

2014-2015
Tigard High School Chemistry
= Scope and Sequence =

	Chemistry	IB Chem 1 <i>Prerequisite: Chemistry</i> ----- Required for Standard Level IB Exam	IB Chem 2 <i>Prerequisite: IB Chem 1</i> ----- Required for Higher Level IB Exam
	Units: 1. Introduction to Chemistry 2. Atomic Theory 3. Periodicity 4. Ionic Bonding 5. Covalent Bonding 6. The Mole 7. Chemical Reactions 8. Stoichiometry 9. Reaction Energy and Kinetics 10. Acids and Bases	Units: 1. Atomic Theory & Periodicity (SL) 2. Bonding (SL) 3. Gas Stoichiometry 4. Solution Stoichiometry 5. Equilibrium 6. Energetics & Kinetics (SL) 7. Oxidation and Reduction (SL) 8. Organic Chemistry (SL) IB Options: 9. Drugs & Medicine (SL) 10. Environmental Chemistry (SL) ----- IB Internal Assessments: • Lab Write-ups • ICT (Information & Communication Technology)	Units: 1. Atomic Theory & Periodicity (HL) 2. Bonding (HL) 3. Energetics (HL) 4. Kinetics (HL) 5. Acids and Bases (HL) 6. Oxidation and Reduction (HL) 7. Organic Chemistry (HL) IB Options: 8. Drugs & Medicine (HL) 9. Environmental Chemistry (HL) ----- IB Internal Assessments: • Lab Write-ups • ICT (Information & Communication Technology)
Note: The assessment standards given below do not necessarily follow the above unit sequence.			

Assessment Standards

- IB Topics 1 - 11: Standard Level (SL)
- IB Topics 12 - 20: Higher Level (HL)
- **Next Generation Science Standards (NGSS)**

Topic 1: Quantitative Chemistry

	Chemistry	IB Chem 1	IB Chem 2
1.1 The Mole Concept	<p>1.1.1 Apply the mole concept to substances.</p> <p>The mole concept applies to all kinds of particles: atoms, molecules, ions, electrons, formula units, and so on. The amount of substance is measured in moles (mol). The approximate value of Avogadro's constant (L), $6.02 \times 10^{23} \text{ mol}^{-1}$, should be known.</p> <p>1.1.2 Determine 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, ions, electrons and formula units.</p>		

	Chemistry	IB Chem 1	IB Chem 2
1.2 Formulas	<p>1.2.1 Define the terms <i>relative atomic mass</i>(A_r) and relative molecular mass (M_r).</p> <p>1.2.2 Calculate the mass of one mole of a species from its formula.</p> <p>The term molar mass (in g mol^{-1}) will be used.</p> <p>1.2.3 Solve problems involving the relationship between the amount of substance in moles, mass and molar mass.</p> <p>1.2.4 Distinguish between the terms empirical formula and molecular formula.</p> <p>1.2.5 Determine the empirical formula from the percentage composition or from other experimental data.</p> <p>1.2.6 Determine the molecular formula when given both the empirical formula and experimental data.</p>		
1.3 Chemical Equations	<p>NGSS HS-PS1-2. Construct and revise an explanation for the outcome of a simple chemical reaction based on the outermost electron states of atoms, trends in the periodic table, and knowledge of the patterns of chemical properties.</p> <p>NGSS HS-PS1-7. Use mathematical representations to support the claim that atoms, and therefore mass, are conserved during a chemical reaction.</p> <p>-----</p> <p>1.3.1 Deduce chemical equations when all reactants and products are given.</p> <p>Students should be aware of the difference between coefficients and subscripts.</p> <p>1.3.2 Identify the mole ratio of any two species in a chemical equation.</p> <p>1.3.3 Apply the state symbols (s), (l), (g) and (aq).</p>		

	Chemistry	IB Chem 1	IB Chem 2
1.4 Mass and gaseous volume relationships in chemical reactions	<p>1.4.1 Calculate theoretical yields from chemical equations.</p> <p>Given a chemical equation and the mass or amount (in moles) of one species, calculate the mass or amount of another species.</p> <p>1.4.2 Determine the limiting reactant and the reactant in excess when quantities of reacting substances are given.</p> <p>1.4.3 Solve problems involving theoretical, experimental and percentage yield.</p>	<p>1.4.4 Apply Avogadro's law to calculate reacting volumes of gases.</p> <p>1.4.5 Apply the concept of molar volume at standard temperature and pressure in calculations.</p> <p>1.4.6 Solve problems involving the relationship between temperature, pressure and volume for a fixed mass of an ideal gas.</p> <p>1.4.7 Solve problems using the ideal gas equation, $PV = nRT$</p> <p>1.4.8 Analyze graphs relating to the ideal gas equation.</p>	
1.5 Solutions	<p>1.5.1 Distinguish between the terms <i>solute</i>, solvent, solution and concentration.</p> <p>Concentration in mol dm⁻³ is often represented by square brackets around the substance under consideration, for example, [HCl].</p> <p>1.5.2 Solve problems involving concentration, amount of solute and volume of solution.</p>		

Topic 2: Atomic Structure

	Chemistry	IB Chem 1	IB Chem 2												
2.1 The Atom	<p>NGSS HS-PS1-8. Develop models to illustrate the changes in the composition of the nucleus of the atom and the energy released during the processes of fission, fusion, and radioactive decay.</p> <p>-----</p> <p>2.1.1 State the position of protons, neutrons and electrons in the atom.</p> <p>2.1.2 State the relative masses and relative charges of protons, neutrons and electrons.</p> <p>The accepted values are:</p> <table> <thead> <tr> <th></th> <th>relative mass</th> <th>relative charge</th> </tr> </thead> <tbody> <tr> <td>proton</td> <td>1</td> <td>+1</td> </tr> <tr> <td>neutron</td> <td>1</td> <td>0</td> </tr> <tr> <td>electron</td> <td>5×10^{-4}</td> <td>-1</td> </tr> </tbody> </table> <p>2.1.3 Define the terms <i>mass number (A)</i>, <i>atomic number (Z)</i> and <i>isotopes of an element</i>.</p>		relative mass	relative charge	proton	1	+1	neutron	1	0	electron	5×10^{-4}	-1		
	relative mass	relative charge													
proton	1	+1													
neutron	1	0													
electron	5×10^{-4}	-1													
	<p>2.1.4 Deduce the symbol for an isotope given its mass number and atomic number.</p> <p>2.1.5 Calculate the number of protons, neutrons and electrons in atoms and ions from the mass number, atomic number and charge.</p> <p>2.1.6 Compare the properties of the isotopes of an element.</p> <p>2.1.7 Discuss the uses of radioisotopes. Examples should include ^{14}C in radiocarbon dating, ^{60}Co in radiotherapy, and ^{131}I and ^{125}I as medical tracers.</p>														

