

**Anglo-Chinese School (Independent)**  
**International Baccalaureate Diploma Programme**  
**Scheme Of Work – Year 5 Chemistry HL**

**Topic 1 Quantitative Chemistry**

**Topic 11 Measurement and Data Processing**

**Topic 9 Redox equation**

1.1 The mole concept and Avogadro's constant

- a) Determine the number of particles and the amount of substance (in moles).

1.2 Formulas

- a) Define the terms relative atomic mass ( $A_r$ ) and relative molecular mass ( $M_r$ ).
- b) Solve problems involving the relationship between the amount of substance in moles, mass and molar mass.
- c) Determine the empirical and molecular formula from the percentage composition or from other experimental data.

1.3 Chemical Equation

- a) Deduce chemical equations when all reactants and products are given.
- b) Identify the mole ratio of any two species in a chemical equation.

1.4 Solutions

- a) Solve problems involving concentration, amount of solute and volume of solution.
- b) Calculate theoretical yields from chemical equations.
- c) Determine the limiting reactant and the reactant in excess when quantities of reacting substances are given.
- d) Solve problems involving theoretical, experimental and percentage yield.

1.5 Mass and gaseous volume relationships in chemical reactions

- a) Apply Avogadro's law to calculate reacting volumes of gases.
- b) Apply the concept of molar volume at standard temperature and pressure in calculations.
- c) Solve problems involving the relationship between temperature, pressure and volume for a fixed mass of an ideal gas.
- d) Solve problems using the ideal gas equation,  $PV = nRT$
- e) Analyse graphs relating to the ideal gas equation.

Topic 11 Data processing and measurement

11.1 Uncertainty and error in measurement

11.2 Uncertainties in calculated results

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9.1 Introduction to oxidation and reduction

- a) Deduce the oxidation number of an element in a compound.
- b) Deduce whether an element undergoes oxidation or reduction in reactions using oxidation numbers.

9.2 Redox equations

- a) Deduce simple oxidation and reduction half-equations given the species involved in a redox reaction.
- b) Deduce redox equations using half equations. The balancing of equations for reactions in alkaline solution will not be assessed but must be covered as enrichment.
- c) Identify the oxidizing and reducing agents in redox equations.

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**Scheme Of Work – Year 5 Chemistry HL**

**Topic 2 and 12 Atomic structure**

2.1 The atom

- a) Define the terms mass number ( $A$ ), atomic number ( $Z$ ) and isotopes of an element.

2.2 The mass spectrometer

- a) Describe and explain the operation of a mass spectrometer.
- b) Describe how the mass spectrometer may be used to determine relative atomic mass using the  $^{12}\text{C}$  scale.
- c) Calculate non-integer relative atomic masses and abundance of isotopes from given data.

2.3 Electron arrangement

- a) Describe the electromagnetic spectrum by describing the variation in wavelength, frequency and energy across the spectrum.
- b) Distinguish between a continuous spectrum and a line spectrum.

2.3 Electron arrangement

- c) Explain how the lines in the emission spectrum of hydrogen are related to electron energy levels. An understanding of convergence is expected. Series should be considered in the ultraviolet
- d) Solve problems using  $E = hv$ , calculate the value of the first ionization energy from spectral data which gives the wavelength or frequency of the convergence limit. (the value of Planck's constant( $h$ ) and  $E = hv$  are given in the data booklet in sections 1 & 2)

12.1 Electron Configuration

- a) Explain how evidence from first ionization energies across periods accounts for the existence of main energy levels and sub-levels in atoms.
- b) Explain how successive ionization energy data is related to the electron configuration of an atom.
- c) State the relative energies of s, p, d and f orbitals in a single energy level.
- d) State the maximum number of orbitals in a given energy level.
- e) Draw the shape of an s orbital and the shapes of the  $p_x$ ,  $p_y$  and  $p_z$  orbitals.
- f) Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to  $Z = 54$ .

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**Topic 4 and 14 Chemical Bonding**

4.4 Metallic bonding

- a) Describe the metallic bond as the electrostatic attraction between a lattice of positive ions and delocalized electrons.
- b) Explain the electrical conductivity and malleability of metals.

4.1 Ionic bonding

- a) Describe the ionic bond as the electrostatic attraction between oppositely charged ions.
- b) 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.
- c) Describe the lattice structure of ionic compounds.

