

AP Chemistry Syllabus

Text

Chemistry and Chemical Reactivity; Kotz and Treichel; 5th ed.; Thomson—Brooks/Cole. ISBN: 0-03-033604-X.

Support Materials

Laboratory Experiments for Advanced Placement Chemistry; Vonderbrink; 1st and 2nd ed.; Flinn Scientific, Inc. ISBN: 978-1-933709-03-01.

Chemical Principles in the Laboratory; Slowinski, Wolsey, and Masterton; 7th ed.; Harcourt College Publishers. ISBN: 0-03-031167-5.

Laboratory Experiments; Nelson and Kemp; 10th Ed.; Pearson—Prentice Hall. ISBN: 0-13-146479-5.

ACS Small-Scale Laboratory Assessment Activities; Silberman and Eubanks; American Chemical Society Division of Chemical Education Examinations Institute.

ACS Test-Item Bank for General Chemistry; Eubanks and Eubanks; American Chemical Society Division of Chemical Education Examinations Institute.

AP Chemistry; Barker; Thomson—Peterson's. ISBN: 0-7689-1828-6.

5 Steps to a 5—AP Chemistry; Moore and Langley; McGraw—Hill. ISBN: 0-07-140075-1.

Cracking the AP Chemistry Exam; Foglino; The Princeton Review. ISBN: 0-375-76527-1.

Cracking the SAT Chemistry Subject Test; Silver; The Princeton Review. ISBN: 0-375-76448-8.

AP Chemistry; Dumas, Fikar, Samples, and Templin, Research and Education Association. ISBN: 0-87891-136-7.

Fast Track to a 5; Zumdahl and Zumdahl; McDougal Littell; ISBN: 0-618-22171-9.

AP Chemistry; Waterman; Pearson—Prentice Hall. ISBN: 0-13-236721-1.

Coarse Goals

- Students will learn to be critical readers and critical thinkers, applying these skills to the conceptual understanding of General Chemistry.
- Students will become independent learners, relying on support materials and one another for primary sources of information.
- Students will function safely and effectively in the Laboratory, with qualitative and quantitative analysis a primary focus. Students will be able to intelligently discuss experimental protocols and results in scientific terms in oral, written, and practicum formats.
- Students will be prepared for the AP Examination in May, and are required to take the exam to receive course credit.
- Students will be prepared for the SAT Chemistry Examination in May, and are required to take the exam to receive course credit.

Topics Covered

As pursuant to the College Board recommendations, the course will cover the following topics:

- Structure of Matter
 - I. Atomic theory and atomic structure.

- II. Chemical bonding.
 - III. Nuclear chemistry.
- States of Matter
 - I. Gases.
 - II. Liquids and solids.
 - III. Solutions.
- Reactions
 - I. Reaction types.
 - II. Stoichiometry.
 - III. Equilibrium.
 - IV. Kinetics.
 - V. Thermodynamics.
- Descriptive Chemistry
 - I. Chemical reactivity and products of chemical reactions.
 - II. Relationships in the periodic table.
 - III. Introduction to organic chemistry.
- Laboratory

Class Format

- AP Chemistry is offered as a double period. Students attend AP Chemistry for 92 minutes four days per week (Monday, Tuesday, Wednesday, and Friday). The following format is offered as a guide to weekly activities.

<u>Day</u>	<u>Designation*</u>	<u>Activities</u>
Monday	1-1	Lecture, small group and self-directed study
Tuesday	1-2	Laboratory
Wednesday	1-3	Laboratory
Friday	1-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

*regarding designation, the first number corresponds to the week of instruction (the above example is for the first week of instruction), and the second number corresponds to the day of week.

- The class will move at a quick pace, covering approximately one chapter per week. Students will have completed the primary text in late February, and will spend March and April reviewing for the AP Exam.
- Students should anticipate spending ~2 hours per week preparing for the assessment component of this course.
- Students will be assessed weekly via multiple choice and constructed response. Students who receive less than a 75% on the assessment will be required to complete all chapter questions, to be turned in the following Monday.
- All labs will be due on the Friday following completion of the lab. All lab reports will be handwritten in student's composition notebook. All labs will be wet. Failure of the lab component of the course will result in failure of the course.
- All lab reports must fulfill the lab requirements noted below.

Lab Requirements:

Labs will be performed nearly every week and consist of a large portion of your grade. The lab assignments will a pre-lab assignment, and a comprehensive

write-up, to be completed in your composition notebook and turned in on Friday of each week. The following guidelines are specific for all lab reports.