	Chemistry	IB Chem 1	IB Chem 2
2.2 The Mass Spectrometer		<p>2.2.1 Describe and explain the operation of a mass spectrometer. 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.</p> <p>2.2.2 Describe how the mass spectrometer may be used to determine relative atomic mass using the ^{12}C scale.</p> <p>2.2.3 Calculate non-integer relative atomic masses and abundance of isotopes from given data.</p>	
2.3 Electron Arrangement 12.1 Electron Configuration	<p>2.3.1 Describe the electromagnetic spectrum. Students should be able to identify the ultraviolet, visible and infrared regions, and to describe the variation in wavelength, frequency and energy across the spectrum.</p> <p>2.3.2 Distinguish between a <i>continuous spectrum</i> and a <i>line spectrum</i>.</p> <p>2.3.3 Explain how the lines in the emission spectrum of hydrogen are related to electron energy levels.</p> <p>2.3.4 Deduce the electron arrangement for atoms and ions up to atomic number (Z) = 20.</p>		

	Chemistry	IB Chem 1	IB Chem 2
	<p>12.1.1 Explain how evidence from first ionization energies across periods accounts for the existence of main energy levels and sub-levels in atoms.</p> <p>12.1.2 Explain how successive ionization energy data is related to the electron configuration of an atom.</p> <p>12.1.3 State the relative energies of s, p, d and f orbitals in a single energy level.</p> <p>12.1.5 Draw the shape of an s orbital and the shapes of the p_x, p_y and p_z orbitals.</p> <p>12.1.6 Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to atomic number (Z) = 54.</p> <p>For $Z = 23$, the full electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$ and the abbreviated electron configuration is $[\text{Ar}]4s^2 3d^3$ or $[\text{Ar}]3d^3 4s^2$. Exceptions to the principle for copper and chromium should be known. Students should be familiar with the representation of the spinning electron in an orbital as an arrow in a box.</p>		

Topic 3: Periodicity

	Chemistry	IB Chem 1	IB Chem 2
3.1 The Periodic Table	<p>NGSS HS-PS1-1. Use the periodic table as a model to predict the relative properties of elements based on the patterns of electrons in the outermost energy level of atoms.</p> <p>-----</p> <p>3.1.1 Describe the arrangement of elements in the periodic table in order of increasing atomic number.</p> <p>3.1.2 Distinguish between the terms <i>group</i> and <i>period</i>.</p> <p>3.1.3 Apply the relationship between the electron arrangement of elements and their position in the periodic table up to atomic number (Z) = 20.</p> <p>3.1.4 Apply the relationship between the number of electrons in the highest occupied energy level for an element and its position in the periodic table.</p>		
3.2 Physical Properties	<p>3.2.1 Define the terms <i>first ionization energy</i> and <i>electronegativity</i>.</p> <p>3.2.2 Describe and explain the trends in atomic radii, ionic radii, first ionization energies, electronegativities and melting points for the alkali metals (Li --> Cs) and the halogens (F --> I).</p> <p>3.2.4 Compare the relative electronegativity values of two or more elements based on their positions in the periodic table.</p>		

	Chemistry	IB Chem 1	IB Chem 2
3.3 Chemical Properties		<p>3.3.1 Discuss the similarities and differences in the chemical properties of elements in the same group.</p> <p>The following reactions should be covered.</p> <ul style="list-style-type: none"> • Alkali metals (Li, Na and K) with water • Alkali metals (Li, Na and K) with halogens (Cl₂, Br₂ and I₂) • Halogens (Cl₂, Br₂ and I₂) with halide ions (Cl⁻, Br⁻ and I⁻) <p>3.3.2 Discuss the changes in nature, from ionic to covalent and from basic to acidic, of the oxides across period 3.</p> <p>Equations are required for the reactions of Na₂O, MgO, P₄O₁₀ and SO₃ with water.</p>	
13.1 Trends Across Period 3	3.2.3 Describe and explain the trends in atomic radii, ionic radii, first ionization energies and electronegativities for elements across period 3.		<p>13.1.1 Explain the physical states (under standard conditions) and electrical conductivity (in the molten state) of the chlorides and oxides of the elements in period 3 in terms of their bonding and structure.</p> <p>Include the following oxides and chlorides.</p> <ul style="list-style-type: none"> • Oxides: Na₂O, MgO, Al₂O₃, SiO₂, P₄O₆ and P₄O₁₀, SO₂ and SO₃, Cl₂O and Cl₂O₇ • Chlorides: NaCl, MgCl₂, Al₂Cl₆, SiCl₄, PCl₃ and PCl₅, and Cl₂ <p>13.1.2 Describe the reactions of chlorine and the chlorides referred to in 13.1.1 with water.</p>

	Chemistry	IB Chem 1	IB Chem 2
13.2 First-row d-block Elements			<p>13.2.1 List the characteristic properties of transition elements.</p> <p>Examples should include variable oxidation number, complex ion formation, existence of colored compounds and catalytic properties.</p> <p>13.2.2 Explain why Sc and Zn are not considered to be transition elements.</p> <p>13.2.3 Explain the existence of variable oxidation number in ions of transition elements.</p> <p>Students should know that all transition elements can show an oxidation number of +2. In addition, they should be familiar with the oxidation numbers of the following: Cr (+3, +6), Mn (+4, +7), Fe (+3) and Cu (+1).</p> <p>13.2.4 Define the term <i>ligand</i>.</p> <p>13.2.5 Describe and explain the formation of complexes of d-block elements.</p> <p>Include $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$, $[\text{Fe}(\text{CN})_6]^{3-}$, $[\text{CuCl}_4]^{2-}$ and $[\text{Ag}(\text{NH}_3)_2]^+$. Only monodentate ligands are required.</p> <p>13.2.6 Explain why some complexes of d-block elements are colored.</p> <p>Students need only know that, in complexes, the d sub-level splits into two sets of orbitals of different energy and the electronic transitions that take place between them are responsible for their colors.</p>

