Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ **Preparing for the IB Chemistry Exam** <https://vhsibchem.weebly.com>

**Highlight** key words for definitions. Circle parts you don’t know/understand. Write notes on what you can remember. Check these notes against your study guide. If you watch any Richard Thornley videos or do the questionbank. Revise your notes.

**Topic 1: Stoichiometric Relationships – 13.5 Hours**

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| **IB Points to Understand** | | **Notes** | **Video?** | **QB?** |
| 1.1 | ***Introduction to the particulate nature of matter and chemical change***   * "Atoms of different elements combine in fixed ratios to form compounds, which have different properties from their component elements." * "Mixtures contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties." * "Mixtures are either homogeneous or heterogeneous."   ***Applications and skills***  • Deduction of chemical equations when reactants and products are specified.  • Application of the state symbols (s), (l), (g) and (aq) in equations.  • Explanation of observable changes in physical properties and temperature  during changes of state.  ***Guidance***  •Balancing of equations should include a variety of types of reactions.  •Names of the changes o state—melting, freezing, vaporization (evaporation and boiling) condensation, sublimation and deposition—should be covered |  |  |  |
| 1.2 | ***The mole concept***   * "The mole is a fixed number of particles and refers to the amount, n, of substance." * "Masses of atoms are compared on a scale relative to 12C and are expressed as relative atomic mass (Ar) and relative formula/molecular mass (Mr)." * "Molar mass (M) has the units g mol-1." * "The empirical formula and molecular formula of a compound give the simplest ratio and the actual number of atoms present in a molecule respectively."   ***Applications and skills***  • Calculation of the molar masses of atoms, ions, molecules and formula units.  • Solution of problems involving the relationships between the number of particles, the amount of substance in moles and the mass in grams.  • Interconversion of the percentage composition by mass and the empirical formula.  • Determination of the molecular formula of a compound from its empirical formula and molar mass.  • Obtaining and using experimental data for deriving empirical formulas from reactions involving mass changes.  ***Guidance***  •The value of the Avogadro’s constant (L or NA) is given in the data booklet in section 2 and will be given for paper 1 questions.  •The generally used unit of molar mass (g mol-1) is a derived SI unit. |  |  |  |
| 1.3 | ***Reacting masses and volumes***   * "Reactants can be either limiting or excess." * "The experimental yield can be different from the theoretical yield." * "Avogadro’s law enables the mole ratio of reacting gases to be determined from volumes of the gases." * "The molar volume of an ideal gas is a constant at specified temperature and pressure." * "The molar concentration of a solution is determined by the amount of solute and the volume of solution." * "A standard solution is one of known concentration."   ***Applications and skills***  • Solution of problems relating to reacting quantities, limiting and excess reactants, theoretical, experimental and percentage yields.  • Calculation of reacting volumes of gases using Avogadro’s law.  • Solution of problems and analysis of graphs involving the relationship between temperature, pressure and volume for a fixed mass of an ideal gas.  • Solution of problems relating to the ideal gas equation.  • Explanation of the deviation of real gases from ideal behaviour at low temperature and high pressure.  • Obtaining and using experimental values to calculate the molar mass of a gas from the ideal gas equation.  • Solution of problems involving molar concentration, amount of solute and volume of solution.  • Use of the experimental method of titration to calculate the concentration of a solution by reference to a standard solution.  ***Guidance***  • Values for the molar volume of an ideal gas are given in the data booklet in section 2.  •The ideal gas equation, PV=nRT, and the value of the gas constant (R) can be found in the databook.  Units of concentration to include: g dm-3, mol dm-3  and parts per million (ppm).  • The use of square brackets to denote molar concentration is required |  |  |  |

