**Curriculum**

**IB Chemistry Y1**

**Course Overview**

This course lays out the foundation for the basics facts and principles of chemistry. It includes a detailed study of the nature and behavior of matter while combining a theoretical foundation with practical experiments. The process of scientific inquiry is demonstrated in required readings and practiced by students through group work and in the laboratory. An emphasis is placed on data interpretation and analysis. Topics covered include atomic structure and chemical bonding, chemical composition, reactions and mass relationships, kinetic theory of gases, periodicity of the elements and the chemistry of the acid and bases. Throughout the course, students develop their analytical and critical thinking skills as well as manipulative skills in the laboratory. \*IB students take part in the collaborative Group Four Science project at the end of the year.

**Department Standards**

**STANDARD 1: THE NATURE OF SCIENCE**

**STANDARD 2: SCIENCE AND TECHNOLOGY**

**STANDARD 3: THE PHYSICAL SETTING**

**STANDARD 4: THE LIVING ENVIRONMENT**

**STANDARD 5: SCIENCE AND SOCIETY**

**Benchmarks**:

[IB Benchmarks](http://acidale.on-rev.com/dante/Science/IBChemistryGuide.pdf)

**Performance Indicators**

**IB Chemistry Year 1**

**Performance Indicators**

The following describe what the IB Year one student should be able to do. These performance indicators are taken directly from the IB Chemistry course objectives, produced by the IBO.

**Topic 1: Quantitative chemistry.**

1.1 The mole concept and Avogadro’s #

1.1.1 Apply the mole concept to substances

1.1.2 Determine the number of particles and the amount of a substance in moles

1.2 Formulas 1.2.1 Define relative atomic mass and relative molecular mass

1.2.2 Calculate molecular mass

1.2.3 Solve problems involving the relationship between the amount of a substance in moles , mass and molar mass

1.2.4 Distinguish molecular and empirical formula

1.2.5 Determine the empirical formula and the percentage composition

1.2.6 Determine the molecular formula when given both the empirical formula and experimental data

1.3 Chemical equations

1.3.1 Deduce chemical equations when all reactants and products are given

1.3.2 Identify the mole ratio of any two species in a chemical reaction

1.3.3 Apply the state symbols (s), (l), (g), and (aq)

1.4 Mass and gaseous volume relationships in chemical reactions.

1.4.1 Calculate theoretical yields from chemical reactions

1.4.2 Determining the limiting reactant and the reactant in excess when quantities of reacting substances are given.

1.4.3 Solve problems involving theoretical, experimental, and percentage yield.

1.4.4 Apply Avogadro’s law to calculate reacting volumes of gases.

1.4.5 Apply the concept of molar volume at standard temperature and pressure in calculations.

1.4.6 Solve problems involving the relationship between temperature, pressure and volume for a fixed mass of an ideal gas.

1.4.7 Solve problems using the ideal gas equation

1.4.8 Analyse graphs relating to the ideal gas equation

1.5 Solutions

1.5.1 Distinguish between the terms solute solvent, solution and concentration (g dm-3 and mole dm-3

1.5.2 Solve problems involving concentration, amount of solute and volume of solution.

**Topic 2: Atomic Structure**

2.1 The atom

2.1.1 State the positions of protons neutrons and electrons in the atom

2.1.2 State the relative masses and relative charges of protons, neutrons and electrons

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

2.1.4 Deduce the symbol for an isotope given its mass number and atomic number.

2.1.5 Calculate the number of protons, neutrons, and electrons in atoms and ions from the mass number atomic number and charge.

2.1.6 Compare the properties of the isotopes of an element.

2.1.7 Discuss the uses of radioisotopes

2.2 The mass spectrometer

2.2.1 Describe and explain the operation of a mass spectrometer.

2.2.2 Describe how the mass spectrophotometer may be used to determine relative atomic mass using the 12C scale.

2.2.3 Calculate non-integer relative atomic masses and abundance of isotopes from the given data.

2.3 Electron arrangement

2.3.1 Describe the electromagnetic spectrum

2.3.2 Distinguish between a continuous spectrum and line spectrum

2.3.3 Explain how the lines in the emission spectrum of hydrogen are related to electron energy levels.

2.3.4 Deduce the electron arrangement for atoms and ions up to z = 20.

**Additional atomic structure topics for HL IB**

12.1 Electron configuration

12.1.1 Explain how evidence from first ionization energies across periods accounts for the existence of main energy levels and sublevels in atoms.

