



COPPELL ISD **SUBJECT** YEAR AT A GLANCE

IB CHEMISTRY YEARS 1 AND 2

GRADES
11-12

21
UNITS

Program Transfer Goals

- Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
- Evaluate scientific claims & the validity of scientific research.
- Apply knowledge of science to make informed decisions.

PACING:

Year 1

1st 9 Weeks		2nd 9 Weeks		3rd 9 Weeks		4th 9 Weeks	
Topic 0 3.5 Weeks	Topic 1 6-7 weeks	Topic 2/12 3 weeks	Topic 3/13 5 Weeks	Topic 4/14 6 Weeks	Topic 5/15 5-6 Weeks	Topic 6/16 6 Weeks	

Year 2

1st 9 Weeks		2nd 9 Weeks		3rd 9 Weeks		4th 9 Weeks	
Topic 7/17 4 Weeks	Topic 8/18 7 Weeks	Topic 9/19 6 weeks		Topic 10/20 5-6 Weeks	Topic 11/21 4 Weeks	Prepare for Subject Exams Independent Study for Options	

** Black box indicates flex time, which could be used to catch up if pacing is behind or in-class work on IA.

Assurances for a Guaranteed and Viable Curriculum

Adherence to this scope and sequence affords every member of the learning community clarity on the knowledge and skills on which each learner should demonstrate proficiency. In order to deliver a guaranteed and viable curriculum, our team commits to and ensures the following understandings:

Shared Accountability: Responding to the Needs of All Learners

- High levels of learning for all students.
- The district and course formative assessments aligned to the standards for this course support educators and learners in monitoring academic achievement and leveraging interventions.

Shared Understanding: Curriculum Design

- The district curriculum design weaves together the elements of content, skills and assessments in order to adhere to curriculum design at the macro and micro level, ensuring vertical alignment.
- The district curriculum incorporates standards, scope and sequence, enduring understandings, essential questions, performance assessments, and recommended resources.

Interdependence: Curriculum Units

Members of the learning community utilize the curriculum units, plan collaboratively, and reflect on results for continuous improvement.

The district curriculum units may be found: <http://tinyurl.com/Coppell-Curriculum>

UNIT 0: INTRODUCTION TO MEASUREMENT AND NOMENCLATURE

TIMELINE: 3.5 WEEKS - 1ST YEAR, FIRST 9 WEEKS

Unit Summary: The purpose of this unit is to introduce learners to key processes and vocabulary used throughout the course. Learners will review general lab safety, discuss measurement techniques and the uncertainties inherent in them, as well as how to actively reduce uncertainty and sources of error. Finally, learners will learn nomenclature of inorganic and organic compounds.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. Qualitative data includes all non-numerical information obtained from observations not from measurement.
2. 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.
3. Propagation of random errors in data processing shows the impact of the uncertainties on the final result.
4. Experimental design and procedure usually lead to systematic errors in measurement, which cause a deviation in a particular direction.
5. Repeat trials and measurements will reduce random errors but not systematic errors.
6. 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.
7. Sketched graphs have labelled but unscaled axes, and are used to show qualitative trends, such as variables that are proportional or inversely proportional.
8. Drawn graphs have labelled and scaled axes, and are used in quantitative measurements.
9. Positive ions (cations) form by metals losing valence electrons. (4.1)
10. Negative ions (anions) form by non-metals gaining electrons. (4.1)
11. 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. (10.1)
12. Structural formulas can be represented in full and condensed format.(10.1)
13. Structural isomers are compounds with the same molecular formula but different arrangements of atoms. (10.1)

Students will be skilled at...

- Distinguishing between random and systematic errors.

- Recording uncertainties in all measurements as a range (+) to an appropriate precision.
- Determining ways to reduce uncertainties in an experiment.
- Propagating uncertainties in processed data, including the use of percentage uncertainties.
- Discussing systematic errors in all experimental work, their impact on the results and how they can be reduced.
- Estimating whether a particular source of error is likely to have a major or minor effect on the final result.
- Calculating percentage error when the experimental result can be compared with a theoretical or accepted result.
- Distinguishing between accuracy and precision in evaluating results
- Drawing graphs of experimental results including the correct choice of axes and scale.
- Interpreting graphs in terms of the relationships of dependent and independent variables.
- Producing and interpreting best-fit lines or curves through data points, including an assessment of when it can and cannot be considered as a linear function.
- Calculating quantities from graphs by measuring slope (gradient) and intercept, including appropriate units.
- Deducing the formula and name of an ionic compound from its component ions, including polyatomic ions. (4.1)
- Using naming conventions to name covalently bonded compounds.(4.2)
- Identifying different classes: alkanes, alkenes, alkynes, halogenoalkanes, alcohols, ethers, aldehydes, ketones, esters, carboxylic acids, amines, amides, nitriles and arenes. (10.1)
- Identifying typical functional groups in molecules eg phenyl, hydroxyl, carbonyl, carboxyl, carboxamide, aldehyde, ester, ether, amine, nitrile, alkyl, alkenyl and alkynyl. (10.1)
- Applying IUPAC rules in the nomenclature of straight-chain and branched-chain isomers. (10.1)
- Identifying primary, secondary and tertiary carbon atoms in halogenoalkanes and alcohols and primary, secondary and tertiary nitrogen atoms in amines. (10.1)