**Topic #2: Atomic Structure - 6 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 2.1 | ***The nuclear atom***   * "Atoms contain a positively charged dense nucleus composed of protons and neutrons (nucleons)." * "Negatively charged electrons occupy the space outside the nucleus." * "The mass spectrometer is used to determine the relative atomic mass of an element from its isotopic composition."   ***Applications and skills***  •Use of the nuclear symbol notation to deduce the number of protons, neutrons and electrons in atoms and ions.  •Calculations involving non-integer relative atomic masses and abundance of isotopes from given data, including mass spectra.  ***Guidance***  •Relative masses and charges of the subatomic particles should be known, actual values are given in section 4 of the data booklet. The mass of the  electron can be considered negligible.  • Specific examples of isotopes need not be learned.  •The operation of the mass spectrometer is not required. |  |  |  |
| 2.2 | ***Electron configuration***   * "Emission spectra are produced when photons are emitted from atoms as excited electrons return to a lower energy level." * "The line emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels, which converge at higher energies." * "The main energy level or shell is given an integer number, n, and can hold a maximum number of electrons, 2n2." * "A more detailed model of the atom describes the division of the main energy level into s, p, d and f sub-levels of successively higher energies." * "Sub-levels contain a fixed number of orbitals, regions of space where there is a high probability of finding an electron." * "Each orbital has a defined energy state for a given electronic configuration and chemical environment and can hold two electrons of opposite spin."   ***Applications and skills***  • Description of the relationship between colour, wavelength, frequency and energy across the electromagnetic spectrum.  • Distinction between a continuous spectrum and a line spectrum  • Description of the emission spectrum of the hydrogen atom, including the relationships between the lines and energy transitions to the first, second and third energy levels.  •  Recognition of the shape of an s atomic orbital and the px, py and pz atomic orbitals.  • Application of the Aufbau principle, Hund’s rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to Z =36.  ***Guidance***  • Details of the electromagnetic spectrum are given in the data booklet in section 3.  •The names of the different series in the hydrogen line emission spectrum are not required.  •Full electron configurations (eg 1s22s22p63s23p4) and condensed electron configurations (eg [Ne] 3s23p4) should be covered. Orbital diagrams should be used to represent the character and relative energy of  orbitals. Orbital diagrams refer to arrow  -in-box diagrams, such as the one given  below.    •The electron configurations of Cr and Cu as exceptions should be covered. |  |  |  |

**Topic #3: Periodicity - 6 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 3.1 | ***Periodic table***   * "The periodic table is arranged into four blocks associated with the four sub- levels—s, p, d, and f." * "The periodic table consists of groups (vertical columns) and periods (horizontal rows)." * "The period number (n) is the outer energy level that is occupied by electrons." * "The number of the principal energy level and the number of the valence electrons in an atom can be deduced from its position on the periodic table." * "The periodic table shows the positions of metals, non-metals and metalloids."   ***Applications and skills***  • Deduction of the electron configuration of an atom from the element’s position on the periodic table, and vice versa.  ***Guidance***  • The terms alkali metals, halogens, noble gases, transition metals, lanthanoids and actinoids should be known.  • The group numbering scheme from group 1 to group 18, as recommended by IUPAC, should be used. |  |  |  |
| 3.2 | ***Periodic trends***   * "Vertical and horizontal trends in the periodic table exist for atomic radius, ionic radius, ionization energy, electron affinity and electronegativity." * "Trends in metallic and non-metallic behaviour are due to the trends above." * "Oxides change from basic through amphoteric to acidic across a period."   ***Applications and skills***  • Prediction and explanation of the metallic and non -metallic behaviour of an element based on its position in the periodic table.  •Discussion of the similarities and differences in the properties of elements in the same group, with reference to alkali metals (group 1) and halogens (group 17).  •Construction of equations to explain the pH changes for reactions of Na2O, MgO, P4O10, and the oxides of nitrogen and sulfur with water.  ***Guidance***  •Only examples of general trends across periods and down groups are required. For ionization energy the discontinuities in the increase across a period should be covered.  •Group trends should include the treatment of the reactions of alkali metals with water, alkali metals with halogens and halogens with halide ions |  |  |  |

