# Syllabus details—Core

## Topic 1: Quantitative chemistry (12.5 hours)

### 1.1 The mole concept and Avogadro's constant

**2 hours**

**TOK:** Assigning numbers to the masses of the chemical elements allowed chemistry to develop into a physical science and use mathematics to express relationships between reactants and products.

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 1.1.1 | Apply the mole concept to substances. | 2 | 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*), $6.02 \times 10^{23}$ $mol^{-1}$, should be known. **TOK:** Chemistry deals with enormous differences in scale. The magnitude of Avogadro's constant is beyond the scale of our everyday experience. |
| 1.1.2 | Determine the number of particles and the amount of substance (in moles). | 3 | Convert between the amount of substance (in moles) and the number of atoms, molecules, ions, electrons and formula units. |

### 1.2 Formulas

**3 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 1.2.1 | Define the terms *relative atomic mass* ($A_r$) and *relative molecular mass* ($M_r$). | 1 | |
| 1.2.2 | Calculate the mass of one mole of a species from its formula. | 2 | The term molar mass (in g $mol^{-1}$) will be used. |
| 1.2.3 | Solve problems involving the relationship between the amount of substance in moles, mass and molar mass. | 3 | |
| 1.2.4 | Distinguish between the terms *empirical formula* and *molecular formula*. | 2 | |
| 1.2.5 | Determine the empirical formula from the percentage composition or from other experimental data. | 3 | **Aim 7:** Virtual experiments can be used to demonstrate this. |
| 1.2.6 | Determine the molecular formula when given both the empirical formula and experimental data. | 3 | |

## 1.3 Chemical equations

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 1.3.1 | Deduce chemical equations when all reactants and products are given. | 3 | Students should be aware of the difference between coefficients and subscripts. |
| 1.3.2 | Identify the mole ratio of any two species in a chemical equation. | 2 | |
| 1.3.3 | Apply the state symbols (s), (l), (g) and (aq). | 2 | **TOK:** When are these symbols necessary in aiding understanding and when are they redundant? |

## 1.4 Mass and gaseous volume relationships in chemical reactions

**4.5 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 1.4.1 | Calculate theoretical yields from chemical equations. | 2 | Given a chemical equation and the mass or amount (in moles) of one species, calculate the mass or amount of another species. |
| 1.4.2 | Determine the limiting reactant and the reactant in excess when quantities of reacting substances are given. | 3 | **Aim 7:** Virtual experiments can be used here. |
| 1.4.3 | Solve problems involving theoretical, experimental and percentage yield. | 3 | |
| 1.4.4 | Apply Avogadro's law to calculate reacting volumes of gases. | 2 | |
| 1.4.5 | Apply the concept of molar volume at standard temperature and pressure in calculations. | 2 | The molar volume of an ideal gas under standard conditions is $2.24 \times 10^{-2}$ $m^3$ $mol^{-1}$ (22.4 $dm^3$ $mol^{-1}$). |
| 1.4.6 | Solve problems involving the relationship between temperature, pressure and volume for a fixed mass of an ideal gas. | 3 | **Aim 7:** Simulations can be used to demonstrate this. |
| 1.4.7 | Solve problems using the ideal gas equation, $PV = nRT$ | 3 | **TOK:** The distinction between the Celsius and Kelvin scales as an example of an artificial and natural scale could be discussed. |
| 1.4.8 | Analyse graphs relating to the ideal gas equation. | 3 | |

## 1.5 Solutions

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 1.5.1 | Distinguish between the terms *solute*, *solvent*, *solution* and *concentration* (g $dm^{-3}$ and mol $dm^{-3}$). | 2 | Concentration in mol $dm^{-3}$ is often represented by square brackets around the substance under consideration, for example, [HCl]. |
| 1.5.2 | Solve problems involving concentration, amount of solute and volume of solution. | 3 | |

# Topic 2: Atomic structure (4 hours)

## 2.1 The atom

**1 hour**

**TOK:** What is the significance of the model of the atom in the different areas of knowledge? Are the models and theories that scientists create accurate descriptions of the natural world, or are they primarily useful interpretations for prediction, explanation and control of the natural world?