UNIT 1: STOICHIOMETRIC RELATIONS

TIMELINE: 6-7 WEEKS - 1ST YEAR, FIRST 9 WEEKS

Unit Summary: Elements interact with each other on a particulate basis. Chemistry is a quantitative science.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. Atoms of different elements combine in fixed ratios to form compounds, which have different properties from their component elements
2. Mixtures contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties.
3. Mixtures are either homogeneous or heterogeneous.
4. The mole is a fixed number of particles and refers to the amount, n , of substance.
5. Masses of atoms are compared on a scale relative to ^{12}C and are expressed as relative atomic mass (A_r) and

- relative formula/molecular mass (M_r).
6. Molar mass (M) has the units g mol^{-1} .
 7. The empirical formula and molecular formula of a compound give the simplest ratio and the actual number of atoms present in a molecule respectively.
 8. Reactants can be either limiting or excess.
 9. The experimental yield can be different from the theoretical yield.
 10. Avogadro's law enables the mole ratio of reacting gases to be determined from volumes of the gases.
 11. The molar volume of an ideal gas is a constant at specified temperature and pressure.
 12. The molar concentration of a solution is determined by the amount of solute and the volume of solution.
 13. A standard solution is one of known concentration.

Students will be skilled at...

- Deducing chemical equations when reactants and products are specified.
- Applying state symbols (s), (l), (g) and (aq) in equations.
- Explaining observable changes in physical properties and temperature during changes of state.
- Calculating the molar masses of atoms, ions, molecules and formula units.
- Solving problems involving the relationships between the number of particles, the amount of substance in moles and the mass in grams.
- Interconverting percentage composition by mass and the empirical formula.
- Determining 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.
- Solving problems relating to reacting quantities, limiting and excess reactants, theoretical, experimental and percentage yields.
- Calculating reacting volumes of gases using Avogadro's law.
- Solving problems and analyzing graphs involving the relationship between temperature, pressure and volume for a fixed mass of an ideal gas.
- Solving problems relating to the ideal gas equation.
- Explaining 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.
- Solving problems involving molar concentration, amount of solute and volume of solution.
- Using the experimental method of titration to calculate the concentration of a solution by reference to a standard solution.

UNIT 2/12: ATOMIC STRUCTURE

TIMELINE: 3 WEEKS -1ST YEAR, 2ND NINE WEEKS

Unit Summary: The structure of an atom can be described and quantified and leads to the observable properties and behaviors of the element.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. Atoms contain a positively charged dense nucleus composed of protons and neutrons (nucleons).
2. Negatively charged electrons occupy the space outside the nucleus.
3. The mass spectrometer is used to determine the relative atomic mass of an element from its isotopic composition.
4. Emission spectra are produced when photons are emitted from atoms as excited electrons return to a lower energy level.
5. The line emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels, which converge at higher energies.
6. The main energy level or shell is given an integer number, n , and can hold a maximum number of electrons, $2n^2$.
7. 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.
8. Sub-levels contain a fixed number of orbitals, regions of space where there is a high probability of finding an electron.
9. Each orbital has a defined energy state for a given electronic configuration and chemical environment and can hold two electrons of opposite spin.
10. **(HL)** In an emission spectrum, the limit of convergence at a higher frequency corresponds to the first ionization energy.
11. **(HL)** Trends in first ionization energy across periods account for the existence of the main energy levels and sub-levels in atoms.
12. **(HL)** Successive ionization energy data for an element give information that shows relations to electron configurations.

Students will be skilled at...

- Use of the nuclear symbol notation A_ZX 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.
- Describing the relationship between color, 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
- Recognizing the shape of an S atomic orbital and the p_x , p_y and p_z 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 $Z = 36$.
- **(HL)** Solving problems using $E = hv$
- **(HL)** Calculating the value of the first ionization energy from spectral data which gives the wavelength or frequency of the convergence limit.
- **(HL)** Deducing the group of an element from its successive ionization energy data.
- **(HL)** Explaining the trends and discontinuities in first ionization energies across a period

UNIT 3/13: PERIODICITY

TIMELINE: 5 WEEKS -1ST YEAR, 2ND NINE WEEKS

Unit Summary: The arrangement of the periodic table is purposeful and can answer many questions in chemistry. Elements exhibit characteristic properties which are reflected by their location in the periodic table.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. The periodic table is arranged into four blocks associated with the four sub-levels—s, p, d, and f
2. The periodic table consists of groups (vertical columns) and periods (horizontal rows).
3. The period number (*n*) is the outer energy level that is occupied by electrons.
4. 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.
5. The periodic table shows the positions of metals, non-metals and metalloids.
6. Vertical and horizontal trends in the periodic table exist for atomic radius, ionic radius, ionization energy, electron affinity and electronegativity.
7. Trends in metallic and non-metallic behaviour are due to the trends above.
8. Oxides change from basic through amphoteric to acidic across a period.
9. Group trends should include the treatment of the reactions of alkali metals with water, alkali metals with halogens and halogens with halide ions
10. **(HL)** Transition elements have variable oxidation states, form complex ions with ligands, have coloured compounds, and display catalytic and magnetic properties.
11. **(HL)** Zn is not considered to be a transition element as it does not form ions with incomplete d-orbitals.
12. **(HL)** Transition elements show an oxidation state of +2 when the s-electrons are removed.
13. **(HL)** The d sub-level splits into two sets of orbitals of different energy in a complex ion.
14. **(HL)** Complexes of d-block elements are coloured, as light is absorbed when an electron is excited between the d-orbitals.
15. **(HL)** The colour absorbed is complementary to the colour observed.

