Syllabus for Advanced Chemistry

Fall Semester 2009

General Description: This Advanced Chemistry Course is designed to be the equivalent of the general chemistry course taken during the first year of college. Students successfully completing this course will be endowed with an exceptional understanding of the fundamentals of chemistry and achieve proficiency in solving chemical problems. This course will contribute to the development of each student's ability to think critically and to express his/her ideas, in both oral and written fashion, with clarity and logic. Students must be disciplined, self-motivated and industrious.

Course Objectives:

- Quantitatively and qualitatively describe matter and its changes by applying concepts of liquids, solids, gases, solutions, chemical reactions, atomic theory, chemical bonding, nuclear chemistry, stoichiometry, equilibrium, kinetics, and thermodynamics.
- Apply and analyze chemical concepts through chemical calculations such as percent composition, molar masses, empirical formulas, gas laws, mole fractions, chemical kinetics, and standard electrode potentials and their use.
- Create, conduct, and analyze the laboratory experiments to engage and reinforce learning of concepts taught throughout the course.
- Demonstrate critical and independent thinking and an appreciation for the natural world.

Grade Scale: The DA grading scale is as follows:

A	100-93	C	76-73
A-	92-90	C-	72-70
B+	89-87	D+	69-67
B	86-83	D	66-63
B-	82-80	D-	62-60
C+	79-77	F	59 or below

Grading: All significant assignments will have written instructions including due dates and clear expectations. Refer to the course overview for further grading breakdown.

Classroom Policies:
This class will follow the DA discipline plan as described in the handbook and discussed during orientation. Students who misbehave will be given a green or red card, be excused from class, and be expected to email their parents about whatever behavior issue they struggled with.

Drop Date: The last day to drop a DA course (without incurring a *W* [withdrawal] on your transcripts) is October 5th. Please pay careful attention to this date.

Tentative Schedule for the First 16 Weeks

Adjustments will be made as necessary to meet the interests of the students and to maintain reasonable expectations. There will be little assigned homework other than finishing any work that isn't completed in class.

Date	Class time	Due Dates
week 1 8/24-8/28	**Matter and Measurement (Ch. 1)** • *Matter & Measurement Quiz* • *Pass out Periodic Table, Solubility Charts, Formula Sheets* • *Ch. 1 Text Problems: 4,9,15,19,21,25,29,31,33,35, 37, 39, 41,45,55*	 *8/27* *9/17*
week 2 8/31-9/4	**Elements, Molecules and Ions (Ch. 2)** • *Atoms, Molecules, & Ions Focus Activity (naming)* • *Nomenclature w/s* • *Naming Compounds (mixed)* • *Atoms, Molecules, & Ions Presentation* • *Ch. 2 Quiz* • *Ch. 2 Text Problems: 2,7,23,25,33,47,49,51,53,55,63*	 *8/27 – 9/2* *9/2 – 9/3* *9/3* *9/8* *9/9* *9/17*
week 3 9/7-9/11	**Stoichiometry: Calculations with Chemical Formulas** • *Functional Groups* • *Stoichiometry Focus Activity II* • *Chemical Equations & Reactions w/s*	 *9/10* *9/9 – 9/10* *9/10 – 9/14*
week 4 9/14-9/18	**Stoichiometry: Calculations with Chemical Formulas** • *Ch. 3 Text Problems: 1,5,11,17,19,33,45a,47a,59,71,79* • *Stoichiometry Quiz* • *Review Ch. 1 – 3* • *Test ch. 1 - 3*	 *9/17* *9/15* *9/16* *9/17*
week 5-6 9/21-10/2	**Aqueous Reactions and Solution Stoichiometry (ch. 4)** • *Net Ionic Equations Tutorial 5* • *Solution Chemistry Focus Activity* • *Ch. 4 Questions w/s* • *Ch. 4 Chemical Reactions Focus Activity* • *Solubility Quiz* • *Redox Assessment* • *Ch. 4 Review* • *Ch. 4 Take Home Test* • *Ch. 4 Text Problems: 5,7,13,15,17,19,21,23,27,29,43,49, 59,61,71,85*	 *9/21* *9/22* *9/22 – 9/23* *9/23* *9/24* *9/25* *9/24 -9/28* *10/1* *10/1*
week 7	**Thermochemistry (Ch. 5)** • *Thermochemistry Problems w/s*	 *10/7*