4.2 Covalent bonding

- a) Describe the covalent bond as the electrostatic attraction between a pair of electrons and positively charged nuclei.
- b) Deduce the Lewis structures of molecules and ions.
- c) State and explain the relationship between the number of bonds, bond length and bond strength.
  
- d) 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.
- e) Predict the relative polarity of bonds from electronegativity values.
- f) Describe and compare the structure and bonding in the three allotropes of carbon (diamond, graphite and C<sub>60</sub> fullerene).
- g) Describe the structure of and bonding in silicon and silicon dioxide.

4.3 and 14.1 Shapes of molecules

- a) Predict the shape and bond angles using the valence shell electron pair repulsion theory (VSEPR).
- b) Formal charge(FC) can be used to decide which Lewis structure is preferred from several.
- c) Predict whether or not a molecule is polar from its molecular shape and bond polarities.

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4.4 Intermolecular forces

- a) Describe the types of intermolecular forces (attractions between molecules that have London dispersion forces, dipole-dipole forces, or hydrogen bonding) and explain how they arise from the structural features of molecules.
- b) Describe and explain how intermolecular forces affect the boiling points of substances.

4.5 Physical properties

- a) Compare and explain the properties of substances resulting from different types of bonding. Examples should include melting and boiling points, volatility, electrical conductivity and solubility in non-polar and polar solvents.

14.2 Hybridization

- a) Describe  $\sigma$  and  $\pi$  bonds.
- b) Explain hybridization in terms of the mixing of atomic orbitals to form new orbitals for bonding.
- c) Identify and explain the relationships between Lewis structures, molecular shapes and types of hybridization ( $sp$ ,  $sp^2$  and  $sp^3$ ).

14.3 Delocalization of electrons

- a) Describe the delocalization of  $\pi$  electrons and explain how this can account for the structures of some species. Examples should include  $\text{NO}_3^-$ ,  $\text{NO}_2^-$ ,  $\text{CO}_3^{2-}$ ,  $\text{O}_3$ ,  $\text{RCOO}^-$  and benzene.

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**Topic 3 and 13 Periodicity**

3.1 The periodic table

- a) Apply the relationship between the electron arrangement of elements and their position in the periodic table up to  $Z = 20$ .
- b) Apply the relationship between the number of electrons in the highest occupied energy level for an element and its position in the periodic table.

3.2 Physical properties

- a) Define the terms first ionization energy and electronegativity.
- b) Describe and explain the trends in atomic radii, ionic radii, first ionization energies, electronegativities and melting points for the alkali metals and the halogens and for elements in period 3.
- c) Compare the relative electronegativity values of two or more elements based on their positions in the periodic table.

3.3 Chemical properties

- a) Discuss the similarities and differences in the chemical properties of elements in the same group. The following reactions should be covered: Alkali metals with water; Alkali metals with halogens; Halogens with halide ions.
- b) Discuss the similarities and differences in the chemical properties of elements in the same group. Equations are required for the reactions of  $\text{Na}_2\text{O}$ ,  $\text{MgO}$ ,  $\text{P}_4\text{O}_{10}$  and  $\text{SO}_3$  with water, may teach more for enrichment.

13.1 Trends across period 3

- a) 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.

13.2 First-row d-block elements

- a) List the characteristic properties of transition elements, including magnetic properties.
- b) Explain the magnetic properties in transition metals in terms of unpaired electrons.
- c) Explain why Zn is not considered to be transition elements. (Scandium is now included as transition metal).
- d) Explain the existence of variable oxidation number in ions of transition elements.
- e) Define the term *ligand*.
- f) Describe and explain the formation of complexes of d-block elements.
- g) Explain why some complexes of d-block elements are coloured.
- h) Explain the effect of the identity of the metal ions, the oxidation number of the metal and the identity of the ligand on the colour of the transition metal ion complexes.
- i) The colour observed and absorbed is illustrated by the colour wheel in the data booklet in section 17.
- j) State examples of the catalytic action of transition elements and their compounds. Examples should include:•  $\text{MnO}_2$  in the decomposition of hydrogen Peroxide,  $\text{V}_2\text{O}_5$  in the Contact process, Fe in the Haber process and in heme, Ni in the conversion of alkenes to alkanes, Co in vitamin B12 and Pd and Pt in catalytic converters

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**Topic 5 and 15 Energetics**

5.1 Exothermic and endothermic reactions

5.2 Calculation of enthalpy changes

- a) Design suitable experimental procedures for measuring the heat energy changes of reactions.
- b) Calculate the enthalpy change for a reaction using experimental data on temperature changes, quantities of reactants and mass of water.
- c) Evaluate the results of experiments to determine enthalpy changes.

