Orange Unified School District

CHEMISTRY AP

Year Course

GRADE LEVEL: 11-12

PREREQUISITES:
1. Completion of High School Chemistry with a B or higher
2. Completion of Integrated Math III or Algebra II with a B or higher and concurrent enrollment in Pre-Calculus with Trigonometry
3. Reading comprehension at 11th grade level or higher

INTRODUCTION TO THE SUBJECT:

Advanced Placement Chemistry is designed to be the equivalent of the general chemistry course usually taken during the first college year. Students will attain a depth of understanding of fundamentals and a reasonable competence in dealing with chemical problems. This course differs qualitatively from the first high school course in chemistry with respect to the kind of textbook used, the topics covered, the emphasis on chemical calculations and the mathematical formulation of principles, and the kind of laboratory work done by students. Quantitative differences appear in the number of topics treated, the time spent on the course by students, and the nature of the variety of experiments done in the laboratory.

Mathematics will be used continually to process and analyze problems both theoretical and practical. The course should contribute to the development of the students= abilities to think clearly and express their ideas, orally and in writing, with clarity and logic. Students will be expected to participate in the reviews and study sessions required by the instructor in preparation for the AP Chemistry exam given in May of the second semester. A minimum of seven hours per week is expected to be spent by the student in unsupervised reading and study, beyond assigned homework and calculations. Additional mandatory labs will be occasionally scheduled beyond the regular class period.

COURSE OBJECTIVES:

BY THE END OF THE COURSE THE STUDENT WILL BE ABLE TO:

Become familiar with the principles of chemistry and its applications.

Predict outcomes of chemical interactions.

Formulate generalizations based on a knowledge of Periodicity and Quantum Mechanics.

Perform various chemical calculations and understand mathematical formulation of principles.

Develop laboratory manipulative skills necessary to conduct chemical investigations.

Use chemical apparatus, glassware, chemicals, and other materials safely and effectively in a lab.

Design appropriate chemical investigations to answer specific chemical questions.

Record laboratory work systematically and maintain a required Laboratory Notebook.

Assess positive and negative consequences of chemical activities on the environment and society.

COURSE OVERVIEW AND APPROXIMATE UNIT TIME ALLOTMENTS:

FIRST SEMESTER	WEEKS
I. Introduction	
A. Scientific measurement	2
1. Foundations of science	
2. Accuracy and precision	
3. Significant figures in measurements and calculations	
4. The International Systems of Units: Metrics	
5. Mathematical manipulation review/unit cancellation review	
B. Chemical quantities and stoichiometric review	2
1. Chemical quantities: atomic structure, molecules, molecular formulas, percent composition, the mole, Gram Formula Mass, molar mass	
2. Reactions:	
a. equations review: writing, balancing, predicting products, net-ionic	
b. types: composition, decomposition, single-replacement, double-replacement	
3. Stoichiometry	
a. solids: mass/mass	
b. solids and gases: volume/volume, mass/volume, mass/volume	
c. liquids: concentrations, molarity of solutions	
d. limiting reagents	
II. Structure of Matter	2
A. Atomic theory and atomic structure	
1. Evidence for the atomic theory	
2. Atomic masses; determination by chemical and physical means	
3. Atomic number and mass number; isotopes	
4. Electron energy levels: atomic spectra, quantum numbers, atomic orbitals	
5. Periodic relationships including, for example, atomic radii, ionization energies, electron affinities, oxidation states	

	WEEKS
B. Chemical bonding	
1. Binding forces	2
a. Types: ionic, covalent, metallic, hydrogen bonding, van der Waals (including London dispersion forces)	
b. Relationships to states, structure, and properties of matter	
c. Polarity of bonds, electronegativities	
2. Molecular models	2
a. Lewis structures	
b. Valence bond: hybridization of orbitals, resonance, sigma and pi bonds	
c. VSEPR	
3. Geometry of molecules and ions, structural isomerism of simple organic molecules and coordination complexes; dipole moments of molecules; relation of properties to structure	
C. Nuclear chemistry: nuclear equations, half-lives, and radioactivity; chemical applications	1
III. States of Matter	
A. Gases	2
1. Laws of ideal gases	
a. Equation of state for an ideal gas	
b. Partial pressures	
2. Kinetic-molecular theory	
a. Interpretation of ideal gas laws on the basis of this theory	
b. Avogadro=s hypothesis and the mole concept	
c. Dependence of kinetic energy of molecules on temperature	
d. Deviations from ideal gas laws	
B. Liquids and solids	2
1. Liquids and solids form the kinetic-molecular viewpoint	
2. Phase diagrams of one-component systems	
3. Changes of state, including critical points and triple points	
4. Structure of solids; lattice energies	
C. Solutions	2
1. Types of solutions and factors affecting solubility	
2. Methods of expressing concentration (The use of normalities is not tested.)	
3. Raoult=s law and colligative properties (nonvolatile solutes); osmosis	
4. Non-ideal behavior (qualitative aspects)	

