**Curriculum**

**Advanced Placement Chemistry**

**Course Overview**

A strong emphasis is placed understanding the basic principles of chemistry. There is considerable emphasis on chemical calculations and the mathematical formulation of principles as well as an emphasis on the experimental nature of chemistry. Students collect data, analyze and interpret it, and draw conclusions. They verify principles, develop hypotheses, and design experiments to test these hypotheses in the context of practical applications of chemistry. Topics covered include knowledge core to chemistry and preparatory for year 2 AP or IB.

**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**:

[Science Department Standards & Benchmarks](http://acidale.on-rev.com/dante/Science/Standards&BenchmarksK-12.docx)

**Performance Indicators**

**AP Chemistry**

**Performance Indicators**

Some of the performance indicators below are taken directly from the IB Chemistry Syllabus, published by the IBO. Other performance indicators have been changed from their IB versions and some new ones added added to reflect the different emphasis placed by the AP College board on various topics in AP chemistry. Four examples of AP vs IB difference are given below:

a. The AP Chemistry performance indicators below include those covering basic knowledge of nuclear chemistry, a topic not included in IB chemistry.

b. Organic chemistry AP indicators are formula and nomenclature based and, unlike IB, include no organic reactions.

c. Much heavier emphasis is placed upon kinetic, electrochemical and equilibrium calculations in the AP than in the IB, and include the use of the Nernst and Arrhenius equations.

d. The AP chemistry student is expected to know a wider range of inorganic reactions than the IB student, thus the indicators below contain statements toward that end.

**First Quarter**

1. Structure of Matter

A. Atomic theory and atomic structure (Sci Standard 3, grades 9-12, #16,17,18)

1. State the positions of protons neutrons and electrons in the atom

2. State the relative masses and relative charges of protons, neutrons and electrons

3. Define mass number (A), atomic number (Z) and isotopes of an element.

4. Define relative atomic mass and relative molecular mass

5. Calculate molecular mass

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

B. Isotopes (Sci Standard 3, grades 9-12, #19,20)

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

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

3. Compare the properties of the isotopes of an element.

4. Discuss the uses of radioisotopes.

C. Molecular and empirical formulas (Sci Standard 3, grades 9-12, #22)

1. Distinguish molecular and empirical formula

2. Determine the empirical formula and the percentage composition

3. Determine the molecular formula when given both the empirical formula and experimental data

D. Mass spectroscopy

1. Describe and explain the operation of a mass spectrometer.

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

3. Calculate non-integer relative at. masses and abundance of isotopes from the given data.

2. Chemical bonding

A. Ionic Bonds

1. Describe ionic bonds as the electrostatic attraction between oppositely charged ions

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

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

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

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

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

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

8. Describe the lattice structure of ionic compounds.

B. Covalent bonds

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

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

C. Describing bonds

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

2. State and explain the relationship between the # of bonds, bond length & bond strength.

3. 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. Predict the relative polarity of bonds from electronegativity values.

D. Molecular shapes, and VSEPR (Sci Standard 2, grades 9-12, #32) (Sci Standard 3, grades 9-12, #27,28,28)

1. Predict the shape and bond angles for species from 6 to 2 negative charge centers on the central atom using the valence shell electron pair repulsion theory (VSEPR).

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

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

4. 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.

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

6. Describe a sigma and pi bond

7. Explain hybridization in terms of the mixing of atomic orbitals to form new orbitals 8. Identify and explain the relationships between Lewis structures, molecular shapes and types of hybridization (sp, sp2, sp3). *(Both organic and inorganic examples)*

9. Describe the delocalization of pi electrons and explain how this could account for the structures of some species.

E. Metallic Bonds (Sci Standard 3, grades 9-12, #55)

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

2. Explain the electrical conductivity and malleability of metals

F. Bonding and properties (Sci Standard 3, grades 9-12, #56,57)

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

3. Quantitative chemistry

A. The mole

1. Conversions to grams, liters (STP), and number of particles

B. Concentrations

1. Distinguish between terms solute solvent, solution and concentration

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

3. Solve problems in which concentrations are expressed in molality and normality.

C. Gas laws

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

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

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

4. Solve problems using the ideal gas equation

5. Solve problems with Graham’s law

6. Solve problems with Dalton’s law

7. Analyse graphs relating to the ideal gas equation

**Second Quarter**

4. Reactions

A. Reaction types (Sci Standard 3, grades 9-12, 6 applied to combustion CO2)

1. Complete a laboratory practical involving double displacement reactions

2. Complete and balance combustion, redox and acid and base reactions

3. Deduce products of a wide range of reactions, including combustion, single and double displacement, REDOX, complex formation, and carbonate and sulfite reactions with acids.