Laboratory notebook

Quantitative and qualitative data should be recorded in the lab notebook regardless of whether you are given supplemental worksheets or sample data tables.

Organization of lab notebook

- 1) Leave pages 1-4 blank. This is your table of contents.
- 2) Each lab should have a title (top line) and date of experiment (right of title). The pages should be numbered in the lower right-hand corner. Fill in the table of contents accordingly.
- 3) Always use black ink. Never use blue ink or pencil.
- 4) Never use white out to correct an error. If you make a mistake, draw a single line through the error and continue the entry. For Example:
Chemistry is ~~a pain in the~~ fun and exciting.
- 5) Information you wish to be graded will be recorded on the right-hand pages only. The left-hand pages are for your own use (notes, rough data and calculations, etc.).
- 6) The format for the lab notebook must include the following.
 - a. Purpose—what is it you intend to investigate.
 - b. Equipment and Reagents.
 - c. Chemical Equations—all that are relevant; this includes state of matter transitions.
 - d. Procedure—In your own words; flow charts are acceptable.
 - e. Data and Calculations—this includes both qualitative and quantitative data; you must also include sample calculations for each type of mathematical operation utilized. Error analysis is mandatory for each lab.
 - f. Discussion—what was the chemistry relevant to the experiment? This is the most critically graded portion of the report. Be specific and speak in terms relative to your scientific education. This is also where you will discuss error analysis, but you must be specific regarding impacts of errors related to protocol and the quantitative impact on your results.
 - g. Conclusion—answers purpose.

Semester 1

Summer Homework

1. Read Chapters 1-3; answer the following problems showing all work.
 - a) page 38-43 #1-77
 - b) page 73-77 #1-68
 - c) page 111-117 #1-113

2. Study for all ion quiz to be given on day 1-1.
3. Study for CH 1-3 Exam to be administered on day 1-5.

Week 1 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 1):

- Identify the name or symbol for an element, given its symbol or name.
- Use the terms atoms, molecule, element, and compound correctly.
- Identify physical properties of matter and give some examples.
- Use density as a way to connect the volume and mass of a substance.
- Convert between temperatures on the Celsius and Kelvin scales.
- Understand the difference between extensive and intensive properties and give examples.
- Explain the difference between chemical and physical change.
- Recognize the different states of matter and give their characteristics.
- Understand the basic ideas of kinetic-molecular theory.
- Understand the difference between matter represented at the macroscopic level and at the particulate level.
- Appreciate the difference between pure substances and mixtures and the difference between homogenous and heterogeneous mixtures.
- Recognize and know how to use the prefixes that modify metric units.
- Use dimensional analysis to carry out unit conversions and other calculations.
- Know the difference between precision and accuracy and how to calculate percent error.
- Understand the use of significant figures.

Students will understand the following items conceptually and with relation to problem solving (Chapter 2):

- Explain the historical development of the atomic theory and identify some of the scientist who made important contributions.
- Describe electrons, protons, and neutrons, and the general structure of the atom.
- Understand the relative mass scale and the atomic mass unit.
- Define isotope and give the mass number and number of neutrons for a specific isotope.
- Calculate the atomic weight of an element from isotope abundances and masses.
- Understand that the molar mass of an element is the mass in grams of Avogadro's number of atoms of an element.
- Know how to use the molar mass of an element in calculations.
- Identify the periodic table location of groups, periods, metals, metalloids, non-metals, alkali metals, alkaline earth metals, halogens, noble gases, and the transition elements.
- Examine horizontal, vertical, and diagonal relationships of the periodic table corresponding to locations of groups, periods, metals, metalloids, non-metals, alkali metals, alkaline earth metals, halogens, noble gases, and the transition elements.

Students will understand the following items conceptually and with relation to problem solving (Chapter 3):