SECOND SEMESTER **WEEKS**

IV. Reactions

A. Reaction types — 2

1. Acid-base reactions; concepts of Arrhenius, Bronsted-Lowry, and Lewis
2. Precipitation reactions
3. Oxidation-reduction reactions
   a. Oxidation number
   b. The role of the electron in oxidation-reduction
   c. Electrochemistry: electrolytic and galvanic cells; Faraday=s laws; standard half-cell potentials; Nernst equation; prediction of the direction redox reactions

B. Stoichiometry — 1

1. Ionic and molecular species present in chemical systems: net ionic equations
2. Balancing of equations including those for redox reactions
3. Mass and volume relations with emphasis on the mole concept, including empirical formulas and limiting reactants

C. Equilibrium — 2

1. Concept of dynamic equilibrium, physical and chemical; Le Chatelier=s principle; equilibrium constants
2. Quantitative treatment
   a. Equilibrium constants for gaseous reactions: Kp and Kc
   b. Equilibrium constants for reactions in solution
      i. Constants for acids and bases; pK; pH
      ii. Solubility product constants and their application to precipitation and the dissolution of slightly soluble compounds
      iii. Common ion effect; buffers; hydrolysis

D. Kinetics — 2

1. Concept of rate of reaction
2. Use of differential rate laws to determine order of reaction and rate constant from experimental data
3. Effect of temperature change on rates
4. Energy of activation; the role of catalysts
5. The relationship between the rate-determining step and a mechanism

E. Thermodynamics — 2

1. State functions
2. First law: change in enthalpy; heat of formation; heat of reaction; Hess=s law; heat of vaporization and fusion; calorimetry

WEEKS

3. Second law: entropy; free energy of formation; free energy of reaction; dependence of change in free energy on enthalpy and entropy changes
4. Relationship of change in free energy to equilibrium constants and electrode potentials

Selected Required Laboratory Experiments: 9

.5 week/lab x 18 labs = 9 week lab total

1. Determination of the formula of a compound
2. Determination of the percentage of water in a hydrate
3. Determination of the ability of Colligative Properties to provide molar mass through utilizing both freezing-point depression and boiling-point elevation
4. Standardization of a solution using a primary standard
5. Determination of concentration by acid-base titration
6. Determination of molar mass through a chemical reaction
7. Determination of the equilibrium constant for a chemical reaction
8. Determination of the solubility of product constant for a given salt
9. Decision by experimentation of appropriate indicators for various pH and pOH ranges
10. Determination of the rate of a reaction and its order. Limited to 2nd order reactions
11. Determination of the enthalpy change associated with a reaction
12. Separation and qualitative analysis of cations and anions
13. Spectrophotometric analysis
14. Chromatographic separation
15. Confirmation of electronegativity/electrochemical series
16. Synthesis of organic compounds I
17. Synthesis of organic compounds II
18. Oxidation/reduction reaction

DATE OF CONTENT REVISION: NEW

DATE OF BOARD APPROVAL: April 15, 1999

APPENDIX

Atomic and Molecular Structure

1. The Periodic Table displays the elements in increasing atomic number and shows how periodicity of the physical and chemical properties of the elements relates to atomic structure. As a basis for understanding this concept, students know:

   a. How to relate the position of an element in the Periodic Table to its atomic number and atomic mass.
   b. How to use the Periodic Table to identify metals, semimetals, nonmetals, and halogens.
   c. How to use the Periodic Table to identify alkali metals, alkaline earth metals and transition metals, and trends in ionization energy, electronegativity, and the relative sizes of ions and atoms.
   d. How to use the Periodic Table to determine the number of electrons available for bonding.
   e. The nucleus is much smaller in size than the atom yet contains most of its mass.
   f.* How to use the Periodic Table to identify the lanthanides and actinides, and transactinide elements, and know that the transuranium elements were man made.
   g.* How to relate the position of an element in the Periodic Table to its quantum electron configuration, and reactivity with other elements in the table.
   h.* The experimental basis for Thomson=s discovery of the electron, Rutherford=s nuclear atom, Millikan=s oil drop experiment, and Einstein=s explanation of the photoelectric effect.
   i.* The experimental basis for the development of the quantum theory of atomic structure and the historical importance of the Bohr model of the atom.
   j.* Spectral lines are a result of transitions of electrons between energy levels. Their frequency is related to the energy spacing between levels using Planck=s relationship ($E=h\,v$).

Chemical Bonds

2. Biological, chemical, and physical properties of matter result from the ability of atoms to form bonds based on electrostatic forces between electrons and protons, and between atoms and molecules. As a basis for understanding this concept, students know:

   a. Atoms combine to form molecules by sharing electrons to form covalent or metallic bonds, or by exchanging electrons to form ionic bonds.

b. Chemical bonds between atoms in molecules such as H_2, CH_4, NH_3, H_2CCH_2, N_2, CI_2, and many large biological molecules are covalent.
c. Salt crystals such as NaCI are repeating patterns of positive and negative ions held together by electrostatic attraction.
d. In a liquid the inter-molecular forces are weaker than in a solid, so that the molecules can move in a random pattern relative to one-another.
e. How to draw Lewis dot structures.
f.* How to predict the shape of simple molecules and their polarity from Lewis dot structures.
g.* How electronegativity and ionization energy relate to bond formation.
h.* How to identify solids and liquids held together by Van der Waals forces or hydrogen bonding, and relate these forces to volatility and boiling/melting point temperatures.

Conservation of Matter and Stoichiometry

3. <u>The conservation of atoms in chemical reactions leads to the principle of conservation of matter and the ability to calculate the mass of products and reactants.</u> As a basis for understanding this concept, students know:

a. How to describe chemical reactions by writing balanced equations.
b. The quantity one mole is defined so that one mole of carbon 12 atoms has a mass of exactly 12 grams.
c. One mole equals 6.02×10^{23} particles (atoms or molecules).
d. How to determine molar mass of a molecule from its chemical formula and a table of atomic masses, and how to convert the mass of a molecular substance to moles, number of particles or volume of gas at standard temperature and pressure.
e. How to calculate the masses of reactants and products in a chemical reaction from the mass of one of the reactants or products, and the relevant atomic masses.
f.* How to calculate percent yield in a chemical reaction.
g.* How to identify reactions that involve oxidation and reduction and how to balance oxidation-reduction reactions.

Gases and Their Properties

4. <u>The Kinetic Molecular theory describes the motion of atoms and molecules and explains the properties of gases</u>. As a basis for understanding this concept, students know:

a. The random motion of molecules and their collisions with a surface create the observable pressure on that surface.
b. The random motion of molecules explains the diffusion of gases.
c. How to apply the gas laws to relations between the pressure, temperature, and volume of any amount of an ideal gas or any mixture of ideal gases.
d. The values and meanings of standard temperature and pressure (STP).
e. How to convert between Celsius and Kelvin temperature scales.
f. There is no temperature lower than 0 Kelvin.

g.* The kinetic theory of gases relates the absolute temperature of a gas to the average kinetic energy of its molecules and atoms.
h.* How to solve problems using the ideal gas law in the form PV=nRT.
i.* How to apply Dalton=s Law of Partial Pressures to describe the composition gases, and Graham=s Law to describe diffusion of gases.

Acids and Bases

5. Acids, bases, and salts are three classes of compounds that form ions in water solutions. As a basis for understanding this concept, students know:

a. The observable properties of acids, bases and salt solutions.
b. Acids are hydrogen-ion-donating and bases are hydrogen-ion-accepting substances.
c. Strong acids and bases fully dissociate and weak acids and bases partially dissociate.
d. How to use pH scale to characterize acid and base solutions.
e.* The Arrhenius, Bronsted-Lowry, and Lewis acid-base definitions.
f.* How to calculate pH from the hydrogen ion concentration.
g.* Buffers stabilize pH in acid-base reactions.