4. Write net ionic equations when given formula and equation names.

5. Deduce whether a given reaction is exothermic or endothermic given the reaction type

B. Stoichiometry

1. Solve problems involving stoichiometry

2. Identify the limiting reactants

3. Solve problems limiting reactants

5. Periodicity (Sci Standard 3, grades 9-12, #21, 22, 23, 24) (Sci Standard 3, grades 9-12, #29,30)

A. Electrons and light

1. Describe the electromagnetic spectrum

2. Distinguish between a continuous spectrum and line spectrum

3. Explain how emission lines of hydrogen are related to electron energy levels.

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

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

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

B. Quantum numbers as electron and atom descriptors

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

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

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

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

5. Represent and elements quantum numbers using the n,l,ml,ms system.

C. The periodic table

1. Describe the arrangement of elements in the periodic table in order of atomic number

2. Distinguish between the terms group and period

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

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.

D. Trends in mp/bp

1. Apply periodic prop. & knowledge of intermolecular forces to predict trends in mp and bp

E. Trends in IE, atomic radii and electronegativity

1. Define the terms first ionization energy and electronegativity

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

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

F. Group names and properties

1. Discuss similarities & differences in chemical properties of elements in the same group.

2. List the characteristic properties of transition elements.

3. Explain why Sc and Zn are not considered transition elements

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

5. Define the term ligand.

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

7. Explain why some complexes of d-block elements are coloured

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

6. Kinetics

A. Rate of reaction (Sci Standard 3, grades 9-12, #37)

1. Define the term rate of reaction

2. Describe suitable experimental procedures for measuring rates of reactions

3. Analyze data from rate experiments.

B Collision theory

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

2. Define the term activation energy

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

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

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.

6. Describe the effect of a catalyst on a chemical reaction

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

C. Use of experimental data and graphical analysis to determine reactant order, rate constants, and reaction rate laws

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

2. Deduce the rate expression for a reaction from experimental data

3. Solve problems involving the rate expression

4. Sketch, identify and analyze graphical representations for zero, first and

second order reactions

D. Reaction mechanism

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

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

E. Activation energy

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

2. Determine the activ. energy (Ea) values from Arrhenius equation by a graphical method.

**Third Quarter**

7. Thermodynamics

A. Exothermic and endothermic reactions

1. Define the terms exothermic reaction, endothermic reaction and standard enthalpy change of reaction (ΔH)

2. State that combustion and neutralization are exothermic processes.

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

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

B. Calculations of enthalpy changes (Sci Standard 2, grades 9-12, #26, 27)

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

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

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

4. Evaluate the results of experiments to determine enthalpy changes.

5. Determine the enthalpy change of a reaction that is the sum of two or three reactions with known enthalpy changes. (Hess’s law)

C. Second law and entropy (Sci Standard 2, grades 9-12, #26, 27)

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

2. Predict whether the entropy change (ΔS) for a given reaction or process is positive or negative

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

D. Gibbs free energy (Sci Standard 2, grades 9-12, #26, 27)

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

2. Calculate ΔG for a reaction using the equation Δ G = ΔH-TΔS.

3. Predict the effect of a change in temperature on the spontaneity of a reaction using standard entropy and enthalpy changes and the equation Δ G = ΔH-TΔS

8. Equilibrium

A. Concept of dynamic equilibrium, physical and chemical

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

B. Le Chatelier’s principle and shifting equilibrium (Sci Standard 1, grades 6-12, #19)

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

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)*

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.