- Interpret the meaning of molecular formulas, condensed formulas, and structural formulas.
- Recognize that metal atoms commonly lose one or more electrons to form positive ions, and nonmetal atoms often gain electrons to form negative ions.
- Recognize that the charge on a metal cation is equal to the number of the group in which the element is found in the periodic table. Transition metal cations are often +2 or +3, but other charges are observed.
- Recognize that the negative charge on a single-atom or monatomic anion, X^{n-} , is given by $n=8 - \text{Group number (on the periodic table)}$.
- Give the names of formulas of polyatomic ions, knowing their formulas or names, respectively.
- Write the formulas for ionic compounds by combining ions in the proper ratio to give no overall charge.
- Understand the importance of Coulomb's law. Electrostatic forces are responsible for the attraction or repulsion of charged species. The magnitude of the force is given by Coulomb's law, which states that the force of attraction between oppositely charged species increase with electric charge and with decreasing distance between the species.
- Name the ionic compounds and simple binary compounds of nonmetals.
- Understand that the molar mass of a molecule compound is the mass in grams of Avogadro's number of molecules and that the molar mass of an ionic compound is the mass in grams of Avogadro's number of the combination of ions shown in the formula.
- Calculate the molar mass of a compound from its formula and a table of atomic weights.
- Calculate the amount of a compound represented by a given mass, and vice versa.
- Express the composition of a compound in terms of percent composition.
- Use the percent composition to determine the empirical formula of a compound.
- Understand how mass spectrometry can be used to find a molar mass.
- Use experimental data to find the number of water molecules in a hydrated compound.

Weekly Schedule:

Day	Designation	Activities
Monday	1-1	All ions quiz
Tuesday	1-2	Laboratory—Copper Reactions and % yield.
Wednesday	1-3	Laboratory—Copper Reactions and % yield
Friday	1-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 2 Objectives

Students will understand the following items conceptually and with relation to problem solving (Chapter 4):

- Interpret information conveyed by a balanced chemical equation.
- Balance simple chemical equations.
- Understand the principle of the conservation of matter, the basis of chemical stoichiometry.
- Calculate the mass of one reactant or product from the mass of another reactant or product by using the balanced chemical equation.
- Understand the impact of a limiting reactant on the outcome of a chemical reaction.
- Determine which of two reactants is the limiting reactant.
- Determine the quantity of a product based on the limiting reactant.
- Explain the difference among actual yield, theoretical yield, and percent yield, and calculate percent yield.
- Use stoichiometry principles to analyze a mixture or to find the empirical formula of an unknown compound.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	2-1	Lecture, small group and self-directed study
Tuesday	2-2	Laboratory—Synthesis of Alum ($\text{AlK}(\text{SO}_4)_2 \cdot 12 \text{H}_2\text{O}$)
Wednesday	2-3	Laboratory—Synthesis of Alum ($\text{AlK}(\text{SO}_4)_2 \cdot 12 \text{H}_2\text{O}$)
Friday	2-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 3 Objectives

Students will understand the following items conceptually and with relation to problem solving (Chapter 5):

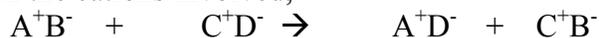
- Explain the difference between electrolytes and nonelectrolytes.
- Predict the solubility of ionic compounds in water.
- Recognize what ions are formed when an ionic compound or acid or base dissolve in water.
- Predict the products of precipitation reactions, the formation of an insoluble reaction product by the exchange of anions between the cations of the reactants.
- Write the net ionic equation for a given reaction.
- Recognize common acids and bases and understand their behavior in aqueous solution.
- Predict the products of acid-base reactions involving common acids and strong bases.
- Understand that the net ionic equation for the reaction of a strong acid with a strong base is $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l})$.
- Predict the products of gas forming reactions, the most common of which are those between a metal carbonate and an acid.

$$\text{NiCO}_3(\text{s}) + 2 \text{HNO}_3(\text{aq}) \rightarrow \text{Ni}(\text{NO}_3)_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$
- Use ideas developed as an aid in recognizing four of the common types of reactions that occur in aqueous solution and write balanced equations for such reactions.

Reaction Type Driving Force

Precipitation	Formation of an insoluble compound.
Acid-strong base	Formation of a salt and water.
Gas-forming	Evolution of a gas such as CO ₂ .
Oxidation-reduction	Transfer of electrons.

Note that the first three of these reaction types involve the exchange of anions between the cations involved,



And so are called exchange reactions. The fourth (redox reactions) involves the transfer of electrons.

