

Advanced Placement Chemistry

2006

COURSE DESCRIPTION: (The course description sets the parameters, scope and sequence for the course:

I. Laboratory Safety

A. Laboratory Safety and Precautions

1. Predict what could occur in various laboratory situations.
2. Apply laboratory safety rules to laboratory situations throughout the year.
3. Plan a course of action for an emergency in the laboratory.
4. Demonstrate consistently the safe use of laboratory equipment.

II. Types of Chemical Reactions and Solution Stoichiometry

A. The characteristics of the common solvent water.

1. Discuss the three interactions involved in the solution process.
2. Distinguish between strong and weak electrolytes.
3. Discuss the composition of solutions.
4. Identify the types of chemical reactions that take place in solution.
5. Explain what is meant by a precipitation reaction and give examples.
6. Describe in detail reactions that occur in solution.
7. Explain the procedure for determining whether or not a precipitate will form.
8. Solve stoichiometric problems involved with precipitation reactions.
9. Identify the characteristics of acid/base reactions.
10. Define an oxidation – reduction reaction
11. Explain oxidation, reduction, oxidizing agent, and reducing agent.
12. Explain the steps in balancing redox reactions in both acidic and basic solutions.
13. Explain how to identify the oxidizing and reducing agent in a redox reaction.

III. Gases

A. Pressure

1. Define pressure
2. Use dimensional analysis to convert pressure.
Explain atmospheric pressure and discuss how it is measured.
4. List the values for standard atmospheric pressure.
Explain how to measure the pressure of a gas using a manometer.

B. Gas Laws

1. Explain and apply Boyle's Law, Charles' Law, Gay-Lussac's Law, and Avogadro's Law.
Derive and apply the Ideal Gas Law and its alternate forms.
2. Derive the Universal Gas Constant and identify its units.

C. Gas Stoichiometry

Solve gas stoichiometry problems using the Ideal Gas Law.

Calculate the molar mass of a substance using the Ideal Gas Law.

D. Dalton's Law of Partial Pressures

Define and explain Dalton's Law of Partial Pressures.

2. Discuss the relationships between partial pressure and the number of moles of gas.

3. Explain how to collect a gas over water.

E. Kinetic Molecular Theory

1. Explain the Kinetic Molecular Theory.

2. Account for Boyle's Law, Charles' Law, Avogadro's Law, Dalton's Law and Gay-Lussac's Law using the Kinetic Molecular Theory.

3. Explain temperature.

3. Define and apply root mean square velocity.

F. Effusion and Diffusion

1. Distinguish between effusion and diffusion.

2. Derive and apply Graham's Law

G. Real Gases

1. Distinguish between real and ideal gases.

2. Explain the conditions of temperature and pressure under which a gas will not behave ideally and why this is so.

3. Explain VanderWaal's Equation and the reasons why temperature and pressure must be corrected.

4. Apply VanderWaal's Equation.

IV. Bonding

A. General Concept of Bonding

1. Define bond energy and explain the significance of its value.

2. Explain and define ionic bond.

3. Discuss Coulomb's Law and explain that it is used to calculate the force of attraction between two oppositely charged particles.

4. Define bond length.

5. Using a graph of potential energy versus bond length, explain that a bond will form when the energy of the aggregate is lower than that of the separate atoms.

6. Define electronegativity and explain how difference in electronegativity determines bond polarity.

7. Distinguish between bond polarity and molecular polarity.

8. Explain and define a dipole moment.

9. Discuss why a molecule that contains polar bonds may not exhibit a dipole moment.

10. Explain and give examples of the fact that atoms in stable compounds usually have a noble gas electron configuration.

11. Discuss the fact that in the solid state of an ionic compound, the ions are relatively close together, and many ions are simultaneously interacting.