	Chemistry	IB Chem 1	IB Chem 2
			<p>13.2.7 State examples of the catalytic action of transition elements and their compounds.</p> <p>Examples should include:</p> <ul style="list-style-type: none"> • MnO₂ in the decomposition of hydrogen peroxide • V₂O₅ in the Contact process • Fe in the Haber process and in heme • Ni in the conversion of alkenes to alkanes • Co in vitamin B₁₂ • Pd and Pt in catalytic converters. <p>The mechanisms of action will not be assessed.</p> <p>13.2.8 Outline the economic significance of catalysts in the Contact and Haber processes.</p>

Topic 4: Bonding

	Chemistry	IB Chem 1	IB Chem 2
4.1 Ionic Bonding	<p>4.1.1 Describe the ionic bond as the electrostatic attraction between oppositely charged ions.</p> <p>4.1.2 Describe how ions can be formed as a result of electron transfer.</p> <p>4.1.3 Deduce which ions will be formed when elements in groups 1, 2 and 3 lose electrons.</p> <p>4.1.4 Deduce which ions will be formed when elements in groups 5, 6 and 7 gain electrons.</p> <p>4.1.5 State that transition elements can form more than one ion.</p> <p>Include examples such as Fe^{2+} and Fe^{3+}.</p> <p>4.1.6 Predict whether a compound of two elements would be ionic from the position of the elements in the periodic table or from their electronegativity values.</p> <p>4.1.7 State the formula of common polyatomic ions formed by nonmetals in periods 2 and 3.</p> <p>Examples include NO_3^-, OH^-, SO_4^{2-}, CO_3^{2-}, PO_4^{3-}, NH_4^+, HCO_3^-.</p> <p>4.1.8 Describe the lattice structure of ionic compounds. Students should be able to describe the structure of sodium chloride as an example of an ionic lattice.</p>		

	Chemistry	IB Chem 1	IB Chem 2
<p>4.2 Covalent Bonding</p> <p>14.2 Hybridization</p> <p>14.3 Delocalization of Electrons</p>	<p>4.2.1 Describe the covalent bond as the electrostatic attraction between a pair of electrons and positively charged nuclei.</p> <p>Single and multiple bonds should be considered. Examples should include O₂, N₂, CO₂, HCN, C₂H₄ (ethene) and C₂H₂ (ethyne).</p> <p>4.2.2 Describe how the covalent bond is formed as a result of electron sharing</p> <p>4.2.3 Deduce the Lewis (electron dot) structures of molecules and ions for up to four electron pairs on each atom.</p> <p>A pair of electrons can be represented by dots, crosses, a combination of dots and crosses or by a line.</p> <p>Coordinate (dative) covalent bonds are required. Examples include NH₄⁺ and H₃O⁺.</p> <p>4.2.4 State and explain the relationship between the number of bonds, bond length and bond strength.</p> <p>The comparison should include the bond lengths and bond strengths of:</p> <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. <p>4.2.5 Predict whether a compound of two elements would be covalent from the position of the elements in the periodic table or from their electronegativity values.</p>		<p>14.1.1 Predict the shape and bond angles for species with five and six negative charge centers using the VSEPR theory.</p> <p>Examples should include PCl₅, SF₆, XeF₄ and PF₆⁻.</p> <p>14.2.1 Describe σ and π bonds. Treatment should include:</p> <ul style="list-style-type: none"> • σ bonds resulting from the axial overlap of orbitals • π bonds resulting from the sideways overlap of parallel p orbitals • double bonds formed by one σ and one π bond • triple bonds formed by one σ and two π bonds. <p>14.2.2 Explain hybridization in terms of the mixing of atomic orbitals to form new orbitals for bonding.</p> <p>Students should consider sp, sp² and sp³ hybridization, and the shapes and orientation of these orbitals.</p> <p>14.2.3 Identify and explain the relationships between Lewis structures, molecular shapes and types of hybridization (sp, sp² and sp³).</p> <p>Students should consider examples from inorganic as well as organic chemistry.</p>

	Chemistry	IB Chem 1	IB Chem 2
	<p>4.2.6 Predict the relative polarity of bonds from electronegativity values</p> <p>4.2.7 Predict the shape and bond angles for species with four, three and two negative charge centers on the central atom using the valence shell electron pair repulsion theory (VSEPR).</p> <p>Examples should include CH₄, NH₃, H₂O, NH₄⁺, H₂O⁺, BF₃, C₂H₄, SO₂, C₂H₂ and CO₂.</p> <p>4.2.8 Predict whether or not a molecule is polar from its molecular shape and bond polarities.</p>	<p>4.2.9 Describe and compare the structure and bonding in the three allotropes of carbon (diamond, graphite and C₆₀ fullerene).</p> <p>4.2.10 Describe the structure of and bonding in silicon and silicon dioxide.</p>	<p>14.3.1 Describe the delocalization of π electrons and explain how this can account for the structures of some species.</p> <p>Examples should include NO₃⁻, NO₂⁻, CO₃²⁻, O₃, RCOO⁻ and benzene.</p>
4.3 Intermolecular Forces	<p>4.3.1 Describe the types of intermolecular forces (attractions between molecules that have temporary dipoles, permanent dipoles or hydrogen bonding) and explain how they arise from the structural features of molecules.</p> <p>The term van der Waals' forces can be used to describe the interaction between non-polar molecules.</p> <p>4.3.2 Describe and explain how intermolecular forces affect the boiling points of substances.</p> <p>The presence of hydrogen bonding can be illustrated by comparing:</p> <ul style="list-style-type: none"> • HF and HCl • H₂O and H₂S • NH₃ and PH₃ • CH₃OCH₃ and CH₃CH₂OH • CH₃CH₂CH₃, CH₃CHO and CH₃CH₂OH. 		
4.4 Metallic Bonding	<p>4.4.1 Describe the metallic bond as the electrostatic attraction between a lattice of positive ions and delocalized electrons.</p> <p>4.4.2 Explain the electrical conductivity and malleability of metals.</p>		

	Chemistry	IB Chem 1	IB Chem 2
4.5 Physical Properties	<p>NGSS HS-PS1-3. Plan and conduct an investigation to gather evidence to compare the structure of substances at the bulk scale to infer the strength of electrical forces between particles.</p> <p>NGSS HS-PS2-6. Communicate scientific and technical information about why the molecular-level structure is important in the functioning of designed materials.</p> <p>-----</p> <p>4.5.1 Compare and explain the properties of substances resulting from different types of bonding.</p> <p>Examples should include melting and boiling points, volatility, electrical conductivity and solubility in non-polar and polar solvents.</p>		