12.1.2 Explain how successive ionization energy data is related to the electron configuration of an atom.

12.1.3 State the relative energies of s, p, d, and f orbitals in every single energy level.

12.1.4 State the maximum number of orbitals in a given energy level.

12.1.5 Draw the shape of an S orbital, and the shapes of the Px, Py and Pz orbitals

12.1.6 Apply the Aufbau principle, Hund’s Rule, and the Pauli exclusion principle to write electron configurations for atoms and ions up to Z = 54

**Topic 3: Periodicity**

3.1 The periodic table

3.1.1 Describe the arrangement of elements in the periodic table in order of increasing atomic number

3.1.2 Distinguish between the terms group and period

3.1.3 Apply the relationship between the electron arrangement of elements and their position in the periodic table

3.1.4 Apply the relationship between the number of electrons in the highest occupied energy level for an element and its position in the periodic table.

3.2 Physical properties

3.2.1 Define the terms first ionization energy and electronegativity

3.2.2 Describe and explain the trends in atomic radii, ionic radii first ionization energies, electronegativities and melting points for the alkali metals and the halogens.

3.2.3 Describe and explain the trends in the atomic radii, ionic radii, first ionization energies and electronegativities for elements across period three.

3.2.4 Compare the relative electronegativity values of two or more elements based upon their positions in the periodic table.

3.3 Chemical properties

3.3.1 Discuss the similarities and differences in the chemical properties of elements in the same group.

3.3.2 Discuss the changes in nature from ionic to covalent and from basic to acidic of the oxides across period 3

**Additional periodicity topics for HL IB**

13.1 Trends across period 3

13.1.1 Explain the physical states (under standard conditions) and electrical conductivity (in the molten state) of the chlorides and the oxides of the elements in period 3 in terms of their bonding and structure.

13.1.2 Describe the reactions of chlorine and the chlorides referred to in 13.1.1 with water.

13.2 First row D-block elements

13.2.1 List the characteristic properties of transition elements.

13.2.2 Explain why Sc and Zn are not considered transition elements

13.2.3 Explain the existence of variable oxidation number in ions of transition elements.

13.2.4 Define the term ligand.

13.2.5 Describe and explain the formation of complexes of d block elements .

13.2.6 Explain why some complexes of d-block elements are coloured

13.2.7 State examples of the catalytic action of transition elements and their compounds.

13.2.8 Outline the economic significance of catalysts in the contact and Haber processes.

**Topic 4: Bonding**

4.1 Ionic Bonding

4.1.1 Describe the ionic bond as the electrostatic attraction between oppositely charged ions

4.1.2 Describe how ions can be formed as the result of electron transfer.

4.1.3 Describe how ions can be formed when elements in groups 1.2. and 3 lose electrons

4.1.4 Deduce which ions will be formed when elements in groups 5, 6 and 7 gain electrons.

4.1.5 State that transition elements can form more than one ion.

4.1.6 Predict whether a compound of two elements would be ionic from the position of the elements in the periodic table or from their electronegativities

4.1.7 State the common polyatomic ions formed by non-metals in periods 2 & 3.

4.1.8 Describe the lattice structure of ionic compounds

4.2 Covalent bonding

4.2.1 Describe the covalent bond as the electrostatic attraction between a pair of electrons and positively charged nuclei.

4.2.2 Describe how the covalent bond is formed as a result of electron sharing.

4.2.3 Deduce the Lewis (electron dot) structures of molecules and ions for up to four electron pairs on each atom.

4.2.4 State and explain the relationship between the number of bonds, bond length and bond strength.

4.2.5 Predict whether a compound of two elements would be covalent from the position of the elements in the periodic table or from their electronegativity values.

4.2.6 Predict the relative polarity of bonds from electronegativity values.

4.2.7 Predict the shape and bond angles for species with four three and two negative charge centers on the central atom using the valence shell electron pair repulsion theory (VSPR).

4.2.8 Predict whether or not a molecule is polar from its molecular shape and bond polarities.

4.2.9 Describe and compare the structure and bonding of the three allotropes of carbon (diamond, graphite and fullerene)

4.2.10 Describe the structure of and bonding in silicon and silicon dioxide.

4.3 Intermolecular Forces

4.3.1 Describe the types of intermolecular forces (attractions between molecules that have temporary dipoles, permanent dipoles or hydrogen bonding) and explain how they arise from the structural features of the molecules.