10/5-10/9	• *Ch. 5 Problems w/s* • *Ch. 5 Text Problems: 4,9,11,19,23,25,29,31,33,37,53,61, 65, 75*	*10/12* *10/15*
week 8 10/12-10/16	**Thermochemistry (Ch. 5)** • *Thermo-chemistry Problems cont. w/s* • *Ch. 5 Review* • *Ch. 5 Test*	 *10/13* *10/15* *10/15*
week 9 10/19-10/23	**Electronic Structure of Atoms (Ch. 6)** • *Ch. 6 Text Problems: 10,12,15,19,26,33,47,53,63, 67, 71, 73* • *Photoelectric Effect* • *Electromagnetic Radiation & Quantization of Energy* • *Ch. 6 Electronic Structure Focus Activity* • *Chapter 6 Questions w/s (with some gas laws)* • *Line Spectra, Bohr Model, & Quantum Theory*	 *10/29* *10/19* *10/20* *10/20* *10/22*
week 10 10/26-10/29	**Electronic Structure of Atoms (Ch. 6)** • *Ch. 6 Orbital Questions* • *Wave Behavior, Quantum Mechanics & Atomic Orbitals* • *Ch. 6 Review* • *Ch. 6 Test: Electronic Structure*	 *10/27* *10/26* *10/29* *10/29*
week 11 11/2-11/6	**Periodic Properties of the Elements (Ch. 7)** • *Ch. 7: Periodic Trends Practice* • *Chemical Periodicity Practice Problems* • *Ch. 7 Practice Problems w/s* • *Ch. 7 Review* • *Ch. 7 Periodic Properties Test* • *Ch. 7 Text Problems: 13,21,23,25,31,33,39,41,45,53,55, 57,59,67,69*	 *11/3* *11/4* *11/5* *11/9* *11/9* *11/9*
week 12 11/9-11/13	**Chemical Bonding (Ch. 8)** • *Ch. 8 Text Problems: 7,11,15,17,19,24,35,37,39,45,55, 57,69* • *Ch. 8 Practice Problems w/s* • *Ch. 8 Review*	 *12/1* *11/12* *12/1*
week 13 11/16-11/20	**Molecular Geometry and Bonding Theories (Ch. 9)** • *Ch. 9 Text Problems: 17,19,21,31,35,43,49,57,59,63,67* • *Bonding Questions Ch. 9 w/s* • *Molecular Geometry Notes* • *Molecular Geometry Practice w/s*	 *12/1* *11/19* *11/23*
week 14 11/23-11/25	**Molecular Geometry and Bonding Theories (Ch. 9)** • *Bonding Questions* • *Molecular Geometry / Bonding Free Response Questions w/s* • *Ch. 9 Review* • *Ch. 8 – 9 Test*	 *11/24* *11/25* *12/1* *12/1*