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 2.1.1 | State the position of protons, neutrons and electrons in the atom. | 1 | **TOK:** None of these particles can be (or will be) directly observed. Which ways of knowing do we use to interpret indirect evidence gained through the use of technology? Do we believe or know of their existence? |
| 2.1.2 | State the relative masses and relative charges of protons, neutrons and electrons. | 1 | The accepted values are:<br><br>relative mass / relative charge<br>proton: 1, +1<br>neutron: 1, 0<br>electron: $5 \times 10^{-4}$, –1 |
| 2.1.3 | Define the terms *mass number* (*A*), *atomic number* (*Z*) and *isotopes of an element.* | 1 | |
| 2.1.4 | Deduce the symbol for an isotope given its mass number and atomic number. | 3 | The following notation should be used: $^{A}_{Z}X$ , for example, $^{12}_{6}C$ |
| 2.1.5 | Calculate the number of protons, neutrons and electrons in atoms and ions from the mass number, atomic number and charge. | 2 | |
| 2.1.6 | Compare the properties of the isotopes of an element. | 3 | |

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 2.1.7 | Discuss the uses of radioisotopes | 3 | Examples should include $^{14}C$ in radiocarbon dating, $^{60}Co$ in radiotherapy, and $^{131}I$ and $^{125}I$ as medical tracers.<br>**Aim 8:** Students should be aware of the dangers to living things of radioisotopes but also justify their usefulness with the examples above. |

## 2.2 The mass spectrometer

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 2.2.1 | Describe and explain the operation of a mass spectrometer. | 3 | 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.<br>**Aim 7:** Simulations can be used to illustrate the operation of a mass spectrometer. |
| 2.2.2 | Describe how the mass spectrometer may be used to determine relative atomic mass using the $^{12}C$ scale. | 2 | |
| 2.2.3 | Calculate non-integer relative atomic masses and abundance of isotopes from given data. | 2 | |

## 2.3 Electron arrangement

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 2.3.1 | Describe the electromagnetic spectrum. | 2 | 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.<br>**TOK:** Infrared and ultraviolet spectroscopy are dependent on technology for their existence. What are the knowledge implications of this? |
| 2.3.2 | Distinguish between a *continuous spectrum* and a *line spectrum*. | 2 | |
| 2.3.3 | Explain how the lines in the emission spectrum of hydrogen are related to electron energy levels. | 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 will not be assessed.<br>**Aim 7:** Interactive simulations modelling the behaviour of electrons in the hydrogen atom can be used. |

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 2.3.4 | Deduce the electron arrangement for atoms and ions up to $Z = 20$. | 3 | For example, 2.8.7 or 2,8,7 for $Z = 17$.<br>**TOK:** In drawing an atom, we have an image of an invisible world. Which ways of knowing allow us access to the microscopic world? |

# Topic 3: Periodicity (6 hours)

**TOK:** The early discoverers of the elements allowed chemistry to make great steps with limited apparatus, often derived from the pseudoscience of alchemy. Lavoisier's work with oxygen, which overturned the phlogiston theory of heat, could be discussed as an example of a paradigm shift.

**Int:** The discovery of the elements and the arrangement of them is a story that exemplifies how scientific progress is made across national boundaries by the sharing of information.

## 3.1 The periodic table

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 3.1.1 | Describe the arrangement of elements in the periodic table in order of increasing atomic number. | 2 | Names and symbols of the elements are given in the *Chemistry data booklet*. The history of the periodic table will not be assessed.<br>**TOK:** The predictive power of Mendeleev's periodic table could be emphasized. He is an example of a "scientist" as a "risk taker". |
| 3.1.2 | Distinguish between the terms *group* and *period*. | 2 | The numbering system for groups in the periodic table is shown in the *Chemistry data booklet*. Students should also be aware of the position of the transition elements in the periodic table. |
| 3.1.3 | Apply the relationship between the electron arrangement of elements and their position in the periodic table up to $Z = 20$. | 2 | |
| 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. | 2 | |

## 3.2 Physical properties

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 3.2.1 | Define the terms *first ionization energy* and *electronegativity*. | 1 | |
| 3.2.2 | Describe and explain the trends in atomic radii, ionic radii, first ionization energies, electronegativities and melting points for the alkali metals ( $Li \rightarrow Cs$ ) and the halogens ( $F \rightarrow I$ ). | 3 | Data for all these properties is listed in the *Chemistry 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. |
| 3.2.3 | Describe and explain the trends in atomic radii, ionic radii, first ionization energies and electronegativities for elements across period 3. | 3 | **Aim 7:** Databases and simulations can be used here. |
| 3.2.4 | Compare the relative electronegativity values of two or more elements based on their positions in the periodic table. | 3 | |