4.3.2 Describe and explain how intermolecular forces affect the boiling points of substances.

4.4 Metallic Bonding

4.4.1 Describe the metallic bond as the electrostatic attraction between a lattice of positive ions and delocalized electrons

4.4.2 Explain the electrical conductivity and malleability of metals

4.5 Physical Properties

4.5.1 Compare and explain the properties of substances resulting from different types of bonds.

**Additional bonding topics for higher level IB**

14.1 Shapes of molecules and ions

14.1.1 Predict the shape and bond angles for species with 5 or 6 negative charge centers using the VSEPR theory. (Examples should include PCl5, SF6, XeF4 and PF6-)

14.2 Hybridization

14.2 Describe a sigma and pi bond

14.2.2 Explain hybridization in terms of the missing of atomic orbitals to form new orbitals for bonding.

14.2.3 Identify and explain the relationships between Lewis structures, molecular shapes and types of hybridization (sp, sp2, sp3). (Both organic and inorganic examples)

14.3 Delocalization of electrons

14.3.1 Describe the delocalization of pi electrons and explain how this could account for the structures of some species. (examples should include NO3-1, NO2-1, CO3-2, O3, RCOO- and Benzene)

**Topic 5: Energetics**

5.1 Exothermic and endothermic reactions

5.1.1 Define the terms exothermic reaction, endothermic reaction and standard enthalpy change of reaction (∆H)

5.1.2 State that combustion and neutralization are exothermic processes.

5.1.3 Apply the relationship between temperature change, enthalpy change and the classification of a reaction as endothermic or exothermic.

5.1.4 Deduce, from an enthalpy level diagram, the relative stabilities of reactants and products, and the sign of the enthalpy change for the reaction.

5.2 Calculations of enthalpy changes

5.2.1 Calculate the heat energy change when the temperature of a pure substance is changed.

5.2.2 Design suitable experimental procedures for measuring the heat changes of reactions.

5.2.3 Calculate the enthalpy change for a reaction using experimental data on temperature changes, quantities of reactants and mass of water.

5.2.4 Evaluate the results of experiments to determine enthalpy changes.

5.3 Hess’s Law

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

5.4 Bond Enthalpies

5.4.1 Define the term average bond enthalpy

5.4.2 Explain, in terms of average bond enthalpies, why some reactions are exothermic and others are endothermic.

**Additional topics for Energetics for HL IB**

15.1 Standard Enthalpy Changes of reaction

15.1.1 Define and apply the terms standard state, standard enthalpy change of formation and standard enthalpy change of combustion.

15.1.2 Determine the enthalpy change of a reaction using standard enthalpy changes of formation and combustion.

15.2 Born Haber Cycle

15.2.1 Define and apply the terms lattice enthalpy and electron affinity

15.2.2 Explain how the relative sizes and the charges of ions affect the lattice enthalpies of different compounds. (The relative value of the theoretical lattice enthalpy increases with higher ionic charge and smaller ionic radius due to attractive forces.)

15.2.3 Construct a Born Haber cycle for group a and 2 oxides and chlorides and use it to calculate an enthalpy change.

15.2.4 Discuss the difference between theoretical and experimental lattice enthalpy values of ionic compounds in terms of their covalent character. (A significant difference between the two values indicates covalent character)

15.3 Entropy

15.3.1 State and explain the factors that increase the entropy of a system.

15.3.2 Predict whether the entropy change (∆S) for a given reaction or process is positive or negative

15.3.3 Calculate the standard entropy change for a reaction using standard entropy values.

15.4 Spontaneity

15.4.1 Predict whether a reaction or process will be spontaneous by using the sign of ∆ G.

15.4.2 Calculate ∆G for a reaction using the equation ∆ G = ∆H-T∆S.

15.4.3 Predict the effect of a change in temperature on the spontaneity of a reaction using standard entropy and enthalpy changes and the equation ∆ G = ∆H-

T∆S

**Topic 6: Kinetics**

6.1 Rates of reactions

6.1.1 Define the term rate of reaction

6.1.2 Describe suitable experimental procedures for measuring rates of reactions

6.1.3 Analyze data from rate experiments. (Students should be familiar with graphs of changes in concentration, volume, and mass against time.)

6.2 Collision theory

6.2.1 Describe the kinetic theory in terms of the movement of particles whose average energy is proportional to the temperature in Kelvin.

6.2.2 Define the term activation energy

6.2.3 Describe the collision theory. (Depends upon frequency, number of particles with energy > Ea and appropriate geometry.)

6.2.4 Predict and explain, using the collision theory, the qualitative effects of particle size, temperature, concentration, and pressure on the rate of reaction.

6.2.5 Sketch and explain qualitatively the Maxwell-Boltzmann energy distribution curve for a fixed amount of gas at different temperatures and its consequences for changes in reaction rate. (Students should be able to explain the area under the curve is constant and does not change with temperature.)