Solutions

6. Solutions are homogenous mixtures of two or more substances. As a basis for understanding this concept, students know:

a. Definitions of solute and solvent.
b. How to describe the dissolving process as a result of random molecular motion.
c. Temperature, pressure, and surface area affect the dissolving process.
d. How to calculate the concentration of a solute in terms of grams per liter, molarity, parts per million and percent composition.
e.* The relationship between the molality of solute in a solution, and the solution=s depressed freezing point or elevated boiling point.
f.* How molecules in solution are separated or purified by the methods of chromatography and distillation.

Chemical Thermodynamics

7. Energy is exchanged or transformed in all chemical reactions and physical changes of matter. As a basis for understanding this concept, students know:

a. How to describe temperature and heat flow in terms of the motion of molecules (or atoms)
b. Chemical processes can either release (exothermic) or absorb (endothermic) thermal energy.
c. Energy is released when a material condenses or freezes and absorbed when a material evaporates or melts.

d. How to solve problems involving heat flow and temperature changes, using known values of specific heat, and latent heat of phase change.
e.* How to apply Hess=s Law to calculate enthalpy change in a reaction.
f.* How to use the Gibbs free energy equation to determine whether a reaction would be spontaneous.

Reaction Rates

8. <u>Chemical reaction rates depend on factors that influence the frequency of collision of reactant molecules.</u> As a basis for understanding this concept, students know:

a. The rate of reaction is the decrease in concentration of reactants or the increase in concentration of products with time.
b. How reaction rates depend on such factors as concentration, temperature, and pressure.
c. The role a catalyst plays in increasing the reaction rate.
d.* The definition and role of activation energy in a chemical reaction.

Chemical Equilibrium

9. <u>Chemical equilibrium is a dynamic process at the molecular level</u>. As a basis for understanding this concept, students know:

a. How to use LeChatelier=s Principle to predict the effect of changes in concentration, temperature and pressure.
b. Equilibrium is established when forward and reverse reaction rates are equal.
c.* How to write and calculate an equilibrium constant expression for a reaction.

Organic and Biochemistry

10. <u>The bonding characteristics of carbon lead to many different molecules with varied sizes, shapes, and chemical properties, providing the biochemical basis of life</u>. As a basis of understanding this concept, students know:

a. Large molecules (polymers) such as proteins, nucleic acids, and starch are formed by repetitive combinations of simple sub-units.
b. The bonding characteristics of carbon lead to a large variety of structures ranging from simple hydrocarbons to complex polymers and biological molecules.
c. Amino acids are the building blocks of proteins.
d.* The system for naming the ten simplest linear hydrocarbons and isomers containing single bonds, simple hydrocarbons with double and triple bonds, and simple molecules containing a benzene ring.
e.* How to identify the functional groups which form the basis of alcohols, ketones, ethers, amines, esters, aldehydes, and organic acids.
f.* The R-group structure of amino acids and how they combine to form the polypeptide backbone structure of proteins.

Nuclear Processes

11. Nuclear processes are those in which an atomic nucleus changes, including radioactive decay of naturally occurring and man-made isotopes, nuclear fission, and nuclear fusion. As a basis for understanding this concept, students know:

   a. Protons and neutrons in the nucleus are held together by strong nuclear forces which are stronger than the electromagnetic repulsion between the protons.
   b. The energy release per gram of material is much larger in nuclear fusion or fission reactions than in chemical reactions: change in mass (calculated by E=mc_) is small but significant in nuclear reactions.
   c. Many naturally occurring isotopes of elements are radioactive, as are isotopes formed in nuclear reactions.
   d. The three most common forms of radioactive decay (alpha, beta, gamma) and how the nucleus changes in each type of decay.
   e. Alpha, beta, and gamma radiation produce different amounts and kinds of damage in matter and have different penetrations.
   f.* How to calculate the amount of a radioactive substance remaining after an integral number of half lives have passed.
   g.* Protons and neutrons have substructure and consist of particles called quarks.

*Standards without asterisks represent those that all students are expected to achieve in the course of their studies. Standards with asterisks represent those that all students should have the opportunity to learn.