**Topic #4: Chemical Bonding and Structure - 13.5 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 4.1 | ***Ionic bonding and structure***   * "Positive ions (cations) form by metals losing valence electrons." * "Negative ions (anions) form by non-metals gaining electrons." * "The number of electrons lost or gained is determined by the electron configuration of the atom." * "The ionic bond is due to electrostatic attraction between oppositely charged ions." * "Under normal conditions, ionic compounds are usually solids with lattice structures."   ***Applications and skills***  • Deduction of the formula and name of an ionic compound from its component ions, including polyatomic ions.  • Explanation of the physical properties of ionic compounds (volatility, electrical conductivity and solubility) in terms of their structure.  ***Guidance***  • Students should be familiar with the names of these polyatomic ions: NH4+, OH-, NO3-, HCO3-, CO32-, SO42- and PO43- |  |  |  |
| 4.2 | ***Covalent bonding***   * "A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei." * "Single, double and triple covalent bonds involve one, two and three shared pairs of electrons respectively." * "Bond length decreases and bond strength increases as the number of shared electrons increases." * "Bond polarity results from the difference in electronegativities of the bonded atoms."   ***Applications and skills***  • Deduction of the polar nature of a covalent bond from electronegativity values.  ***Guidance***  • Bond polarity can be shown either with partial charges, dipoles or vectors.  • Electronegativity values are given in the data booklet in section 8. |  |  |  |
| 4.3 | ***Covalent structures***   * "Lewis (electron dot) structures show all the valence electrons in a covalently bonded species." * "The “octet rule” refers to the tendency of atoms to gain a valence shell with a total of 8 electrons." * "Some atoms, like Be and B, might form stable compounds with incomplete octets of electrons." * "Resonance structures occur when there is more than one possible position for a double bond in a molecule." * "Shapes of species are determined by the repulsion of electron pairs according to VSEPR theory." * "Carbon and silicon form giant covalent/network covalent structures."   ***Applications and skills***  • Deduction of Lewis (electron dot) structure of molecules and ions showing all valence electrons for up to four electron pairs on each atom.  • The use of VSEPR theory to predict the electron domain geometry and the molecular geometry for species with two, three and four electron domains.  • Prediction of bond angles from molecular geometry and presence of non-bonding pairs of electrons.  • Prediction of molecular polarity from bond polarity and molecular geometry.  • Deduction of resonance structures, examples include but are not limited to C6H6, CO32-and O3.  • Explanation of the properties of giant covalent compounds in terms of their structures.  ***Guidance***  • The term “electron domain” should be used in place of “negative charge centre”.  • Electron pairs in a Lewis (electron dot) structure can be shown as dots, crosses, a dash or any combination.  • Allotropes of carbon (diamond, graphite, graphene, C60 buckminsterfullerene) and SiO2 should be covered.  • Coordinate covalent bonds should be covered. |  |  |  |
| 4.4 | ***Intermolecular forces***   * "Intermolecular forces include London (dispersion) forces, dipole-dipole forces and hydrogen bonding." * "The relative strengths of these interactions are London (dispersion) forces < dipole-dipole forces < hydrogen bonds."   ***Applications and skills***  • Deduction of the types of intermolecular force present in substances, based on their structure and chemical formula.  • Explanation of the physical properties of covalent compounds (volatility, electrical conductivity and solubility) in terms of their structure and intermolecular forces.  ***Guidance***  • The term “London (dispersion) forces” refers to instantaneous induced dipole-induced dipole forces that exist between any atoms or groups of atoms and should be used for non-polar entities. The term “van der Waals” is an inclusive term, which includes dipole–dipole, dipole-induced dipole and London (dispersion) forces. |  |  |  |
| 4.5 | ***Metallic bonding***   * "A metallic bond is the electrostatic attraction between a lattice of positive ions and delocalized electrons." * "The strength of a metallic bond depends on the charge of the ions and the radius of the metal ion." * "Alloys usually contain more than one metal and have enhanced properties."   ***Applications and skills:***  • Explanation of electrical conductivity and malleability in metals.  • Explanation of trends in melting points of metals. • Explanation of the properties of alloys in terms of non-directional bonding.  ***Guidance***  • Trends should be limited to s- and p-block elements.  • Examples of various alloys should be covered. |  |  |  |

