal ($\text{CH}_3\text{CH}_2\text{CHO}$) and propanone (CH_3COCH_3).

A.S.		Obj
11.3.3	Outline the existence of optical isomers. Restrict this to the fact that, if a carbon atom has four different substituents, the molecule exists in two enantiomeric forms that rotate the plane of polarized light in opposite directions. Students should be able to identify a chiral (asymmetric) centre.	2
11.3.4	Discuss the volatility, solubility in water and acid–base behaviour of the functional groups aldehyde, ketone, carboxylic acid, alcohol, amide, amine, ester and halogenoalkane. For example, use functional groups to explain the higher boiling point of methanol compared with methane. Cross reference with 4.3.	3
	Note: Mechanisms are not required for 11.3.5 to 11.3.10. Conditions are also not required unless specified.	
11.3.5	Outline the reaction of symmetrical alkenes with hydrogen, bromine, hydrogen halides and water. A double bond is relatively reactive, therefore molecules such as $\text{H}_2\text{C}=\text{CH}_2$ are important starting materials in organic synthesis.	2
11.3.6	Outline the uses of reactions of alkenes. Hydrogenation is used in the production of margarine, hydration of ethene is used in the manufacture of ethanol, and bromination can be used to distinguish between alkanes and alkenes.	2
11.3.7	Outline the polymerization of alkenes. Polyethene and polyvinyl chloride should be used as examples of addition polymers. Students should be able to draw the structures of the monomer and the repeating unit of the polymer.	2
11.3.8	Outline the condensation reaction of an alcohol with a carboxylic acid to form an ester, and state the uses of esters. Esters are used as flavouring agents, in plasticizers, as solvents and in perfumes.	1, 2
11.3.9	Describe the partial and complete oxidation of ethanol. A suitable oxidizing agent is acidified potassium dichromate(VI). Both oxidation products (ethanal and ethanoic acid) can be obtained by altering the conditions, eg ethanal by distilling off the product as it is formed, and ethanoic acid by heating under reflux.	2
11.3.10	Deduce the condensation polymers formed by amines and by carboxylic acids. Emphasize the need for two functional groups on the monomers: <ul style="list-style-type: none">• polyamides (nylons)—eg hexanedioic acid and 1,6-diaminohexane• polyesters—eg benzene-1,4-dicarboxylic acid and ethane-1,2-diol.	3

A.S.

Obj

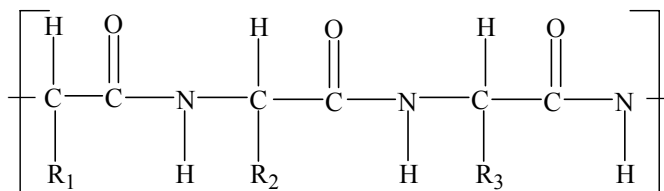
11.3.11

Outline the formation of peptides and proteins from 2-amino acids.

2

All 2-amino acids (α -amino acids), except aminoethanoic acid (glycine), can show optical activity (see 11.3.3). Peptides are formed from amino acids, and two functional groups allow for the formation of macromolecules.

Students should be familiar with simple primary structures (order of amino acids) for peptides containing up to three amino acids, eg:



Topic 12: Atomic Theory

A.S.		Obj
	12.1 The Mass Spectrometer (1h)	
12.1.1	State the principles of a mass spectrometer and outline the main stages in its operation. A simple diagram of a single beam mass spectrometer is required. The following stages of operation should be considered: vaporization, ionization, acceleration, deflection and detection.	1, 2
12.1.2	Describe how the mass spectrometer may be used to determine relative isotopic, atomic and molecular masses using the ^{12}C scale. Students should be able to calculate the relative atomic mass from the abundance of the isotopes (see 2.1.6). Interpretation of fragmentation patterns is not required.	2
	12.2 Electron Configuration of Atoms (3h)	
12.2.1	State and explain how evidence from first and successive ionization energies accounts for the existence of the main energy levels and sub-levels. Interpretation of graphs of first ionization and successive ionization energies versus atomic number provides evidence for the existence of the main energy levels and sub-levels.	1, 3
12.2.2	State how orbitals are labelled. Limit this to $n < 5$.	1
12.2.3	State the relative energies of s, p, d and f orbitals.	1
12.2.4	State the number of orbitals at each energy level.	1
12.2.5	Draw the shape of an s orbital and the shapes of the p_x , p_y and p_z orbitals.	1
12.2.6	State the Aufbau principle. Reference should be made to Hund's rule.	1
12.2.7	Apply the Aufbau principle to electron configurations. Apply the Aufbau principle for an atom up to $Z = 54$, eg for $Z = 23$ the electronic configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$ or $[\text{Ar}] 4s^2 3d^3$ or $[\text{Ar}] 3d^3 4s^2$. Exceptions to this rule are not expected.	2
12.2.8	Relate the electron configuration of an atom to its position in the periodic table. Students should be able to label the s, p, d and f blocks of the periodic table.	2

Topic 13: Periodicity

A.S.

Obj

13.1 Periodic Trends Na → Ar (the third period) (2h)

13.1.1 Explain the physical properties of the chlorides and oxides of the elements in the third period (Na → Ar) in terms of their bonding and structure. **3**

Refer to the following oxides and chlorides:

- Oxides: Na₂O, MgO, Al₂O₃, SiO₂, P₄O₆ and P₄O₁₀, SO₂ and SO₃, Cl₂O and Cl₂O₇.
- Chlorides: NaCl, MgCl₂, Al₂Cl₆, SiCl₄, PCl₃ and PCl₅, and Cl₂ (sulfur chloride is not required).

Limit the explanation to the physical states of the compounds under standard conditions and electrical conductivity in the molten state only.

13.1.2 Describe the chemical trends for the chlorides and oxides referred to in 13.1.1. Include relevant equations. **2**

Limit this to acid–base properties of the oxides and the reactions of the chlorides and oxides with water.

13.2 d-block Elements (first row) (2h)

13.2.1 List the characteristic properties of transition elements. **1**

Restrict this to variable oxidation states, complex ion formation, coloured compounds and catalytic properties.

13.2.2 Identify which elements are considered to be typical of the d-block elements. **2**

Sc and Zn are not typical.

13.2.3 Describe the existence of variable oxidation states in d-block elements. **2**

The 4s and 3d sub-levels are close in energy. Students should know that all d-block elements can show an oxidation state of +2. In addition, they should be familiar with the oxidation states of the following: Cr (+3, +6), Mn (+4, +7), Fe (+3) and Cu (+1).

13.2.4 Define the term *ligand*. **1**

13.2.5 Describe how complexes of d-block elements are formed. **2**

Suitable examples are: [Fe(H₂O)₆]³⁺, [Fe(CN)₆]³⁻, [CuCl₄]²⁻, [Cu(NH₃)₄]²⁺, [Ag(NH₃)₂]⁺. Only monodentate ligands are required.

13.2.6 Explain why some complexes of d-block elements are coloured. **3**

Students need only know that in complexes the d orbitals are split into two sets at different energy levels and the electronic transitions that take place between them are responsible for their colours.

A.S.		Obj
13.2.7	Outline the catalytic behaviour of d-block elements and their compounds. Limit this to: <ul style="list-style-type: none">• MnO_2 in the decomposition of hydrogen peroxide• V_2O_5 in the Contact process• Fe in the Haber process• Ni in the conversion of alkenes to alkanes. The mechanisms of action are not required.	2

Topic 14: Bonding

A.S. Obj

14.1 Shapes of Molecules and Ions (1h)

14.1.1 State and predict the shape and bond angles using the VSEPR theory for 5- and 6-negative charge centres. 1, 3

The shape of the molecules/ions and bond angles if all pairs of electrons are shared, and the shape of the molecules/ions if one or more lone pairs surround the central atom, should be considered. Examples such as PCl_5 , SF_6 and XeF_4 can be used.

14.2 Hybridization (2h)

14.2.1 Describe σ and π bonds. 2

Treatment should be restricted to:

- σ bonds—electron distribution has axial symmetry around the axis joining the two nuclei
- π bonds resulting from the combination of parallel p orbitals
- double bonds formed by a σ and a π bond
- triple bonds formed by a σ and two π bonds.

14.2.2 State and explain the meaning of the term *hybridization*. 1, 3

Hybridization should be explained in terms of the mixing of atomic orbitals to form new orbitals for bonding. Students should consider sp, sp² and sp³ hybridization, and the shapes and orientation of these orbitals.

14.2.3 Discuss the relationships between Lewis structures, molecular shapes and types of hybridization (sp, sp² and sp³). 3

Using examples from inorganic as well as organic chemistry, students should write the Lewis structure, deduce the shape of the molecule and recognize the type of hybridization.

14.3 Delocalization of Electrons (2h)

14.3.1 State what is meant by the delocalization of π electrons and explain how this can account for the structures of some substances. 1, 3

Examples such as NO_3^- , NO_2^- , CO_3^{2-} , O_3 , RCOO^- and benzene can be used. (These could also be dealt with through the resonance approach.)

14.4 Structures of Allotropes of Carbon (1h)

14.4.1 Describe and explain the structures and properties of diamond, graphite and fullerene. 2, 3

Students should recognize the type of hybridization present in each allotrope and the delocalization of electrons in graphite and C₆₀ fullerene.