- Recognize oxidation numbers of elements in a compound and understand that these numbers represent the charge an atom has, or appears to have, when the electrons of the compound are counted according to periodic table location.
- Recognize oxidation-reductions reactions (often called redox reactions).
- Calculate the concentration of a solute in a solution in units of moles per liter (mol/L) (molarity), and use concentrations in calculations.
- Describe how to prepare a solution of a given molarity from the solute and a solvent or by dilution from a more concentrated solution.
- Calculate the pH of a solution containing a strong acid or base and know what this means in terms of the relative amount of hydrogen ion in solution. Calculate the hydrogen ion concentration of a solution from the pH.
- Solve stoichiometry problems using solution volume and concentrations.
- Explain how a titration is carried out, explain the procedure in standardization and calculate concentrations or amounts of reactants from titration data.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	3-1	Lecture, small group and self-directed study
Tuesday	3-2	Laboratory—Analysis of Alum (AlK(SO ₄) ₂ *12 H ₂ O)
Wednesday	3-3	Laboratory—Analysis of Alum (AlK(SO ₄) ₂ *12 H ₂ O)
Friday	3-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 4 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 6):

- Describe various forms of energy and the nature of heat and thermal energy transfer.
- Use the most common energy unit, the joule, and convert other energy units to joules.
- Recognize and use the language of thermodynamics: the system and its surrounding; exothermic and endothermic reactions.
- Use specific heat capacity in calculations of heat transfer and temperature changes.
- Understand the sign conventions in thermodynamics.
- Use heat of fusion and heat of vaporization to find the quantity of thermal energy involved in changes of state.
- Understand the basis of the first law of thermodynamics.

- Recognize state functions whose value is determined only by the state of the system and not by the pathway taken by the process.
- Recognize that when a process is carried out under constant pressure conditions, the heat transferred is the enthalpy change, ΔH .
- Describe how to measure the quantity of heat energy transferred in a reaction by using calorimetry.
- Apply Hess's law to find the enthalpy change for the reaction.
- Know how to draw and interpret energy level diagrams.
- Use standard molar enthalpy of formation, ΔH_f° , to calculate the enthalpy change for a reaction, ΔH°_f .
- Discuss uses of energy in our economy and problems and opportunities for developing new energy resources.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	4-1	Lecture, small group and self-directed study
Tuesday	4-2	Laboratory—Unknown Carbonate Determination
Wednesday	4-3	Laboratory—Unknown Carbonate Determination
Friday	4-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 5 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 7):

- Use the terms wavelength, frequency, amplitude, and node.
- Use Equation 7.1 ($c = \lambda \times \nu$), the relationship between wavelength (λ) and frequency (ν) of electromagnetic radiation and the speed of light (c).
- Recognize the relative wavelength (or frequency) of the various types of electromagnetic radiation.
- Understand that the energy of a photon, a massless particle of radiation, is proportional to its frequency (Planck's equation, Equation 7.2). This is an extension of Planck's idea that energy at the atomic level is quantized.
- Describe the Bohr model of the atom, how it can account for the emission line spectra and excited hydrogen atoms, and the limitations of the model.
- Understand that, in the Bohr model of the H atom, the electron can occupy only certain energy levels, each with an energy proportional to $1/n^2$ ($E = -Rhc/n^2$), where n is the principal quantum number. If an electron moves from one energy state to another; the amount of energy absorbed or emitted in the process is equal to the difference in energy between the two states.
- Understand that in the modern view of the atom, electrons are described by the physics of waves. The wavelength of an electron or any subatomic particle is given by the Broglie's equation.
- Recognize the significance of quantum mechanics in describing the modern view of atomic structure.
- Understand that an orbital for an electron in an atom corresponds to the allowed energy of that electron.

- Understand that the position of the electron is not known with certainty; only the probability of the electron being at a given point of space can be calculated. This is the interpretation of the quantum mechanical model and embodies the postulate called the uncertainty principle.
- Describe the allowed energy states of the electron in an atom using three quantum numbers n , l , and m_l which are assigned based on periodic table location.
- Describe the shapes of the orbitals, which are based on periodic table location.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	5-1	Lecture, small group and self-directed study
Tuesday	5-2	Laboratory—Thermochemistry and Hess's Law
Wednesday	5-3	Laboratory—Thermochemistry and Hess's Law
Friday	5-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 6 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 8):