**Topic #5: Energetics/Thermochemistry - 9 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 5.1 | ***Measuring energy changes***   * "Heat is a form of energy." * "Temperature is a measure of the average kinetic energy of the particles." * "Total energy is conserved in chemical reactions." * "Chemical reactions that involve transfer of heat between the system and the surroundings are described as endothermic or exothermic." * "The enthalpy change (∆H) for chemical reactions is indicated in kJ mol-1." * "∆H values are usually expressed under standard conditions, given by ∆H°, including standard states."   ***Applications and skills***  • Calculation of the heat change when the temperature of a pure substance is changed using 𝑞=cmΔT.  • A calorimetry experiment for an enthalpy of reaction should be covered and the results evaluated.  ***Guidance***  • Enthalpy changes of combustion (ΔHc0 ) and formation (ΔHf0)should be covered.  • Consider reactions in aqueous solution and combustion reactions.  • Standard state refers to the normal, most pure stable state of a substance measured at 100 kPa. Temperature is not a part of the definition of standard state, but 298 K is commonly given as the temperature of interest.  • The specific heat capacity of water is provided in the data booklet in section 2.  • Students can assume the density and specific heat capacities of aqueous solutions are equal to those of water but should be aware of this limitation.  • Heat losses to the environment and the heat capacity of the calorimeter in experiments should be considered, but the use of a bomb calorimeter is not required. |  |  |  |
| 5.2 | ***Hess’s Law***  “enthalpy change for a reaction that is carried out in a series of steps is equal to the sum of the enthalpy changes for the individual steps."  ***Applications and skills***  • Application of Hess’s Law to calculate enthalpy changes.  • Calculation of  • Determination of the enthalpy change of a reaction that is the sum of multiple reactions with known enthalpy changes.  ***Guidance***  • Enthalpy of formation data can be found in the data booklet in section 12. |  |  |  |
| 5.3 | • Bond-forming releases energy and bond-breaking requires energy.  • Average bond enthalpy is the energy needed to break one mol of a bond in a gaseous molecule averaged over similar compounds.  **Applications and skills:**  • Calculation of the enthalpy changes from known bond enthalpy values and comparison of these to experimentally measured values.  • Sketching and evaluation of potential energy profiles in determining whether reactants or products are more stable and if the reaction is exothermic or endothermic.  • Discussion of the bond strength in ozone relative to oxygen in its importance to the atmosphere.  **Guidance:**  • Bond enthalpy values are given in the data booklet in section 11.  • Average bond enthalpies are only valid for gases and calculations involving bond enthalpies may be inaccurate because they do not take into account intermolecular forces. |  |  |  |

**Topic #6: Chemical Kinetics - 7 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 6.1 | ***Collision theory and rates of reaction***   * "Species react as a result of collisions of sufficient energy and proper orientation." * "The rate of reaction is expressed as the change in concentration of a particular reactant/product per unit time." * "Concentration changes in a reaction can be followed indirectly by monitoring changes in mass, volume and colour." * "Activation energy (Ea) is the minimum energy that colliding molecules need in order to have successful collisions leading to a reaction." * "By decreasing Ea, a catalyst increases the rate of a chemical reaction, without itself being permanently chemically changed."   ***Applications and skills***  • Description of the kinetic theory in terms of the movement of particles whose average kinetic energy is proportional to temperature in Kelvin.  • Analysis of graphical and numerical data from rate experiments.  • Explanation of the effects of temperature, pressure/concentration and particle size on rate of reaction.  • Construction of Maxwell–Boltzmann energy distribution curves to account for  the probability of successful collisions and factors affecting these, including the effect of a catalyst.  • Investigation of rates of reaction experimentally and evaluation of the results.  • Sketching and explanation of energy profiles with and without catalysts.  ***Guidance***  • Calculation of reaction rates from tangents of graphs of concentration, volume or mass vs time should be covered.  • Students should be familiar with the interpretation of graphs of changes in concentration, volume or mass against time. |  |  |  |