Topic 15: Energetics

A.S.		Obj
	15.1 Standard Enthalpy Changes of Reaction (1h)	
15.1.1	Define and use the terms <i>standard state</i> and <i>standard enthalpy change of formation</i> (ΔH_f^\ominus).	1
15.1.2	Calculate the enthalpy change of a reaction using standard enthalpy changes of formation.	2
	15.2 Lattice Enthalpy (2h)	
15.2.1	Define the term <i>lattice enthalpy</i> . The sign of $\Delta H_{\text{lattice}}$ indicates whether the lattice is being formed or broken.	1
15.2.2	Compare the effect of both the relative sizes and the charges of ions on the lattice enthalpies of different ionic compounds. The relative value of the theoretical lattice enthalpy increases with higher ionic charge and smaller ionic radius due to increased attractive forces.	2
15.2.3	Construct a Born–Haber cycle and use it to calculate an enthalpy change.	2, 3
15.2.4	Analyse theoretical and experimental lattice enthalpy values. A significant difference between the two values indicates covalent character.	3
	15.3 Spontaneity of a Reaction (1h)	
15.3.1	Calculate the standard entropy change for a reaction (ΔS^\ominus) using values of absolute entropies.	2
15.3.2	Calculate ΔG^\ominus for a reaction using the equation $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$ or by using values of the standard free energy change of formation, ΔG_f^\ominus .	2

Topic 16: Kinetics

A.S.		Obj
	16.1 Rate Expression (3h)	
16.1.1	Define the terms <i>rate constant</i> and <i>order of reaction</i> .	1
16.1.2	Derive the rate expression for a reaction from data. $\text{Rate} = k[\text{A}]^m[\text{B}]^n$ where k = rate constant, $[\text{A}]$ = concentration of A in mol dm ⁻³ etc, m and n = integers, $m + n$ = overall order of the reaction.	3
16.1.3	Draw and analyse graphical representations for zero-, first- and second-order reactions.	1, 3
16.1.4	Define the term <i>half-life</i> and calculate the half-life for first-order reactions only. The half-life should be calculated from graphs and by using the integrated form of the rate equation. The integrated rate equation for second-order reactions is not required.	1, 2
	16.2 Reaction Mechanism (1h)	
16.2.1	Define the terms <i>rate-determining step</i> , <i>molecularity</i> and <i>activated complex</i> .	1
16.2.2	Describe the relationship between mechanism, order, rate-determining step and activated complex. Limit examples to one- or two-step reactions where the mechanism is known. Students should understand what an activated complex (transition state) is and how the order of a reaction relates to the mechanism.	2
	16.3 Activation Energy (2h)	
16.3.1	Describe qualitatively the relationship between the rate constant (k) and temperature (T).	2
16.3.2	Describe how the Arrhenius equation can be used to determine the activation energy and the Arrhenius constant (A). Arrhenius equation: $k = Ae^{(-E_a/RT)}$ A relates to the geometric requirements of the collisions (see 7.2). Direct substitution using simultaneous equations and a graphical method can be used. The logarithmic form of the Arrhenius equation is: $\ln k = -\frac{E_a}{RT} + \ln A$ Both methods should be explained, but actual calculations are not needed.	2

A.S.		Obj
16.3.3	Draw and explain enthalpy level diagrams for reactions with and without catalysts.	1, 3
16.3.4	Distinguish between <i>homogeneous catalysts</i> and <i>heterogeneous catalysts</i> . Homogeneous catalyst—reactants and catalyst are in the same phase. Heterogeneous catalyst—reactants and catalyst are in different phases.	2
16.3.5	Outline the use of homogeneous and heterogeneous catalysts. Examples include hydrogenation using metals (see 13.2.7) and acid-catalysed formation of esters.	2

Topic 17: Equilibrium

A.S.		Obj
	17.1 Phase Equilibrium (2h)	
17.1.1	State and explain the equilibrium established between a liquid and its own vapour. Liquid–vapour equilibrium is a dynamic equilibrium established when the rate of condensation equals the rate of vaporization. The vapour pressure is independent of the volume of the container, liquid or vapour.	1, 3
17.1.2	State and explain the qualitative relationship between vapour pressure and temperature. Students should be able to show the relationship graphically and explain it in terms of kinetic theory.	1, 3
17.1.3	State and explain the relationship between enthalpy of vaporization, boiling point and intermolecular forces. Students should be able to predict the relative strength of intermolecular forces of different liquids when given the physical properties, or vice versa. Cross reference with 4.3.	1, 3
	17.2 The Equilibrium Law (2h)	
17.2.1	Solve homogeneous equilibrium problems using the expression for K_c . Calculate K_c given all equilibrium concentrations. Given K_c and other appropriate concentrations, find an equilibrium concentration. K_p and K_{sp} are not required, nor is use of the quadratic expression.	3

Topic 18: Acids and Bases

A.S.		Obj
	18.1 Brønsted–Lowry Acids and Bases (2h)	
18.1.1	Define acids and bases according to the Brønsted–Lowry theory.	1
18.1.2	Identify whether or not a compound could act as a Brønsted–Lowry acid or base.	2
18.1.3	Identify the conjugate acid–base pairs in a given acid–base reaction.	2
18.1.4	Determine the structure for the conjugate acid (or base) of any Brønsted–Lowry base (or acid). The members of a conjugate acid–base pair always differ by a single proton (H^+). Structures of conjugate acid–base pairs should always make clear the approximate location of the proton transferred, eg $\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$ rather than $\text{C}_2\text{H}_4\text{O}_2/\text{C}_2\text{H}_3\text{O}_2^-$.	3
	18.2 Lewis Theory (1h)	
18.2.1	Define and apply the terms <i>Lewis acid</i> and <i>Lewis base</i> . A Lewis acid–base reaction involves the formation of a new covalent bond in which both electrons are provided by one species. Such bonds are called dative covalent bonds. The formation of complexes (see 13.2.4 and 13.2.5) is usually a Lewis acid–base reaction.	1, 2
	18.3 Calculations Involving Acids and Bases (5h)	
	Note: A proton in water can be written as $\text{H}^+(\text{aq})$ or $\text{H}_3\text{O}^+(\text{aq})$; the former is adopted here.	
18.3.1	State the expression for the ionic product constant of water (K_w). $K_w = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$ at 298 K, but this varies with temperature.	1
18.3.2	Deduce $[\text{H}^+(\text{aq})]$ and $[\text{OH}^-(\text{aq})]$ for water at different temperatures given K_w values.	3
18.3.3	Define <i>pH</i> , <i>pOH</i> and <i>pK_w</i> .	1
18.3.4	Calculate $[\text{H}^+(\text{aq})]$, $[\text{OH}^-(\text{aq})]$, pH and pOH from specified concentrations. The values of $[\text{H}^+(\text{aq})]$ or $[\text{OH}^-(\text{aq})]$ are directly related to the concentration of the acid or base.	2

A.S.		Obj
18.3.5	<p>State the equation for the reaction of any weak acid or weak base with water, and hence derive the ionization constant expression.</p> <p>In general $\text{HA(aq)} \rightleftharpoons \text{H}^{\text{+}}(\text{aq}) + \text{A}^{-}(\text{aq})$ $\text{B(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{BH}^{\text{+}}(\text{aq}) + \text{OH}^{-}(\text{aq})$ (base hydrolysis)</p> <p>Then $K_{\text{a}} = \frac{[\text{H}^{\text{+}}(\text{aq})][\text{A}^{-}(\text{aq})]}{[\text{HA(aq)}]}$ and $K_{\text{b}} = \frac{[\text{BH}^{\text{+}}(\text{aq})][\text{OH}^{-}(\text{aq})]}{[\text{B(aq)}]}$</p> <p>Examples used should involve the transfer of only one proton.</p>	1, 3
18.3.6	Derive the expression $K_{\text{a}} \times K_{\text{b}} = K_{\text{w}}$ and use it to solve problems for any weak acid and its conjugate base and for any weak base and its conjugate acid.	3
18.3.7	State and explain the relationship between K_{a} and $\text{p}K_{\text{a}}$ and between K_{b} and $\text{p}K_{\text{b}}$.	1, 3
18.3.8	Determine the relative strengths of acids or their conjugate bases from K_{a} or $\text{p}K_{\text{a}}$ values.	3
18.3.9	<p>Apply K_{a} or $\text{p}K_{\text{a}}$ in calculations.</p> <p>Calculations can be performed using various forms of the acid ionization constant expression (see 18.3.5). Students should state when approximations are used in equilibrium calculations. Use of the quadratic expression is not required.</p>	2
18.3.10	<p>Calculate the pH of a specified buffer system.</p> <p>Calculations will involve the transfer of only one proton. Cross reference with 9.4.</p>	2

18.4 Salt Hydrolysis (1h)

18.4.1	<p>State and explain whether salts form acidic, alkaline or neutral aqueous solutions.</p> <p>Examples should include salts formed from the four possible combinations of strong and weak acids and bases. The effect of the charge density of the cations in groups 1, 2, 3 and d-block elements should also be considered, eg $[\text{Fe}(\text{H}_2\text{O})_6]^{3+} \rightleftharpoons [\text{Fe}(\text{OH})(\text{H}_2\text{O})_5]^{2+} + \text{H}^{\text{+}}$.</p>	1, 3
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18.5 Acid–base Titrations (1h)

18.5.1	<p>Draw and explain the general shapes of graphs of pH against volume of titrant for titrations involving monoprotic acids and bases.</p> <p>All combinations should be covered: strong acid + strong base, strong acid + weak base, weak acid + strong base and weak acid + weak base.</p>	1, 3
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A.S.		Obj
	18.6 Indicators (1h)	
18.6.1	Describe qualitatively how an acid–base indicator works. Use $\text{HIn(aq)} \rightleftharpoons \text{H}^+(\text{aq}) + \text{In}^-(\text{aq})$, or similar. colour A colour B	2
18.6.2	State and explain how the pH range of an acid–base indicator relates to its $\text{p}K_{\text{a}}$ value.	1, 3
18.6.3	Determine an appropriate indicator for a titration, given the equivalence point of the titration and K_{a} (or $\text{p}K_{\text{a}}$) values for possible indicators.	3

Topic 19: Oxidation and Reduction

A.S.		Obj
	19.1 Redox Equations (2h)	
19.1.1	Balance redox equations in acid solution. Half-equations and oxidation numbers may be used. $\text{H}^+(\text{aq})$ and H_2O should be used where necessary to balance half-equations.	2
	19.2 Standard Electrode Potentials (3h)	
	Note: In 19.2.1 to 19.2.4 half-equations can be used to introduce redox couples, including H^+/H_2 and a selection of common couples from the electrochemical series. The Daniell cell provides a good illustration of the principles under consideration here.	
19.2.1	Describe the standard hydrogen electrode. Laboratory work using the standard hydrogen electrode is not required.	2
19.2.2	Define the term <i>standard electrode potential</i> and explain the measurement of standard electrode potentials to produce the electrochemical series.	1, 3
19.2.3	Define the term <i>cell potential</i> and calculate cell potentials using standard electrode potentials.	1, 3
19.2.4	Predict whether a reaction will be spontaneous using standard electrode potential (E^\ominus) values. Students should be able to predict the direction of electron flow in an external circuit and the reaction taking place in a cell. Relate positive E^\ominus values for spontaneous reactions to negative ΔG^\ominus values (see 6.6).	3
	19.3 Electrolysis (2h)	
19.3.1	List and explain the factors affecting the products formed in the electrolysis of aqueous solutions. Factors to be considered are position in the electrochemical series, nature of the electrode and concentration. Suitable examples for electrolysis include water, aqueous sodium chloride and aqueous copper(II) sulfate.	1, 3
19.3.2	List the factors affecting the amount of product formed during electrolysis. Factors are charge on the ion, current and duration of electrolysis.	1
19.3.3	Determine the relative amounts of the products formed during the electrolysis of aqueous solutions.	3

Topic 20: Organic Chemistry

A.S.