## 3.3 Chemical properties

**3 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 3.3.1 | Discuss the similarities and differences in the chemical properties of elements in the same group. | 3 | The following reactions should be covered.<br>• Alkali metals (Li, Na and K) with water<br>• Alkali metals (Li, Na and K) with halogens ($Cl_2$, $Br_2$ and $I_2$)<br>• Halogens ($Cl_2$, $Br_2$ and $I_2$) with halide ions ($Cl^-$, $Br^-$ and $I^-$) |
| 3.3.2 | Discuss the changes in nature, from ionic to covalent and from basic to acidic, of the oxides across period 3. | 3 | Equations are required for the reactions of $Na_2O$, $MgO$, $P_4O_{10}$ and $SO_3$ with water.<br>**Aim 8:** Non-metal oxides are produced by many large-scale industrial processes and the combustion engine. These acidic gases cause large-scale pollution to lakes and forests, and localized pollution in cities. |

# Topic 4: Bonding (12.5 hours)

## 4.1 Ionic bonding

**2 hours**

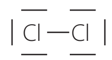
| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 4.1.1 | Describe the ionic bond as the electrostatic attraction between oppositely charged ions. | 2 | |
| 4.1.2 | Describe how ions can be formed as a result of electron transfer. | 2 | |
| 4.1.3 | Deduce which ions will be formed when elements in groups 1, 2 and 3 lose electrons. | 3 | |
| 4.1.4 | Deduce which ions will be formed when elements in groups 5, 6 and 7 gain electrons. | 3 | |
| 4.1.5 | State that transition elements can form more than one ion. | 1 | Include examples such as $Fe^{2+}$ and $Fe^{3+}$. |
| 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. | 3 | |
| 4.1.7 | State the formula of common polyatomic ions formed by non-metals in periods 2 and 3. | 1 | Examples include $NO_3^-$, $OH^-$, $SO_4^{2-}$, $CO_3^{2-}$, $PO_4^{3-}$, $NH_4^+$, $HCO_3^-$. |
| 4.1.8 | Describe the lattice structure of ionic compounds. | 2 | Students should be able to describe the structure of sodium chloride as an example of an ionic lattice. |

## 4.2 Covalent bonding

**6 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 4.2.1 | Describe the covalent bond as the electrostatic attraction between a pair of electrons and positively charged nuclei. | 2 | Single and multiple bonds should be considered. Examples should include $O_2$, $N_2$, $CO_2$, HCN, $C_2H_4$ (ethene) and $C_2H_2$ (ethyne). |
| 4.2.2 | Describe how the covalent bond is formed as a result of electron sharing. | 2 | Dative covalent bonds are required. Examples include CO, $NH_4^+$ and $H_3O^+$. |

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 4.2.3 | Deduce the Lewis (electron dot) structures of molecules and ions for up to four electron pairs on each atom. | 3 | 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:  or  or  or  Note: Cl–Cl is not a Lewis structure. |
| 4.2.4 | State and explain the relationship between the number of bonds, bond length and bond strength. | 3 | The comparison should include the bond lengths and bond strengths of:<br>• two carbon atoms joined by single, double and triple bonds<br>• the carbon atom and the two oxygen atoms in the carboxyl group of a carboxylic acid. |
| 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. | 3 | |
| 4.2.6 | Predict the relative polarity of bonds from electronegativity values | 3 | **Aim 7:** Simulations may be used here. |
| 4.2.7 | Predict the shape and bond angles for species with four, three and two negative charge centres on the central atom using the valence shell electron pair repulsion theory (VSEPR). | 3 | Examples should include $CH_4$, $NH_3$, $H_2O$, $NH_4^+$, $H_3O^+$, $BF_3$, $C_2H_4$, $SO_2$, $C_2H_2$ and $CO_2$.<br>**Aim 7:** Simulations are available to study the three-dimensional structures of these and the structures in 4.2.9 and 4.2.10. |
| 4.2.8 | Predict whether or not a molecule is polar from its molecular shape and bond polarities. | 3 | |
| 4.2.9 | Describe and compare the structure and bonding in the three allotropes of carbon (diamond, graphite and $C_{60}$ fullerene). | 3 | |
| 4.2.10 | Describe the structure of and bonding in silicon and silicon dioxide. | 2 | |