15.1 Standard enthalpy changes of reaction

- a) Define and apply the terms standard state, standard enthalpy change of formation and standard enthalpy change of combustion.
- b) Determine the enthalpy change of a reaction using standard enthalpy changes of formation and combustion.

5.4 Bond enthalpies

- a) Define the term average bond enthalpy.
- b) Explain, in terms of average bond enthalpies, why some reactions are exothermic and others are endothermic.

5.3 Hess's Law

- a) Determine the enthalpy change of a reaction that is the sum of two or three reactions with known enthalpy changes.

15.2 Born-Haber cycle

- a) Define and apply the terms lattice enthalpy and electron affinity.
- b) Explain how the relative sizes and the charges of ions affect the lattice enthalpies of different ionic compounds.
- c) Construct a Born–Haber cycle for group 1 and 2 oxides and chlorides, and use it to calculate an enthalpy change, include hydration enthalpy and enthalpy of solution.
- d) Discuss the difference between theoretical and experimental lattice enthalpy values of ionic compounds in terms of their covalent character.

15.3 Entropy

- a) State and explain the factors that increase the entropy in a system.
- b) Predict whether the entropy change ( $\Delta S$ ) for a given reaction or process is positive or negative.
- c) Calculate the standard entropy change for a reaction using standard entropy values .

15.4 Spontaneity

- a) Predict whether a reaction or process will be spontaneous by using the sign of  $\Delta G$  .
- b) Calculate  $\Delta G$  for a reaction using the equation  $\Delta G = \Delta H - T\Delta S$  and by using values of the standard free energy change of formation,  $\Delta G_f^\ominus$
- c) Predict the effect of a change in temperature on the spontaneity of a reaction using standard entropy and enthalpy changes and the equation:  $\Delta G = \Delta H - T\Delta S$



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**Topic 6 and 16 Kinetics**

6.1 Rates of reaction

- a) Describe suitable experimental procedures for measuring rates of reactions.
- b) Analyse data from rate experiments.

6.2 Collision theory

- a) Define the term activation energy,  $E_a$ .
- b) Describe the collision theory.
- c) Predict and explain, using the collision theory, the qualitative effects of particle size, temperature, concentration and pressure on the rate of a reaction.
- d) Describe the effect of a catalyst on a chemical reaction.
- e) Sketch and explain qualitatively the Maxwell–Boltzmann energy distribution curve for a fixed amount of gas at different temperatures or with catalyst added and its consequences for changes in reaction rate

16.1 Rate expression

- a) Distinguish between the terms rate constant, overall order of reaction and order of reaction with respect to a particular reactant.
- b) Deduce and solve the rate expression for a reaction from experimental data.
- c) Sketch, identify and analyse graphical representations for zero-, first- and second-order reactions.

16.2 Reaction mechanism

- a) Explain that reactions can occur by more than one step and that the slowest step determines the rate of reaction (rate-determining step).
- b) Describe the relationship between reaction mechanism, order of reaction and rate-determining step.

16.3 Activation energy

- a) Describe qualitatively the relationship between the rate constant ( $k$ ) and temperature ( $T$ ).
- b) Determine activation energy ( $E_a$ ) values from the Arrhenius equation by a graphical method.
- c) Use energy level diagrams to illustrate multi-step reactions showing the RDS in the diagram.

To include teaching of heterogeneous and homogeneous catalysts

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**Topic 7 and 17 Equilibrium**

7.1 Dynamic equilibrium

- a) Outline the characteristics of chemical and physical systems in a state of equilibrium.

7.2 The position of equilibrium

- a) Deduce the equilibrium constant expression ( $K_c$ ) from the equation for a homogeneous reaction.
- b) Deduce the extent of a reaction from the magnitude of the equilibrium constant.
- c) 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.
- d) State and explain the effect of a catalyst on an equilibrium reaction.
- e) Apply the concepts of kinetics and equilibrium to industrial processes. Suitable examples include the Haber and Contact processes.

17.1 The equilibrium law

- a) Solve homogeneous equilibrium problems using the expression for  $K_c$
- b) Calculations using the equation  $\Delta G = -RT \ln K$

17.2 Reaction quotient ( $Q$ )

- a)  $Q$  measures the relative amount of products and reactants present during a reaction.