week 15 11/30-12/4	**Gases (Ch. 10)** • *Ch. 10 Text Problems: 4,9,19,21,23,25,31,35,43,45,47, 57,71,75,81* • *Ch. 10 Gases Problems Worksheet 1*	*TBA*
week 16 12/7-12/11	**Gases (Ch. 10)** • *Free Response Gas Questions w/s* • *Ch. 10 Review* • *Chapter 10 Test*	*TBA*
week 17 12/14-12/16	• *Semester Review* • *Semester Exam*	
	Intermolecular Forces, Liquids, and Solids (Ch. 11) • *Ch. 11 Text Problems: 9,11,13,15,19,25,29,39,47, 51, 57, 71,77* • *Ch. 11 Notes / Problems* • *Intermolecular Bonding – Hydrogen Bonds* • *Ch. 11 Key Terms* • *Problem Set 11.2: How Bonding affects Physical Properties* • *Ch. 11 Questions (handout)* • *Structures of Solids (notes)* • *Ch. 11 Review* • *Ch. 11 Test*	*TBA*
	Properties of Solutions (Ch. 13) • *Ch. 13 Text Problems: 6,11,13,17,27,29,33,35,37,39, 45, 47, 55,59,63,67,69,79* • *Ch. 13 Properties of Solutions (notes)* • *Colligative Properties of Solutions (notes)* • *Ch. 13 Properties of Solutions Quiz* • *Ch. 13 Questions: Solubility w/s* • *Ch. 13 Review*	*TBA*
	Chemical Kinetics (Ch. 14) • *Ch. 14 Text Problems: 5, 6, 11, 13, 17, 19, 21, 23, 25, 27, 31, 33, 35, 39, 43, 45, 51, 59, 61, 63, 67, 69* • *Ch. 14 ppt. presentation: Sect. 14.1 – 14.3* • *Chemical Kinetics Notes & Practice Problems pkt.* • *Factors Affecting Reaction Rates (Ch. 14) pkt* • *Reaction Rates (ch. 14)* • *Ch. 14 Questions pkt.* • *Determining Reaction Rates* • *Chemical Kinetic w/s + Free Response Questions* • *Ch. 14 Review Packet*	

Course Overview
Advanced Chemistry

The Advanced Chemistry Course is designed to be the equivalent of the general chemistry course taken during the first year of college. Students successfully completing this course will be endowed with an exceptional understanding of the fundamentals of chemistry and achieve proficiency in solving chemical problems. This course will contribute to the development of each student's ability to think critically and to express his/her ideas, in both oral and written fashion, with clarity and logic. Students must be disciplined, self-motivated and industrious. Upon completion of Advanced Chemistry students will be able to

- Quantitatively and qualitatively describe matter and its changes by applying concepts of liquids, solids, gases, solutions, chemical reactions, atomic theory, chemical bonding, nuclear chemistry, stoichiometry, equilibrium, kinetics, and thermodynamics.
- Apply and analyze chemical concepts through chemical calculations such as percent composition, molar masses, empirical formulas, gas laws, mole fractions, chemical kinetics and standard electrode potentials and their use.
- Create, conduct, and analyze the laboratory experiments to engage and reinforce learning of concepts taught throughout the course.
- Attain an acceptable score on the College Board AP Chemistry Examination
- Demonstrate critical and independent thinking and an appreciation for the natural world.

Laboratory

Laboratory investigation is a central pillar of the AP Chemistry course. Labs will include an emphasis on experimental procedures. Each week students will spend approximately two hours devoted to lab work outside of normal class time. Much of the laboratory work requires the use of experimental apparatus including volumetric glassware, such as pipettes and burettes. Students will gain experience with filtrations, titrations, collection and handling of gases, colorimetry, potentiometric measurements, and synthesis of compounds. Some of this laboratory work will also involve the analysis of unknown compounds either by individually devised schemes or by systematic qualitative analysis.

Students are required to complete a report for each lab experiment, including an a priori hypothesis, lab procedure, documentation of observations/data, demonstration of calculations, and a conclusion. All reports are kept in a well-organized, easily reviewed, quad-ruled, permanently bound notebook (no spiral notebooks, please!)

Materials

a. Text: Chemistry: The Central Science 10^{th} edition Brown, LeMay, Bursten
b. Lab Manual: Laboratory Experiments Chemistry The Central Science 10^{th} ed. Nelson, Kemp
c. Laboratory Notebook quad ruled
d. A graphing calculator will be **extremely** useful however not required

Homework
Homework will be assigned every night and on weekends. Homework will typically be checked at the beginning of class. For full credit you should show all work including formula units, significant figures (where applicable) and answers must be boxed.