12. Demonstrate the criss-cross method of predicting formulas of ionic compounds.
 13. Compare the sizes of metallic ions versus nonmetal ions.
- B. The Formation of Binary Ionic Compounds
1. Define lattice energy and explain the steps involved in the formation of a binary ion compound from its ions in the gaseous phase.
 2. Perform lattice energy calculations to relate lattice energy values to strength of bond.
 3. Explain how to calculate the percent ionic character of a covalent bond.
- C. The Covalent Chemical Bond
1. Explain why chemical bonds occur.
 2. Explain the model of bond energies to explain the change in enthalpy of a reaction.
 3. Discuss how bonding is a model proposed to explain molecular stability.
 4. Use covalent bond energies to calculate the heat of reaction of a reaction.
- D. The Localized Electron Bonding Model
1. Explain the three parts of the localized electron model.
- E. Lewis Structures
1. Define the characteristics of a Lewis Structure.
 2. Explain the rules and demonstrate the procedures for writing Lewis Structures.
 3. Explain how to determine multiple bonding in structures.
 4. List the exceptions to the octet rule.
 5. Explain what is meant by resonance and determine resonance structures.
 6. Define formal charges and demonstrate how to calculate them.
 7. Discuss the rules governing formal charges.
- F. Molecular Structure: The Valence Shell Electron Pair Repulsion Theory (VSEPR)
1. Explain the assumptions of the VSEPR Theory.
 2. Relate the VSEPR theory to molecular geometry.
 3. Discuss the shape, bond angle, and polarity of bent, linear, trigonal planar, tetrahedral, trigonal bipyramidal, and octahedral species.
 4. Explain the possible structures of structures containing multiple bonds.
 5. Discuss how to determine the structure of molecules containing no central atom.

V. Covalent Bonding and Hybridization

A. Hybridization and the Localized Electron Model

1. Explain the characteristics of the orbitals formed from hybridization.
2. Discuss the geometry, bonding patterns and bond angles of a hybridized atom.

3. Explain the assumptions of the Localized Electron Model.
- B. The Molecular Orbital Model and Bond Order
1. Explain how Molecular Orbitals are formed.
 2. Discuss how bond order is determined and what it signifies.
 3. Explain how to draw an orbital diagram representing the bonding in homonuclear diatomic molecules.
 4. Define paramagnetism and explain the particular molecular orbital diagram that is responsible for this phenomena.
 5. Explain the bonding in heteronuclear diatomic molecules.
 6. Discuss the combining of the Localized Electron and Molecular Orbital Model.
 7. Explain what is meant by delocalized pi bonding.
 8. Discuss how delocalized pi bonding occurs in a molecule.
 9. Emphasize the equality of all carbon-carbon bonds in molecules.

VI. Properties of Solutions

- A. Solution Composition-Concentration Expressions
1. Define mass percent and explain how to calculate it.
 2. Define mole fraction and explain how to calculate it.
 3. Define molarity and explain how to calculate it in acid-base and redox reactions.
 4. Define normality and explain how to calculate it acid-base and redox reactions.
 5. Explain how to interconvert concentration expressions.
- B. The Energies of Solutions
1. Explain the three interactions of a solute and a solvent involved in solution formation.
 2. Discuss whether these interactions are exothermic or endothermic.
 3. Explain that the enthalpy of solution is equal to the sum of these three interaction.
 4. Distinguish between the enthalpy of solution and the enthalpy of hydration.
 5. Explain the expression like dissolves like and explain how it relates to enthalpy of solution.
- C. Factors Affecting Solubility
1. Explain the structure factors that affect solubility.
 2. Explain how pressure affects the solubility of a gas and explain Henry's Law.
 3. Perform calculations using Henry's Law.
 4. Discuss the temperature affects for aqueous solutions.
 5. Explain what conditions exist when thermal pollution has occurred.
 6. Compare and contrast graphs of solubility of solids and gases as a function of temperature at constant pressure of one atmosphere.
- D. The Vapor Pressure of Solutions
1. Define equilibrium vapor pressure of a liquid.
 2. Explain that the presence of a non-volatile solute reduces the tendency

of solvent molecules to escape and lowers the vapor pressure of a solvent.

3. Calculate the vapor pressure of a solution containing an ionic solute.
 4. Distinguish between an ideal solution and a non-ideal solution.
 5. Calculate the vapor pressure of a solution containing two liquids.
 6. Calculate the vapor pressure of a solution containing an ionic solute.
 7. Define colligative property and identify the four colligative properties of solutions.
 8. Derive the formula used to calculate the boiling point elevation and the freezing point depression of a solution.
 9. Explain how to determine the molar mass of a solute by freezing point depression.
 10. Define osmotic pressure and use diagrams to explain this concept.
 11. Determine molar mass from osmotic pressure.
 12. Define what is meant by reverse osmosis and explain how it is used in the desalination process
- E. Colligative Properties of Electrolyte Solutions
1. Explain how to use the Van't Hoff factor in colligative property calculations.