Topic 5: Energetics

	Chemistry	IB Chem 1	IB Chem 2
5.1 Exothermic and Endothermic Reactions	<p>5.1.1 Define the terms exothermic reaction, endothermic reaction and standard enthalpy change of reaction (ΔH°).</p> <p>5.1.3 Apply the relationship between temperature change, enthalpy change and the classification of a reaction as endothermic or exothermic.</p>	<p>5.1.2 State that combustion and neutralization are exothermic processes.</p> <p>5.1.4 Deduce, from an enthalpy level diagram, the relative stabilities of reactants and products, and the sign of the enthalpy change for the reaction.</p>	
5.2 Calculation of Enthalpy Changes		<p>5.2.1 Calculate the heat energy change when the temperature of a pure substance is changed.</p> <p>Students should be able to calculate the heat energy change for a substance given the mass, specific heat capacity and temperature change using $q = mc\Delta T$.</p> <p>5.2.2 Design suitable experimental procedures for measuring the heat energy changes of reactions.</p> <p>Students should consider reactions in aqueous solution and combustion reactions.</p> <p>5.2.3 Calculate the enthalpy change for a reaction using experimental data on temperature changes, quantities of reactants and mass of water.</p> <p>5.2.4 Evaluate the results of experiments to determine enthalpy changes. Students should be aware of the assumptions made and errors due to heat loss.</p>	

	Chemistry	IB Chem 1	IB Chem 2
5.3 Hess's Law 15.1 Standard Enthalpy Changes of Reaction		5.3.1 Determine the enthalpy change of a reaction that is the sum of two or three reactions with known enthalpy changes. Students should be able to use simple enthalpy cycles and enthalpy level diagrams and to manipulate equations. Students will not be required to state Hess's law.	15.1.1 Define and apply the terms <i>standard state</i> , <i>standard enthalpy change of formation</i> (ΔH_f^\ominus) and <i>standard enthalpy change of combustion</i> (ΔH_c^\ominus). 15.1.2 Determine the enthalpy change of a reaction using standard enthalpy changes of formation and combustion.
5.4 Bond Enthalpy 15.2 Born-Haber Cycle	NGSS HS-PS1-4. Develop a model to illustrate that the release or absorption of energy from a chemical reaction system depends upon the changes in total bond energy. ----- 5.4.1 Define the term <i>average bond enthalpy</i> . 5.4.2 Explain, in terms of average bond enthalpies, why some reactions are exothermic and others are endothermic.		15.2.1 Define and apply the terms <i>lattice enthalpy</i> and <i>electron affinity</i> . 15.2.2 Explain how the relative sizes and the charges of ions affect 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. 15.2.3 Construct a Born–Haber cycle for group 1 and 2 oxides and chlorides, and use it to calculate an enthalpy change. 15.2.4 Discuss the difference between theoretical and experimental lattice enthalpy values of ionic compounds in terms of their covalent character. A significant difference between the two values indicates covalent character.

	Chemistry	IB Chem 1	IB Chem 2
15.3 Entropy			<p>15.3.1 State and explain the factors that increase the entropy in a system.</p> <p>15.3.2 Predict whether the entropy change (ΔS) for a given reaction or process is positive or negative.</p> <p>15.3.3 Calculate the standard entropy change for a reaction (ΔS^\ominus) using standard entropy values (S^\ominus).</p>
15.4 Spontaneity			<p>15.4.1 Predict whether a reaction or process will be spontaneous by using the sign of ΔG^\ominus.</p> <p>15.4.2 Calculate (ΔG^\ominus) for a reaction using the equation $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$ and by using values of the standard free energy change of formation, ΔG_f^\ominus.</p> <p>15.4.3 Predict the effect of a change in temperature on the spontaneity of a reaction using standard entropy and enthalpy changes and the equation $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$.</p>

Topic 6: Kinetics

	Chemistry	IB Chem 1	IB Chem 2
<p>6.1 Rates of Reaction</p> <p>16.1 Rate Expression</p>	<p>NGSS HS-PS1-5. Apply scientific principles and evidence to provide an explanation about the effects of changing the temperature or concentration of the reacting particles on the rate at which a reaction occurs.</p> <p>NGSS HS-PS1-6. Refine the design of a chemical system by specifying a change in conditions that would produce increased amounts of products at equilibrium.</p> <p>-----</p> <p>6.1.1 Define the term <i>rate of reaction</i>.</p> <p>6.1.2 Describe suitable experimental procedures for measuring rates of reactions.</p>	<p>6.1.3 Analyze data from rate experiments.</p> <p>Students should be familiar with graphs of changes in concentration, volume and mass against time.</p>	<p>16.1.1 Distinguish between the terms <i>rate constant</i>, <i>overall order of reaction</i> and <i>order of reaction</i> with respect to a particular reactant.</p> <p>16.1.2 Deduce the rate expression for a reaction from experimental data.</p> <p>16.1.3 Solve problems involving the rate expression.</p> <p>16.1.4 Sketch, identify and analyze graphical representations for zero-, first- and second-order reactions.</p> <p>Students should be familiar with both concentration–time and rate–concentration graphs.</p>

	Chemistry	IB Chem 1	IB Chem 2
<p>6.2 Collision Theory</p> <p>16.2 Reaction Mechanism</p> <p>16.3 Activation Energy</p>	<p>6.2.1 Describe the kinetic theory in terms of the movement of particles whose average energy is proportional to temperature in kelvins.</p>	<p>6.2.5 Sketch and explain qualitatively the Maxwell–Boltzmann energy distribution curve for a fixed amount of gas at different temperatures and its consequences for changes in reaction rate.</p> <p>Students should be able to explain why the area under the curve is constant and does not change with temperature.</p> <p>6.2.6 Describe the effect of a catalyst on a chemical reaction.</p> <p>6.2.7 Sketch and explain Maxwell–Boltzmann curves for reactions with and without catalysts.</p>	<p>16.2.1 Explain that reactions can occur by more than one step and that the slowest step determines the rate of reaction (rate-determining step).</p> <p>16.2.2 Describe the relationship between reaction mechanism, order of reaction and rate-determining step.</p> <p>Only examples with one- or two-step reactions where the mechanism is given will be assessed.</p> <p>TOK: Agreement between rate equation and a suggested mechanism only provides evidence to support a reaction mechanism. Disagreement disproves the mechanism.</p> <p>16.3.1 Describe qualitatively the relationship between the rate constant (k) and temperature (T).</p> <p>16.3.2 Determine activation energy (E_a) values from the Arrhenius equation by a graphical method.</p> <p>The Arrhenius equation and its logarithmic form are provided in the <i>Chemistry data booklet</i>. The use of simultaneous equations will not be assessed.</p>