## 4.3 Intermolecular forces

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 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. | 3 | The term van der Waals' forces can be used to describe the interaction between non-polar molecules. |
| 4.3.2 | Describe and explain how intermolecular forces affect the boiling points of substances. | 3 | The presence of hydrogen bonding can be illustrated by comparing:<br>• HF and HCl<br>• $H_2O$ and $H_2S$<br>• $NH_3$ and $PH_3$<br>• $CH_3OCH_3$ and $CH_3CH_2OH$<br>• $CH_3CH_2CH_3$, $CH_3CHO$ and $CH_3CH_2OH$. |

## 4.4 Metallic bonding

**0.5 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 4.4.1 | Describe the metallic bond as the electrostatic attraction between a lattice of positive ions and delocalized electrons. | 2 | |
| 4.4.2 | Explain the electrical conductivity and malleability of metals. | 3 | **Aim 8:** Students should appreciate the economic importance of these properties and the impact that the large-scale production of iron and other metals has made on the world. |

## 4.5 Physical properties

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 4.5.1 | Compare and explain the properties of substances resulting from different types of bonding. | 3 | Examples should include melting and boiling points, volatility, electrical conductivity and solubility in non-polar and polar solvents. |

# Topic 5: Energetics (8 hours)

## 5.1 Exothermic and endothermic reactions

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 5.1.1 | Define the terms *exothermic reaction*, *endothermic reaction* and *standard enthalpy change of reaction* ( $\Delta H^{\ominus}$ ) . | 1 | Standard enthalpy change is the heat energy transferred under standard conditions—pressure 101.3 kPa, temperature 298 K. Only $\Delta H$ can be measured, not $H$ for the initial or final state of a system. |
| 5.1.2 | State that combustion and neutralization are exothermic processes. | 1 | |
| 5.1.3 | Apply the relationship between temperature change, enthalpy change and the classification of a reaction as endothermic or exothermic. | 2 | |
| 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. | 3 | |

## 5.2 Calculation of enthalpy changes

**3 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 5.2.1 | Calculate the heat energy change when the temperature of a pure substance is changed. | 2 | Students should be able to calculate the heat energy change for a substance given the mass, specific heat capacity and temperature change using $q = mc\Delta T$. |
| 5.2.2 | Design suitable experimental procedures for measuring the heat energy changes of reactions. | 3 | Students should consider reactions in aqueous solution and combustion reactions.<br>Use of the bomb calorimeter and calibration of calorimeters will not be assessed.<br>**Aim 7:** Data loggers and databases can be used here. |
| 5.2.3 | Calculate the enthalpy change for a reaction using experimental data on temperature changes, quantities of reactants and mass of water. | 2 | |
| 5.2.4 | Evaluate the results of experiments to determine enthalpy changes. | 3 | Students should be aware of the assumptions made and errors due to heat loss.<br>**TOK:** What criteria do we use in judging whether discrepancies between experimental and theoretical values are due to experimental limitations or theoretical assumptions? |

## 5.3 Hess's law

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 5.3.1 | Determine the enthalpy change of a reaction that is the sum of two or three reactions with known enthalpy changes. | 3 | 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.<br>**TOK:** As an example of the conservation of energy, this illustrates the unification of ideas from different areas of science. |

## 5.4 Bond enthalpies

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 5.4.1 | Define the term *average bond enthalpy*. | 1 | |
| 5.4.2 | Explain, in terms of average bond enthalpies, why some reactions are exothermic and others are endothermic. | 3 | |

# Topic 6: Kinetics (5 hours)

## 6.1 Rates of reaction

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 6.1.1 | Define the term *rate of reaction*. | 1 | |
| 6.1.2 | Describe suitable experimental procedures for measuring rates of reactions. | 2 | **Aim 7:** Data loggers can be used to collect data and produce graphs.<br>**TOK:** The empirical nature of the topic should be emphasized. Experimental results can support the theory but cannot prove it. |
| 6.1.3 | Analyse data from rate experiments. | 3 | Students should be familiar with graphs of changes in concentration, volume and mass against time. |

## 6.2 Collision theory

**3 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 6.2.1 | Describe the kinetic theory in terms of the movement of particles whose average energy is proportional to temperature in kelvins. | 2 | |
| 6.2.2 | Define the term *activation energy*, $E_a$. | 1 | |
| 6.2.3 | Describe the collision theory. | 2 | Students should know that reaction rate depends on:<br>• collision frequency<br>• number of particles with $E \geq E_a$<br>• appropriate collision geometry or orientation. |
| 6.2.4 | Predict and explain, using the collision theory, the qualitative effects of particle size, temperature, concentration and pressure on the rate of a reaction. | 3 | **Aim 7:** Interactive simulations can be used to demonstrate this. |
| 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. | 3 | Students should be able to explain why the area under the curve is constant and does not change with temperature.<br>**Aim 7:** Interactive simulations can be used to demonstrate this. |
| 6.2.6 | Describe the effect of a catalyst on a chemical reaction. | 2 | |
| 6.2.7 | Sketch and explain Maxwell–Boltzmann curves for reactions with and without catalysts. | 3 | |