VII. Chemical Kinetics

- A. Measurement of Reaction Rate
1. Examine and interpret concentration versus time data for a reaction.
 2. Distinguish between average rate and instantaneous rate.
 3. Explain how coefficients in a balanced equation affect the value of rate.
- B. Rate Laws
1. Discuss the form of the differential rate law.
 2. Distinguish between the orders of reaction.
 3. Explain how rate constant and order of reaction are determined experimentally.
 4. Distinguish between the differential rate law and the integrated rate law.
- C. Determining the Form of the Rate Law
1. Explain the method of initial rates for determining the form of the differential rate law
 2. Demonstrate the procedure for determining a rate law for a reaction with three reactants.
- D. The Integrated Rate Law
1. Explain that the integrated rate law expresses the concentration of a reactant as a function of time.
 2. Explain the form of the integrated first-order rate law and its differential rate expression.
 3. Interpret the natural log of the concentration versus time graph.
 4. Describe and explain the equation for determining the half-life of a first-order reaction.
 5. Explain that half-life of a first-order reaction is independent of initial

concentration.

6. Discuss the form of the integrated second-order rate law.
7. Derive the half-life formula for a second-order reaction.
8. Explain how to determine the order of a reaction by plotting the natural log of the concentration, or the inverse concentration or the concentration versus time.
9. Define a zero-order reaction.
10. Give the form of the integrated and differential rate laws for a zero order.
11. Give an overall summation of the important points and equations developed throughout the chapter.

E. Reaction Mechanism

1. Explain a reaction mechanism.
2. Explain the two requirements that a series of elementary steps must be satisfied.
3. Define an intermediate.
4. Distinguish between unimolecular, bimolecular and termolecular elementary steps.
5. Explain why the slow step is the rate determining steps.
6. Identify the intermediate in various reaction mechanism.

F. A Model For Chemical Kinetics

1. Explain a potential energy diagram.
2. Identify the activation energy and the activated complex.
3. Interpret a plot showing the number of collisions with a particular energy at temperature one and temperature two, where temperature 2 is greater than temperature 1.
4. Explain the Arrhenius equation.
5. Derive the equation, which is a straight line for the plot of natural log of the rate constant as a function of the inverse of the temperature.
6. Calculate activation energy of a reaction knowing the value of the rate constant at two different temperatures.

G. Catalysis

1. Define the terms catalyst and catalysis.
2. Distinguish between homogeneous and heterogeneous catalysis.

VIII. Chemical Equilibrium

A. The Equilibrium Condition

Explain the characteristics of a chemical equilibrium.

B. The Equilibrium Constant

1. Examine a concentration profile for a reaction at equilibrium.
2. Explain Guldberg and Waage's postulate of the law of mass action.
3. Calculate the equilibrium constant of a reaction given equilibrium concentrations.
4. Explain why each set of equilibrium concentrations is called an equilibrium position.

C. Equilibrium Expressions Involving Pressures.

1. Explain why partial pressures are directly related to number of moles.
2. Set up a mass action expression using partial pressures of gases at equilibrium.
3. Derive the formula which represents the relationship between K_{eq} and K_p .
4. Solve problems calculating K_{eq} from K_p
5. Distinguish between homogeneous and heterogeneous equilibria.
6. Explain that the position of a heterogeneous equilibrium does not depend on the amounts of pure solids and liquids present
7. Write equilibrium expressions for heterogeneous equilibria.

D. Applications of the Equilibrium Constant

1. Given the equilibrium constant and the initial concentration of the reactant calculate the equilibrium concentration for a reaction.
3. Explain the significance of the magnitude of an equilibrium constant.
4. Define initial reaction quotient.
5. Discuss the procedure of comparing the value of the reaction quotient to the equilibrium constant to determine in which way a system will shift to reach equilibrium.
5. Explain how to find the equilibrium concentration (or pressure) of reactants and products, given the value of the equilibrium constant and the initial concentrations (or pressure).

E. Solving Equilibrium Problems

1. List and apply the procedure for solving all types of equilibrium problems.
2. Explain that systems having a very small equilibrium constant results in a small shift to the right to reach equilibrium.
3. Demonstrate how to mathematically simplify these equilibrium calculations.

F. Le Chatelier's Principle

1. Identify the factors that are stresses to an equilibrium system.
2. State Le Chatelier's Principle.
3. Explain the effect of a change in concentration on equilibrium system and predict in which direction the equilibrium shifts.
4. Explain the effect of a change in pressure (due to the change in volume) on an equilibrium system and predict in which direction the equilibrium shifts.
5. Discuss the fact that if an inert gas is added to an equilibrium system, the total pressure increases, but this has no effect on the concentration or partial pressures of reactants and products.
6. Explain the effect of a change in temperature on a system at equilibrium.
7. Discuss the fact that temperature is the only factor that can change the value of the equilibrium constant.
8. State the rules for predicting the shift in equilibrium that occurs when temperature is changed.