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

5. Apply the concepts of kinetics and equilibrium to industrial processes.

C. Heterogeneous vs homogeneous equilibrium

1. Calculate Ksp from data given

2. Read equations to determine the phases involved in the equilibrium expression

3. Solve homogeneous equilibrium problems using the expression for Kc

D. Kc, Kp, Ksp

1. Apply knowledge of state and units to determine the most appropriate equilibrium expression

2. Apply the use of estimation to avoid the use of quadratics in equilibrium expressions

E. Reaction quotients

1. Calculate reaction quotients

2. Predict the changes in concentration when given Q

9. Acid base

A. Properties and definitions of acids and bases

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

2. Deduce whether a species could act as a Bronsted-Lowry and/or a Lewis acid or base.

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.)*

4. Outline the characteristic properties of acids and bases in aqueous solution.

B. Strong and weak acids and bases

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

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.)*

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

C. pH

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

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

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

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

D. Acid base equilibrium

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

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

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

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

5. Solve problems involving solutions of weak acids and bases using expressions:

Ka x Kb = Kw

pKa + pKb = pKw

pH + pOH = pKw

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

7. Describe the composition of a buffer solution and explain its action

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

E. Titrations and indicators

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

2. Describe qualitatively the action of an acid-base indicator

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

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

10. Reduction oxidation

A. Theory

1. Define oxidation and reduction in terms of electrons loss of gain

2. Deduce the oxidation number of an element in a compound

3. State the names of compounds using oxidation numbers.

4. Deduce whether an element undergoes oxidation or reduction in reactions.

B. Writing redox equations

1. Deduce simple oxidation and reduction half-reactions given the species involved in a redox reaction.

2. Deduce redox equations using half-equations

3. Define the terms oxidizing agent and reducing agent

4. Identify the oxidizing and reducing agents in redox equations.

B. Application to cells

1. Explain how a redox reaction is used to produce electricity in a voltaic cell.

2. State that oxidation occurs at the negative electrode(anode) and reduction occurs at the positive electrode (cathode)

3. Describe, using a diagram, the essential components of an electrolytic cell

4. State that oxidation occurs at the positive electrode (anode) and reduction occurs at the negative electrode (cathode).

5. Describe how current is conducted in an electrolytic cell.

6. Deduce the products of the electrolysis of a molten salt.

7. Predict and explain the products of electrolysis of aqueous solutions.

8. Determine the relative amounts of the products formed during electrolysis. 9. Describe the use of electrolysis in electroplating.

C. Standard cells

19.1.1 Describe the standard hydrogen electrode

19.1.2 Define the term standard electrode potential (Eo)

D. Calculations of potential and determination of spontaneity

1. Calculate cell potentials using standard electrode potentials

2. Predict whether a reaction will be spontaneous using standard electrode potential values.

3. Compare ∆G and Eo values

4. Calculate ∆G or Eo values from each other.

E. Nernst equation

1. Define standard conditions

2. Calculate the Eo of a cell given non standard conditions

**Fourth Quarter**

11. Organic chemistry

A. Isomerism and formulas

1. Distinguish between empirical, molecular and structural formulas.

2. Describe structural isomers as compounds with the same molecular formula but with different arrangements of atoms

3. Deduce structural formulas for isomers of the non-cyclic alkanes up to C6

4. Apply IUPAC rules for naming isomers of straight chain alkenes up to C6.

5. Deduce structural formulas for compounds containing up to 6 carbon atoms with one of the following functional groups: alcohol, aldehyde, ketone, carboxylic acid and halide.

6. Deduce structural formulas for the isomers of the straight-chain alkenes up to C6.

B. Nomenclature

1. Apply IUPAC rules for naming isomers of the non-cyclic alkanes up to C6.

2. Apply IUPAC rules for naming compounds containing up to six carbons with one of the following functional groups: alcohol, aldehyde, ketone carboxylic acid and halide.

C. Functional groups

1. Identify the following functional groups when present in structural formulas: amino (NH2), benzene ring and esters (RCOOR).