Topic 7: Equilibrium

	Chemistry	IB Chem 1	IB Chem 2
<p>7.1 Dynamic Equilibrium</p> <p>17.1 Liquid-Vapor Equilibrium</p>		<p>7.1.1 Outline the characteristics of chemical and physical systems in a state of equilibrium.</p> <p>17.1.1 Describe the equilibrium established between a liquid and its own vapor and how it is affected by temperature changes.</p> <p>17.1.2 Sketch graphs showing the relationship between vapor pressure and temperature and explain them in terms of the kinetic theory.</p> <p>17.1.3 State and explain the relationship between enthalpy of vaporization, boiling point and intermolecular forces.</p>	

	Chemistry	IB Chem 1	IB Chem 2
<p>7.2 The Position of Equilibrium</p> <p>17.2 The Equilibrium Constant</p>		<p>7.2.1 Deduce the equilibrium constant expression (K_c) from the equation for a homogeneous reaction.</p> <p>Consider gases, liquids and aqueous solutions.</p> <p>7.2.2 Deduce the extent of a reaction from the magnitude of the equilibrium constant.</p> <p>When $K_c \gg 1$, the reaction goes almost to completion. When $K_c \ll 1$, the reaction hardly proceeds.</p> <p>7.2.3 Apply Le Chatelier's principle to predict the qualitative effects of changes of temperature, pressure and concentration on the position of equilibrium and on the value of the equilibrium constant.</p> <p>Students will not be required to state Le Chatelier's principle.</p> <p>7.2.4 State and explain the effect of a catalyst on an equilibrium reaction.</p> <p>7.2.5 Apply the concepts of kinetics and equilibrium to industrial processes.</p> <p>Suitable examples include the Haber and Contact processes.</p> <p>17.2.1 Solve homogeneous equilibrium problems using the expression for K_c.</p>	

Topic 8: Acids and Bases

	Chemistry	IB Chem 1	IB Chem 2
8.1 Theories of Acids and Bases	<p>8.1.1 Define <i>acids</i> and <i>bases</i> according to the Brønsted–Lowry theory.</p> <p>8.1.2 Deduce whether or not a species could act as a Brønsted–Lowry acid or base.</p> <p>8.1.3 Deduce the formula of the conjugate acid (or base) of any Brønsted–Lowry base (or acid).</p> <p>Students should make clear the location of the proton transferred, for example, CH₃COOH/CH₃COO⁻ rather than C₂H₄O₂/C₂H₃O₂⁻.</p>		
8.2 Properties of Acids and Bases	<p>8.2.1 Outline the characteristic properties of acids and bases in aqueous solution.</p> <p>Bases that are not hydroxides, such as ammonia, soluble carbonates and hydrogencarbonates, should be included.</p> <p>Alkalis are bases that dissolve in water. Students should consider the effects on indicators and the reactions of acids with bases, metals and carbonates.</p>		

	Chemistry	IB Chem 1	IB Chem 2
8.3 Strong and Weak Acids and Bases	<p>8.3.1 Distinguish between <i>strong</i> and <i>weak</i> acids and bases in terms of the extent of dissociation, reaction with water and electrical conductivity.</p> <p>8.3.2 State whether a given acid or base is strong or weak.</p> <p>Students should consider hydrochloric acid, nitric acid and sulfuric acid as examples of strong acids, and carboxylic acids and carbonic acid (aqueous carbon dioxide) as weak acids.</p> <p>Students should consider all group 1 hydroxides and barium hydroxide as strong bases, and ammonia and amines as weak bases.</p> <p>8.3.3 Distinguish between <i>strong</i> and <i>weak</i> acids and bases, and determine the relative strengths of acids and bases, using experimental data.</p>		

	Chemistry	IB Chem 1	IB Chem 2
<p>8.4 The pH Scale</p> <p>18.1 Calculations Involving Acids and Bases</p>	<p>8.4.1 Distinguish between aqueous solutions that are <i>acidic</i>, <i>neutral</i> or <i>alkaline</i> using the pH scale.</p> <p>8.4.2 Identify which of two or more aqueous solutions is more acidic or alkaline using pH values.</p> <p>Students should be familiar with the use of a pH meter and universal indicator.</p> <p>8.4.3 State that each change of one pH unit represents a 10-fold change in the hydrogen ion concentration $[H^+(aq)]$.</p> <p>Relate integral values of pH to $[H^+(aq)]$ expressed as powers of 10.</p> <p>8.4.4 Deduce changes in $[H^+(aq)]$ when the pH of a solution changes by more than one pH unit.</p> <p>18.1.1 State the expression for the ionic product constant of water (K_w).</p> <p>18.1.2 Deduce $[H^+(aq)]$ and $[OH^-(aq)]$ for water at different temperatures given K_w values.</p> <p>18.1.3 Solve problems involving $[H^+(aq)]$, $[OH^-(aq)]$, pH and pOH.</p>		<p>18.1.4 State the equation for the reaction of any weak acid or weak base with water, and hence deduce the expressions for K_a and K_b.</p> <p>Only examples involving the transfer of one proton will be assessed.</p> <p>18.1.5 Solve problems involving solutions of weak acids and bases using the expressions: $K_a \times K_b = K_w$ $pK_a + pK_b = pK_w$ $pH + pOH = pK_w$.</p> <p>Students should state when approximations are used in equilibrium calculations.</p> <p>The use of quadratic equations will not be assessed.</p> <p>18.1.6 Identify the relative strengths of acids and bases using values of K_a, K_b, pK_a and pK_b.</p>

	Chemistry	IB Chem 1	IB Chem 2
18.2 Buffer Solutions			<p>18.2.1 Describe the composition of a buffer solution and explain its action.</p> <p>18.2.2 Solve problems involving the composition and pH of a specified buffer system.</p> <p>Only examples involving the transfer of one proton will be assessed. Examples should include ammonia solution/ammonium chloride and ethanoic acid/sodium ethanoate.</p> <p>Students should state when approximations are used in equilibrium calculations. The use of quadratic equations will not be assessed.</p>
18.3 Salt Hydrolysis			<p>18.3.1 Deduce 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 and 3 and d-block elements should also be considered. For example,</p> $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}(\text{aq}) \longrightarrow [\text{Fe}(\text{OH})(\text{H}_2\text{O})_5]^{2+}(\text{aq}) + \text{H}^+(\text{aq})$
18.4 Acid-Base Titrations			<p>18.4.1 Sketch the general shapes of graphs of pH against volume for titrations involving strong and weak acids and bases, and explain their important features.</p> <p>Only examples involving the transfer of one proton will be assessed. Important features are:</p> <ul style="list-style-type: none"> • intercept with pH axis • equivalence point • buffer region • points where $\text{p}K_{\text{a}} = \text{pH}$ or $\text{p}K_{\text{b}} = \text{pOH}$.