IX. Acids and Bases

A. The Nature of Acids and Bases

1. Explain the Arrhenius concept and the Brønsted-Lowry model of acids and bases.
2. Draw the Lewis Structure for the hydronium ion.
3. Define conjugate acid and base and identify conjugate pairs in Brønsted-Lowry acid-base reaction.
4. Define acid ionization constant and how the mass action expression for the equilibrium constant.
5. Demonstrate the procedure for writing acid ionization reaction.

B. Acid Strength

1. List the six strong acids.
2. Discuss the conjugate base of a strong acid is weak.
3. Explain why the conjugate base of a weak acid is strong.
4. List the four ways to describe acid strength.

C. Relative Base Strength

1. List a series of bases in order of increasing strength.
2. Give examples of reactions in which water can act as an acid or a base.
3. Define amphoteric.
4. Distinguish between neutral, acidic or basic solution in terms of the concentration of the hydronium ion and the concentration of the hydroxide ion.
5. Explain the autoionization of water and determine the concentration of hydronium ion and hydroxide ion in pure water knowing that the equilibrium constant of water is 1×10^{-14} .

D. The pH Scale

1. Explain how the pH scale is organized.
2. Derive the formulas to define and calculate pH, H_3O^+ , OH^- , and pOH.
3. Discuss the inverse relationship between H_3O^+ and OH^- .
4. Using the K_w equation and the relationship between pH and pOH, calculate H_3O^+ , OH^- , and the pOH of solutions with whole number pH values.

E. Calculating the pH of Strong Acid Solutions.

1. Solve problems involving strong acids and explain how the H_3O^+ from the acid drives the autoionization for to the left.

F. Calculating the pH of Weak Acid Solution

1. Distinguish between initial concentration and equilibrium concentration.
2. Set up a concentration chart to determine representation for the equilibrium concentration to be put on mass actions expression.
3. Explain the 5% rule.
4. Discuss how to determine the pH of a mixture of weak acids
5. Explain how to calculate percent ionization.
6. Discuss how the percent ionization increases as acid becomes more dilute.
7. Calculate the ionization constant of a weak acid to give the percent ion.

G. Bases

1. Calculate the pH of a solution of a strong base.
2. Explain that the K_b of a weak base always refers to the reaction of the base with water to form the conjugate acid and hydroxide ion.
3. Discuss why amines can act as weak bases.
4. Calculate the pH of a solution of a weak base.

H. Polyprotic Acids

1. Define and give examples of polyprotic acids.
2. Explain the stepwise ionization of polyprotic acids.
3. Discuss why the equilibrium constants of successive steps become increasingly smaller.
4. Calculate the pH of a weak polyprotic acid given the K of each step.

I. Acid-Base Properties of Salt

1. Discuss the structure of salts that produce neutral, basic or acidic solutions.
2. Explain what determines the pH of a solution of a salt, which contains only ions, which hydrolyze.

J. The Effect of Structure on Acid-Base Properties

1. Explain the effect of the number of attached oxygens on the H-O in a series of chlorine oxyacids.

K. Acid-Base Properties of Oxides

1. Discuss that non-metallic oxides form acidic solutions and most metallic oxides form basic solutions.
2. List examples of amphoteric oxides.

L. The Lewis Acid-Base Model

1. Define and give an example of a Lewis acid-base reaction.
2. Identify in a reaction the Lewis acid and the Lewis base.

M. Strategy for Solving Acid-Base Problems

1. Summarize the steps in solving acid-base problems.

X. Application of Aqueous Equilibrium

A. Acid-Base Equilibria

1. Explain the equilibrium calculations involving solutions of acid or bases containing a common ion.
2. Discuss the common-ion effect and account for this phenomena using Le Chatelier's Principle.

B. Buffered Solutions.

1. Define buffered solution.
2. Explain the mechanism by which pH remains relatively constant.
3. Explain calculations involved in determining pH changes in buffered solutions.
4. Discuss the two-step procedure, stoichiometry calculation first, then do equilibrium calculations.
5. Derive the Henderson-Hasselbalch equation and demonstrate its value in solving buffer problems.