Students will be skilled at...

- Deduction of the electron configuration of an atom from the element's position on the periodic table, and vice versa.
- 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 chemical equations to explain the pH changes for reactions of Na_2O , MgO , P_4O_{10} , and the oxides of nitrogen and sulfur with water.
- **(HL)** Explanation of the ability of transition metals to form variable oxidation states from successive ionization energies.
- **(HL)** Explanation of the nature of the coordinate bond within a complex ion.
- **(HL)** Deduction of the total charge given the formula of the ion and ligands present.
- **(HL)** Explanation of the magnetic properties in transition metals in terms of unpaired electrons.
- **(HL)** Explanation of the effect of the identity of the metal ion, the oxidation number of the metal and the identity of the ligand on the colour of transition metal ion complexes.
- **(HL)** Explanation of the effect of different ligands on the splitting of the d-orbitals in transition metal complexes and colour observed using the spectrochemical series.

UNIT 4/14: CHEMICAL BONDING AND STRUCTURE

TIMELINE: 6 WEEKS - 1ST YEAR, 3RD 9 WEEKS

Unit Summary: The type of bonding in a compound is determined by the movement/activity of electrons between atoms. Bonding leads to molecular geometry. Physical properties of substances result from the bonding and molecular and intermolecular forces present. **(HL)** Electrons can be delocalized and shared between all atoms in a molecule. This can be represented using resonance structures and understood through the hybridization model.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. Positive ions (cations) form by metals losing valence electrons (Topic 0).
2. Negative ions (anions) form by non-metals gaining electrons (Topic 0).
3. The number of electrons lost or gained is determined by the electron configuration of the atom.
4. The ionic bond is due to electrostatic attraction between oppositely charged ions.
5. Under normal conditions, ionic compounds are usually solids with lattice structures.
6. A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei.
7. Single, double and triple covalent bonds involve one, two and three shared pairs of electrons respectively.
8. Bond length decreases and bond strength increases as the number of shared electrons increases.
9. Bond polarity results from the difference in electronegativities of the bonded atoms.
10. Lewis (electron dot) structures show all the valence electrons in a covalently bonded species.
11. The "octet rule" refers to the tendency of atoms to gain a valence shell with a total of 8 electrons.
12. Some atoms, like Be and B, might form stable compounds with incomplete octets of electrons.
13. Resonance structures occur when there is more than one possible position for a double bond in a molecule.
14. Shapes of species are determined by the repulsion of electron pairs according to VSEPR theory.
15. Carbon and silicon form giant covalent/network covalent structures.
16. Intermolecular forces include London (dispersion) forces, dipole-dipole forces and hydrogen bonding.
17. The relative strengths of these interactions are London (dispersion) forces < dipole-dipole forces < hydrogen bonds.
18. A metallic bond is the electrostatic attraction between a lattice of positive ions and delocalized electrons.
19. The strength of a metallic bond depends on the charge of the ions and the radius of the metal ion.
20. Alloys usually contain more than one metal and have enhanced properties.
21. **(HL)** Covalent bonds result from the overlap of atomic orbitals. A sigma bond (σ) is formed by the direct head-on/end-to-end overlap of atomic orbitals, resulting in electron density concentrated between the nuclei of the bonding atoms. A pi bond (π) is formed by the sideways overlap of atomic orbitals, resulting in electron density above and below the plane of the nuclei of the bonding atoms.
22. **(HL)** Formal charge can be used to decide which Lewis structure is preferred from several. The FC is the charge an atom would have if all atoms in the molecule had the same electronegativity.
23. **(HL)** Exceptions to the octet rule include some species having incomplete octets and expanded octets.
24. **(HL)** Delocalization involves electrons that are shared by/between all atoms in a molecule or ion as opposed to being localized between a pair of atoms.
25. **(HL)** Resonance involves using two or more Lewis (electron dot) structures to represent a particular molecule or ion. A resonance structure is one of two or more alternative Lewis (electron dot) structures for a molecule or ion.

that cannot be described fully with one Lewis (electron dot) structure alone.

Students will be skilled at...

- Deducing the formula and name of an ionic compound from its component ions, including polyatomic ions (Topic 0).
- Using naming conventions to name covalently bonded compounds (Topic 0).
- Explaining the physical properties of ionic compounds (volatility, electrical conductivity and solubility) in terms of their structure.
- Deducing the polar nature of a covalent bond from electronegativity values.
- Deducing Lewis (electron dot) structure of molecules and ions showing all valence electrons for up to four electron pairs on each atom.
- Using VSEPR theory to predict the electron domain geometry and the molecular geometry for species with two, three and four electron domains.
- Predicting bond angles from molecular geometry and presence of non-bonding pairs of electrons.
- Predicting molecular polarity from bond polarity and molecular geometry.
- Deducing resonance structures, including but not limited to C_6H_6 , CO_3^{2-} and O_3 .
- Explaining the properties of giant covalent compounds in terms of their structures.
- Deducing the types of intermolecular forces present in substances, based on their structure and chemical formula.
- Explaining the physical properties of covalent compounds (volatility, electrical conductivity and solubility) in terms of their structure and intermolecular forces.
- Explaining electrical conductivity and malleability in metals.
- Explaining trends in melting points of metals.
- Explaining the properties of alloys in terms of non-directional bonding.
- **(HL)** Predicting whether sigma (σ) or pi (π) bonds are formed from the linear combination of atomic orbitals.
- **(HL)** Deducing the Lewis (electron dot) structures of molecules and ions showing all valence electrons for up to six electrons pairs on each atom.
- **(HL)** Applying formal charge to ascertain which Lewis (electron dot) structure is preferred from different Lewis (electron dot) structures.
- **(HL)** Deducing using VSEPR theory of the electron domain geometry and molecular geometry with five and six electron domains and associated bond angles.
- **(HL)** Explaining the wavelength of light required to dissociate oxygen and ozone.
- **(HL)** Describing the mechanism of catalysis of ozone depletion when catalyzed by CFCs and NO_x .
- **(HL)** Explaining the formation of sp^3 , sp^2 and sp hybrid orbitals in methane, ethene and ethyne.
- **(HL)** Identifying and explaining the relationships between Lewis (electron dot) structures, electron domains, molecular geometries and types of hybridization.