**Topic #7: Equilibrium - 4.5 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 7.1 | * "A state of equilibrium is reached in a closed system when the rates of the forward and reverse reactions are equal." * "The equilibrium law describes how the equilibrium constant (Kc) can be determined for a particular chemical reaction." * "The magnitude of the equilibrium constant indicates the extent of a reaction at equilibrium and is temperature dependent." * "The reaction quotient (Q) measures the relative amount of products and reactants present during a reaction at a particular point in time. Q is the equilibrium expression with non-equilibrium concentrations. The position of the equilibrium changes with changes in concentration, pressure, and temperature." * "A catalyst has no effect on the position of equilibrium or the equilibrium constant."   ***Applications and skills***  • The characteristics of chemical and physical systems in a state of equilibrium.  • Deduction of the equilibrium constant expression (*Kc*) from an equation for a homogeneous reaction.  • Determination of the relationship between different equilibrium constants (*K*c) for the same reaction (at the same temperature) when represented by equations written in different ways.  • Application of Le Châtelier’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.  ***Guidance***  • Physical and chemical systems should be covered.  • Relationship between *Kc* values for reactions that are multiples or inverses of one another should be covered.  • Specific details of any industrial process are not required. |  |  |  |

**Topic #8: Acids and Bases - 6.5 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 8.1 | ***Theories of acids and bases***   * "A Brønsted–Lowry acid is a proton/H+ donor and a Brønsted–Lowry base is a proton/H+ acceptor." * "Amphiprotic species can act as both Brønsted–Lowry acids and bases." * "A pair of species differing by a single proton is called a conjugate acid-base pair."   ***Applications and skills***  • Deduction of the Brønsted–Lowry acid and base in a chemical reaction.  • Deduction of the conjugate acid or conjugate base in a chemical reaction.  ***Guidance***  • Lewis theory is not required here.  • The location of the proton transferred should be clearly indicated. For example, CH3COOH/CH3COO– rather than C2H4O2/C2H3O2-  • Students should know the representation of a proton in aqueous solution as both H+ (aq) and H3O+ (aq).  • The difference between the terms amphoteric and amphiprotic should be covered. |  |  |  |
| 8.2 | ***Properties of acids and bases***   * "Most acids have observable characteristic chemical reactions with reactive metals, metal oxides, metal hydroxides, hydrogen carbonates and carbonates." * "Salt and water are produced in exothermic neutralization reactions."   ***Applications and skills***  • Balancing chemical equations for the reaction of acids.  • Identification of the acid and base needed to make different salts.  • Candidates should have experience of acid-base titrations with different indicators.  ***Guidance***  • Bases which are not hydroxides, such as ammonia, soluble carbonates and hydrogen carbonates should be covered.  • The colour changes of different indicators are given in the data booklet in section 22. |  |  |  |
| 8.3 | ***The pH scale***   * "pH = − log[H+(aq)] and [H+] = 10−pH." * "A change of one pH unit represents a 10-fold change in the hydrogen ion concentration [H+]." * "pH values distinguish between acidic, neutral and alkaline solutions." * "The ionic product constant, K𝑤 = [H+][OH−] = 10−14 at 298 K."   ***Applications and skills***  • Solving problems involving pH, [H+] and [OH−].  • Students should be familiar with the use of a pH meter and universal indicator.  ***Guidance***  • Students will not be assessed on pOH values. • Students should be concerned only with strong acids and bases in this sub-topic.  • Knowing the temperature dependence of 𝐾𝑤 is not required.  • Equations involving H3O+ instead of H+ may be applied. |  |  |  |
| 8.4 | ***Strong and weak acids and bases***   * "Strong and weak acids and bases differ in the extent of ionization." * "Strong acids and bases of equal concentrations have higher conductivities than weak acids and bases." * "A strong acid is a good proton donor and has a weak conjugate base." * "A strong base is a good proton acceptor and has a weak conjugate acid."   ***Applications and skills***  • Distinction between strong and weak acids in terms of the rates of their reactions with metals, metal oxides, metal hydroxides, metal hydrogen carbonates and metal carbonates. Strong and weak acids and bases also differ in their electrical conductivities for solutions of equal concentrations.  ***Guidance***  • The terms ionization and dissociation can be used interchangeably.  • See section 21 in the data booklet for a list of weak acids and bases. |  |  |  |
| 8.5 | ***Acid deposition***   * "Rain is naturally acidic because of dissolved CO2 and has a pH of 5.6. Acid deposition has a pH below 5.6." * "Acid deposition is formed when nitrogen or sulfur oxides dissolve in water to form HNO3, HNO2, H2SO4 and H2SO3." * "Sources of the oxides of sulfur and nitrogen and the effects of acid deposition should be covered."   ***Applications and skills***  • Balancing the equations that describe the combustion of sulfur and nitrogen to their oxides and the subsequent formation of H2SO3, H2SO4, HNO2 and HNO3.  • Distinction between the pre-combustion and post-combustion methods of reducing sulfur oxides emissions.  • Deduction of acid deposition equations for acid deposition with reactive metals and carbonates. |  |  |  |