6.2.6 Describe the effect of a catalyst on a chemical reaction

6.2.7 Sketch and explain Maxwell-Boltzmann curves for reactions without catalysts.

**(Additional HL topics for Kinetics)**

16.1 Rate Expression

16.1.1 Distinguish between the terms rate constant, overall order of reaction and order of reaction with respect to a particular reactant

16.1.2 Deduce the rate expression for a reaction from experimental data

16.1.3 Solve problems involving the rate expression

16.1.4 Sketch, identify and analyze graphical representations for zero, first and second order reactions (Both concentration-time and rate-concentrations graphs)

16.2 Reaction mechanism

16.2.1 Explain that reactions can occur by more than one step and that the slowest step determines the rate of reaction. (Rate determining step)

16.2.2 Describe the relationship between reaction mechanism, order of reaction and rate determining step

16.3 Activation energy

16.3.1 Describe qualitatively the relationship between the rate constant (k) and temperature.

16.3.2 Determine the activation energy (Ea) values from the Arrhenius equation by a graphical method. (The Arrhenius equation and its logarithmic form are provided in the Chemistry data booklet. The use of simultaneous equations will not be assessed.)

**Topic 7: Equilibrium**

7.1 Dynamic Equilibrium

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

7.2 The position of equilibrium

7.2.1 Deduce the equilibrium constant expression (Kc) from the equation for a homogeneous reaction.

7.2.2 Deduce the extent of the of a reaction from the magnitude of the equilibrium constant. (When Kc>>1 the reaction goes almost to completion. When Kc<<1 the reactions hardly proceeds)

7.2.3 Apply LeChatelier’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. (Not required to state LeChatelier’s principle)

7.2.4 State and explain the effect of a catalyst on an equilibrium reaction.

7.2.5 Apply the concepts of kinetics and equilibrium to industrial processes. (Know Haber and contact process)

**Additional HL topics for Equilibrium**

17.1 Liquid-Vapour equilibrium

17.1.1 Describe the equilibrium established between a liquid and its own vapour and how it is affected by temperature.

17.1.2 Sketch graphs showing the relationship between vapour pressure and temperature and explain them in terms of kinetic theory.

17.1.3 State and explain the relationship between enthalpy of vaporization boiling point and intermolecular forces.

17.2 The equilibrium law

17.2.1 Solve homogeneous equilibrium problems using the expression for Kc

**Topic 8: Acids and Bases**

8.1 Theories of acids and bases

8.1.1 Define acids and bases according to the Bronsted-Lowry and Lewis theories

8.1.2 Deduce whether or not s species could act as a Bronsted-Lowry and/or a Lewis acid or base.

8.1.3 Deduce the formula of the conjugate acid (or Base) of any Bronsted-Lowry base (or acid). (Make clear the location of the proton transferred.)

8.2 Properties of acids and bases

8.2.1 Outline the characteristic properties of acids and bases in aqueous solution. (Bases not hydroxides such as ammonia and soluble carbonates and hydrogen carbonates should be included. Alkalis are bases that dissolve in water. Students should consider the effects on indicators and the reactions

of acids with bases, metals and carbonates.)

8.3 Strong and weak acids and bases

8.3.1 Distinguish between strong and weak acids and bases in terms of the extent of dissociation, reaction with water and electrical conductivity

8.3.2 State whether a given acid or base is strong or weak. (Hydrochloric, nitric and sulfuric strong acids, carboxylic and carbonic acid weak acids. Group 1 hydroxides and Barium hydroxide strong bases, ammonia and amines weak bases.)

8.3.3 Distinguish between strong and weak acids and bases, and determine the relative strengths of acids and bases using experimental data

8.4 The pH scale

8.4.1 Distinguish between aqueous solutions that are acidic, neutral or alkaline using the pH scale.

8.4.2 Identify which of two or more aqueous solutions is more acidic or alkaline using pH values.

8.4.3 State that each change of one pH unit represents a 10 fold change in the hydrogen ion concentration. [H+(aq)]

8.4.4 Deduce changes in [H+(aq)] when the pH of a solution changes by more than one pH unit.

**(Additional HL topics for Acids and Bases)**

18.1 Calculations involving acids and bases

18.1.1 State the expression for the ionic product constant of water.

18.1.2 Deduce [H+(aq)] and [OH- (aq)] for water at different temperatures given Kw values

18.1.3 Solve problems involving [H+(aq)], [OH- (aq)], pH and pOH.

18.1.4 State the equation for the reaction of any weak acid or weak base with water, and hence deduce the expressions for Ka and Kb

18.1.5 Solve problems involving solutions of weak acids and bases using expressions: Ka x Kb = Kw

pKa + pKb = pKw

pH + pOH = pKw

18.1.6 Identify the relative strengths of acids and bases using values of Ka, Kb pKa and pKb

18.2 Buffer solutions

18.2.1 Describe the composition of a buffer solution and explain its action

18.2.2 Solve problems involving the composition and pH of a specified buffer system.

18.3 Salt Hydrolysis

18.3.1 Deduce whether salts from acidic alkaline or neutral aqueous solutions.

18.4 Acid Base Titrations

18.4.1 Sketch the general shapes of graphs of pH against volume for titrations involving strong and weak acids and bases and explain important features.