Tests and Quizzes
There will be numerous short quizzes on topics covered in class. There will also be essay exams and unit tests that go along with each unit covered in class. Essay exams (or free response exams) will be derived from College Board free response exam questions and graded according to the College Board standards.

Grading

The grading points are divided up into three categories: Formative and Summative assignments each weighted at 37.5% of the student's grade and the semester exam which is weighted at 25% of the student grade. The point values are as follows:

Assignment type	point value
Homework (formative)	5-30 points
Quizzes (formative)	5-30 points
Tests (summative)	100 points
Labs (formative & summative)	30 points
Semester Exam	100 points

Course Outline
AP Chemistry

I. Matter and Measurement (1.0 week)

a. Matter, its classification and properties
b. Units of measurement
c. Uncertainty in measurement
d. Dimensional Analysis

The student will:

1. classify matter according to a classification scheme among the properties of homogeneous, heterogeneous, pure substance, mixture, solution, element and compound
2. understand the difference between measured numbers and exact numbers and that uncertainties always exist in measured numbers
3. solve problems using various units of measurement including those with length, mass, temperature, volume, and density
4. distinguish between accuracy and precision in measurements
5. solve problems using the metric system and dimensional analysis and proper significant figures
6. name common polyatomic ions given the formulas and vice versa

II. Elements, Molecules, and Ions (1.0 week)

a. Atomic structure, isotopes, atomic numbers, mass numbers
b. The periodic table
c. Molecules and molecular compounds
d. Ions and ionic compounds
e. Naming inorganic compounds

The student will:

1. relate atomic theory with atomic structure based on indirect evidence
2. describe atomic structure and the properties of atoms, molecules and matter
3. define and describe key terms such as isotopes, atomic number, mass number, chemical and empirical formulas
4. use the periodic table to accurately predict trends within the families and periods
5. distinguish among metals, nonmetals, and metalloids on the periodic table
6. compare empirical formulas from molecular formulas and be able to calculate empirical and molecular formulas from experimental data
7. discuss differences between ionic and molecular compounds
8. name inorganic compounds, including acid using a set of systematic rules

III. Stoichiometry: Calculations with Chemical Formulas (2.5 weeks)

a. Chemical equations
b. Atomic and molecular weights; the mole
c. Masses of atoms and molecules
d. Empirical formulas from chemical analysis
e. Quantitative information from balanced equations
f. Limiting reagents
g. % composition

The student will:

1. write balanced chemical equations to describe a chemical reaction for synthesis, decomposition, single replacement, metathesis, redox, combustion, and acid-base reactions
2. calculate the molar mass of a substance, use the molar mass and Avogadro's number to interconvert among mass, moles and number of particles of a substance
3. work problems involving mole concepts, molarity, percent composition, empirical formulas, and molecular formulas
4. solve stoichiometric problems involving percent yield, and limiting reagents

IV. Aqueous Reactions and Solution Stoichiometry (1.0 weeks)

a. Solution composition and concentration
b. Properties of solutes in solution
c. Solutions of acids, bases, and salts; neutralization
d. Ionic equations
e. Metathesis reactions
f. Solution stoichiometry and chemical analysis

The student will:

1. describe the nature of aqueous solutions through water as a solvent and strong and weak electrolytes as solutes
2. identify common strong and weak acids
3. determine the solubility of ionic compounds from general solubility rules
4. write molecular, ionic, and net ionic equations
5. identify metathesis reactions that go to completion (formation of a gas, precipitate or molecular product)
6. predict the products for reactions that are redox, neutralization, and precipitation reactions
7. perform stoichiometric calculations on acid-base volumetric (titrations), precipitation, and redox reactions