**Topic #9: Redox Processes - 8 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 9.1 | ***Oxidation and reduction***   * "Oxidation and reduction can be considered in terms of oxygen gain/hydrogen loss, electron transfer or change in oxidation number." * "An oxidizing agent is reduced and a reducing agent is oxidized." * "Variable oxidation numbers exist for transition metals and for most main-group non-metals." * "The activity series ranks metals according to the ease with which they undergo oxidation." * "The Winkler Method can be used to measure biochemical oxygen demand (BOD), used as a measure of the degree of pollution in a water sample."   ***Applications and skills***  *• Deduction* of the oxidation states of an atom in an ion or a compound.  • Deduction of the name of a transition metal compound from a given formula, applying oxidation numbers represented by Roman numerals.  • Identification of the species oxidized and reduced and the oxidizing and reducing agents, in redox reactions.  • Deduction of redox reactions using half-equations in acidic or neutral solutions.  • Deduction of the feasibility of a redox reaction from the activity series or reaction data.  • Solution of a range of redox titration problems.  • Application of the Winkler Method to calculate BOD.  ***Guidance***  • Oxidation number and oxidation state are often used interchangeably, though IUPAC does formally distinguish between the two terms. Oxidation numbers are represented by Roman numerals according to IUPAC.  • Oxidation states should be represented with the sign given before the number, e.g. +2 not 2+.  • The oxidation state of hydrogen in metal hydrides (-1) and oxygen in peroxides (-1) should be covered.  • A simple activity series is given in the data booklet in section 25. |  |  |  |
| 9.2 | ***Electrochemical cells***  *Voltaic (Galvanic) cells*   * "Voltaic cells convert energy from spontaneous, exothermic chemical processes to electrical energy." * "Oxidation occurs at the anode (negative electrode) and reduction occurs at the cathode (positive electrode) in a voltaic cell."   *Electrolytic cells*   * "Electrolytic cells convert electrical energy to chemical energy, by bringing about non-spontaneous processes." * "Oxidation occurs at the anode (positive electrode) and reduction occurs at the cathode (negative electrode) in an electrolytic cell."   ***Applications and skills***  • Construction and annotation of both types of electrochemical cells.  • Explanation of how a redox reaction is used to produce electricity in a voltaic cell and how current is conducted in an electrolytic cell.  • Distinction between electron and ion flow in both electrochemical cells.  • Performance of laboratory experiments involving a typical voltaic cell using two metal/metal-ion half-cells.  • Deduction of the products of the electrolysis of a molten salt.  ***Guidance***  • For voltaic cells, a cell diagram convention should be covered. |  |  |  |