UNIT 5/15: ENERGETICS AND THERMOCHEMISTRY

TIMELINE: 5-6 WEEKS - YEAR 1, 3RD-4TH 9 WEEKS

UNIT SUMMARY: The spontaneity of chemical reactions is dependent enthalpy and entropy. Changes in enthalpy can be calculated in multiple ways.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. Heat is a form of energy.
2. Temperature is a measure of the average kinetic energy of the particles.
3. Total energy is conserved in chemical reactions.
4. Chemical reactions that involve transfer of heat between the system and the surroundings are described as endothermic or exothermic.
5. The enthalpy change (ΔH) for chemical reactions is indicated in kJ mol^{-1} .
6. ΔH values are usually expressed under standard conditions, given by ΔH° , including standard states.
7. The 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.
8. Bond-forming releases energy and bond-breaking requires energy.
9. Average bond enthalpy is the energy needed to break one mol of a bond in a gaseous molecule averaged over similar compounds.
10. **(HL)** Representative equations (eg $\text{M}^+(\text{g}) \rightarrow \text{M}^+(\text{aq})$) can be used for enthalpy/energy of hydration, ionization, atomization, electron affinity, lattice, covalent bond and solution.
11. **(HL)** Enthalpy of solution, hydration enthalpy and lattice enthalpy are related in an energy cycle.
12. **(HL)** Entropy (S) refers to the distribution of available energy among the particles. The more ways the energy can be distributed the higher the entropy.
13. **(HL)** Gibbs free energy (G) relates the energy that can be obtained from a chemical reaction to the change in enthalpy (ΔH), change in entropy (ΔS), and absolute temperature (T).
14. **(HL)** Entropy of $\text{gas} > \text{liquid} > \text{solid}$ under same conditions.

Students will be skilled at...

- Calculation of the heat change when the temperature of a pure substance is changed using $q = mC\Delta T$.
- Application of Hess's Law to calculate enthalpy changes.
- Calculation of ΔH reactions using ΔH_f° data.
- Determination of the enthalpy change of a reaction that is the sum of multiple reactions with known enthalpy changes.
- 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.
- **(HL)** Construction of Born-Haber cycles for group 1 and 2 oxides and chlorides.
- **(HL)** Construction of energy cycles from hydration, lattice and solution enthalpy. For example dissolution of solid NaOH or NH_4Cl in water.
- **(HL)** Calculation of enthalpy changes from Born-Haber or dissolution energy cycles.
- **(HL)** Relating size and charge of ions to lattice and hydration enthalpies.
- **(HL)** Performing lab experiments which could include single replacement reactions in aqueous solutions
- **(HL)** Prediction of whether a change will result in an increase or decrease in entropy by considering the states of the reactants and products.
- **(HL)** Calculation of entropy changes (ΔS) from given standard entropy values (S°).
- **(HL)** Application of $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ in predicting spontaneity and calculation of various conditions of enthalpy and temperature that will affect this.
- **(HL)** Relation of ΔG to position of equilibrium.

UNIT 6/16: CHEMICAL KINETICS

TIMELINE: 6 WEEKS - 1ST YEAR, 4TH 9 WEEKS

UNIT SUMMARY: In order for reactions to occur, molecules must collide at the proper orientation with sufficient energy. The probability of this occurring determines the rate at which the reaction occurs. Activation energy is the minimum kinetic energy for particles to have a successful collision and must be present for the reaction to begin.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. Species react as a result of collisions of sufficient energy and proper orientation.
2. The rate of reaction is expressed as the change in concentration of a particular reactant/product per unit time.
3. Concentration changes in a reaction can be followed indirectly by monitoring changes in mass, volume and color.
4. Activation energy (E_a) is the minimum energy that colliding molecules need in order to have successful collisions leading to a reaction.
5. By decreasing E_a , a catalyst increases the rate of a chemical reaction, without itself being permanently chemically changed.
6. **(HL)** Reactions may occur by more than one step and the slowest step determines the rate of reaction (rate determining step/RDS).
7. **(HL)** The molecularity of an elementary step is the number of reactant particles taking part in that step.
8. **(HL)** The order of a reaction can be either integer or fractional in nature. The order of a reaction can describe, with respect to a reactant, the number of particles taking part in the rate-determining step.
9. **(HL)** Rate equations can be determined experimentally.
10. **(HL)** The value of the rate constant (k) is affected by temperature and its units are determined from the overall order of reaction.
11. **(HL)** Catalysts alter a reaction mechanism, introducing a step with lower activation energy.
12. **(HL)** The Arrhenius equation uses the temperature dependence of the rate constant to determine the activation energy.
13. **(HL)** A graph of $1/T$ against $\ln k$ is a linear plot with gradient $-E_a/R$ and intercept, $\ln A$.
14. **(HL)** The frequency factor (or pre-exponential factor) (A) takes into account the frequency of collisions with proper orientations.