Obj

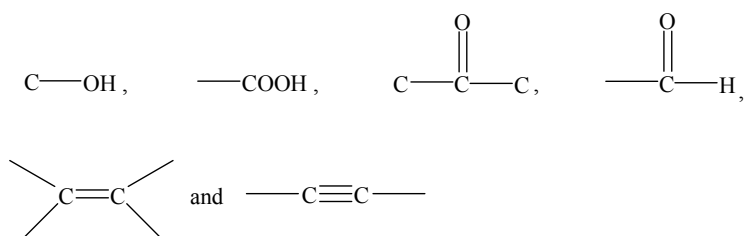
20.1 Determination of Structure (4h)

20.1.1 State that the structure of a compound can be determined using information from a variety of spectroscopic and chemical techniques. **1**

Students should realize that information from only one technique is usually insufficient to determine or confirm a structure.

20.1.2 Describe and explain how information from an infrared spectrum can be used to identify functional groups in a compound. **2, 3**

Restrict this to using infrared spectra to show the presence of the functional groups:



and to match the fingerprint region to a known spectrum.

20.1.3 Describe and explain how information from a mass spectrum can be used to determine the structure of a compound. **2, 3**

Restrict this to using mass spectra to determine the relative molecular mass of a compound and to identify simple fragments, for example:

- $(M_r - 15)^+$ loss of CH_3
- $(M_r - 29)^+$ loss of C_2H_5 or CHO
- $(M_r - 31)^+$ loss of CH_3O
- $(M_r - 45)^+$ loss of COOH .

20.1.4 Describe and explain how information from a ^1H NMR spectrum can be used to determine the structure of a compound. **2, 3**

Restrict this to using NMR spectra to determine the number of different environments in which hydrogen is found and the number of hydrogen atoms in each environment. Splitting patterns are not required.

20.2 Hydrocarbons (2h)

20.2.1 State and explain the low reactivity of alkanes in terms of the inertness of C–H and C–C bonds. **1, 3**

A.S.		Obj
20.2.2	<p>State that alkanes can react with halogens and distinguish between <i>homolytic</i> and <i>heterolytic fission</i>.</p> <p>Students should be able to define and recognize a <i>free radical</i>. Mechanisms are not required.</p>	1, 2
20.2.3	<p>Describe and explain the structure of benzene using chemical and physical evidence.</p> <p>Consider the special stability of the ring system (heat of combustion or hydrogenation of C₆H₆ in comparison to that of cyclohexene, cyclohexadiene and cyclohexatriene), as well as benzene's tendency to undergo substitution rather than addition reactions.</p>	2, 3

20.3 Nucleophilic Substitution Reactions (2h)

20.3.1	Distinguish between <i>primary</i> , <i>secondary</i> and <i>tertiary halogenoalkanes</i> .	2
20.3.2	<p>Describe and explain the S_N1 and S_N2 mechanisms in nucleophilic substitution.</p> <p>Students must be able to draw a stepwise mechanism. Examples of nucleophiles should include -CN, -OH and NH₃ for each reaction type.</p>	2, 3
20.3.3	<p>Describe and explain the molecularity for the S_N1 and S_N2 mechanisms.</p> <p>The predominant mechanism for tertiary halogenoalkanes is S_N1 and for primary halogenoalkanes it is S_N2. Both mechanisms occur for secondary halogenoalkanes.</p>	2, 3
20.3.4	Describe how the rate of nucleophilic substitution in halogenoalkanes depends on both the identity of the halogen and whether the halogenoalkane is primary, secondary or tertiary.	2

20.4 Alcohols (1h)

20.4.1	Describe the dehydration reaction of alcohols to form alkenes.	2
20.4.2	Determine the products formed by the oxidation of primary, secondary and tertiary alcohols using acidified potassium dichromate(VI) solution.	3

Options Outline

Options Standard Level

- A Higher physical organic chemistry

Options Standard Level/Higher Level

- B Medicines and drugs
- C Human biochemistry
- D Environmental chemistry
- E Chemical industries
- F Fuels and energy

Note: **Standard level candidates are required to study the core of each option chosen. Higher level candidates study both the core and the extension material.**

Options Higher Level

- G Modern analytical chemistry
- H Further organic chemistry

Standard level candidates are required to study any **two** options from A–F.
The duration of each option is 15 hours.

Higher level candidates are required to study any **two** options from B–H.
The duration of each option is 22 hours.

Option A: Higher Physical Organic Chemistry

This option builds on some of the key ideas in both physical and organic chemistry that were introduced in the core.

A.S. Obj

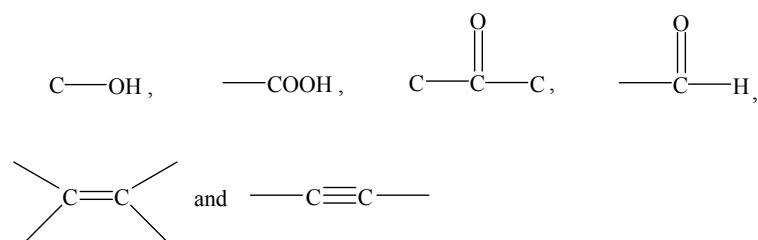
A.1 Determination of Structure (5h)

A.1.1 State that the structure of a compound can be determined by using information from a variety of spectroscopic and chemical techniques. 1

Students should realize that information from only one technique is usually insufficient to determine or confirm a structure.

A.1.2 Describe and explain how information from an infrared spectrum can be used to identify functional groups in a compound. 2, 3

Restrict this to using infrared spectra to show the presence of the functional groups:



and to match the fingerprint region to a known spectrum.

A.1.3 Describe and explain how information from a mass spectrum can be used to determine the structure of a compound. 2, 3

Restrict this to using mass spectra to determine the relative molecular mass of a compound and to identify simple fragments, for example:

- $(M_r - 15)^+$ loss of CH_3
- $(M_r - 29)^+$ loss of C_2H_5 or CHO
- $(M_r - 31)^+$ loss of CH_3O
- $(M_r - 45)^+$ loss of COOH .

A.1.4 Describe and explain how information from a ^1H NMR spectrum can be used to determine the structure of a compound. 2, 3

Restrict this to using NMR spectra to determine the number of different environments in which hydrogen is found and the number of hydrogen atoms in each environment. Splitting patterns are not required.

A.1.5 Describe and explain the structure of benzene using chemical and physical evidence. 2, 3

Consider the special stability of the ring system (heat of combustion or hydrogenation of C_6H_6 in comparison to that of cyclohexene, cyclohexadiene and cyclohexatriene), as well as benzene's tendency to undergo substitution rather than addition reactions.

A.S.		Obj
	A.2 Rate Expression (3h)	
A.2.1	Define the terms <i>rate constant</i> and <i>order of reaction</i> .	1
A.2.2	Derive the rate expression for a reaction from data. $\text{Rate} = k[\text{A}]^m[\text{B}]^n$ where k = rate constant, $[\text{A}]$ = concentration of A in mol dm ⁻³ etc, m and n = integers, $m + n$ = overall order of the reaction.	3
A.2.3	Draw and analyse graphical representations for zero-, first- and second-order reactions.	1, 3
A.2.4	Define the term <i>half-life</i> and calculate the half-life for first-order reactions only. The half-life should be calculated from graphs and by using the integrated form of the rate equation. The integrated rate equation for second-order reactions is not required.	1, 2
	A.3 Reaction Mechanism (1h)	
A.3.1	Define the terms <i>rate-determining step</i> , <i>molecularity</i> and <i>activated complex</i> .	1
A.3.2	Describe the relationship between mechanism, order, rate-determining step and activated complex. Limit examples to one- or two-step reactions where the mechanism is known. Students should understand what an activated complex (transition state) is and how the order of a reaction relates to the mechanism.	2
	A.4 Nucleophilic Substitution Reactions (2h)	
A.4.1	Distinguish between <i>primary</i> , <i>secondary</i> and <i>tertiary halogenoalkanes</i> .	2
A.4.2	Describe and explain the S _N 1 and S _N 2 mechanisms in nucleophilic substitution. Students must be able to draw a stepwise mechanism. Examples of nucleophiles should include -CN, -OH and NH ₃ for each reaction type.	2, 3
A.4.3	Describe and explain the molecularity for the S _N 1 and S _N 2 mechanisms. The predominant mechanism for tertiary halogenoalkanes is S _N 1 and for primary halogenoalkanes it is S _N 2. Both mechanisms occur for secondary halogenoalkanes.	2, 3
A.4.4	Describe how the rate of nucleophilic substitution in halogenoalkanes depends on both the identity of the halogen and whether the halogenoalkane is primary, secondary or tertiary.	2

A.S.		Obj
	A.5 Acids, Bases and Buffers (4h)	
A.5.1	State the expression for the ionic product constant of water (K_w). $K_w = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$ at 298 K, but this varies with temperature.	1
A.5.2	Deduce $[\text{H}^+(\text{aq})]$ and $[\text{OH}^-(\text{aq})]$ for water at different temperatures given K_w values.	3
A.5.3	Define pH , pOH and pK_w .	1
A.5.4	Calculate $[\text{H}^+(\text{aq})]$, $[\text{OH}^-(\text{aq})]$, pH and pOH from specified concentrations. The values of $[\text{H}^+(\text{aq})]$ or $[\text{OH}^-(\text{aq})]$ are directly related to the concentration of the acid or base.	2
A.5.5	State the equation for the reaction of any weak acid or weak base with water and hence derive the ionization constant expression. In general $\text{HA}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{A}^-(\text{aq})$ $\text{B}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{BH}^+(\text{aq}) + \text{OH}^-(\text{aq})$ (base hydrolysis) Then $K_a = \frac{[\text{H}^+(\text{aq})][\text{A}^-(\text{aq})]}{[\text{HA}(\text{aq})]}$ and $K_b = \frac{[\text{BH}^+(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{B}(\text{aq})]}$ Examples used should involve the transfer of only one proton.	1, 3
A.5.6	State and explain the relationship between K_a and pK_a .	1, 3
A.5.7	Determine the relative strengths of acids or their conjugate bases from K_a or pK_a values.	3
A.5.8	Apply K_a or pK_a in calculations. Calculations can be performed using various forms of the acid ionization constant expression (see A.5.5). Students should state when approximations are used in equilibrium calculations. Use of the quadratic expression is not required.	2
A.5.9	Calculate the pH of a specified buffer system. Calculations will involve the transfer of only one proton. Cross reference with 9.4.	2

Option B: Medicines and Drugs

The aim of this option is to give students an understanding of how drugs and medicines can influence the functioning of the body. Students should be able to recognize the fundamental structures and relevant functional groups of several classes of drugs and medicines (as listed below or in 11.3.1), and should be able to distinguish between them. Memorizing of complex formulas is not required. Throughout the option, stress the contribution that science has made (and continues to make) toward maintaining and improving the health and well-being of the world's population.