12. Nuclear chemistry: (Sci Standard 2, grades 9-12, #23) (Sci Standard 3, grades 9-12, #60)

A. nuclear equations

1. Deduce the isotopic products from an alpha decay reaction

2. Deduce the isotopic products from a beta decay

B. Half-lives

1. Define half life as a first order reaction

2. Calcualate ½ life

C. radioactivity; chemical applications

1. List the most common uses of radioisotopes

13. Chemistry research project

A. audio visual

B. topic of interest and relevance

**Over all four quarters the following laboratory performance objectives are assessed.** (Sci Standard 1, grades 6-12, many)

A. Make observations of chemical reactions and substances

B. Gather experimental data from the best devices possible

C. Operate safely in the laboratory

D. Operate with confidence and independence in the laboratory

E. Record and presenting data using spreadsheet programs

F. Design experiments to answer lab questions.

G. Calculate and interpreting results based on the quantitative data obtained

H. Communicate effectively the results of experimental work

**Assessments**

**AP Chemistry**

**Assessments**

**First Quarter**

AP free response over VSEPR

AP free response over periodic trends

AP free response over gas laws and stoichiometry

Modeling lab

Several small reaction types labs, single displacement, double displacement etc.

**Second Quarter**

AP mc question sets over energetics, kinetics and equilibria

AP free response question sets over energetics, kinetics and equilibria

Thiosulfate kinetics lab

Enthalpy lab

Equilibria lab (TBD)

**Third Quarter**

AP mc question sets over acids bases and redox

AP free response questions over Acids/bases and redox

Electrochemistry lab

Titration lab

Indicator lab

Synthesis of acids/bases lab

**Fourth Quarter**

AP mc question set containing organic chemistry

AP nuclear chemistry question set

AP free response containing organic reactions or calculations

Ester synthesis

Nitration of an aromatic ring lab

**Core Topics**

**AP Chemistry**

**Core Topics**

**First Quarter**

1. Structure of Matter

A. Atomic theory and atomic structure

B. Isotopes

C. Molecular and empirical formulas

D. Mass spectroscopy

2. Chemical bonding

A. Ionic Bonds

B. Covalent bonds

C. Describing covalent bonds, dots, lines etc

D. Molecular shapes, and VSEPR

E. Metallic bonds

F. Bonding and properties

3. Quantitative chemistry

A. The mole

B. Concentrations

C. Gas laws

**Second Quarter**

4. Reactions

A. Reaction types

B. Stoichiometry

5. Periodicity

A. electrons and light

B. Quantum numbers as electron and atom descriptors

C. The periodic table

D. Trends in mp/bp

E. Trends in atomic radii, IE and electronegativity

F. Group names and properties

6. Kinetics

A. Rate of reaction

B. Collision theory

C. Use of experimental data and graphical analysis to determine reactant order, rate constants, and reaction rate laws

D. Reaction mechanism

E. Activation energy

**Third Quarter**

7. Thermodynamics

A. Exothermic and endothermic reactions

B. Calculations of enthalpy changes

C. Second law and entropy

D. Gibbs free energy

8. Equilibrium

A. Concept of dynamic equilibrium, physical and chemical

B. Le Chatelier’s principle and shifting equilibrium

C. Heterogeneous vs homogeneous equilibrium

D. Kc, Kp, Ksp

E. Reaction quotients

9. Acid base

A. Properties and definitions of acids and bases

B. Strong and weak acids and bases

C. pH

D. Acid base equilibrium

E. Titrations and indicators

10. Reduction oxidation

A. Theory

B. Writing redox reactions

C. Application to cells

C. Standard cells

D. Calculations of potential

E. Nernst equation

**Fourth Quarter**

11. Organic chemistry

A. Isomerism and formulas

B. Nomenclature

C. Functional groups

12. Nuclear chemistry:

A. Nuclear equations

B. Half-lives

C. radioactivity; chemical applications

13. Chemistry research project

A. audio visual

B. topic of interest and relevance

**Over all four quarters the following laboratory performance objectives are assessed.**

A. Make observations of chemical reactions and substances

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D. Operate with confidence and independence in the laboratory