**Topic #10: Organic Chemistry - 11 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 10.1 | ***Fundamentals of organic chemistry***   * "A homologous series is a series of compounds of the same family, with the same general formula, which differ from each other by a common structural unit." * "Structural formulas can be represented in full and condensed format." * "Structural isomers are compounds with the same molecular formula but different arrangements of atoms." * "Functional groups are the reactive parts of molecules." * "Saturated compounds contain single bonds only and unsaturated compounds contain double or triple bonds."      * "Benzene is an aromatic, unsaturated hydrocarbon."   ***Applications and skills***  • Explanation of the trends in boiling points of members of a homologous series.  • Distinction between empirical, molecular and structural formulas.  Identification of different classes: alkanes, alkenes, alkynes, halogenoalkanes, alcohols, ethers, aldehydes, ketones, esters, carboxylic acids, amines, amides, nitriles and arenes.  • Identification of typical functional groups in molecules eg phenyl, hydroxyl, carbonyl, carboxyl, carboxamide, aldehyde, ester, ether, amine, nitrile, alkyl, alkenyl and alkynyl.  • Construction of 3-D models (real or virtual) of organic molecules.  • Application of IUPAC rules in the nomenclature of straight-chain and branched-chain isomers.  • Identification of primary, secondary and tertiary carbon atoms in halogenoalkanes and alcohols and primary, secondary and tertiary nitrogen atoms in amines.  • Discussion of the structure of benzene using physical and chemical evidence.  ***Guidance***  • Skeletal formulas should be discussed in the course.  • The general formulas (eg CnH2n+2) of alkanes, alkenes, alkynes, ketones, alcohols, aldehydes and carboxylic acids should be known.  • The distinction between class names and functional group names needs to be made. Eg for OH, hydroxyl is the functional group whereas alcohol is the class name.  • The following nomenclature should be covered: – non-cyclic alkanes and halogenoalkanes up to halohexanes. – alkenes up to hexene and alkynes up to hexyne. – compounds up to six carbon atoms (in the basic chain for nomenclature purposes) containing only one of the functional groups: hydroxyl, ether, carbonyl (from aldehydes or ketones), ester and carboxyl. |  |  |  |
| 10.2 | ***Functional group chemistry***  Alkanes:   * "Alkanes have low reactivity and undergo free-radical substitution reactions."   Alkenes:   * "Alkenes are more reactive than alkanes and undergo addition reactions. Bromine water can be used to distinguish between alkenes and alkanes."   Alcohols:   * "Alcohols undergo nucleophilic substitution reactions with acids (also called esterification or condensation) and some undergo oxidation reactions."   Halogenoalkanes:   * "Halogenoalkanes are more reactive than alkanes. They can undergo (nucleophilic) substitution reactions. A nucleophile is an electron-rich species containing a lone pair that it donates to an electron-deficient carbon."   Polymers:   * "Addition polymers consist of a wide range of monomers and form the basis of the plastics industry."   Benzene:   * "Benzene does not readily undergo addition reactions but does undergo electrophilic substitution reactions."   ***Applications and skills***  *Alkanes:*  • Writing equations for the complete and incomplete combustion of hydrocarbons.  • Explanation of the reaction of methane and ethane with halogens in terms of a free-radical substitution mechanism involving photochemical homolytic fission.  *Alkenes:*  • Writing equations for the reactions of alkenes with hydrogen and halogens and of symmetrical alkenes with hydrogen halides and water.  • Outline of the addition polymerization of alkenes.  • Relationship between the structure of the monomer to the polymer and repeating unit.  *Alcohols:*  • Writing equations for the complete combustion of alcohols.  • Writing equations for the oxidation reactions of primary and secondary alcohols (using acidified potassium dichromate(VI) or potassium manganate(VII) as oxidizing agents). Explanation of distillation and reflux in the isolation of the aldehyde and carboxylic acid products.  • Writing the equation for the condensation reaction of an alcohol with a carboxylic acid, in the presence of a catalyst (eg concentrated sulfuric acid) to form an ester.  *Halogenoalkanes:*  • Writing the equation for the substitution reactions of halogenoalkanes with aqueous sodium hydroxide.  ***Guidance***  • Reference should be made to initiation, propagation and termination steps in free-radical substitution reactions. Free radicals should be represented by a single dot.  •The mechanisms of SN1 and SN2 and electrophilic substitution reactions are not required. |  |  |  |