# Topic 7: Equilibrium (5 hours)

## 7.1 Dynamic equilibrium

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 7.1.1 | Outline the characteristics of chemical and physical systems in a state of equilibrium. | 2 | **Aim 7:** Spreadsheets and simulations can be used here. |

## 7.2 The position of equilibrium

**4 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 7.2.1 | Deduce the equilibrium constant expression ($K_c$) from the equation for a homogeneous reaction. | 3 | Consider gases, liquids and aqueous solutions. |
| 7.2.2 | Deduce the extent of a reaction from the magnitude of the equilibrium constant. | 3 | When $K_c >> 1$, the reaction goes almost to completion.<br>When $K_c << 1$, the reaction hardly proceeds. |
| 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. | 2 | Students will not be required to state Le Chatelier's principle.<br>**Aim 7:** Simulations are available that model the behaviour of equilibrium systems. |
| 7.2.4 | State and explain the effect of a catalyst on an equilibrium reaction. | 3 | |
| 7.2.5 | Apply the concepts of kinetics and equilibrium to industrial processes. | 2 | Suitable examples include the Haber and Contact processes.<br>**Aim 8:** A case study of Fritz Haber could be included to debate the role of scientists in society. |

# Topic 8: Acids and bases (6 hours)

## 8.1 Theories of acids and bases

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 8.1.1 | Define *acids* and *bases* according to the Brønsted–Lowry and Lewis theories. | 1 | **TOK:** Discuss the value of using different theories to explain the same phenomenon. What is the relationship between depth and simplicity? |
| 8.1.2 | Deduce whether or not a species could act as a Brønsted–Lowry and/or a Lewis acid or base. | 3 | |
| 8.1.3 | Deduce the formula of the conjugate acid (or base) of any Brønsted–Lowry base (or acid). | 3 | Students should make clear the location of the proton transferred, for example, $CH_3COOH$/$CH_3COO^-$ rather than $C_2H_4O_2$/$C_2H_3O_2^-$. |

## 8.2 Properties of acids and bases

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 8.2.1 | Outline the characteristic properties of acids and bases in aqueous solution. | 2 | Bases that are not hydroxides, such as ammonia, soluble carbonates and hydrogencarbonates, should be included.<br>Alkalis are bases that dissolve in water.<br>Students should consider the effects on indicators and the reactions of acids with bases, metals and carbonates. |

## 8.3 Strong and weak acids and bases

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 8.3.1 | Distinguish between *strong* and *weak* acids and bases in terms of the extent of dissociation, reaction with water and electrical conductivity. | 2 | **Aim 8:** Although weakly acidic solutions are relatively safe, they still cause damage over long periods of time. Students could consider the effects of acid deposition on limestone buildings and living things. |
| 8.3.2 | State whether a given acid or base is strong or weak. | 1 | 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.<br>Students should consider all group 1 hydroxides and barium hydroxide as strong bases, and ammonia and amines as weak bases. |
| 8.3.3 | Distinguish between *strong* and *weak* acids and bases, and determine the relative strengths of acids and bases, using experimental data. | 2 | |

## 8.4 The pH scale

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 8.4.1 | Distinguish between aqueous solutions that are *acidic*, *neutral* or *alkaline* using the pH scale. | 2 | |
| 8.4.2 | Identify which of two or more aqueous solutions is more acidic or alkaline using pH values. | 2 | Students should be familiar with the use of a pH meter and universal indicator. |
| 8.4.3 | State that each change of one pH unit represents a 10-fold change in the hydrogen ion concentration $[H^+(aq)]$. | 1 | Relate integral values of pH to $[H^+(aq)]$ expressed as powers of 10.<br>Calculation of pH from $[H^+(aq)]$ is not required.<br>**TOK:** The distinction between artificial and natural scales could be discussed. |
| 8.4.4 | Deduce changes in $[H^+(aq)]$ when the pH of a solution changes by more than one pH unit. | 3 | **Aim 8:** A study of the effects of small pH changes in natural environments could be included. |

# Topic 9: Oxidation and reduction (7 hours)

**Aim 8:** The Industrial Revolution was the consequence of the mass production of iron by a reduction process. However, iron spontaneously reverts back to an oxidized form. What price do we continue to pay in terms of energy and waste for choosing a metal so prone to oxidation and why was it chosen?