6. Explain that the capacity of a buffer depends on the concentration of the acid and base whereas; the pH of the buffer depends on the value of the acid-base ratio.
7. Explain what calculations are required in order to prepare a buffer of a particular pH.
8. Explain what the value of the acid-base ratio should be in order for the buffer to be almost effect.

C. Titrations and pH Curves

1. Explain what is meant by a pH curve (or a titration curve).
2. Examine the characteristics of the pH curve for a titration between a strong acid and a strong base and determine the pH at the equivalence point and at the midpoint.
3. Explain the calculations involved in determining the pH values at seven different points in the titration.
4. Discuss why the pH at the equivalence point in this type of titration is 7.
5. Examine the characteristics of the pH curve for a titration between a weak acid and a strong base and determine the pH at the equivalence point and at the midpoint.
6. Explain the calculations involved in determining the pH values at seven different points in the titration.
7. Explain why the pH at the equivalence point of a titration of a weak acid with a strong base is always greater than 7.
8. Interpret a graph, which reinforces the concept that the equivalence point is defined by the stoichiometry not the pH.
9. Compare and contrast the pH curve representing the titration of five different weak acids with varying K_a values with 0.1 M NaOH.
10. Examine the characteristics of the pH curve for a titration between a weak base and a strong acid and determine the pH at the equivalence point and at the midpoint.
11. Explain the calculations involved in determining the pH values at seven different points in the titration.
12. Explain why the pH at the equivalence point of a titration of a weak base and a strong acid is always less than 7.

D. Acid-Base Indicators

1. Define the role of an acid-base indicator in an acid-base titration.
2. Distinguish between the end point and the equivalence point in a titration.
3. Explain the mechanism by which an indicator changes color by a shift in equilibrium.
4. Discuss what is meant by an acid-base indicator's interval change.
5. Using the Henderson-Hasselbalch equation, demonstrate the procedure for determining the interval of change for an indicator.
6. Examine a graph and determine the useful pH ranges for several common indicators.

7. Identify the indicators that change in each of the following ranges: acidic, basic, and neutral.

E. Solubility Equilibria

1. Explain solubility equilibria.
2. Write the dissociation equations representing the solubility equilibrium established when slightly soluble substance dissolves.
3. Explain how to set up a mass-action expression for the above equilibria.
4. Explain that evaluation of the mass-action expression using equilibrium concentrations is equal to the solubility product constant or K_{sp} .
5. Solve problems involving calculating K_{sp} from solubility.
6. Solve problems involving calculating solubility from K_{sp} .
7. Discuss the significance of the magnitude of the K_{sp} .
8. Explain how to compare relative solubilities of a group of salts.
9. Describe how the common ion effect can affect the solubility of a substance.
10. Identify why the pH of a solution can affect the solubility of some substances.

F. Precipitation and Qualitative Analysis

1. Distinguish between the value of the K_{sp} , which is calculated using equilibrium concentrations and the ion product (Q), which is calculated using initial concentrations.
2. Explain the multi-step procedure to predict whether or not a precipitate will form when two solutions of known volume and molarity are mixed.
3. Discuss the need to compare the value of K_{sp} to that of the ion product in order to predict whether precipitation will occur.
4. Define what is meant by selective precipitation.
5. Explain the process in which a mixture of metal ions in aqueous solution can be separated by using a reagent whose anion forms a precipitate with one or only a few of the metal ions in the mixture.
6. Given the K_{sp} values, calculate which substance will precipitate first from a mixture of metal ions when the proper anion is added to the mixture.
7. Determine the concentration of a specific anion needed to precipitate each compound in a mixture of metallic ions.
8. Qualitative Analysis
 - a. Examine and interpret a schematic diagram of the classic method for separating the common cations by selective precipitation.
 - b. Explain the scheme for separating the insoluble chlorides, Group I cations, by selective precipitation.
9. Define complex ion.
10. Write equilibrium reactions representing complex ion formation.
11. Define ligand and coordination number.

12. Explain the mechanism of complex ion formation in a Lewis acid-base reaction.
13. Explain that the formation of a complex ion is the stepwise mechanism and that the complex ion formation can increase solubility of a slightly soluble solid.
14. Discuss how complex ion formation is an important part in the separation of ions in the scheme of qualitative analysis.