V. Thermochemistry (2.0 weeks)

a. The first law of thermodynamics
b. State Functions
c. Enthalpy
d. Enthalpy changes
e. Calorimetry including working problems with calories and specific heat
f. Hess's Law
g. Enthalpy of formation

The student will:

1. describe the energy flow between a system and its surroundings
2. Explain the significance of the first law of thermodynamics and use the law to calculate ΔE, q and w
3. define and distinguish among heat, temperature, work, energy, kinetic and potential energy
4. calculate the enthalpy change associated with phase changes
5. determine the enthalpy change or stoichiometric quantities for thermochemical equations
6. use Hess's Law to calculate the enthalpy change for a reaction
7. describe a state function
8. use standard enthalpies of formation to calculate ΔH for a reaction
9. solve calorimetry problems using $q = mc\Delta T$
10. interconvert among calories, Calories, and Joules

VI. Electronic Structure of Atoms (2.0 weeks)

a. Light and quanta
b. The Bohr model of the atom and electron energies
c. Wave behavior of matter
d. The quantum mechanical model of the atom
e. Atomic orbitals
f. Electron configurations

The student will:

1. determine from the electromagnetic spectrum: relative frequencies, wavelengths and energies
2. quantitatively and Qualitatively relate frequency, wavelength and speed of a wave
3. describe Planck's concept of quantized energy and calculate the energy of a photon using the relationship $\lambda = h\nu$
4. relate Bohr's model of the atom to the quantum theory
5. calculate the energy difference resulting from the change in energy levels of an electron
6. state the meaning and possible values of the quantum numbers and assign the quantum numbers to a given sublevel or orbital
7. use the quantum numbers, Aufbau Principle, and Hund's Rule to assign an electron configuration for a given element or ion

VII. Periodic Properties of the Elements (1.0 weeks)

a. Atomic sizes
b. Ionization energies
c. Electron affinities
d. Metals, nonmetals, and semimetals
e. Group trends for Groups 1, 2, 16, 17, and 18

The student will:

1. interpret trends within the periodic table in terms of: atomic radii, ionization energy, electron affinity, and ionic radii
2. distinguish between meals and nonmetals and semimetals
3. describe how effective nuclear charge varies with position on the periodic table
4. compare the relative energies of atomic energy levels and of sublevels

VIII. Basic Concepts of Chemical Bonding (2.5 weeks)

a. Ionic bonding and energetics of ionic bonding
b. Ionic sizes
c. Covalent Bonding
d. Lewis Structures
e. Bond polarity and Electronegativity
f. Covalent bond strength
g. Oxidation numbers and formal charge

The student will:

1. use periodic trends and electronegativity to predict bond types
2. compare and contrast different types of bonding
3. compare bond strength with ionic sizes of elements on the periodic table
4. relate the enthalpy dissociation of ionic bonding to bond strength
5. draw Lewis structures for various atoms, ions, and molecules
6. draw resonance structures for various molecules
7. use formal charges to determine the most likely resonance structure
8. compare oxidation numbers and formal charges for atoms in a molecule
9. relate electronegativity values to bond polarity
10. compare and contrast bond distance and bond energy for single and multiple bonds

IX. Molecular Geometry and Bonding Theories (2.0 weeks)

a. The VSEPR model of molecular structure
b. Polarity of molecules
c. Hybridization
d. Sigma and pi bonds

The student will:

1. use the VSEPR model to predict molecular geometry
2. determine molecular polarity using dipole moments of individual bonds
3. compare VSEPR structures to the hybridization of orbitals
4. compare and contrast sigma and pi bonds
5. predict the number of sigma and pi bonds in a structure
6. compare and contrast valence bond theory with molecular orbital theory
7. contrast molecular orbitals (delocalized) and orbitals derived from the valence-bond theory (localized)