Students will be skilled at...

- Describing the kinetic theory in terms of the movement of particles whose average kinetic energy is proportional to temperature in kelvin.
- Analyzing graphical and numerical data from rate experiments.
- Explaining the effects of temperature, pressure/concentration and particle size on rate of reaction.
- Constructing Maxwell-Boltzmann energy distribution curves to account for the probability of successful collisions and factors affecting these, including the effect of a catalyst.
- Investigating rates of reactions experimentally and evaluating the results.
- Sketching and explaining energy profiles with and without catalysts.
- **(HL)** Deducing the rate expression for an equation from experimental data and solving problems involving the rate expression.
- **(HL)** Sketching, identifying, and analyzing graphical representations for zero, first, and second order reactions.
- **(HL)** Evaluating proposed reaction mechanisms to be consistent with kinetic and stoichiometric data.
- **(HL)** Analyzing graphical representation of the Arrhenius equation in its linear form
- $\ln k = \frac{-E_a}{RT} + \ln A$
- **(HL)** Using the Arrhenius equation $k = Ae^{\frac{-E_a}{RT}}$
- **(HL)** Describing the relationships between temperature and rate constant; frequency factor and complexity of molecules colliding.
- **(HL)** Determining and evaluating values of activation energy and frequency factors from data.

UNIT 7/17: EQUILIBRIUM

TIMELINE: 4 WEEKS - 2ND YEAR, 1ST 9 WEEKS

Unit Summary: Chemical reactions reach equilibrium. This equilibrium can be altered by changing conditions, and can be used to quantitatively describe reaction states.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. A state of equilibrium is reached in a closed system when the rates of the forward and reverse reactions are equal.
2. The equilibrium law describes how the equilibrium constant (K_c) can be determined for a particular chemical reaction.
3. The magnitude of the equilibrium constant indicates the extent of a reaction at equilibrium and is temperature dependent.
4. 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.
5. A catalyst has no effect on the position of equilibrium or the equilibrium constant. Le Châtelier's principle for changes in concentration can be explained by the equilibrium law.
6. **(HL)** Le Châtelier's principle for changes in concentration can be explained by the equilibrium law.
7. **(HL)** The position of equilibrium corresponds to a maximum value of entropy and a minimum in the value of the Gibbs free energy.
8. **(HL)** The Gibbs free energy change of a reaction and the equilibrium constant can both be used to measure the position of an equilibrium reaction and are related by the equation, $\Delta G^\circ = -R \ln K$.

Students will be skilled at...

- The characteristics of chemical and physical systems in a state of equilibrium.
- Deduction of the equilibrium constant expression (K_c) 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.
- **(HL)** Solution of homogeneous equilibrium problems using the expression for K_c .
- **(HL)** Relationship between ΔG° and the equilibrium constant.
- **(HL)** Calculations using the equation $\Delta G^\circ = -RT \ln K$.

TOPIC 8/18: ACIDS AND BASES

TIMELINE: 7 WEEKS - YEAR 2, 1ST 9 WEEKS

Unit Summary: Acids and bases can be defined in terms of their structural characteristics and behavior, which in turn explain the characteristic properties and reactions of acids and bases. pH is a measure of how acidic a substance is and is based on the concentration of the hydrogen ion in solution. The strength of acids and bases is dependent on how much dissociation takes place (equilibrium). Acids/Bases/Acid-base systems hold significant places in our everyday lives.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. A Bronsted-Lowry acid is a proton/ H^+ donor and a Bronsted-Lowry base is a proton/ H^+ acceptor.
2. Amphiprotic species can act as both Bronsted-Lowry acids and bases.
3. A pair of species differing by a single proton is called a conjugate acid-base pair.
4. Most acids have observable characteristic chemical reactions with reactive metals, metal oxides, metal hydroxides, hydrogen carbonates and carbonates.
5. Salt and water are produced in exothermic neutralization reactions.
6. $pH = -\log [H^+(aq)]$ and $[H^+] = 10^{-pH}$.
7. A change of one pH unit represents a 10-fold change in the hydrogen ion concentration $[H^+]$
8. pH values distinguish between acidic, neutral and alkaline solutions
9. The ionic product constant, $K_w = [H^+][OH^-] = 10^{-14}$ at 298 K.
10. Strong and weak acids and bases differ in the extent of ionization.
11. Strong acids and bases of equal concentrations have higher conductivities than weak acids and bases.
12. A strong acid is a good proton donor and has a weak conjugate base.
13. A strong base is a good proton acceptor and has a weak conjugate acid.
14. Rain is naturally acidic because of dissolved CO_2 and has a pH of 5.6. Acid deposition has a pH below 5.6.
15. Acid deposition is formed when nitrogen or sulfur oxides dissolve in water to form HNO_3 , HNO_2 , H_2SO_4 and H_2SO_3 .
16. Sources of the oxides of sulfur and nitrogen and the effects of acid deposition should be covered.
17. **(HL)** A Lewis acid is a lone pair acceptor and a Lewis base is a lone pair donor.
18. **(HL)** When a Lewis base reacts with a Lewis acid a coordinate bond is formed.
19. **(HL)** A nucleophile is a Lewis base and an electrophile is a Lewis acid.
20. **(HL)** The expression for the dissociation constant of a weak acid (K_a) and a weak base (K_b).
21. **(HL)** For a conjugate acid base pair, $K_a \times K_b = K_w$.
22. **(HL)** The relationship between K_a and pK_a is ($pK_a = -\log K_a$), and between K_b and pK_b is ($pK_b = -\log K_b$).
23. **(HL)** The characteristics of the pH curves produced by the different combinations of strong and weak acids and bases.
24. **(HL)** An acid-base indicator is a weak acid or a weak base where the components of the conjugate acid-base pair have different colors.
25. **(HL)** The relationship between the pH range of an acid-base indicator, which is a weak acid, and its pK_a value.
26. **(HL)** The buffer region on the pH curve represents the region where small additions of acid or base result in little or no change in pH.
27. **(HL)** The composition and action of a buffer solution.