A.S.	<i>Core Material—SL and HL</i>	Obj
	B.1 Pharmaceutical Products (2h)	
B.1.1	<p>List the effects of drugs and medicines.</p> <p>Generally a drug or medicine is any chemical which does one or more of the following:</p> <ul style="list-style-type: none"> • alters incoming sensory sensations • alters mood or emotions • alters physiological state, including consciousness, activity level or coordination. <p>Stress the importance of the body's natural healing processes and the placebo effect.</p>	1
B.1.2	<p>Outline the stages involved in research, development and testing of new pharmaceutical products.</p> <p>Refer to the Thalidomide case as an example of what can go wrong. The use of combinatorial chemistry is not required here, but is covered in B.8.4.</p>	2
B.1.3	<p>Describe the different methods of administering drugs.</p> <p>The four main methods are oral, rectal, inhalation and parenteral (by injection). Injections may be intravenous, intramuscular or subcutaneous.</p>	2
B.1.4	<p>Discuss the terms <i>lethal dosage</i> (LD_{50}), <i>tolerance</i> and <i>side effects</i>.</p> <p>LD_{50} is the lethal dose required for 50% of the population.</p> <p>A person who develops tolerance requires a larger dose of a drug in order to achieve the effect originally obtained by a smaller dose. Stress that the difference between the main effect and side effects is relative. For example, morphine is often used as a pain killer with intestinal constipation being a side effect. For a person with diarrhoea the constipation induced becomes the main effect, with the pain relief a side effect. The risk:benefit ratio should be considered.</p>	3

A.S. *Core Material—SL and HL* Obj

B.2 Antacids (1h)

B.2.1 State and explain how excess acidity in the stomach can be reduced by the use of different bases. **1, 3**

Examples should include aluminium and magnesium compounds and sodium hydrogencarbonate. Students should be able to write balanced equations for neutralization reactions and know that antacids are often combined with alginates (which produce a neutralizing layer preventing acid in the stomach from rising into the esophagus and causing heartburn), and with anti-foaming agents (such as dimethicone).

B.3 Analgesics (3h)

B.3.1 Describe and explain the different ways that analgesics prevent pain. **2, 3**

Mild analgesics function by intercepting the pain stimulus at the source, often by interfering with the production of substances (eg prostaglandins) that cause pain, swelling or fever. Strong analgesics work by temporarily bonding to receptor sites in the brain, preventing the transmission of pain impulses without depressing the central nervous system.

B.3.2 Describe the use of derivatives of salicylic acid as mild analgesics and compare the advantages and disadvantages of using aspirin and paracetamol (acetaminophen). **2**

Aspirin has been found to be useful in preventing the recurrence of heart attacks. The disadvantages of aspirin include ulceration and stomach bleeding, allergic reactions and Reye's syndrome in children (a potentially fatal liver and brain disorder). Paracetamol is very safe in the correct dose but can, rarely, cause blood disorders and kidney damage. Overdosage can lead to serious liver damage, brain damage and even death.

B.3.3 Compare the structures of morphine, codeine and the semi-synthetic opiate, heroin. **2**

Stress the simple modification to the structure of morphine which results in the semi-synthetic drug, heroin.

B.3.4 Discuss the advantages and disadvantages of using morphine and its derivatives as strong analgesics. **3**

Include the social as well as physiological effects of both short- and long-term use.

B.4 Depressants (3h)

B.4.1 Describe the effects of depressants. **2**

At low doses a depressant may exert little or no effect. At moderate doses the compound may induce sedation (soothing, reduction of anxiety). At higher doses it may induce sleep and at extremely high doses it may cause death. Depressants are often described as anti-depressants because they relieve depression.

A.S.	<i>Core Material—SL and HL</i>	Obj
B.4.2	Discuss the social and physiological effects of the use and abuse of ethanol. Include effects on the family, cost to society and the short- and long-term health effects.	3
B.4.3	Describe and explain the techniques used for the detection of ethanol in the breath and in the blood or urine. Include potassium dichromate(VI) in the breathalyser, analysis of blood or urine by chromatography and absorption of infra-red radiation in the intoximeter.	2, 3
B.4.4	Describe the synergistic effects of ethanol with other drugs. Examples include increased risk of stomach bleeding with aspirin, and increased risk of heavy sedation with any drug that has a sedative effect on the central nervous system.	2
B.4.5	List other commonly used depressants and describe their structures. Limit this to a brief mention of the use of diazepam (Valium®), nitrazepam (Mogadon®) and fluoxetine hydrochloride (Prozac®).	1, 2

B.5 Stimulants (2.5h)

B.5.1	List the physiological effects of stimulants.	1
B.5.2	Compare amphetamines and adrenaline. Amphetamines and adrenaline are chemically similar in that both derive from the phenylethylamine structure. Amphetamines mimic the effects of adrenaline and are known as sympathomimetic drugs.	2
B.5.3	Discuss the short- and long-term effects of nicotine consumption. <ul style="list-style-type: none"> • Short-term effects: increased heart rate and blood pressure and reduction in urine output, as well as stimulating effects. • Long-term effects: increased risk of heart disease, coronary thrombosis and peptic ulcers. Discuss also the addictive properties of nicotine and the further risks associated with smoking tobacco. 	3
B.5.4	Describe the effects of caffeine and compare its structure with that of nicotine. Caffeine is a respiratory stimulant. When consumed in large amounts it can cause anxiety, irritability and sleeplessness. It is a weak diuretic. Both caffeine and nicotine contain a tertiary amine group.	2

B.6 Antibacterials (2h)

B.6.1	Outline the historical development of penicillins. Include the discovery by Fleming and the development by Florey and Chain.	2
B.6.2	Compare <i>broad-spectrum</i> and <i>narrow-spectrum</i> antibiotics.	2

A.S.	<i>Core Material—SL and HL</i>	Obj
B.6.3	Explain how penicillins work and discuss the effects of modifying the side chain. Penicillins work by interfering with the chemicals that bacteria need to form normal cell walls. Modifying the side chain results in penicillins which are more resistant to the penicillinase enzyme.	3
B.6.4	Discuss and explain the effect overprescription of penicillins has, and the use of penicillins in animal feedstock.	3
B.7 Antivirals (1.5h)		
B.7.1	State how viruses differ from bacteria.	1
B.7.2	Describe the different ways in which antiviral drugs work. Antiviral drugs may work by altering the cell's genetic material so that the virus cannot use it to multiply. Alternatively they may prevent the viruses from multiplying by blocking enzyme activity within the host cell.	2
B.7.3	Discuss the difficulties associated with solving the AIDS problem. Specific proteins on the HIV virus bind to a receptor protein on certain white blood cells (T cells). Because of the ability of the HIV viruses to mutate and because their metabolism is linked closely with that of the cell, effective treatment with antiviral drugs is very difficult, as is vaccine development.	3

A.S. *Extension Material—HL only* Obj

B.8 Stereochemistry in Drug Action and Design (3h)

- B.8.1** Describe the importance of geometrical isomerism in drug action. **2**
 Students should be aware that *cis*- and *trans*-isomerism can occur in inorganic complexes and that the two different isomers can have different pharmacological effects. The anti-cancer drug cisplatin is a good example.
- B.8.2** Discuss the importance of chirality in drug action. **3**
 The two enantiomers in a racemic mixture of a drug may have very different effects, eg Thalidomide. One enantiomer of Thalidomide alleviates morning sickness in pregnant women, whilst the other enantiomer causes deformities in the limbs of the fetus.
- B.8.3** Describe the use of chiral auxiliaries to form the desired enantiomer. **2**
 A chiral auxiliary is used to convert a non-chiral molecule into just the desired enantiomer, thus avoiding the need to separate enantiomers from a racemic mixture. It works by attaching itself to the non-chiral molecule to create the stereochemical conditions necessary to force the reaction to follow a certain path. Once the new molecule has been formed the auxiliary can be taken off (recycled) to leave the desired enantiomer. An example is the synthesis of Taxol, an anti-cancer drug.
- B.8.4** Explain the use of combinatorial chemistry to synthesize new drugs. **3**
 Combinatorial chemistry is used to synthesize a large number of different compounds and screen them for biological activity, resulting in a “combinatorial library” (for example the “mix and split” process whereby polypeptides can be made by every combination of amino acids, using polystyrene resin beads). Stress the importance of solid phase chemistry.

B.9 Anesthetics (2h)

- B.9.1** Compare local and general anesthetics in terms of their mode of action. **2**
- B.9.2** Compare the structures and effects of cocaine, procaine and lidocaine. **2**
- B.9.3** Discuss the advantages and disadvantages of nitrous oxide, ethoxyethane, trichloromethane, cyclopropane and halothane. **3**
 Nitrous oxide is not very potent, trichloromethane leads to liver damage, ethoxyethane and cyclopropane are highly flammable. Halothane (2-bromo-2-chloro-1,1,1-trifluoroethane) is widely used but is potentially harmful to the ozone layer.
- B.9.4** Calculate the partial pressures of component gases in an anesthetic mixture. **2**
 Knowledge of how to use Dalton's law of partial pressures is required. Students are not expected to state the law.