X. Gases (2.0 weeks)

a. The ideal gas equation
b. Gas densities and molar masses
c. Partial pressures
d. Kinetic-molecular theory
e. Diffusion and effusion
f. Real gases

The student will:

1. examine the relationship between pressure, volume and temperature of ideal gases
2. apply Charles', Boyles', Gay-Lusaac, Dalton's, and the ideal gas laws quantitatively and qualitatively
3. analyze the kinetic molecular theory
4. use Grahams Law to relate the molar masses of gases to their rates or times of effusion
5. describe how real gases deviate from ideal behavior, show how van der Waals' equation allows for real conditions
6. use the ideal gas law equation to calculate the density or molar mass of a gas and solve stoichiometric calculations at standard and non-standard conditions
7. use the molar volume at STP conditions in calculations

XI. Intermolecular Forces, Liquids, and Solids (1.5 weeks)

a. Intermolecular forces (London dispersion, hydrogen, dipole-dipole, etc.)
b. Properties of liquids
c. Phase changes, phase diagrams (H_2O and CO_2)
d. Vapor pressure
e. Structures of solids

The student will:

1. describe the intermolecular forces such as dipole-dipole, hydrogen bonding, and London dispersion forces
2. describe the effects that IM forces have on the properties of liquids and solids such as melting point, boiling point, vapor pressure, viscosity, state of matter, phase changes and solubility
3. characterize the processes of evaporation, condensation, sublimation, fusion at the particle level
4. distinguish among ionic, molecular, network covalent and metallic solids with regard to particle structure, physical properties, and inter- and intra-molecular forces
5. apply the concepts of unit cells and crystal lattices for solids to calculations involving atomic radii, volume, density or identity
6. explain the relationship of boiling point to vapor pressure
7. using phase diagrams, be able to calculate energy of various phase changes for water and carbon dioxide

XII. Properties of Solutions (1.5 weeks)

a. Solution Formation, Energy changes, solvation, hydration
b. Saturated Solutions and Solubility
c. Factors Affecting Solubility
d. Ways of Expressing Concentration
e. Colligative Properties
f. Mole fractions, molar and molal solution calculations
g. Colloids

The student will:

1. define and describe solution formation, energy changes, salvation and hydration as they relate to solutions
2. describe the unique characteristics of water due to its extensive hydrogen bonding
3. compare and contrast saturated, unsaturated and supersaturated solutions; and be able to interpret graphs and charts of solubility
4. make calculations involving molarity, molality, mass percent and mole fractions as a means of expressing concentration
5. analyze the effects of colligative characteristics on the properties of solutions such as electrolytes vs. non-electrolytes
6. solve problems involving freezing point depression, boiling point elevation, vapor pressure lowering, and increase in osmotic pressure
7. use Raoult's Law to relate vapor pressure lowering to solute mole fraction
8. explain different properties of colloidal systems such as size of particles, Tyndall effect, and Brownian motion

XIII. Kinetics (2.5 weeks)

a. Reaction Rates
b. The Dependence of Rate on Concentration
c. Change of Concentration with Time
d. Temperature and Rate
e. Reaction Mechanisms
f. Catalysts
g. calculations based on kinetics and reaction rates

The student will:

1. describe the collision theory and the requirements for an effective collision
2. list factors that effect the rate of a reaction
3. use experimental data to determine the rate law and rate order of a reaction and to predict a reaction mechanism
4. interpret graphs of endothermic and exothermic reactions identifying the activation energy, enthalpies, and the reaction course with and without catalyst
5. determine a zero, first or second order reaction from graphical analysis of concentration vs. time plots
6. explain the role of a catalyst in a reaction and distinguish between homogeneous and heterogeneous catalysts
7. predict how temperature and concentration affect the rate of a reaction over time
8. use data to calculate the half life of a reaction
9. generally describe the meaning and use of the Arrhenius equation and be able to solve problems involving activation energy and the Arrhenius equation