Students will be skilled at...

- Deducing the Bronsted-Lowry acid and base in a chemical reaction.
- Deducing the conjugate acid or conjugate base in a chemical reaction.
- Balancing chemical equations for the reaction of acids.
- Identifying acids and bases needed to make different salts.
- Completing acid-base titrations with different indicators.
- Solving problems involving pH, $[H^+]$ and $[OH^-]$
- Using a pH meter and universal indicator
- Distinguishing between strong and weak acids and bases in terms of the rates of their reactions with metals, metal oxides, metal hydroxides, metal hydrogen carbonates and metal carbonates and their electrical conductivities for solutions of equal concentrations.

- Balancing the equations that describe the combustion of sulfur and nitrogen to their oxides and the subsequent formation of H_2SO_3 , H_2SO_4 , HNO_2 , HNO_3 .
- Distinguishing between the pre-combustion and post-combustion methods of reducing sulfur oxide emissions.
- Deducing acid deposition equations for acid deposition with reactive metals and carbonates.
- **(HL)** Applying Lewis' acid-base theory to inorganic and organic chemistry to identify the role of the reacting species.
- **(HL)** Solving problems involving $[\text{H}^+(\text{aq})]$, $[\text{OH}^-(\text{aq})]$, pH, pOH, K_a , $\text{p}K_a$, K_b , and $\text{p}K_b$.
- **(HL)** Discussing the relative strengths of acids and bases using values of K_a , $\text{p}K_a$, K_b , and $\text{p}K_b$.
- **(HL)** Discussing the general shapes of graphs of pH against volume for titrations involving strong and weak acids and bases with an explanation of their important features.
- **(HL)** Selecting an appropriate indicator for a titration, given the equivalence point of the titration and the end point of the indicator.
- **(HL)** Preparing a buffer solution by mixing a weak acid/base with a solution of a salt containing its conjugate, or by partial neutralization of a weak acid/base with a strong acid/base.
- **(HL)** Predicting the relative pH of aqueous salt solutions formed by the different combinations of strong and weak acid and base.

TOPIC 9/19: REDOX PROCESSES

TIMELINE: 6 WEEKS - YEAR 2, 2ND NINE WEEKS

Unit Summary: Redox processes play a broad role in chemistry.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. Oxidation and reduction can be considered in terms of oxygen gain/hydrogen loss, electron transfer or change in oxidation number.
2. An oxidizing agent is reduced and a reducing agent is oxidized.
3. Variable oxidation numbers exist for transition metals and for most main-group non-metals.
4. The activity series ranks metals according to the ease with which they undergo oxidation.
5. 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.
6. Voltaic cells:
 - a. Voltaic cells convert energy from spontaneous, exothermic chemical processes to electrical energy.
 - b. Oxidation occurs at the anode (negative electrode) and reduction occurs at the cathode (positive electrode) in a voltaic cell.
7. Electrolytic cells:
 - a. Electrolytic cells convert electrical energy to chemical energy, by bringing about non-spontaneous processes.
 - b. Oxidation occurs at the anode (positive electrode) and reduction occurs at the cathode (negative electrode) in an electrolytic cell.
8. **(HL)** A voltaic cell generates an electromotive force (EMF) resulting in the movement of electrons from the anode (negative electrode) to the cathode (positive electrode) via the external circuit. The EMF is termed the cell potential (E°).

9. **(HL)** The standard hydrogen electrode (SHE) consists of an inert platinum electrode in contact with 1 mol dm⁻³ hydrogen ion and hydrogen gas at 100 kPa and 298 K. The standard electrode potential (E°) is the potential (voltage) of the reduction half-equation under standard conditions measured relative to the SHE. Solute concentration is 1 mol dm⁻³ or 100 kPa for gases. E° of the SHE is 0 V.
10. **(HL)** When aqueous solutions are electrolysed, water can be oxidized to oxygen at the anode and reduced to hydrogen at the cathode.
11. **(HL)** $\Delta G^\circ = -nFE^\circ$. When E° is positive, ΔG° is negative indicative of a spontaneous process. When E° is negative, ΔG° is positive indicative of a non-spontaneous process. When E° is 0, then ΔG° is 0.
12. **(HL)** Current, duration of electrolysis and charge on the ion affect the amount of product formed at the electrodes during electrolysis.
13. **(HL)** Electroplating involves the electrolytic coating of an object with a metallic thin layer.

Students will be skilled at...