## 9.1 Introduction to oxidation and reduction

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 9.1.1 | Define *oxidation* and *reduction* in terms of electron loss and gain. | 1 | |
| 9.1.2 | Deduce the oxidation number of an element in a compound. | 3 | Oxidation numbers should be shown by a sign (+ or –) and a number, for example, +7 for Mn in $KMnO_4$.<br>**TOK:** Are oxidation numbers "real"? |
| 9.1.3 | State the names of compounds using oxidation numbers. | 1 | Oxidation numbers in names of compounds are represented by Roman numerals, for example, iron(II) oxide, iron(III) oxide.<br>**TOK:** Chemistry has developed a systematic language that has resulted in older names becoming obsolete. What has been gained and lost in this process? |
| 9.1.4 | Deduce whether an element undergoes oxidation or reduction in reactions using oxidation numbers. | 3 | |

## 9.2 Redox equations

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 9.2.1 | Deduce simple oxidation and reduction half-equations given the species involved in a redox reaction. | 3 | |
| 9.2.2 | Deduce redox equations using half-equations. | 3 | $H^+$ 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. |
| 9.2.3 | Define the terms *oxidizing agent* and *reducing agent*. | 1 | |
| 9.2.4 | Identify the oxidizing and reducing agents in redox equations. | 2 | |

## 9.3 Reactivity

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 9.3.1 | Deduce a reactivity series based on the chemical behaviour of a group of oxidizing and reducing agents. | 3 | Examples include displacement reactions of metals and halogens. Standard electrode potentials will not be assessed. |
| 9.3.2 | Deduce the feasibility of a redox reaction from a given reactivity series. | 3 | Students are not expected to recall a specific reactivity series. |

## 9.4 Voltaic cells

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 9.4.1 | Explain how a redox reaction is used to produce electricity in a voltaic cell. | 3 | 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. |
| 9.4.2 | State that oxidation occurs at the negative electrode (anode) and reduction occurs at the positive electrode (cathode). | 1 | |

## 9.5 Electrolytic cells

**2 hours**

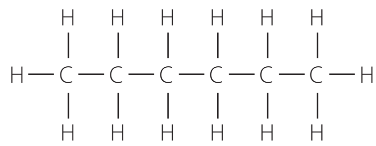
| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 9.5.1 | Describe, using a diagram, the essential components of an electrolytic cell. | 2 | The diagram should include the source of electric current and conductors, positive and negative electrodes, and the electrolyte. |
| 9.5.2 | State that oxidation occurs at the positive electrode (anode) and reduction occurs at the negative electrode (cathode). | 1 | |
| 9.5.3 | Describe how current is conducted in an electrolytic cell. | 2 | |
| 9.5.4 | Deduce the products of the electrolysis of a molten salt. | 3 | Half-equations showing the formation of products at each electrode will be assessed.<br>**Aim 8:** This process (which required the discovery of electricity) has made it possible to obtain reactive metals such as aluminium from their ores. This in turn has enabled subsequent steps in engineering and technology that increase our quality of life. Unlike iron, aluminium is not prone to corrosion and is one material that is replacing iron in many of its applications. |

# Topic 10: Organic chemistry (12 hours)

**Int, aim 8:** Today, we may be starting to experience the consequences of using fossil fuels as our main source of energy. There is a vast range of products that can be derived from fossil fuels as a result of carbon's rich chemistry. This raises the question "are they too valuable to burn?".