XIV. Chemical Equilibria (4.0 weeks)

a. Concept of Equilibria
b. Equilibrium Constants – and the applications of K_p, K_c, K_w
c. Heterogeneous Equilibrium Constants
d. Calculating Equilibrium Constants and their Application
e. Le Chatelier's Principle
f. Acid and Base Equilibria
g. Bronsted-Lowry, Arrhenius, and Lewis Acids and Bases
h. Autoionization of Water
i. pH Scale
j. Strong and Weak Acids and Bases
k. Acid-Base Properties of Salt Solutions
l. Relationship between K_a and K_b
m. The Common Ion Effect
n. Buffered Solutions
o. Acid-Base Titrations
p. Solubility Equilbria
q. Factors that Affect Solubility
r. Precipitation and Separation of Ions (K_{sp})

The student will:

1. discuss the concept of equilibrium
2. write the equilibrium expression for a given equilibrium system in terms of concentrations or pressures
3. calculate values for any of the equilibrium constants
4. given Kc or Kp, calculate the equilibrium concentrations for the species in the system
5. predict the changes in equilibrium that will occur when various stresses are placed on the system (Le Chatelier's Principle): concentration change, temperature change, pressure change, and addition of a catalyst
6. calculate pH, pOH, pK, K_a, K_b, ionization constant, percent ionization, K_{sp}
7. use the reaction quotient, Q, to determine the initial direction of a reaction needed to establish equilibrium
8. write the K_w expression for water
9. explain the common ion effect
10. identify strong and weak acids and bases and write dissociation equations for each
11. predict the direction of equilibrium from knowledge of the strength of the acid-base conjugate pair in water
12. solve problems involving concentrations of substances necessary to produce a precipitate, and concentrations of ions involved in simultaneous equilibrium
13. graphically determine pKa for a weak acid from a titration curve
14. given the composition of a buffer system, determine its pH before and after the addition of known amounts of strong acid or base
15. determine the proportions in which a weak acid and its conjugate base should be mixed to give a buffer of specified pH
16. use the Henderson-Hasselbach equation in equilibrium (buffered) reactions

XV. Thermodynamics (1.5 weeks)

a. Spontaneous Processes
b. Entropy and the Second Law of Thermodynamics including thermochemical calculations
c. Molecular Interpretation of Entropy
d. Calculation of Entropy Changes
e. Gibbs Free Energy
f. Free Energy and Temperature
g. Free Energy and the Equilibrium Constant

The student will:

1. discuss the laws of thermodynamics
2. define entropy, second law of thermodynamics, PV work, enthalpy and free energy
3. use Hess's Law to solve problems of energy, entropy and free energy
4. relate the signs of ΔH and ΔS to determine the direction of a reaction / determine the spontaneity of a reaction
5. predict the sign of entropy change for a given reaction
6. apply the relationship between ΔS, surroundings, ΔH, and Temperature
7. describe how the signs of ΔH, ΔS, and ΔG relate to the spontaneity of a reaction
8. calculate the free energy change for a given reaction
9. calculate ΔS for reactions or phase changes from Absolute Entropy values

XVI. Electrochemistry (1.5 weeks)

a. Oxidation-Reduction Reactions
b. Balancing Oxidation-Reduction Equations
c. Nernst Equations and Standard electrode potentials
d. Voltaic Cells
e. Cell EMF
f. Spontaneity of Redox Reactions
g. Effect of Concentration on Cell EMF
h. Batteries
i. Corrosion
j. Electolysis
k. Faraday's constant

The student will:

1. assign oxidation states to various elements in compounds and molecules
2. recognize reactions that undergo redox reactions by comparing their oxidation states
3. recognize oxidation / reducing agents in various reactions
4. balance redox reactions using the half reaction method
5. balance redox reactions in acidic or basic solutions
6. draw and label parts of an e-cell showing electron flow
7. differentiate between galvanic and electrolytic cells
8. use the Nernst equation to calculate the EMF at non-standard conditions
9. apply Faraday's Law to electrolytic cells in calculating amount of products formed, time required, or current required
10. use the table of standard reduction potentials to determine cell voltages
11. explain the electrochemical nature of lead storage batteries, corrosion (anode and cathode protection), and fuel cells
12. define terms such as redox, anode, cathode, oxidizing agent, reducing agent, emf, electrode, Faraday's Law, voltaic cells, galvanic cells, etc.