- 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.
- 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.
- **(HL)** Calculation of cell potentials using standard electrode potentials.
- **(HL)** Prediction of whether a reaction is spontaneous or not using E° values.
- **(HL)** Determination of standard free-energy changes (ΔG°) using standard electrode potentials.
- **(HL)** Explanation of the products formed during the electrolysis of aqueous solutions.
- **(HL)** Perform lab experiments that could include single replacement reactions in aqueous solutions.
- **(HL)** Determination of the relative amounts of products formed during electrolytic processes.
- **(HL)** Explanation of the process of electroplating.

TOPIC 10/20: ORGANIC CHEMISTRY

TIMELINE: 5-6 WEEKS - YEAR 2, 3RD NINE WEEKS

Unit Summary: Organic chemistry deals with carbon compounds. Organic compounds are numerous and varied but can be divided into subgroups based upon functional groups which themselves give rise to representative physical and chemical properties. Reaction pathways can be classified and used to synthesise specific organic compounds.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. 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 (Topic 0).
2. Structural formulas can be represented in full and condensed format (Topic 0).
3. Structural isomers are compounds with the same molecular formula but different arrangements of atoms.
4. Functional groups are the reactive parts of molecules.
5. Saturated compounds contain single bonds only and unsaturated compounds contain double or triple bonds.
6. Benzene is an aromatic, unsaturated hydrocarbon.
7. Alkanes have low reactivity and undergo free-radical substitution reactions.
8. Alkenes are more reactive than alkanes and undergo addition reactions. Bromine water can be used to distinguish between alkenes and alkynes.
9. Alcohols undergo esterification (or condensation) reactions with acids and some undergo oxidation reactions.
10. Halogenoalkanes are more reactive than alkenes. 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.
11. Addition polymers consist of a wide range of monomers and form the basis of the plastics industry.
12. Benzene does not readily undergo addition reactions but does undergo electrophilic substitution reactions.
13. **(HL) Nucleophilic Substitution Reactions**
 - a. S_N1 represents a nucleophilic unimolecular substitution reaction and S_N2 represents a nucleophilic bimolecular substitution reaction. S_N1 involves a carbocation intermediate. S_N2 involves a concerted reaction with a transition state.
 - b. For tertiary halogenoalkanes the predominant mechanism is S_N1 and for primary halogenoalkanes it is S_N2 . Both mechanisms occur for secondary halogenoalkanes.
 - c. The rate determining step (slow step) in an S_N1 reaction depends only on the concentration of the halogenoalkane, rate = $k[\text{halogenoalkane}]$. For S_N2 , rate = $k[\text{halogenoalkane}][\text{nucleophile}]$. S_N2 is stereospecific with an inversion of configuration at the carbon.
 - d. S_N2 reactions are best conducted using aprotic, polar solvents and S_N1 reactions are best conducted using protic, polar solvents. *Electrophilic Addition Reactions:* • An electrophile is an electron-deficient species that can accept electron pairs from a nucleophile. Electrophiles are Lewis acids.
14. **(HL) Electrophilic Addition Reactions:**
 - a. An electrophile is an electron-deficient species that can accept electron pairs from a nucleophile. Electrophiles are Lewis acids.
 - b. Markovnikov's rule can be applied to predict the major product in electrophilic addition reactions of unsymmetrical alkenes with hydrogen halides and interhalogens. The formation of the major product can be explained in terms of the relative stability of possible carbocations in the reaction mechanism.
15. **(HL) Electrophilic Substitution Reactions:**
 - a. Benzene is the simplest aromatic hydrocarbon compound (or arene) and has a delocalized structure of π bonds around its ring. Each carbon to carbon bond has a bond order of 1.5. Benzene is susceptible to attack by electrophiles.
16. **(HL) Reduction Reactions:**
 - a. Carboxylic acids can be reduced to primary alcohols (via the aldehyde). Ketones can be reduced to secondary alcohols. Typical reducing agents are lithium aluminium hydride (used to reduce carboxylic acids) and sodium borohydride.
17. **(HL) The synthesis of an organic compound stems from a readily available starting material via a series of discrete steps. Functional group interconversions are the basis of such synthetic routes.**
18. **(HL) Retro-synthesis of organic compounds.**
19. **(HL) Stereoisomers are subdivided into two classes - conformational isomers, which interconvert by rotation about a σ bond and configurational isomers that interconvert only by breaking and reforming a bond.**
20. **(HL) Configurational isomers are further subdivided into cis-trans and E/Z isomers and optical isomers.**
21. **(HL) Cis-trans isomers can occur in alkenes or cycloalkanes (or heteroanalogues) and differ in the positions of atoms (or groups) relative to a reference plane. According to IUPAC, E/Z isomers refer to alkenes of the form $R_1R_2C=CR_3R_4$ ($R_1 \neq R_2$, $R_3 \neq R_4$) where neither R_1 nor R_2 need be different from R_3 or R_4 .**
22. **(HL) A chiral carbon is a carbon joined to four different atoms or groups.**

23. **(HL)** An optically active compound can rotate the plane of polarized light as it passes through a solution of the compound. Optical isomers are enantiomers. Enantiomers are non-superimposable mirror images of each other. Diastereomers are not mirror images of each other.
24. **(HL)** A racemic mixture (or racemate) is a mixture of two enantiomers in equal amounts and is optically inactive.

Students will be skilled at...