## 10.1 Introduction

**4 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 10.1.1 | Describe the features of a homologous series. | 2 | Include the same general formula, neighbouring members differing by $CH_2$, similar chemical properties and gradation in physical properties. |
| 10.1.2 | Predict and explain the trends in boiling points of members of a homologous series. | 3 | |
| 10.1.3 | Distinguish between *empirical*, *molecular* and *structural* formulas. | 2 | A structural formula is one that shows unambiguously how the atoms are arranged together.<br>A full structural formula (sometimes called a graphic formula or displayed formula) shows every atom and bond, for example, for hexane:<br><br>A condensed structural formula can omit bonds between atoms and can show identical groups bracketed together, for example, for hexane:<br>$CH_3CH_2CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_4CH_3$.<br>The use of R to represent an alkyl group and  to represent the benzene ring can be used in condensed structural formulas.<br>Although skeletal formulas are used for more complex structures in the *Chemistry data booklet*, such formulas will not be accepted in examination answers.<br>**TOK:** The use of the different formulas illustrates the value of different models with different depths of detail. |
| 10.1.4 | Describe structural isomers as compounds with the same molecular formula but with different arrangements of atoms. | 2 | No distinction need be made between different types of structural isomerism, such as chain and position isomerism and functional group isomerism. Knowledge of stereoisomerism is not required in the core. |
| 10.1.5 | Deduce structural formulas for the isomers of the non-cyclic alkanes up to $C_6$. | 3 | Include both straight-chain and branched-chain isomers. |
| 10.1.6 | Apply IUPAC rules for naming the isomers of the non-cyclic alkanes up to $C_6$. | 2 | **TOK:** This could be discussed as an example of the use of the language of chemistry as a tool to classify and distinguish between different structures. |

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 10.1.7 | Deduce structural formulas for the isomers of the straight-chain alkenes up to $C_6$. | 3 | |
| 10.1.8 | Apply IUPAC rules for naming the isomers of the straight-chain alkenes up to $C_6$. | 2 | The distinction between *cis* and *trans* isomers is not required. |
| 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. | 3 | Condensed structural formulas can use OH, CHO, CO, COOH and F/Cl/Br/I. |
| 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. | 2 | |
| 10.1.11 | Identify the following functional groups when present in structural formulas: amino ($NH_2$), benzene ring (  ) and esters (RCOOR). | 2 | |
| 10.1.12 | Identify primary, secondary and tertiary carbon atoms in alcohols and halogenoalkanes. | 2 | The terms primary, secondary and tertiary can also be applied to the molecules containing these carbon atoms. |
| 10.1.13 | Discuss the volatility and solubility in water of compounds containing the functional groups listed in 10.1.9. | 3 | |

## 10.2 Alkanes

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 10.2.1 | Explain the low reactivity of alkanes in terms of bond enthalpies and bond polarity. | 3 | |
| 10.2.2 | Describe, using equations, the complete and incomplete combustion of alkanes. | 2 | |
| 10.2.3 | Describe, using equations, the reactions of methane and ethane with chlorine and bromine. | 2 | |
| 10.2.4 | Explain the reactions of methane and ethane with chlorine and bromine in terms of a free-radical mechanism. | 3 | Reference should be made to homolytic fission and the reaction steps of initiation, propagation and termination.<br>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, $Cl\bullet$ . |

## 10.3 Alkenes

**2 hours**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 10.3.1 | Describe, using equations, the reactions of alkenes with hydrogen and halogens. | 2 | |
| 10.3.2 | Describe, using equations, the reactions of symmetrical alkenes with hydrogen halides and water. | 2 | |
| 10.3.3 | Distinguish between *alkanes* and *alkenes* using bromine water. | 2 | |
| 10.3.4 | Outline the polymerization of alkenes. | 2 | Include the formation of poly(ethene), poly(chloroethene) and poly(propene) as examples of addition polymers.<br>Include the identification of the repeating unit, for example, $-(-CH_2-CH_2-)_n-$ for poly(ethene). |
| 10.3.5 | Outline the economic importance of the reactions of alkenes. | 2 | **Aim 8:** Include the hydrogenation of vegetable oils in the manufacture of margarine, the hydration of ethene in the manufacture of ethanol, and polymerization in the manufacture of plastics. |

## 10.4 Alcohols

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 10.4.1 | Describe, using equations, the complete combustion of alcohols. | 2 | |
| 10.4.2 | Describe, using equations, the oxidation reactions of alcohols. | 2 | A suitable oxidizing agent is acidified potassium dichromate(VI).<br>Equations may be balanced using the symbol [O] to represent oxygen supplied by the oxidizing agent.<br>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. |
| 10.4.3 | Determine the products formed by the oxidation of primary and secondary alcohols. | 3 | Assume that tertiary alcohols are not oxidized by potassium dichromate(VI). |