18.5 Indicators

18.5.1 Describe qualitatively the action of an acid-base indicator

18.5.2 State and explain how the pH range of an acid base indicator relates to its pKa value.

18.5.3 Identify an appropriate indicator for a titration, given the equivalence point of the titration and the pH range of the indicator.

**Assessments**

**IB Chemistry Year One**

**Assessments**

**First Quarter**

De Rarum Natura introduction to atomism poem

Separations lab

Separation research question

Molar mass of a gas lab

PHET gas lab, gas laws

Thin layer chromatography

Quantitative chemistry MC question set from IB

Stoichiometry Quiz

Clicker quiz over quant chem mc questions

Generation of CO2 lab

Generation of Hydrogen lab

PHET Neon lab, light and spectra

**Second Quarter**

Atomic theory MC set from IB

Clicker quiz over atomic theory mc questions

Periodicity MC questions from the IB

Molarity of Unknown lab, DCP/CE prototype

Clicker quiz over Periodicity mc questions

Isotope review and practice

Electron dot practice

VSEPR drawing and modelling

Bonding MC questions from the IB

Polyatomic ion quizzes

Mid term exam

**Third Quarter**

Ion exchange lab

∆H solution lab

∆T lab quiz

Est. mass by using ∆T lab quiz

IB MC and short answer set for Energetics

∆Hsol preparation for IA one. Including excel and vernier technology practice

Hess’s law lab official IA number 1

IB MC and short answer set for Kinetics

Reaction rate on line lab

Reaction rate design lab IA number 2

Quiz kinetics

Kinetics mc on study wiz

Equilibrium packet from the IB

**Fourth Quarter**

Equilibrium free response from AP

Equilibrium lab

IB MC question set for Acids and Bases

Acids titration lab

Make a base lab

Acids bases quiz

Redox mc questions

Redox packet

Redox battery and electroplating lab

**Core Topics**

**IB Chemistry Year One**

**Core Topics**

**First Quarter**

Quantitative chemistry

Atomic Structure

Lab work

**Second Quarter**

Periodicity

BondingLab work

**Third Quarter**

Energetics

KineticsLab Work

**Fourth Quarter**

Equilibrium

Acids and Bases

Lab work

**Specific Content**

**IB Chemistry Year One**

**Specific Content**

**First Quarter**

1. Quantitative chemistry

1.1 The mole concept and Avogadro’s number

1.2 Formulas

1.3 Chemical Equations

1.4 Mass and gaseous volume relationships in chemical reactions

1.5 Solutions

Related Labs:

1. Separation of mixture

2. TLC lab

3. Glass blowing, burner use

4. Molar volume of gas

5. PHET gas laws

6. Molarity of unknown/titration

2. Atomic structure

2.1 The atom

2.2 The mass spectrometer

2.3 Electron arrangement Related Labs:

7. PHET – neon lights and other discharge devices.

**Second Quarter**

3. Periodicity

3.1 The periodic table

3.2 Physical properties

3.3 Chemical properties

4. Bonding

4.1 Ionic bonding

4.2 Covalent bonding

4.3 intermolecular forces

4.4 Metallic bonding

4.5 Physical properties

**Third Quarter**

5. Energetics

5.1 Exothermic and endothermic reactions

5.2 Calculation of enthalpy changes

5.3 Hess's law

5.4 Bond Enthalpies

6. Kinetics

6.1 Rates of reaction

6.2 Collision theory

**Fourth Quarter**

7. Equilibrium

7.1 Dynamic equilibrium

7.2 The position of equilibrium

8. Acids and bases

8.1 Theories of acids and bases

8.2 Properties of acids and bases

8.3 Strong and weak acids and bases

8.4 the pH scale

**Resources**

**IB Year One Chemistry**

**Chemistry Resources**

IB question bank

AP question bank, teacher made.

IB and AP question banks

IB book, SL Chemistry by Catrin Brown, Mike Ford

IB/AP question banks

IB/AP MC question banks

Collection of unit specific jeopardy games, self made.