E. Record and presenting data using spreadsheet programs

F. Design experiments to answer lab questions.

G. Calculate and interpreting results based on the quantitative data obtained

H. Communicate effectively the results of experimental work

**Specific Content**

**AP Chemistry**

**Specific Content**

**First Quarter**

1. Structure of Matter

A. Atomic theory and atomic structure

1. Atomic mass and number

2. Relative masses of subatomic particles

3. Calculating molecular mass and molar mass

B. Isotopes

C. Molecular and empirical formulas

D. Mass spectroscopy

2. Chemical bonding

A. Ionic Bonds

B. Covalent bonds

C. Describing covalent bonds, dots, lines etc

D. Molecular shapes, and VSEPR

E. Metallic bonds

F. Bonding and properties

3. Quantitative chemistry

A. The mole

1. Conversions to grams, liters (STP), and number of particles

B. Concentrations

1. Molarity

2. Molality

3. Normality

C. Gas laws

1. Boyles, Charles, Gay Lussac

2. Ideal gas law

3. Grahams law of diffusion

4. Dalton’s law of partial pressure

**Second Quarter**

4. Reactions

A. Reaction types

1. Displacement

2. Combustion

3. Redox

4. Acid base

5. Exothermic vs endothermic

B. Stoichiometry

1. Methods for solving stoichiometry problems

2. Limiting reactants

5. Periodicity

A. electrons and light

B. Quantum numbers as electron and atom descriptors

C. The periodic table

D. Trends in mp/bp

E. Trends in atomic radii, IE and electronegativity

F. Group names and properties

6. Kinetics

A. Rate of reaction

1. How to measure in lab

2. Graphically determining rate

B. Collision theory

C. Use of experimental data and graphical analysis to determine reactant order, rate constants, and reaction rate laws

1. Zero, 1st and 2nd order reactions

2. Calculating rate constants and rates

3. Graphical analysis of 0, 1st, 2nd order reactions.

D. Reaction mechanism

E. Activation energy

**Third Quarter**

7. Thermodynamics

A. Exothermic and endothermic reactions

B. Calculations of enthalpy changes

C. Second law and entropy

1. Predictions of entropy changes

D. Gibbs free energy

1. free energy of reaction; dependence of change in free energy on enthalpy and entropy changes

2. Spontaneity

8. Equilibrium

A. Concept of dynamic equilibrium, physical and chemical

B. Le Chatelier’s principle and shifting equilibrium

1. Effects of pressure changes

2. Effects of temperature changes

3. Effects of volume changes

4. Effects of concentration changes

C. Heterogeneous vs homogeneous equilibrium

1. Solubility products

2. Phases and equations

D. Kc, Kp, Ksp

1. Interchangeability of Kc and Kp

2. Estimation and avoiding quadratics

E. Reaction quotients

1. Q vs Kc

2. Predicting equilibrium shifts from Q

9. Acid base

A. Properties and definitions of acids and bases

B. Strong and weak acids and bases

C. pH

1. Logarithmic

2. Predicting pH

3. Buffers

D. Acid base equilibrium

1. Ka, Kb

2. Determination of pH form Ka, Kb

E. Titrations and indicators

10. Reduction oxidation

A. Theory

B. Writing redox reactions

C. Application to cells

1. voltaic

2. electrolytic

C. Standard cells

1. SHE

2. Standard reduction potentials

D. Calculations of potential

1. determination of Eo of a cell

2. Spontaneity and Eo

3. Spontaneity and Eo.

4. G vs Eo

E. Nernst equation

1. defining standard conditions

2. Use of the Nernst equation

**Fourth Quarter**

11. Organic chemistry

A. Isomerism and formulas

B. Nomenclature

1. Alkanes and alkenes

3. Alcohols, aldehydes, ketones , halides and acids

C. Functional groups

1. Alkanes

2. Alkenes

3. Alcohols

4. Aldehydes and ketones

5. Acids

12. Nuclear chemistry:

A. Nuclear equations

1. alpha decay

2. Beta decay

B. Half-lives

C. radioactivity; chemical applications

1. dating

2. medicine

13. Chemistry research project

A. audio visual

B. topic of interest and relevance

**Over all four quarters the following laboratory performance objectives are assessed.**

A. Make observations of chemical reactions and substances

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C. Operate safely in the laboratory

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E. Record and presenting data using spreadsheet programs

F. Design experiments to answer lab questions.

G. Calculate and interpreting results based on the quantitative data obtained

H. Communicate effectively the results of experimental work

**Resources**

**AP Chemistry**

**Resources**

AP question banks.

Released AP exams

Free responses from all AP exams

Laboratories from various sources including recommended labs from the AP

Zumdahl “Chemistry” 2000, 5th edition.