	Chemistry	IB Chem 1	IB Chem 2
18.5 Indicators			<p>18.5.1 Describe qualitatively the action of an acid–base indicator.</p> <p>Use $\text{HIn}(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{In}^-(\text{aq})$</p> <p style="text-align: center;">colour A colour B</p> <p>18.5.2 State and explain how the pH range of an acid–base indicator relates to its $\text{p}K_{\text{a}}$ value.</p> <p>18.5.3 Identify an appropriate indicator for a titration, given the equivalence point of the titration and the pH range of the indicator.</p> <p>Examples of indicators are listed in the <i>Chemistry data booklet</i>.</p>

Topic 9: Oxidation and Reduction

	Chemistry	IB Chem 1	IB Chem 2
9.1 Introduction to Oxidation and Reduction		<p>9.1.1 Define <i>oxidation</i> and <i>reduction</i> in terms of electron loss and gain.</p> <p>9.1.2 Deduce the oxidation number of an element in a compound.</p> <p>Oxidation numbers should be shown by a sign (+ or -) and a number, for example, +7 for Mn in KMnO_4.</p> <p>9.1.3 State the names of compounds using oxidation numbers.</p> <p>Oxidation numbers in names of compounds are represented by Roman numerals, for example, iron(II) oxide, iron(III) oxide.</p> <p>9.1.4 Deduce whether an element undergoes oxidation or reduction in reactions using oxidation numbers.</p>	
9.2 Redox Equations		<p>9.2.1 Deduce simple oxidation and reduction half-equations given the species involved in a redox reaction.</p> <p>9.2.2 Deduce redox equations using half-equations.</p> <p>H^+ and H_2O should be used where necessary to balance half-equations in acid solution. The balancing of equations for reactions in alkaline solution will not be assessed.</p> <p>9.2.3 Define the terms <i>oxidizing agent</i> and <i>reducing agent</i>.</p> <p>9.2.4 Identify the oxidizing and reducing agents in redox equations.</p>	

	Chemistry	IB Chem 1	IB Chem 2
9.3 Reactivity		<p>9.3.1 Deduce a reactivity series based on the chemical behavior of a group of oxidizing and reducing agents.</p> <p>Examples include displacement reactions of metals and halogens. Standard electrode potentials will not be assessed.</p> <p>9.3.2 Deduce the feasibility of a redox reaction from a given reactivity series.</p> <p>Students are not expected to recall a specific reactivity series.</p>	
9.4 Voltaic Cells 19.1 Standard Electrode Potentials		<p>9.4.1 Explain how a redox reaction is used to produce electricity in a voltaic cell.</p> <p>This should include a diagram to show how two half-cells can be connected by a salt bridge. Examples of half-cells are Mg, Zn, Fe and Cu in solutions of their ions.</p> <p>9.4.2 State that oxidation occurs at the negative electrode (anode) and reduction occurs at the positive electrode (cathode).</p>	<p>19.1.1 Describe the standard hydrogen electrode.</p> <p>19.1.2 Define the term <i>standard electrode potential</i> (E^\ominus).</p> <p>19.1.3 Calculate cell potentials using standard electrode potentials.</p> <p>19.1.4 Predict whether a reaction will be spontaneous using standard electrode potential values.</p>

	Chemistry	IB Chem 1	IB Chem 2
<p>9.5 Electrolytic Cells</p> <p>19.2 Electrolysis</p>		<p>9.5.1 Describe, using a diagram, the essential components of an electrolytic cell.</p> <p>The diagram should include the source of electric current and conductors, positive and negative electrodes, and the electrolyte.</p> <p>9.5.2 State that oxidation occurs at the positive electrode (anode) and reduction occurs at the negative electrode (cathode).</p> <p>9.5.3 Describe how current is conducted in an electrolytic cell.</p> <p>9.5.4 Deduce the products of the electrolysis of a molten salt.</p> <p>Half-equations showing the formation of products at each electrode will be assessed.</p>	<p>19.2.1 Predict and explain the products of electrolysis of aqueous solutions.</p> <p>Explanations should refer to E° values, nature of the electrode and concentration of the electrolyte. Examples include the electrolysis of water, aqueous sodium chloride and aqueous copper(II) sulfate.</p> <p>19.2.2 Determine the relative amounts of the products formed during electrolysis.</p> <p>The factors to be considered are charge on the ion, current and duration of electrolysis.</p> <p>19.2.3 Describe the use of electrolysis in electroplating.</p>

Topic 10: Organic Chemistry

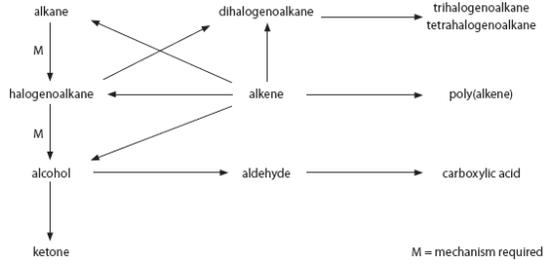
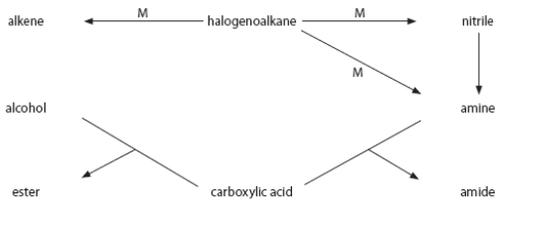
	Chemistry	IB Chem 1	IB Chem 2
<p>10.1 & 20.1 Introduction</p> <p>20.6 Stereoisomerism</p>		<p>10.1.1 Describe the features of a homologous series.</p> <p>Include the same general formula, neighboring members differing by CH_2, similar chemical properties and gradation in physical properties.</p> <p>10.1.2 Predict and explain the trends in boiling points of members of a homologous series.</p> <p>10.1.3 Distinguish between <i>empirical</i>, <i>molecular</i> and <i>structural</i> formulas.</p> <p>A structural formula is one that shows unambiguously how the atoms are arranged together.</p> <p>A full structural formula (sometimes called a graphic formula or displayed formula) shows every atom and bond</p> <p>A condensed structural formula can omit bonds between atoms and can show identical groups bracketed together, for example, for hexane: $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$ or $\text{CH}_3(\text{CH}_2)_4\text{CH}_3$.</p> <p>The use of R to represent an alkyl group and to represent the benzene ring can be used in condensed structural formulas.</p> <p>Although skeletal formulas are used for more complex structures in the <i>Chemistry data booklet</i>, such formulas will not be accepted in examination answers.</p> <p>10.1.4 Describe structural isomers as compounds with the same molecular formula but with different arrangements of atoms.</p>	<p>20.1.1 Deduce structural formulas for compounds containing up to six carbon atoms with one of the following functional groups: amine, amide, ester and nitrile.</p> <p>Condensed structural formulas can use NH_2, CONH_2, $-\text{COOC}-$ and CN for these groups.</p> <p>20.1.2 Apply IUPAC rules for naming compounds containing up to six carbon atoms with one of the following functional groups: amine, amide, ester and nitrile.</p>