- 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 (Topic 0).
- Identification of typical functional groups in molecules eg phenyl, hydroxyl, carbonyl, carboxyl, carboxamide, aldehyde, ester, ether, amine, nitrile, alkyl, alkenyl and alkynyl (Topic 0).
- 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 (Topic 0).
- Identification of primary, secondary and tertiary carbon atoms in halogenoalkanes and alcohols and primary, secondary and tertiary nitrogen atoms in amines (Topic 0).
- Discussion of the structure of benzene using physical and chemical evidence.
- 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
- **(HL)** Nucleophilic Substitution Reactions:
 - Explanation of why hydroxide is a better nucleophile than water.
 - Deduction of the mechanism of the nucleophilic substitution reactions of halogenoalkanes with aqueous sodium hydroxide in terms of S_N1 and S_N2 mechanisms. Explanation of how the rate depends on the identity of the halogen (ie the leaving group), whether the halogenoalkane is primary, secondary or tertiary and the choice of solvent.
 - Outline of the difference between protic and aprotic solvents.
- **(HL)** Electrophilic Addition Reactions:
 - Deduction of the mechanism of the electrophilic addition reactions of alkenes with halogens/interhalogens and hydrogen halides
- **(HL)** Electrophilic Substitution Reactions:
 - Deduction of the mechanism of the nitration (electrophilic substitution) reaction of benzene (using a

mixture of concentrated nitric acid and sulfuric acid).

- **(HL) Reduction Reactions:**
 - Writing reduction reactions of carbonyl containing compounds: aldehydes and ketones to primary and secondary alcohols and carboxylic acids to alcohols, using suitable reducing agents.
 - Conversion of nitrobenzene to phenylamine via a two-stage reaction.
- **(HL) Deduction of multi-step synthetic routes given starting reagents and the product(s).**
- **(HL) Construction of 3-D models (real or virtual) of a wide range of stereoisomers.**
- **(HL) Explanation of stereoisomerism in non-cyclic alkenes and C3 and C4 cycloalkanes.**
- **(HL) Comparison between the physical and chemical properties of enantiomers.**
- **(HL) Description and explanation of optical isomers in simple organic molecules.**
- **(HL) Distinction between optical isomers using a polarimeter.**

TOPIC 11/21: MEASUREMENT AND DATA PROCESSING AND ANALYSIS

TIMELINE: 4 WEEKS - GRADING PERIOD

Unit Summary: Data includes numerical and non-numerical information from observations. There is always a degree of random and systematic error in measurements and that error must be accounted for in successive calculations. There are ways in which to reduce systematic and random error. Graphs can be used qualitatively and quantitatively. Various techniques are available to characterize and identify substances. Many times multiple techniques are necessary for a substance's molecular structure to be fully known.

Transfer Goal:

1. Conduct independent investigation, analyze data statistically, identify trends & make valid conclusions.
2. Evaluate scientific claims & the validity of scientific research.
3. Apply knowledge of science to make informed decisions.

Students will know...

1. Qualitative data includes all non-numerical information obtained from observations not from measurement (Topic 0).
2. 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 (Topic 0).
3. Propagation of random errors in data processing shows the impact of the uncertainties on the final result (Topic 0).
4. Experimental design and procedure usually lead to systematic errors in measurement, which cause a deviation in a particular direction (Topic 0).
5. Repeat trials and measurements will reduce random errors but not systematic errors (Topic 0).
6. 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 (Topic 0).
7. Sketched graphs have labelled but unscaled axes, and are used to show qualitative trends, such as variables that are proportional or inversely proportional (Topic 0).
8. Drawn graphs have labelled and scaled axes, and are used in quantitative measurements. (Topic 0).
9. 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.
10. 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.
11. **(HL)** Structural identification of compounds involves several different analytical techniques including IR, ^1H NMR

and MS.

12. **(HL)** In a high resolution ^1H NMR spectrum, single peaks present in low resolution can split into further clusters of peaks.
13. **(HL)** The structural technique of single crystal X-ray crystallography can be used to identify the bond lengths and bond angles of crystalline compounds.

Students will be skilled at...

- Distinguishing between random errors and systematic errors (Topic 0).
- Recording uncertainties in measurements as a +/- range to an appropriate precision (Topic 0).
- Discussing ways to reduce uncertainties in an experiment (Topic 0).
- Propagating uncertainties of errors in processed data, including the use of percentage uncertainties (Topic 0).
- Discussing systematic errors in all experimental work, their impact on the results and how they can be reduced (Topic 0).
- Estimating whether a particular source of error is likely to have a major or minor effect on the final result (Topic 0).
- Calculating percentage error when the experimental result can be compared with a theoretical or accepted result (Topic 0).
- Distinguishing between accuracy and precision in evaluating results (Topic 0).
- Drawing graphs of experimental results including the correct choice of axes and scale (Topic 0).
- Interpreting graphs in terms of the relationships of dependent and independent variables (Topic 0).
- Producing and interpreting best-fit lines or curves through data points, including an assessment of when it can and cannot be considered as a linear function (Topic 0).
- Calculating quantities from graphs by measuring slope (gradient) and intercept, including appropriate units (Topic 0).
- Determining the IHD from a molecular formula.
- Deducing information of compound structural features from percentage composition data, MS, ^1H NMR, or IR.
- **(HL)** Explaining the use of tetramethylsilane (TMS) as the reference standard.
- **(HL)** Deducing the structure of a compound given information from a range of analytical characterization techniques (X-ray crystallography, IR, ^1H NMR and MS).