XVII. Nuclear Chemistry (.5 weeks)

a. Radioactivity
b. Patterns of Nuclear Stability
c. Nuclear Transmutations
d. Rates of Radioactive Decay
e. Fission vs. Fusion

The student will:

1. for alpha, beta, and gamma radiation, describe each type of radiation in aspect of: mass, charge, penetrating power, and symbol
2. write balanced nuclear equations for radioactive decay and nuclear transformations
3. given a table of nuclear masses, calculate Δ mass for a nuclear reaction and relate it to the energy change, ΔE (binding energy)
4. generally describe the functioning, reactions, and positive and negative aspects of fission and fusion reactors
5. generally discuss the rates of radioactive decay

XVIII. Exam Review (2.5 to 3.5 weeks)

a. Review previously released AP Exams

b. Review net ionic equations
c. Revisit Solubility Rules
d. Practice Equilibria problems from previously released AP exams

Laboratory Guidelines

An essential part of laboratory technique is documentation of your work. A Laboratory notebook is to be purchased for that purpose. The pages of this book must be permanently bound, not a loose leaf or spiral notebook, but with perforated pages. It should be quad-ruled, with four or five squares to the inch. Also, the notebook should come with some means of making duplicate copies, either carbonless or with carbon paper inserts. The purpose of such a notebook is to provide you with a record of your work, to serve as the write-up from which your grade for each experiment will be determined, and for future reference. It may also serve as a means of communicating results among your lab partners. Therefore, its entries should be completely self-explanatory.

For all laboratory reports, you will turn in the carbon copy for evaluation. The notebook is to be organized as follows:

1. The pages must be numbered consecutively
2. The first two pages are reserved for a table of contents which lists the name of the experiment, the dates started and completed and the pages on which the data may be found.
3. For all experiments, reserve ample room for the write-up. Five pages should be sufficient
4. Use the following lab report format for your notebook:
   The title of the experiment, the date, your name (right hand page heading)
   I. **Introduction**: (Describe the nature of the lab. State the problem to be studied. Develop a hypothesis related to the problem.)
   II. **Materials**: (Copy from the text of the laboratory procedure only those materials you will use for that particular experiment.)
   III. **Methods** [or Procedures]: (May be copied from text, or rephrased in own words. Summarize!)
   IV. **Results**: **(**Set up the tables, graphs, lists needed to organize and document laboratory results.)

NOTE: THE ABOVE PORTION OF THE LABORATORY REPORT IS TO BE COMPLETED BEFORE THE LABORATORY EXERCISE HAS BEGUN. MAKE SURE THE TEACHER CHECKS (STAMPS) YOUR LABORATORY NOTEBOOK. POINTS WILL BE DEDUCTED IF THE NOTEBOOK IS NOT IN ORDER.

   V. **Results**: (Write data obtained during the exercise in the tables, charts, lists and graphs above.)
   VI. **Conclusions**: (Evaluate the experiment. What is its significance? Did circumstances produce results that were not expected or different from theoretical expectations? Was the hypothesis supported or refuted?)
   VII. DO NOT CONCLUDE “THIS WAS A GOOD LAB”, “I LEARNED A LOT FROM THIS LAB”, OR “THIS LAB DID NOT WORK”.

5. Lab notebooks must be complete at all times. All notebooks must be turned in for grading at the scheduled completion of each experiment.

6. Exam questions both in lab and in lecture will be taken from the lab experiments.