XI. The Driving Force of Reactions

- A. The First Law of Thermodynamics
 1. Interpret a potential energy diagram.
- B. Spontaneity and Reaction Rate
 1. Distinguish between spontaneity and reaction rate.
- C. Entropy
 1. Discuss the fact that the driving force for a spontaneous process is an increase in the entropy of the universe.
 2. Explain what's meant by positional probability and how it relates to phase changes.
- D. The Second Law of Thermodynamics
 1. Distinguish between the entropy of a system, of the surroundings, and of the universe.
- E. Temperature and Spontaneity
 1. Explain the effect of temperature on spontaneity.
 2. Discuss how entropy changes in the surroundings are primarily determined by heat flow.
 3. Explain that the magnitude of ΔS of the surroundings depends on the temperature at which the heat is transferred.
 4. Derive the formula for calculating ΔS of the surroundings.
- F. Free Energy
 1. Define free energy.
 2. Derive the formula for calculating ΔS of the universe.
 3. Explain what the spontaneity is favored by a decrease in enthalpy ($-\Delta H$) and an increase in entropy.
- G. Gibbs Free Energy
 1. Derive Gibbs free energy equation.
 2. Explain that the free energy of a system must be negative in order for a process to be spontaneous.
 3. Using Gibbs free energy equation, determine the sign of ΔG degree by considering the four possible combinations of signs of ΔH degree and ΔS .
 4. Identify which combinations are temperature dependent.
 5. Explain that when a process established an equilibrium; the values of ΔG degree is zero.
- H. Entropy Changes in Chemical Reactions
 1. Explain how to determine entropy changes in chemical reactions by examination of the equation.

2. Discuss how randomness of phase affects entropy changes.
 3. Identify reactions that have an entropy change determined by positional probability.
 4. Explain the effects of molecular complexity on change in entropy.
- I. Third Law of Thermodynamics
1. Explain the Third Law of Thermodynamics.
 2. Demonstrate the procedure in calculating the ΔS of a reaction using entropy values of substances.
 3. Explain that standard entropy values represent the increase in entropy when a substance is heated from 0° Kelvin to 298°K .
 4. Discuss how molecular complexity relates to the standard entropy values.
- J. Standard Free Energy
1. Define standard free energy change (ΔG°) of a chemical reaction.
 2. Explain that the more negative the value of ΔG° , the further the reaction will go to the right to reach equilibrium.
 3. Demonstrate how to calculate ΔH° , ΔS° , and ΔG° of a chemical reaction.
 4. Explain a second procedure for calculating ΔG° using Hess's Law.
 5. Discuss why the free energy change of a chemical reaction is an extensive property.
 6. Define the standard free energy of formation (ΔG_f°).
 7. Calculate the ΔG° of a reaction using ΔG_f° and Hess' Law.
 8. Solve problems determining whether a reaction is spontaneous.
- K. Free Energy Change of a Reaction
1. Explain why the free energy change of a reaction dependent on pressure.
 2. Solve problems involving calculating the free energy change of a reaction at non-standard pressures.
 3. Discuss the significance of the value of ΔG° of a reaction.
 4. Identify the value of the free energy change of a chemical reaction at equilibrium.
 5. Demonstrate how to calculate the equilibrium constant for a chemical reaction given its standard free energy change.
 6. Derive the formula, which indicates the temperature dependence of the equilibrium constant (K).
 7. Explain the relationship between free energy and work.

XII. Electrochemistry

A. Galvanic Cells

1. Distinguish between an electrochemical cell and an electrolytic cell.
2. Draw a galvanic cell and label all of its parts.
3. Explain the mechanism by which a galvanic cell operates.
 - a. Write half reactions representing this mechanism.
4. Discuss what is meant by a cell's potential.

5. Explain how to read and use a reduction potential chart to calculate the standard voltage of a galvanic cell.
 - a. Distinguish between an oxidizing agent and a reducing agent.
 - b. Calculate the cell potential and a full description of various galvanic cells.
 6. Explain how to write a line notation for a galvanic cell.
- B. The relationship between cell potential, electrical work and free energy.
1. Explain how cell potential is affected by the concentration of the half-cells.
 2. Derive the Nerst equation and use it to calculate cell potentials at non-standard concentration.
 3. Demonstrate the procedure for calculating equilibrium constants from cell potentials.
 4. Relate corrosion of iron to an oxidation-reduction reaction.
- C. Electrolytic Cells
1. Describe the reaction in an electrolytic cell as non-spontaneous, energy absorbing reactions.
 2. Discuss the stoichiometry of electrolytic processes, that is, how much product is produced with the flow of a given current for a specified time.
 3. Explain the process of electroplating and do stoichiometric calculations involving it.
 4. In the electrolysis of a mixture of metallic ions, explain that the one with the highest reduction potential void be produced first.
 5. Arrange species in order of oxidizing ability and predict which one will be reduced at the cathode of a electrolytic cell at lowest voltage.
 6. List the rules for determining what reaction goes on in electrolytic all made of pure water, an acidic solution, a basic solution or a salt solution.