**Topic #11: Measurement and Data Processing - 10 Hours**

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| IB Points to Understand | | **Notes** | **Video?** | **QB?** |
| 11.1 | ***Uncertainties and errors in measurement and results***   * “Qualitative data includes all non-numerical information obtained from observations not from measurement." * "Quantitative data are obtained from measurements, and are always associated with random errors/uncertainties, determined by the apparatus, and by human limitations such as reaction times." * "Propagation of random errors in data processing shows the impact of the uncertainties on the final result." * "Experimental design and procedure usually lead to systematic errors in measurement, which cause a deviation in a particular direction." * "Repeat trials and measurements will reduce random errors but not systematic errors."   ***Applications and skills***  • Distinction between random errors and systematic errors.  • Record uncertainties in all measurements as a range (±) to an appropriate precision.  • Discussion of ways to reduce uncertainties in an experiment.  • Propagation of uncertainties in processed data, including the use of percentage uncertainties.  • Discussion of systematic errors in all experimental work, their impact on the results and how they can be reduced.  • Estimation of whether a particular source of error is likely to have a major or minor effect on the final result.  • Calculation of percentage error when the experimental result can be compared with a theoretical or accepted result.  • Distinction between accuracy and precision in evaluating results.  ***Guidance***  • The number of significant figures in a result is based on the figures given in the data. When adding or subtracting, the final answer should be given to the least number of decimal places. When multiplying or dividing the final answer is given to the least number of significant figures.  • Note that the data value must be recorded to the same precision as the random error.  • SI units should be used throughout the programme. |  |  |  |
| 11.2 | ***Graphical techniques***   * "Graphical techniques are an effective means of communicating the effect of an independent variable on a dependent variable, and can lead to determination of physical quantities." * "Sketched graphs have labeled but unscaled axes, and are used to show qualitative trends, such as variables that are proportional or inversely proportional." * "Drawn graphs have labeled and scaled axes, and are used in quantitative measurements."   **Applications and skills:**  • Drawing graphs of experimental results including the correct choice of axes and scale.  • Interpretation of graphs in terms of the relationships of dependent and independent variables.  • Production and interpretation of best-fit lines or curves through data points, including an assessment of when it can and cannot be considered as a linear function.  • Calculation of quantities from graphs by measuring slope (gradient) and intercept, including appropriate units. |  |  |  |
| 11.3 | ***Spectroscopic identification of organic compounds***   * "The degree of unsaturation or index of hydrogen deficiency (IHD) can be used to determine from a molecular formula the number of rings or multiple bonds in a molecule." * "Mass spectrometry (MS), proton nuclear magnetic resonance spectroscopy (1H NMR) and infrared spectroscopy (IR) are techniques that can be used to help identify compounds and to determine their structure."   ***Applications and skills***  • Determination of the IHD from a molecular formula.  • Deduction of information about the structural features of a compound from percentage composition data, MS, 1H NMR or IR.  ***Guidance***  • The electromagnetic spectrum (EMS) is given in the data booklet in section 3. The regions employed for each technique should be understood.  • The operating principles are not required for any of these methods.  • The data booklet contains characteristic ranges for IR absorptions (section 26), 1H NMR data (section 27) and specific MS fragments (section 28). For 1H NMR, only the ability to deduce the number of different hydrogen (proton) environments and the relative numbers of hydrogen atoms in each environment is required. Integration traces should be covered but splitting patterns are not required. |  |  |  |

**Past Paper Practice**

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| **Paper** | **2017** | **2016** | **2015** | **2014** | **2013** | **2012** |
| Paper 1 –  Time (45mins) |  |  |  |  |  |  |
| Paper 1 –  Score (30) |  |  |  |  |  |  |
| Paper 2 –  Time (1hr 15mins) |  |  |  |  |  |  |
| Paper 2 –  Score (30) |  |  |  |  |  |  |