A.S.	<i>Extension Material—HL only</i>	Obj
B.10 Mind-altering drugs (2h)		
B.10.1	Describe the effects of lysergic acid diethylamide (LSD), mescaline, psilocybin and tetrahydrocannabinol (THC).	2
B.10.2	Discuss the structural similarities and differences between LSD, mescaline and psilocybin. Stress the similarity of all three drugs and compare them to the indole ring.	3
B.10.3	Discuss the arguments for and against the legalization of cannabis. Arguments for legalization include the ability of cannabis to offer relief for certain diseases. Arguments against legalization include the possible harmful effects and the possibility of cannabis users moving on to harder drugs.	3

Option C: Human Biochemistry

The aim of this option is to give students an understanding of the chemistry of important molecules found in the human body and the need for a balanced and healthy diet. Although the role these molecules play in the body should be appreciated, the emphasis is placed on their chemistry, and students who have not followed a course in biology will not be at a disadvantage. Students will not be required to memorize complex structures but will be expected to recognize functional groups and types of bonding within molecules. Structures of some important biological molecules are given in the data booklet.

A.S.	<i>Core Material—SL and HL</i>	Obj
	C.1 Diet (2h)	
C.1.1	Describe what the human body requires for a healthy diet. Students should recognize the importance of a balanced diet, including minimum requirements and the need for essential minerals.	2
C.1.2	Calculate the calorific value of a food from enthalpy of combustion data.	2
C.1.3	Discuss the benefits and concerns of using genetically modified (GM) foods. Crops and animals can be modified to provide more food, be more resistant to disease and be more tolerant to heavy metals. Concerns include the release of genetically modified organisms into the environment where they could spread and compete with the naturally occurring varieties.	3
	C.2 Proteins (3h)	
C.2.1	State the basic structure of 2-amino acids. There are approximately 20 common 2-amino acids (α -amino acids) found in organisms. 2-amino acids have the following formula.	1
	$ \begin{array}{c} \text{H} \\ \\ \text{H}_2\text{N} - \text{C} - \text{COOH} \\ \\ \text{R} \end{array} $	
C.2.2	Describe the condensation reaction of amino acids to form polypeptides.	2
C.2.3	Explain how proteins can be analysed by chromatography and electrophoresis. To use either of these techniques the peptide bonds in the proteins must first be hydrolysed to release individual amino acids. Include the use of R_f values in paper chromatography. Given isoelectric points, students should be able to determine a suitable pH to achieve good separation in electrophoresis.	3

A.S.	<i>Core Material—SL and HL</i>	Obj
C.2.4	Describe and explain the primary, secondary, tertiary and quaternary structure of proteins.	2, 3
C.2.5	List the major functions of proteins in the body. These are structure, biological catalysts (enzymes) and energy sources.	1
C.3 Carbohydrates (2.5h)		
C.3.1	Describe the structural features of monosaccharides. Monosaccharides contain a carbonyl group (C=O) and at least two -OH groups, and have the empirical formula CH_2O .	2
C.3.2	Describe the straight-chain formula of glucose and the structural difference between α -glucose and β -glucose.	2
C.3.3	Describe the condensation of monosaccharides to form disaccharides and polysaccharides. Limit examples to: <ul style="list-style-type: none">• disaccharides—lactose and sucrose• polysaccharides—starch.	2
C.3.4	List the major functions of polysaccharides in the body. These are energy sources, energy reserves (eg glycogen) and precursors for other biologically important molecules.	1
C.4 Fats (2.5h)		
C.4.1	Describe the composition of fats and oils.	2
C.4.2	Describe the difference in structure between saturated and unsaturated fats, and explain the difference in their melting points.	2, 3
C.4.3	Calculate the number of C=C double bonds in an unsaturated fat using addition reactions. The number of C=C bonds can be determined from the number of moles of I_2 which add to one mole of fat.	2
C.4.4	Describe the hydrolysis of fats to form soaps and the action of soaps.	2
C.4.5	List the major functions of fats in the body. These are energy sources, insulation and cell membranes.	1

A.S. *Core Material—SL and HL* Obj

C.5 Vitamins (2.5h)

- C.5.1** Define the term *vitamin*. 1
- C.5.2** Deduce whether a vitamin is water or fat soluble from its structure. 3
- C.5.3** Describe the structures and major functions of retinol (vitamin A), calciferol (vitamin D) and ascorbic acid (vitamin C). 2
- Vitamin A—required for the production of rhodopsin (light-sensitive material in the rods of the retina). Deficiency can cause night blindness and xerophthalmia.
 - Vitamin D—required for the uptake of calcium from food. Deficiency can cause weak bones (rickets).
 - Vitamin C—essential in the production of collagen: the protein of connective tissue. Deficiency can cause scorbutus (scurvy).
- C.5.4** Describe the effects of food processing on the vitamin content of food. 2
- Most vitamins are unstable at higher temperatures so will be affected by prolonged cooking.

C.6 Hormones (2.5h)

- C.6.1** Outline the production and roles of hormones in the body. 2
- Hormones are chemical messengers produced in glands controlled by the pituitary gland, which in turn is controlled by the hypothalamus. Limit examples of production and roles to adrenalin, thyroxine, insulin and sex hormones.
- C.6.2** Compare the structures of cholesterol and the sex hormones. 2
- Stress the common steroid backbone but the difference in functional groups (see the data booklet).
- C.6.3** Describe the mode of action of oral contraceptives. 2
- C.6.4** Outline the use and abuse of steroids. 2

A.S.	<i>Extension Material—HL only</i>	Obj
C.7 Enzymes (3h)		
C.7.1	Outline the basic characteristics of enzymes. Include: <ul style="list-style-type: none"> • enzymes are proteins • activity depends on tertiary and quaternary structure • the specificity of enzyme action. 	2
C.7.2	Determine V_{\max} and the value of the Michaelis constant (K_m) by graphical means.	3
C.7.3	Describe the significance of V_{\max} and K_m .	2
C.7.4	Describe the concept of the active site in enzyme structure.	2
C.7.5	Explain competitive inhibition and non-competitive inhibition.	3
C.7.6	State and explain the effects of heavy metal ions, extremes of temperature and pH changes on enzyme activity.	1, 3
C.7.7	Describe the uses of enzymes in biotechnology. Possible examples include proteases in biological detergents, glucose isomerase converting glucose to fructose and streptokinase in breaking down blood clots.	2
C.8 Nucleic Acids (2h)		
C.8.1	Describe the structure of nucleotides and their condensation polymers (nucleic acids). A nucleotide contains a phosphate group, a pentose sugar group and an organic base. Students should be able to recognize, but need not recall, the structures of the five nucleotide bases: adenine, cytosine, guanine, thymine and uracil.	2
C.8.2	Describe the double helical structure of DNA. Students should be able to describe the hydrogen bonding between specific pairs of nucleotide bases.	2
C.8.3	Outline the role of DNA as the repository of genetic information, including the triplet code.	2
C.8.4	Describe the principles and uses of DNA profiling. Include forensic uses and paternity cases.	2

A.S.	<i>Extension Material—HL only</i>	Obj
C.9 Metal Ions in Biological Systems (2h)		
C.9.1	Explain that different metal ions fulfill different roles in the body due to their different chemical properties. Emphasize differences in charge density, redox properties and complex ion formation.	3
C.9.2	Describe the importance of the difference in Na ⁺ and K ⁺ concentrations across the cell membrane. Explain active transport using the Na ⁺ /K ⁺ pump as an example.	2
C.9.3	Outline the importance of copper ions in electron transport and iron ions in oxygen carriers. Use cytochromes and hemoglobin as examples.	2

Option D: Environmental Chemistry

The effect of human activity on the environment has become increasingly global, with the effects of chemicals in air and water spanning political and natural borders. An understanding of this impact is essential within and beyond the study of chemistry.

A.S. *Core Material—SL and HL* **Obj**

D.1 Primary Air Pollution (3h)

D.1.1 Describe the sources of carbon monoxide, oxides of nitrogen and sulfur, particulates and hydrocarbons in the atmosphere. **2**

Include both natural and man-made sources. Balanced equations should be used where possible.

D.1.2 Outline the effects of primary air pollution on health. **2**

Students should be familiar with at least one harmful effect of each of the substances in D.1.1.

D.1.3 Discuss methods for the reduction of primary air pollution. **3**

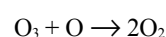
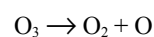
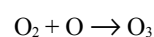
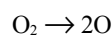
Limit this to the following methods:

- CO—catalytic converters
- NO_x—catalytic converters, lean burn engines, recirculation of exhaust gases
- SO_x—alkaline scrubbing, removal of sulfur-containing compounds from coal and oil, limestone-based fluidized beds
- Particulates—electrostatic precipitation
- Hydrocarbons—catalytic converters.

D.2 Ozone Depletion (2h)

D.2.1 Describe the formation and depletion of ozone by natural processes. **2**

Refer to the following equations.



D.2.2 List the pollutants, and their sources, that cause the lowering of ozone concentration. **1**

Consider chlorofluorocarbons (CFCs) and nitrogen oxides.

D.2.3 State the environmental effects of ozone depletion. **1**

Include the increased incidence of skin cancer and eye cataracts, and the suppression of plant growth.

A.S. *Core Material—SL and HL* **Obj**

- D.2.4** Discuss the alternatives to CFCs in terms of their properties. **3**
 Alternatives include hydrocarbons, fluorocarbons and hydrofluorocarbons (HFCs). Include toxicity, flammability, the relative weakness of the C–Cl bond and the ability to absorb infrared radiation.

D.3 Greenhouse Effect and Global Warming (2h)

- D.3.1** Describe the *greenhouse effect*. **2**
 Greenhouse gases allow the passage of incoming solar radiation but absorb the heat radiation from the Earth, maintaining a mean global temperature. The greenhouse effect is a normal and necessary condition for life on Earth.

- D.3.2** List the main greenhouse gases and their sources, and discuss their relative effects. **1, 3**
 The greenhouse gases to be considered are CH₄, H₂O, CO₂ and N₂O, which have natural and man-made origins. Their effects depend on their abundance and their ability to absorb heat radiation.

- D.3.3** Discuss the influence of increasing amounts of greenhouse gases on global warming. **3**
 Effects include climate change, thermal expansion of the oceans and melting of the polar ice caps.

- D.3.4** Outline the influence of particulates on the Earth's surface temperature. **2**
 Particulates can lower the temperature by reflecting sunlight.