XIII. Transition Metals and Coordination Chemistry

A. Complex Ion Formation

1. List appropriate ligands and transition metals that form complex ions.
 - a. Explain how to determine the coordination number of the ligand.
2. Define what a coordination number is and determine names for them.
3. Explain how complex ion formation increases the solubility of some only slightly soluble substances.
 - a. Describe how complex ion formation is used in the scheme for separating ions in qualitative analysis.

XIV. Introduction to Organic Chemistry

A. The Study of Organic Chemistry

1. Explain the structure of Alkanes: the saturated hydrocarbons.
 - a. Discuss Isomerism in alkanes.
 - b. List the rules for nomenclature of alkanes.

- c. Explain and give examples of reactions of alkanes, i.e. combustion.
 - d. Identify the structure of cyclic alkanes.
 2. Explain the structure of alkenes and alkynes: the unsaturated hydrocarbons.
 - a. Discuss and give examples of the reactions of alkenes and alkynes.
 3. Explain the structural characteristics of aromatic hydrocarbons.
 - a. Discuss the structure and names of important substituted Aromatic hydrocarbons.
 4. Define and give examples of “functional groups.”

CORE CURRICULUM CONTENT STANDARDS:

Standard 5.1 (Scientific processes) All students will develop problem-solving, decision-making and inquiry skills, reflected by formulating usable questions and hypotheses, planning experiments, conducting systematic observations, interpreting and analyzing data, drawing conclusions, and communicating results.

Standard 5.2 (Science and Society) All students will develop an understanding of how people of various cultures have contributed to the advancement of science and technology, and how major discoveries and events have advanced science and technology.

Standard (Mathematical Applications) All students will integrate mathematics as a tool for problem-solving in science, and as a means of expressing and/or modeling scientific theories.

Standard 5.4 (Nature and Process of Technology) All students will understand the interrelationships between science and technology and develop a conceptual understanding of the nature and process of technology.

Standard (Chemistry) All students will gain an understanding of the structure and behavior of matter.

Standard 5.7 (Physics) All students will gain an understanding of natural laws as they apply to motion, forces, and energy transformations.

CUMULATIVE PROGRESS INDICATORS:

Standard 5.1

A. Habits of Mind

1. When making decisions, evaluate conclusions, weigh evidence, and recognize that arguments may not have equal merit.
2. Assess the risks and benefits associated with alternative solutions.
3. Engage in collaboration, peer review, and accurate reporting of findings.
4. Explore cases that demonstrate the interdisciplinary nature of the scientific enterprise.

B. Inquiry and Problem Solving

1. Select and use appropriate instrumentation to design and conduct investigations.
2. Show that experimental results can lead to new questions and further investigations.

C. Safety

1. Understand, evaluate and practice safe procedures for conducting science investigations.

Standard 5.2

A. Cultural Contributions

1. Recognize the role of the scientific community in responding to changing social and political conditions and how scientific and technological achievement effect historical events.

B. Historical Perspectives

1. Examine the lives and contributions of important scientists who effected major breakthroughs in our understanding of the natural and designed world.
2. Discuss significant technological achievements in which science has played an important part as well as technological advances that have contributed directly to the advancement of scientific knowledge.
3. Describe the historical origin of important scientific developments such as atomic theory, genetics, plate tectonics, etc., showing how scientific theories develop, are tested, and can be replaced or modified in light of new information and improved investigative techniques.

Standard 5.3

A. Numerical Operations

1. Reinforce indicators from previous grade level.

B. Geometry and Measurement

1. When performing mathematical operations with measured quantities, express answers to reflect the degree of precision and accuracy of the input data.

C. Patterns and Algebra

1. Apply mathematical models that describe physical phenomena to predict real world events.

D. Data Analysis and Probability

1. Construct and interpret graphs of data to represent inverse and non-linear relationships, and statistical distributions.

Standard 5.4

A. Science and Technology

1. Know that scientific inquiry is driven by the desire to understand the natural world and seeks to answer questions that may or may not directly influence humans, while technology is driven by the need to meet human needs and solve human problems.