## 10.5 Halogenoalkanes

**2 hours**

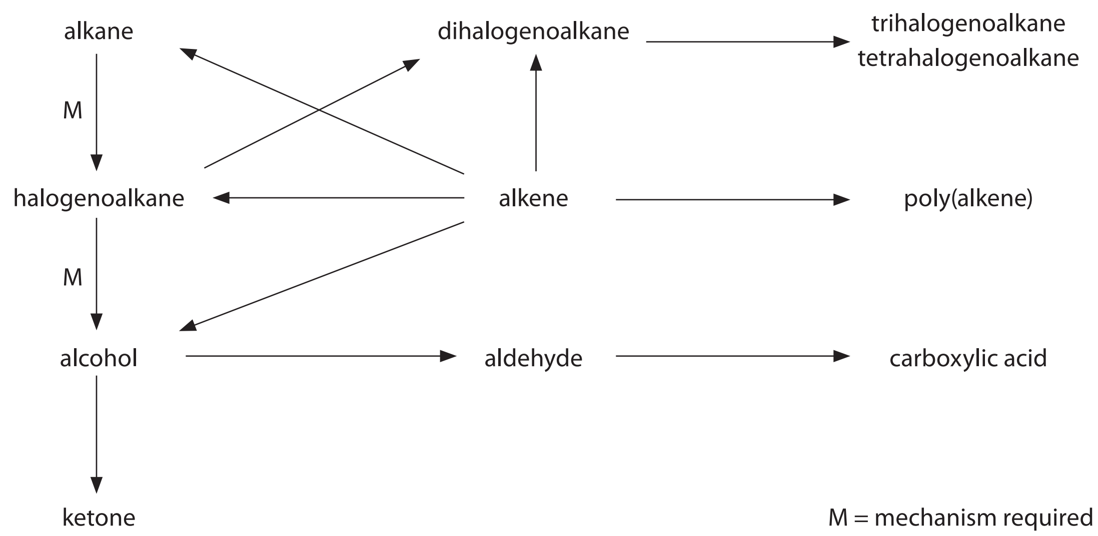
| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 10.5.1 | Describe, using equations, the substitution reactions of halogenoalkanes with sodium hydroxide. | 2 | **Aim 7:** Simulations are available for this. |
| 10.5.2 | Explain the substitution reactions of halogenoalkanes with sodium hydroxide in terms of $S_N1$ and $S_N2$ mechanisms. | 3 | Reference should be made to heterolytic fission. Curly arrows should be used to represent the movement of electron pairs. For tertiary halogenoalkanes the predominant mechanism is $S_N1$, and for primary halogenoalkanes it is $S_N2$. Both mechanisms occur for secondary halogenoalkanes. |

## 10.6 Reaction pathways

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 10.6.1 | Deduce reaction pathways given the starting materials and the product. | 3 | Conversions with more than two stages will not be assessed. Reagents, conditions and equations should be included. 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. |

The compound and reaction types in this topic are summarized in the following scheme:



# Topic 11: Measurement and data processing (2 hours)

## 11.1 Uncertainty and error in measurement

**1 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 11.1.1 | Describe and give examples of random uncertainties and systematic errors. | 2 | |
| 11.1.2 | Distinguish between *precision* and *accuracy*. | 2 | 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). |
| 11.1.3 | Describe how the effects of random uncertainties may be reduced. | 2 | Students should be aware that random uncertainties, but not systematic errors, are reduced by repeating readings. |
| 11.1.4 | State random uncertainty as an uncertainty range (±). | 1 | |
| 11.1.5 | State the results of calculations to the appropriate number of significant figures. | 1 | The number of significant figures in any answer should reflect the number of significant figures in the given data. |

## 11.2 Uncertainties in calculated results

**0.5 hour**

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 11.2.1 | State uncertainties as absolute and percentage uncertainties. | 1 | |
| 11.2.2 | Determine the uncertainties in results. | 3 | 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. |

## 11.3 Graphical techniques

**0.5 hour**

**TOK:** Why are graphs helpful in providing powerful interpretations of reality.

| | Assessment statement | Obj | Teacher's notes |
|---|---|---|---|
| 11.3.1 | Sketch graphs to represent dependences and interpret graph behaviour. | 3 | Students should be able to give a qualitative physical interpretation of a particular graph, for example, the variables are proportional or inversely proportional. |
| 11.3.2 | Construct graphs from experimental data. | 3 | This involves the choice of axes and scale, and the plotting of points.<br>**Aim 7:** Software graphing packages could be used. |
| 11.3.3 | Draw best-fit lines through data points on a graph. | 1 | These can be curves or straight lines. |
| 11.3.4 | Determine the values of physical quantities from graphs. | 3 | Include measuring and interpreting the slope (gradient), and stating the units for these quantities. |