D.4 Acid Rain (1.5h)

- D.4.1** State what is meant by *acid rain* and outline its origins. **1, 2**
 Rain is naturally acidic because of dissolved CO₂; acid rain has a pH of less than 5.6.
 Acid rain is caused by oxides of sulfur and nitrogen. Students should know the equations for the burning of sulfur and nitrogen and for the formation of H₂SO₃ and H₂SO₄.

- D.4.2** Discuss the environmental effects of acid rain and possible methods to counteract them. **3**

D.5 Water Suitable for Drinking (3h)

- D.5.1** Discuss the demand for fresh water and reasons for the inadequacy of its supply. **3**
 Only a small fraction of the Earth's water supply is fresh water. Of this fresh water, over 80% is in the form of ice caps and glaciers. Water is mainly used for agriculture and industry.

A.S.	<i>Core Material—SL and HL</i>	Obj
D.5.2	Compare the advantages and disadvantages of treating drinking water with chlorine and ozone. Include cost, retention time and formation of chlorinated organic compounds.	2
D.5.3	Discuss ways to obtain fresh water from sea water using distillation, reverse osmosis and ion exchange.	3
D.5.4	Discuss ways to reduce the amount of water used and to recycle water.	3
D.6 Dissolved Oxygen in Water (2h)		
D.6.1	Outline the importance of dissolved oxygen in water.	2
D.6.2	Outline biological oxygen demand (BOD) as a measure of oxygen-demanding wastes in water. Refer to the amount of oxygen needed to decompose waste matter over a definite period of time. No distinction between biological and biochemical oxygen demand will be made.	2
D.6.3	Distinguish between aerobic and anaerobic decomposition of organic material in water.	2
D.6.4	Describe the influence of sewage, detergents and fertilizers on the growth of aquatic plants, and the effect of their subsequent decomposition on oxygen concentration (eutrophication). The additional nitrogen and phosphorus compounds encourage growth of aquatic plants often in the form of “algal blooms” or, in coastal areas, “red tides”.	2
D.6.5	Discuss the effect of heat on dissolved oxygen and metabolism in water.	3
D.7 Waste Water Treatment (1.5h)		
D.7.1	Outline the primary and secondary stages of sewage treatment and state what is removed during each stage. For primary treatment filtration, flocculation and sedimentation should be covered. For secondary treatment mention the use of oxygen and bacteria (eg the activated sludge process).	2
D.7.2	Discuss the increasing use of tertiary treatment. Include removal of heavy metals and phosphates by chemical precipitation and nitrates by chemical or biological processes.	3

A.S. *Extension Material—HL only* **Obj**

D.8 Smog (2h)

- D.8.1** Compare reducing and photochemical smog. **2**
- D.8.2** Describe the catalytic effect of particulates and nitrogen oxides on the oxidation of sulfur dioxide. **2**
- Particulates and SO₂—heterolytic catalysis to form SO₃
 - NO_x + SO₂—free radical catalysis to form SO₃
- D.8.3** Outline the formation of secondary pollutants in photochemical smog. **2**
- Treatment should be restricted to the formation of radicals from the reaction of nitrogen oxides with sunlight and the reaction of these radicals with hydrocarbons, leading to the formation of aldehydes and peroxyacylnitrates (PANs).
- D.8.4** Discuss the formation of thermal inversions and their effects on air quality. **3**

D.9 Ozone Depletion (2h)

- D.9.1** Explain the dependence of O₂ and O₃ dissociation on the wavelength of light. **3**
- $\lambda = 242 \text{ nm}$ $\lambda = 330 \text{ nm}$
 $\text{O}_2 \rightarrow 2\text{O}$ $\text{O}_3 \rightarrow \text{O}_2 + \text{O}$
- The energy needed should be related to the bonding in O₂ and O₃.
- D.9.2** Describe the steps in the catalysis of O₃ depletion by CFCs and NO_x. **2**
- For example:
- $\text{CCl}_2\text{F}_2 \rightarrow \text{CClF}_2 + \text{Cl}$
 $\text{Cl} + \text{O}_3 \rightarrow \text{ClO} + \text{O}_2$
 $\text{ClO} + \text{O} \rightarrow \text{O}_2 + \text{Cl}$
 NO_x similar pathway
- D.9.3** Outline the reasons for greater ozone depletion in polar regions. **2**
- Consider the seasonal variation in temperature in the upper atmosphere. Refer to surface catalysis on ice particles.
- D.9.4** Describe the properties required for sun-screening compounds. **2**
- Such compounds should contain conjugated double bonds, eg para-aminobenzoic acid (PABA), so that absorption of ultraviolet light is possible.

D.10 Toxic Substances in Water (3h)

- D.10.1** Discuss the different approaches to expressing toxicity. **3**
- Include the advantages and disadvantages of LD₅₀ (lethal dose in 50% of the population) and maximum daily tolerance.

A.S.	<i>Extension Material—HL only</i>	Obj
D.10.2	State the principal toxic types of chemicals that may be found in polluted water. Include heavy metals, pesticides, dioxins and polychlorinated biphenyls (PCBs).	1
D.10.3	Outline the sources, health and environmental effects of cadmium, mercury and lead compounds. <ul style="list-style-type: none">• Cadmium—metal plating, some rechargeable batteries, pigments• Mercury—seed dressing to prevent mould, batteries• Lead—some kinds of paint, as tetraethyl lead in gasoline	2
D.10.4	Describe the sources and possible health effects of nitrates in drinking water. Include the formation of carcinogenic nitrosamines and a possible link to the formation of nitrites leading to oxygen depletion in the body.	2

Option E: Chemical Industries

Chemical industries have a major place in the world economy. The industrial revolution, which commenced in the 18th century, was also a materials revolution that started with the large-scale extraction of iron. This continues to the present day with the extraction of other metals, the production of bulk chemicals such as fertilizers, the oil and plastics industries, and the speciality chemicals industry producing medicines, enzymes, catalysts and additives. When teaching this option, emphasize chemical reactions and their relevant equations.

A.S. *Core Material—SL and HL* **Obj**

E.1 Initial Overview (2h)

- | | | |
|--------------|---|----------|
| E.1.1 | Outline the abundance, occurrence and availability of sources of materials.
A qualitative picture only is intended. Include minerals (especially metals and their ores), petroleum and biotechnology. | 2 |
| E.1.2 | Identify the factors that influence the establishment of a chemical industry in a particular location.
Include sources of suitable raw materials, energy supply, transport to and from the site, availability of investment, skills and labour and existence of markets for the product. | 2 |
| E.1.3 | Outline the division of the industry into both intermediates and consumer products. | 2 |
| E.1.4 | State the increasing importance of biotechnology in chemical manufacture.
Examples can be found in the manufacture of insulin and vitamins. | 1 |

E.2 Principles of Extraction and Production of Metals (2h)

- | | | |
|--------------|---|----------|
| E.2.1 | Outline the principles used in the physical separation of materials. | 2 |
| E.2.2 | Discuss the chemical principles involved in the extraction of metals from their ores.
Relate this to chemical reactions based on the reactivity series (see 10.2), but also include factors such as chemical conversion (iron), electrolytic conversion (aluminium) and energy requirements. | 3 |

A.S. *Core Material—SL and HL* Obj

E.3 Iron and Aluminium (4h)

- E.3.1** State the main sources of iron. **1**
Include iron ores and scrap (recycled) iron.
- E.3.2** Explain the reactions that occur in the blast furnace. **3**
Include the role of coke, limestone and the formation of slag. The relevant equations should be considered.
- E.3.3** Explain the conversion of iron into steel using the basic oxygen converter. **3**
- E.3.4** Describe the properties and uses of steel as an alloy of iron. **2**
Consider carbon steels and alloy steels (including stainless steel).
- E.3.5** Discuss the production of aluminium by electrolysis of alumina in molten cryolite. **3**
Explain the need for cryolite as a solvent because of the very high melting point of Al_2O_3 . Account for the materials used in the construction of the cell and the choice of electrodes.
- E.3.6** Describe the main properties and uses of aluminium. **2**
Include the properties of the aluminium oxide coating and the resulting resistance to corrosion. Compare with the properties and uses of iron and steel.
- E.3.7** Discuss the environmental impact of iron and aluminium production. **3**
Include the effects of mining the ore, siting the plant, energy costs and recycling the metals.

E.4 The Oil Industry (4h)

- E.4.1** Outline the importance of oil as a source of chemical feedstock. **2**
Although only about 10% of the refined products of crude oil are used as chemical feedstock, it is still the most significant source of organic chemicals. Compare the use of oil as an energy source and as a chemical feedstock.
- E.4.2** Outline the removal of sulfur from crude oil. **2**
Refer to the need for this removal and the use of sulfur in the manufacture of sulfuric acid.
- E.4.3** Describe the fractional distillation of oil. **2**
Compare simple distillation with fractional distillation. Students should understand that the vapour phase is always richer in the more volatile component. No calculations using Raoult's law or Dalton's law are required.
- E.4.4** Describe cracking and its products. **2**
Include thermal cracking (both steam and catalytic) and hydrocracking.