B. Nature of Technology

1. Assess the impacts of introducing a new technology in terms of alternative solutions, costs, tradeoffs, risks, benefits and environmental impact.

Standard 5.6

A. Structure and Properties of Matter

1. Know that atoms are made of a positive nucleus surrounded by negative electrons and that the nucleus, a tiny fraction of the volume of an atom, is composed of protons and neutrons, each almost 2,000 times more massive than an electron.

2. Know that the number of protons in the nucleus defines the element.
3. Know that an atom's electron arrangement, particularly the outermost electrons, determines how the atom can interact with other atoms.
4. Explain that atoms form bonds (ionic and covalent) with other atoms by transferring or sharing electrons.
5. Explain how the Periodic Table of Elements reflects the relationship between the properties of elements and their atomic structure.
6. Know that many biological, chemical and physical phenomena can be explained by changes in the arrangement and motion of atoms and molecules.
7. Recognize that the properties of matter are related to the structure and arrangement of their molecules and atoms, such as in metallic and nonmetallic crystals and carbon compounds.
8. Know that different levels of energy of an atom are associated with different configurations of its electrons.

B. Chemical Reactions

1. Explain that the rate of reactions among atoms and molecules depends on how often they encounter one another and that the rate is affected by nature of reactants, concentration, pressure, temperature, and the presence of a catalyst.
2. Show that some changes in chemical bonds require a net input or net release of energy.

Standard 5.7

A. Motion and Forces

4. Recognize that electrically charged bodies can attract or repel each other with a force that depends upon the size and nature of the charges and the distance between them and know that electric forces play an important role in explaining the structure and properties of matter.
5. Know that there are strong forces that hold the nucleus of an atom together and that significant amounts of energy can be released in nuclear reactions (fission, fusion, and nuclear decay) when these binding forces are disrupted.
6. Explain how electromagnetic, gravitational, and nuclear forces can be used to produce energy by causing chemical, physical, or nuclear changes and relate the amount of energy produced to the nature and relative strength of the force.
7. Demonstrate that moving electric charges can produce magnetic forces and moving magnets can produce electric forces.

B. Energy Transformations

1. Explain how the various forms of energy (heat, electricity, sound, light) move through materials and identify the factors that affect that movement.
2. Explain that while energy can be transformed from one form to another, the total energy of a closed system is constant.
3. Recognize that whenever mechanical energy is transformed, some heat is dissipated and is therefore unavailable for use.
4. Explain the nature of electromagnetic radiation and compare the components of the electromagnetic spectrum from radio waves to gamma rays.

SUGGESTED ACTIVITIES THAT ADDRESS THESE STANDARDS MAY INCLUDE BUT ARE NOT LIMITED TO:

Standard 5.1

Lab Safety Contract
Lab Safety Lecture
All laboratory experiments throughout the year

Standard 5.2

Notes and discussion on the evolution of the acid and base theory

Standard 5.3

Gas Law Problems
Equilibrium Problems
Reading and interpreting titration curves

Standard 5.4

Determination of pH with probeware
Titration using probeware

Standard 5.6

Notes and discussion of Gibbs Free Energy (ΔG , ΔH , ΔS)
Laboratory on Investigating Redox Reactions

Standard 5.7

Hess's Law Lab
Gibbs Free Energy Lab

INSTRUCTIONAL STRATEGIES:

Direct Instruction
Cooperative Learning

EVALUATION/ASSESSMENT OF STUDENTS:

Laboratory Reports
Quizzes
Tests

EVALUATION/ASSESSMENT OF CURRICULUM:

This course of study will be evaluated/assessed by instructional staff during the first year of implementation for the purpose of necessary revision at the end of the first year. In addition, this course of study will be reviewed according to the Five-Year Curriculum Review schedule (see attached).

RESOURCES/BIBLIOGRAPHY:

Course Book:

Chemistry 4th ed. Steven Zumdahl, Houghton Mifflin Company,
1997, ISBN 0-669-41794-7

Teacher Supplements:

Chemistry 6th ed. Raymond Chang, McGraw Hill, 1998,
ISBN 0-07-011644-X

General Chemistry 4th ed. Darrell Ebbing, Houghton Mifflin
Company, 1993, ISBN 0-395-61353-1

Chemistry and Chemical Reactivity 2nd ed. John Kotz and Keith
Purcell, Holt Rinehart Winston, 1991, ISBN 0-03-072569-0