	Chemistry	IB Chem 1	IB Chem 2
		<p>10.1.5 Deduce structural formulas for the isomers of the non-cyclic alkanes up to C₆.</p> <p>Include both straight-chain and branched-chain isomers.</p> <p>10.1.6 Apply IUPAC rules for naming the isomers of the non-cyclic alkanes up to C₆.</p> <p>10.1.7 Deduce structural formulas for the isomers of the straight-chain alkenes up to C₆.</p> <p>10.1.8 Apply IUPAC rules for naming the isomers of the straight-chain alkenes up to C₆.</p> <p>The distinction between <i>cis</i> and <i>trans</i> isomers is not required.</p> <p>10.1.9 Deduce structural formulas for compounds containing up to six carbon atoms with one of the following functional groups: alcohol, aldehyde, ketone, carboxylic acid and halide.</p> <p>Condensed structural formulas can use OH, CHO, CO, COOH and F/Cl/Br/I.</p> <p>10.1.10 Apply IUPAC rules for naming compounds containing up to six carbon atoms with one of the following functional groups: alcohol, aldehyde, ketone, carboxylic acid and halide.</p>	<p>20.6.1 Describe stereoisomers as compounds with the same structural formula but with different arrangements of atoms in space.</p> <p>20.6.2 Describe and explain geometrical isomerism in non-cyclic alkenes.</p> <p>Include the prefixes <i>cis</i>- and <i>trans</i>- and the term restricted rotation.</p> <p>20.6.3 Describe and explain geometrical isomerism in C₃ and C₄ cycloalkanes.</p> <p>Include the dichloro derivatives of cyclopropane and cyclobutane.</p> <p>20.6.4 Explain the difference in the physical and chemical properties of geometrical isomers.</p> <p>Include <i>cis</i>- and <i>trans</i>-1,2-dichloroethene as examples with different boiling points, and <i>cis</i>- and <i>trans</i>-but-2-ene-1,4-dioic acid as examples that react differently when heated.</p> <p>20.6.5 Describe and explain optical isomerism in simple organic molecules.</p> <p>Include examples such as butan-2-ol and 2-bromobutane.</p> <p>The term asymmetric can be used to describe a carbon atom joined to four different atoms or groups.</p> <p>The term chiral can be used to describe a carbon atom joined to four different atoms or groups, and also as a description of the molecule itself.</p> <p>Include the meanings of the terms enantiomer and racemic mixture.</p>

	Chemistry	IB Chem 1	IB Chem 2
		<p>10.1.11 Identify the following functional groups when present in structural formulas: amino (NH_2), benzene ring (⊙) and esters (RCOOR).</p> <p>10.1.12 Identify primary, secondary and tertiary carbon atoms in alcohols and halogenoalkanes.</p> <p>The terms primary, secondary and tertiary can also be applied to the molecules containing these carbon atoms.</p> <p>10.1.13 Discuss the volatility and solubility in water of compounds containing the functional groups listed in 10.1.9.</p>	<p>TOK: The existence of optical isomers provided indirect evidence of a tetrahedrally bonded carbon atom. This is an example of the power of reasoning in allowing us access to the molecular scale. Do we know or believe those carbon atoms are tetrahedrally coordinated? The use of conventions in representing three-dimensional molecules in two dimensions could also be discussed.</p> <p>20.6.6 Outline the use of a polarimeter in distinguishing between optical isomers.</p> <p>Include the meaning of the term plane-polarized light.</p> <p>20.6.7 Compare the physical and chemical properties of enantiomers.</p>
10.2 Alkanes		<p>10.2.1 Explain the low reactivity of alkanes in terms of bond enthalpies and bond polarity.</p> <p>10.2.2 Describe, using equations, the complete and incomplete combustion of alkanes.</p> <p>10.2.3 Describe, using equations, the reactions of methane and ethane with chlorine and bromine.</p> <p>10.2.4 Explain the reactions of methane and ethane with chlorine and bromine in terms of a free-radical mechanism.</p> <p>Reference should be made to homolytic fission and the reaction steps of initiation, propagation and termination.</p> <p>The use of the half-arrow to represent the movement of a single electron is not required. The formulas of free radicals should include the radical symbol, for example, $\text{Cl}\cdot$.</p>	

	Chemistry	IB Chem 1	IB Chem 2
10.3 Alkenes		<p>10.3.1 Describe, using equations, the reactions of alkenes with hydrogen and halogens.</p> <p>10.3.2 Describe, using equations, the reactions of symmetrical alkenes with hydrogen halides and water.</p> <p>10.3.3 Distinguish between <i>alkanes</i> and <i>alkenes</i> using bromine water.</p> <p>10.3.4 Outline the polymerization of alkenes.</p> <p>Include the formation of poly(ethene), poly(chloroethene) and poly(propene) as examples of addition polymers.</p> <p>Include the identification of the repeating unit, for example, $-(\text{CH}_2-\text{CH}_2-)_n-$ for poly(ethene).</p> <p>10.3.5 Outline the economic importance of the reactions of alkenes.</p>	