A.S.	<i>Core Material—SL and HL</i>	Obj
E.4.5	Describe reforming processes and their products. Include isomerization, cyclization and aromatization. Reforming is important in some countries as a source of hydrogen for the Haber process (see 8.2.5).	2
E.4.6	State the uses of refinery products as feedstock for the organic chemical industry. Refinery products are used as raw materials in the manufacture of solvents, plastics, pesticides, food additives, pharmaceuticals, detergents, cosmetics and dyes.	12
 E.5 Polymers (3h) 		
Cross reference with 11.3.		
E.5.1	Describe how the properties of polymers depend on their structural features. Include the: <ul style="list-style-type: none"> • different amounts of branching in low- and high-density polyethene • different positions of the methyl groups in isotactic and atactic polypropene • formation of cross-links in phenol-methanal plastics (compare thermoplastics and thermosets). 	2
E.5.2	Describe ways of modifying the properties of polymers. Include the use of: <ul style="list-style-type: none"> • plasticizers in polyvinyl chloride • volatile hydrocarbons in the formation of expanded polystyrene • air in the manufacture of polyurethane foams. 	2
E.5.3	Discuss the advantages and disadvantages of polymer use. Consider strength, density, insulation, lack of reactivity, use of natural resources, disposal and biodegradability. Use polyethene, polyurethane foams, polyvinyl chloride and phenol-methanal plastics as examples.	3

A.S.	<i>Extension Material—HL only</i>	Obj
	E.6 Silicon (1.5h)	
E.6.1	Describe the extraction and purification of silicon. Include zone refining.	2
E.6.2	Compare the electrical conductivity of a semiconductor with that of metals and non-metals. Relate this to the ionization energies of semiconductors compared to metals and non-metals.	2
E.6.3	Explain the doping of silicon to produce n-type and p-type semiconductors. In p-type semiconductors, electron holes in the crystal are created by introducing a small percentage of a group 3 element (eg In, Ga). In n-type semiconductors inclusion of a group 5 element (eg As) provides extra electrons.	3
	E.7 Ellingham Diagrams (1.5h)	
E.7.1	Analyse Ellingham diagrams to predict the feasibility of reducing metal oxides. Refer also to the Ellingham diagram in the data booklet.	3
	E.8 Mechanisms in the Organic Chemicals Industry (2h)	
E.8.1	Compare and discuss the mechanisms of thermal and catalytic cracking. Thermal cracking involves a free-radical mechanism whereas catalytic cracking has an anionic mechanism.	2, 3
E.8.2	Describe the mechanism involved in the manufacture of low-density polyethene. This is a free-radical mechanism.	2
E.8.3	Outline the use of Ziegler–Natta catalysts in the manufacture of high-density polyethene. The mechanism is ionic but details are not required.	2
	E.9 The Chlor-alkali Industry (2h)	
E.9.1	Discuss the production of chlorine by the electrolysis of sodium chloride. Explain why an electrolytic process is required and why the diaphragm cell is preferable to the mercury cathode cell.	3

A.S.	<i>Extension Material—HL only</i>	Obj
E.9.2	Outline the importance of the products of this process. The process produces sodium hydroxide with chlorine and hydrogen as by-products.	2
E.9.3	Discuss the environmental impact of this process. Include reasons why the diaphragm cell has replaced the mercury-cathode cell in many parts of the world, and that knowledge of the effect on the ozone layer has led to reservations about the use of chlorine-containing solvents.	3

Option F: Fuels and Energy

The development of human society has been directly related to the ability to use and manipulate fuels for energy production. This option considers the chemical principles and environmental issues associated with the use of fossil fuels, and nuclear and solar energy.

A.S.	<i>Core Material—SL and HL</i>	Obj
	F.1 Energy Sources (1h)	
F.1.1	State desirable characteristics of energy sources. These include energy released at reasonable rates (neither too fast nor too slow) and minimal pollution.	1
F.1.2	Outline current and potential energy sources. Consider fossil fuels, nuclear (fission and fusion), electrochemical cells, solar energy and alternative sources (eg wind, tidal, geothermal).	2
	F.2 Fossil Fuels (4h)	
F.2.1	Describe the formation and characteristics of coal, oil and natural gas.	2
F.2.2	Determine and compare the enthalpies of combustion of coal, oil and natural gas. Calculations could be made using enthalpies of formation or from experimental data. Cross reference with 15.1.	3, 2
F.2.3	Outline the composition and characteristics of the crude oil fractions used for fuel. Students should have general, rather than specific, knowledge about the types of compounds found in each fraction, the boiling point range and the uses of the fractions.	2
F.2.4	Describe how the components of a hydrocarbon fuel relate to its octane rating. Octane rating is a measure of the ability of a fuel to resist “knocking” when burnt in a standard test engine. A fuel is rated relative to heptane (rating of 0) and 2,2,4-trimethylpentane (rating of 100). The role of lead additives in fuels and the role of aromatic compounds in unleaded fuels should be mentioned.	2
F.2.5	Explain the processes of coal gasification and liquification. Gasification produces synthesis gas and liquification produces liquid hydrocarbons. Relevant equations should be used. Advantages include the elimination of SO ₂ pollution and the ease of transportation. The main disadvantage is the energy cost of the processes.	3

A.S.	<i>Core Material—SL and HL</i>	Obj
F.2.6	Describe how the burning of fossil fuels produces pollutants. The primary pollutants are CO, CO ₂ , SO ₂ , NO _x , particulates (fly ash) and hydrocarbons.	2
F.2.7	Discuss the advantages and disadvantages of the different fossil fuels. Consider the cost of production and availability (reserves) as well as pollution.	3
F.3 Nuclear Energy (4h)		
F.3.1	Distinguish between nuclear reactions and chemical reactions. Emphasize that in nuclear reactions nuclei are converted to other nuclei, while in chemical reactions only valence electrons are involved and atoms do not change into other atoms.	2
F.3.2	Write balanced nuclear equations. Both the atomic number and mass number must be balanced.	2
F.3.3	Describe the nature of α , β and γ radiation. Compare the charge, mass, penetrating power and behaviour in an electric field.	2
F.3.4	State the concept of half-life. Half-life is independent of the amount of a radioactive sample.	1
F.3.5	Apply the concept of half-life in calculations. Restrict this to whole numbers of half-lives.	2
F.3.6	Compare nuclear fission and nuclear fusion.	2
F.3.7	Explain the functions of the main components of a nuclear power plant. Include the fuel, moderator, control rods, coolant and shielding. The materials used for the different components should be considered.	3
F.3.8	Discuss the differences between conventional power generation and nuclear reactors.	3
F.3.9	Discuss the concerns about safety in nuclear power plants. Consider the effects of: <ul style="list-style-type: none"> • escape of radioactive materials at various stages • radioactive contamination of ordinary materials • escape of sodium from a breeder reactor • loss of control of a nuclear reaction • nuclear waste. 	3

A.S. *Core Material—SL and HL* Obj

F.4 Solar Energy (3h)

- F.4.1** State how solar energy can be converted to other forms of energy. **1**
Include chemical energy (biomass), thermal energy (passive and active methods) and electricity generation (direct and indirect methods).
- F.4.2** Describe the role of photosynthesis in converting solar energy to other forms of energy. **2**
Products of photosynthesis are used for food, primary fuels and conversion to other fuels, eg ethanol. The equation for photosynthesis is required.
- F.4.3** Discuss how biomass can be converted to energy. **3**
Include:
 - direct combustion
 - combustion of waste materials from other processes
 - production of biogas
 - production and use of ethanol.Mention the advantages and disadvantages of each method. The equation for the formation of ethanol from glucose is required.
- F.4.4** Outline the principles of using solar energy for space heating. **2**
Examples should include storage of heat by water and rocks.
- F.4.5** Discuss the methods for converting solar energy into electricity. **3**
Include parabolic mirrors and photovoltaic cells. Consider the advantages and disadvantages of each method.

F5 Electrochemical Energy (3h)

- F.5.1** Explain the workings of lead–acid storage batteries and dry cell (zinc–carbon and alkaline) batteries. **3**
Include the relevant half-equations.
- F.5.2** Identify the factors that affect the voltage and power available from a battery. **2**
Voltage depends primarily on the nature of the materials used while power depends on their quantity.
- F.5.3** Explain how a hydrogen–oxygen fuel cell works. **3**
Include the relevant half-equations.

A.S. *Extension Material—HL only* Obj

F.6 Storage of Energy and Limits of Efficiency (1h)

F.6.1 Discuss the advantages and disadvantages of energy storage schemes. 3
 Include both pumped storage and conversion to hydrogen.

F.7 Nuclear Stability (2h)

F.7.1 Predict nuclear stability and mode of decay from neutron to proton ratios. 3
 Students should be familiar with the belt of stability in the graph of number of neutrons against number of protons for various stable nuclei.

F.7.2 Calculate the energy released in a nuclear reaction. 2

F.7.3 Define and determine *mass defect* and *nuclear binding energy*. 1, 3
 Nuclear binding energy is a quantitative measure of nuclear stability. The graph of nuclear binding energy per nucleon against mass number should be used to explain why the products are more stable than the reactants in both nuclear fission and nuclear fusion, and consequently why both processes are exothermic.

F.8 Radioactive Decay (2h)

F.8.1 Calculate the change in activity over a period of time. 2
 See the data booklet for the integrated form of the rate equation.

F.8.2 Describe the different types of nuclear waste, their characteristics and their sources. 2

F.8.3 Compare the storage and disposal methods for different types of nuclear waste. 2

F.9 Photovoltaics (2h)

F.9.1 State that silicon and germanium are semiconductors. 1

F.9.2 Compare the electrical conductivity of a semiconductor with the conductivity of metals and non-metals. 2
 Relate this to the ionization energies of semiconductors compared to metals and non-metals.

F.9.3 Explain the doping of silicon to produce n-type and p-type semiconductors. 3
 In p-type semiconductors, electron holes in the crystal are created by introducing a small percentage of a group 3 element (eg In, Ga). In n-type semiconductors inclusion of a group 5 element (eg As) provides extra electrons.

F.9.4 Describe how sunlight interacts with semiconductors. 2
 Photons interact with crystals to release electrons.

Option G: Modern Analytical Chemistry

Emphasize problem solving and using the information gained from one or more techniques throughout this option. Students should understand the chemical principles behind each analytical technique but are not expected to have a detailed knowledge of the instruments themselves.