	Chemistry	IB Chem 1	IB Chem 2
10.4 Alcohols		<p>10.4.1 Describe, using equations, the complete combustion of alcohols.</p> <p>10.4.2 Describe, using equations, the oxidation reactions of alcohols.</p> <p>A suitable oxidizing agent is acidified potassium dichromate(VI).</p> <p>Equations may be balanced using the symbol [O] to represent oxygen supplied by the oxidizing agent. Include the different conditions needed to obtain good yields of different products, that is, an aldehyde by distilling off the product as it is formed, and a carboxylic acid by heating under reflux.</p> <p>10.4.3 Determine the products formed by the oxidation of primary and secondary alcohols.</p> <p>Assume that tertiary alcohols are not oxidized by potassium dichromate(VI).</p>	
10.5 Halogenoalkanes		<p>10.5.1 Describe, using equations, the substitution reactions of halogenoalkanes with sodium hydroxide.</p> <p>10.5.2 Explain the substitution reactions of halogenoalkanes with sodium hydroxide in terms of S_N1 and S_N2 mechanisms.</p> <p>Reference should be made to heterolytic fission. Curly arrows should be used to represent the movement of electron pairs.</p> <p>For tertiary halogenoalkanes the predominant mechanism is S_N1, and for primary halogenoalkanes it is S_N2. Both mechanisms occur for secondary halogenoalkanes.</p>	

	Chemistry	IB Chem 1	IB Chem 2
10.6 & 20.5 Reaction Pathways		<p>10.6.1 Deduce reaction pathways given the starting materials and the product.</p> <p>Conversions with more than two stages will not be assessed.</p> <p>Reagents, conditions and equations should be included.</p> <p>For example, the conversion of but-2-ene to butanone can be done in two stages: but-2-ene can be heated with steam and a catalyst to form butan-2-ol, which can then be oxidized by heating with acidified potassium dichromate(VI) to form butanone.</p> <p>The compound and reaction types in this topic are summarized in the following scheme:</p>  <p>M = mechanism required</p>	<p>20.5.1 Deduce reaction pathways given the starting materials and the product.</p> <p>Conversions with more than two stages will not be assessed.</p> <p>Reagents, conditions and equations should be included.</p> <p>For example, the conversion of 1-bromopropane to 1-butylamine can be done in two stages: 1-bromopropane can be reacted with potassium cyanide to form propanenitrile, which can then be reduced by heating with hydrogen and a nickel catalyst.</p> <p>The compound and reaction types in this topic are summarized in the following scheme:</p>  <p>M = mechanism required</p>

	Chemistry	IB Chem 1	IB Chem 2
20.2 Nucleophilic Substitution Reactions			<p>20.2.1 Explain why the hydroxide ion is a better nucleophile than water.</p> <p>20.2.2 Describe and explain how the rate of nucleophilic substitution in halogenoalkanes by the hydroxide ion depends on the identity of the halogen.</p> <p>20.2.3 Describe and explain how the rate of nucleophilic substitution in halogenoalkanes by the hydroxide ion depends on whether the halogenoalkane is primary, secondary or tertiary.</p> <p>20.2.4 Describe, using equations, the substitution reactions of halogenoalkanes with ammonia and potassium cyanide.</p> <p>20.2.5 Explain the reactions of primary halogenoalkanes with ammonia and potassium cyanide in terms of the S_N2 mechanism.</p> <p>20.2.6 Describe, using equations, the reduction of nitriles using hydrogen and a nickel catalyst.</p>
20.3 Elimination Reactions			<p>20.3.1 Describe, using equations, the elimination of HBr from bromoalkanes.</p> <p>20.3.2 Describe and explain the mechanism for the elimination of HBr from bromoalkanes.</p>

	Chemistry	IB Chem 1	IB Chem 2
20.4 Condensation Reactions			<p>20.4.1 Describe, using equations, the reactions of alcohols with carboxylic acids to form esters, and state the uses of esters.</p> <p>Mechanisms will not be assessed.</p> <p>20.4.2 Describe, using equations, the reactions of amines with carboxylic acids.</p> <p>Mechanisms will not be assessed.</p> <p>20.4.3 Deduce the structures of the polymers formed in the reactions of alcohols with carboxylic acids.</p> <p>Emphasize the need for two functional groups on each monomer. Include the polyester formed from ethane-1,2-diol and benzene-1,4-dicarboxylic acid. Include the identification of the repeating unit.</p> <p>20.4.4 Deduce the structures of the polymers formed in the reactions of amines with carboxylic acids.</p> <p>Emphasize the need for two functional groups on each monomer. Include the polyamide formed from 1,6-diaminohexane and hexanedioic acid. Include the identification of the repeating unit.</p> <p>20.4.5 Outline the economic importance of condensation reactions.</p>

Topic 11: Measurement and Data Processing

	Chemistry	IB Chem 1	IB Chem 2
<p>11.1 Uncertainty and Error In Measurement</p>		<p>11.1.1 Describe and give examples of random uncertainties and systematic errors.</p> <p>11.1.2 Distinguish between <i>precision</i> and accuracy.</p> <p>It is possible for a measurement to have great precision yet be inaccurate (for example, if the top of a meniscus is read in a pipette or a measuring cylinder).</p> <p>11.1.3 Describe how the effects of random uncertainties may be reduced.</p> <p>Students should be aware that random uncertainties, but not systematic errors, are reduced by repeating readings.</p> <p>11.1.4 State random uncertainty as an uncertainty range (\pm).</p> <p>11.1.5 State the results of calculations to the appropriate number of significant figures.</p> <p>The number of significant figures in any answer should reflect the number of significant figures in the given data.</p>	
<p>11.2 Uncertainty in Calculated Result</p>		<p>11.2.1 State uncertainties as absolute and percentage uncertainties.</p> <p>11.2.2 Determine the uncertainties in results.</p> <p>Only a simple treatment is required. For functions such as addition and subtraction, absolute uncertainties can be added. For multiplication, division and powers, percentage uncertainties can be added. If one uncertainty is much larger than others, the approximate uncertainty in the calculated result can be taken as due to that quantity alone.</p>	

	Chemistry	IB Chem 1	IB Chem 2
11.3 Graphical Techniques		<p>11.3.1 Sketch graphs to represent dependences and interpret graph behavior.</p> <p>Students should be able to give a qualitative physical interpretation of a particular graph, for example, the variables are proportional or inversely proportional.</p> <p>11.3.2 Construct graphs from experimental data.</p> <p>This involves the choice of axes and scale, and the plotting of points.</p> <p>11.3.3 Draw best-fit lines through data points on a graph. These can be curves or straight lines.</p> <p>11.3.4 Determine the values of physical quantities from graphs.</p> <p>Include measuring and interpreting the slope (gradient), and stating the units for these quantities.</p>	