A.S.		Obj
	G.1 Analytical Techniques (2h)	
G.1.1	State the reasons for using analytical techniques. Analytical techniques are used in structure determination, in analysis of composition of substances and to determine purity.	1
G.1.2	Outline the information that can be obtained from analytical techniques, singly or in combination. Students should be able to draw upon a range of contexts to illustrate the information obtained by using a technique or range of techniques. <ul style="list-style-type: none"> • Visible and ultraviolet (uv) spectroscopy—assaying of metal ions, organic structure determination and detection of drug metabolites • Infrared (ir) spectroscopy—organic structure determination, information on strengths of bonds, secondary structure of proteins and measuring degree of unsaturation of oils and fats • Mass spectrometry—organic structure determination and isotopic dating (eg ^{14}C dating) • Gas chromatography-mass spectrometry (GC-MS)—drug and food testing and forensic science 	2
	G.2 Principles of Spectroscopy (2h)	
G.2.1	Describe the electromagnetic spectrum. X-ray, uv, visible, ir and radio (including microwave) should be identified. Highlight the variation in wavelength, frequency and energy across the spectrum.	2
G.2.2	Distinguish between absorption and emission spectra and how each is produced. Cross reference with 2.2.1.	2
G.2.3	Describe the atomic and molecular processes in which absorption of energy takes place. Cross reference with 2.2. The description should cover vibrations, rotation and electronic transitions only.	2
G.2.4	Describe the operating principles of a double-beam infrared spectrometer. A schematic diagram of a simple double-beam spectrometer is sufficient. This example is chosen to illustrate the general principles of how spectrometers operate. Mention could be made of modern methods of processing signals by Fourier transformation.	2

A.S.		Obj
	G.3 Visible and Ultraviolet Spectroscopy (4h)	
G.3.1	Describe the factors that affect the colour of transition metal complexes. Cross reference with 13.2.6. The factors are the identity of the metal (eg Mn^{2+} , Fe^{2+}), oxidation number (eg Fe^{2+} , Fe^{3+}) and the identity of the ligand. Limit this to octahedral complexes in aqueous solution.	2
G.3.2	Describe the effect of different ligands on the splitting of the d orbitals in transition metal complexes. The ligands should be limited to NH_3 , H_2O and Cl^- .	2
G.3.3	State that organic molecules containing a double bond absorb ultraviolet radiation. Refer to conjugated and delocalized systems: arenes, alkenes and natural products, eg chlorophyll.	1
G.3.4	Describe the effect of the conjugation of double bonds in organic molecules on the wavelength of the absorbed light. Retinol and phenolphthalein are suitable examples.	2
G.3.5	Predict whether or not a particular molecule will absorb ultraviolet or visible radiation.	3
G.3.6	State the Beer–Lambert law. $\log_{10} \frac{I_0}{I} = \epsilon lc$	1
G.3.7	Construct a calibration curve and use the Beer–Lambert law to determine the concentration of an unknown solution.	3
	G.4 Infrared Spectroscopy (3h)	
G.4.1	Describe what occurs at a molecular level during the absorption of infrared radiation by molecules. H_2O , $-\text{CH}_2-$, SO_2 and CO_2 are suitable examples. Stress the change in bond polarity as the vibrations (stretching and bending) occur.	2
G.4.2	State the relationship between wavelength and wavenumber. An inverse relationship exists (ie the wavenumber is the number of wavelengths that make up one cm). High wavenumber implies high energy.	1
G.4.3	Deduce the functional groups in an organic molecule from its infrared spectrum. Examples should contain up to three functional groups. Students are not required to learn the characteristic absorption frequencies of functional groups, but must be familiar with the relevant information in the data booklet. The precise wavenumber of the absorption depends upon neighbouring atoms.	3

A.S.		Obj
	G.5 Nuclear Magnetic Resonance (NMR) Spectroscopy (4h)	
G.5.1	State that atoms with an odd mass number can be detected by NMR spectroscopy.	1
G.5.2	Analyse simple NMR spectra. The emphasis is on ^1H spectra. Interpretation should include the: <ul style="list-style-type: none">• number of peaks• chemical shift (with tetramethylsilane (TMS) as the reference standard)• area under each peak• splitting patterns (treatment of spin-spin coupling constants is not required, but students should be familiar with splitting patterns for simple molecules ie doublet, triplet and quartet).	3
G.5.3	Outline how NMR is used in body scanners. Protons in water in human cells can be detected by magnetic resonance imaging (MRI), giving a three-dimensional view of organs in the human body.	2
	G.6 Mass Spectrometry (3h)	
G.6.1	Discuss how the molecular mass and molecular formula of a compound may be obtained from the molecular ion peak. Spectrometers have sufficient accuracy to allow identification of the molecular formula from the molecular mass using the masses of the commonest isotopes of C, H, N and O.	3
G.6.2	Analyse molecular mass spectra. Stress the importance of isotopes and relate these to the $(M + 1)^+$ peak for ^{13}C and the $(M + 2)^+$ and $(M + 4)^+$ peaks for chlorine and bromine. Include recognition of molecular fragments (see 20.1.3).	3
	G.7 Chromatography (4h)	
G.7.1	State the reasons for using chromatography. Chromatography can be used to separate substances for analysis and to determine purity. Highlight the coupling of chromatography with other techniques.	1
G.7.2	State that all chromatographic techniques require a stationary phase and a mobile phase. Components in a mixture have different tendencies to adsorb onto a surface or dissolve in a solvent. This provides a means of separating the components of a mixture.	1

A.S.		Obj
G.7.3	Explain how the phenomena of adsorption and partition can be used in chromatographic techniques. Each of these phenomena gives rise to different chromatographic techniques. Molecular exclusion is not required.	3
G.7.4	Outline the use of paper chromatography, thin-layer chromatography (TLC), column chromatography, gas-liquid chromatography (GLC) and high performance liquid chromatography (HPLC). An outline of the operation for each technique is all that is required. This should include an understanding of R_f values where relevant.	2
G.7.5	Deduce which chromatographic technique is most appropriate for separating the components in a particular mixture.	3

Option H: Further Organic Chemistry

A.S.		Obj
	H.1 Stereoisomerism (3h)	
H.1.1	Describe and explain geometrical isomerism in non-cyclic alkenes. The existence of geometric (<i>cis-trans</i>) isomers is the result of restricted rotation around the C=C bond.	2, 3
H.1.2	Explain the difference in physical and chemical properties of geometrical isomers. Include different boiling points, eg <i>cis</i> -1,2-dichloroethene and <i>trans</i> -1,2-dichloroethene and different reactions when heated, eg <i>cis</i> - and <i>trans</i> -but-2-ene-1,4-dioic acid.	3
H.1.3	Describe geometrical isomerism in C ₃ and C ₄ cyclo-alkanes. Use dichloro-derivatives of cyclopropane and cyclobutane as examples. Rotation is restricted because the C-C bond is now part of a cyclic system.	2
H.1.4	Define <i>plane-polarized light</i> and describe how it interacts with enantiomers. Cross reference with 11.3.3. Include the use of a polarimeter.	1, 2
H.1.5	Define the term <i>racemic mixture</i> .	1
H.1.6	Compare the physical and chemical properties of enantiomers.	2
	H.2 Free Radical Substitution Reactions (3h)	
H.2.1	Describe the gas phase reactions of alkanes and methylbenzene with halogens. Consider the free radical mechanism for the reactions of methane and methylbenzene with chlorine, and identify the initiation, propagation and termination steps.	2
H.2.2	Describe how the gas phase reactions of chloroalkanes affect the level of ozone in the atmosphere.	2
	H.3 Electrophilic Addition Reactions (4h)	
H.3.1	Describe and explain the electrophilic addition reactions of symmetrical alkenes. A stepwise mechanistic approach is required. Reacting species should include halogens, mixed halogens and hydrogen halides.	2, 3
H.3.2	Apply Markovnikov's rule to predict the outcome of the electrophilic addition reactions of asymmetrical alkenes. A stepwise mechanistic approach is required. Reacting species should include mixed halogens and hydrogen halides.	2

A.S.		Obj
H.3.3	State and explain the relative stabilities of carbocations in order to account for Markovnikov's rule. Consider the stability of primary, secondary and tertiary carbocations, eg CH_3^+ , $(\text{CH}_3)_2\text{CH}^+$ and $(\text{CH}_3)_3\text{C}^+$.	1, 3
H.4 Electrophilic Substitution Reactions (4h)		
H.4.1	Describe and explain the mechanism for the nitration of benzene. A stepwise electrophilic substitution mechanism is required. The formation of NO_2^+ should also be shown.	2, 3
H.4.2	Describe and explain the chlorination and alkylation reactions of benzene and methylbenzene in the presence of a halogen carrier.	2, 3
H.4.3	Describe and explain the directing effects and relative rates of reaction of different substituents on a benzene ring. Examples should be restricted to $-\text{CH}_3$, $-\text{OH}$, $-\text{Cl}$, $-\text{NO}_2$ and $-\text{CO}_2\text{CH}_3$. The reaction of phenol with chlorine to form trichlorophenol (TCP) should be covered.	2, 3
H.5 Nucleophilic Addition Reactions (1h)		
H.5.1	Describe and explain the mechanism for the addition of hydrogen cyanide to aldehydes and ketones, followed by hydrolysis to give carboxylic acids.	2, 3
H.6 Nucleophilic Substitution Reactions (2h)		
H.6.1	Outline how the relative rate of nucleophilic substitution is affected by different nucleophiles. Consider, for example, the relative rate of reaction using hydroxide or water (polarity differences).	2
H.6.2	Describe and explain inductive and steric effects of substituents on substitution reaction.	2, 3
H.6.3	Describe and explain the relative rates of hydrolysis of halogenated benzene compounds. Compare the inertness of halogenated benzene compounds towards substitution relative to halogenoalkanes. Cross reference with 20.3.	2, 3
H.7 Elimination Reactions (2h)		
H.7.1	Describe the mechanism for the elimination of water from alcohols.	2
H.7.2	Describe and explain the mechanism for the elimination of HBr from bromoalkanes. Under different conditions, the same reactants can undergo either nucleophilic substitution or elimination reactions, eg 1-bromobutane with OH^- .	2, 3

A.S.		Obj
	H.8 Addition–elimination Reactions (1h)	
H.8.1	Describe and explain the reactions of 2,4-dinitrophenylhydrazine with aldehydes and ketones. A detailed mechanism is not required.	2, 3
	H.9 Acid–base Reactions (2h)	
H.9.1	Describe and explain the acidic properties of phenol and substituted phenols in terms of bonding. Compare the acidity of phenol with alcohols, and the acidity of 2,4,6-trinitrophenol with phenol.	2, 3
H.9.2	Describe and explain the acidic properties of substituted carboxylic acids in terms of bonding. Relative acidities can be explained in terms of the ease of dissociation of the hydrogen ion or in terms of the relative stability of the conjugate base.	2, 3
H.9.3	Compare and explain the relative basicities of ammonia, amines and amides. Include the formation of salts and the liberation of the amine with NaOH from the salt. Cross reference with topic 18.	2, 3

MATHEMATICAL REQUIREMENTS

All Diploma Programme chemistry students should be able to:

- perform the basic arithmetic functions: addition, subtraction, multiplication and division
- carry out calculations involving means, decimals, fractions, percentages, ratios, approximations and reciprocals
- use standard notation (eg 3.6×10^6)
- use direct and inverse proportion
- solve simple algebraic equations
- plot graphs (with suitable scales and axes) and sketch graphs
- interpret graphs, including the significance of gradients, changes in gradients, intercepts and areas
- interpret data presented in various forms (eg bar charts, histograms, pie charts).