- Classify substances as paramagnetic (attracted to a magnetic field; characterized by unpaired electron spins) or diamagnetic (repelled by a magnetic field; all electrons paired).
- Recognize that each electron in an atom has a different set of the four quantum numbers, n , l , m_l , and m_s , the spin quantum number, has values of $+\frac{1}{2}$ or $-\frac{1}{2}$, all of which are assigned based on periodic table location.
- Understand that the Pauli exclusion principle leads to the conclusion that no atomic orbital can be assigned more than two electrons and that the two electrons in an orbital must have opposite spins (different values of m_s).
- Understand effective nuclear charge, Z^* , and how it can be used to explain why different subshells in the same shell have different energies. Also, understand the role of Z^* in determining the properties of atoms, which is related to periodic table location.
- Use the periodic table as a guide to depict electron configurations of elements and monatomic ions using an orbital box notation or an *spdf* notation. (In both cases, configurations can be abbreviated with the noble gas notation).
- Recognize that electrons are assigned to the subshells of an atom in order of increasing subshell energy. In the H atom the subshell energies increase with increasing n , but in a many-electron atom, the energies depend on both n and l .
- Apply the Pauli exclusion principle and Hund's rule when assigning electrons to atomic orbitals.
- Predict how properties of atoms — size, ionization energy (*IE*), and electron affinity (*EA*) — change on moving down a group or across a period of the periodic table. The general periodic trends for these properties are
 - (a) Atomic size decreases across the period and increases down a group.
 - (b) *IE* increases across the period and decreases down a group.

(c) The affinity for an electron increases generally across a period (the value of EA becomes more negative) and decreases down a group.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	6-1	Lecture, small group and self-directed study
Tuesday	6-2	Laboratory—Analysis of Commercial Bleach
Wednesday	6-3	Laboratory—Analysis of Commercial Bleach
Friday	6-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 7 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 9):

- Describe the basic forms of chemical bonding, ionic and covalent, and the differences between them.
- Predict from the formula whether a compound has ionic or covalent bonding, based on whether a metal is part of the formula.
- Write Lewis symbols for atoms.
- Describe the basic ideas of ionic bonding and how such bonds are affected by the sizes and charges of the ions.
- Understand lattice energy and know how lattice energies are calculated (Born-Haber cycle); recognize trends in lattice energy and how melting points of ionic compounds are correlated with lattice energy.
- Draw Lewis structures for molecular compounds and ions.
- Understand and apply the octet rule; recognize exceptions to the octet rule.
- Write resonance structures, understand what resonance means, and how and when to use this method of representing bonding.
- Calculate formal charges for atoms in a molecule based on the Lewis structure.
- Define electronegativity based on periodic table location, and understand how it is used to describe the unequal sharing of electrons between atoms in a bond.
- Combine formal charge and electronegativity to gain a perspective on the charge distribution in covalent molecules and ions.
- Define and predict trends in bond order, bond length, and bond dissociation energy.
- Use bond dissociation energies, D , in calculations.
- Predict the shape or geometry of molecules and ions of main group elements using VSEPR theory, and understand the relationships among valence electron pairs and electron-pair, molecular geometry, and molecular polarity.
- Understand why some molecules are polar and others are nonpolar.
- Predict the polarity of a molecule.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	7-1	Lecture, small group and self-directed study
Tuesday	7-2	Laboratory—The Geometric Structure of Molecules.

Wednesday 7-3
Friday 7-5

Laboratory—The Geometric Structure of Molecules.
Study (1st 46 minutes) and Assessment (2nd 46 minutes)

Week 8 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 10):

- Describe the main features of valence bond theory and molecular orbital theory, the two commonly used theories for covalent bonding.
- Recognize that the premise for valence bond theory is that bonding results from the overlap of atomic orbitals. By virtue of the overlap of orbitals, electrons are concentrated (or localized) between two atoms.
- Distinguish how sigma (σ) or pi (π) bonds arise. For σ bonding, orbitals overlap in head-to-head fashion, concentrating electrons along the bond axis. Sideways overlap of p atomic orbitals results in π bond formation, with electrons above and below the molecular plane.
- Use the concept of hybridization to rationalize molecular structure.

Hybrid Orbitals	Atomic Orbitals Used	Number of Hybrid Orbitals	Electron-Pair Geometry
Sp	s+p	2	Linear
sp ²	s+p+p	3	Trigonal-planar
sp ³	s+p+p+p	4	Tetrahedral
sp ³ d	s+p+p+p+d	5	Trigonal-bipyramid
sp ³ d ²	s+p+p+p+d+d	6	Octahedral

- Understand molecular orbital theory, in which atomic orbitals are combined to form bonding orbitals, nonbonding orbitals, or antibonding orbitals, that are delocalized over several atoms. In this description, the electrons of the molecule or ion are assigned to the orbitals beginning with the one at lowest energy, according to the Pauli exclusion principle and Hund's rule.
- Use molecular orbital theory to explain the properties of O₂ and other diatomic molecules.
- Appreciate band theory and how it applies to solids, especially metals.
- Understand the difference among an electric conductor, a semiconductor, and an insulator.
- Recognize how dopants such as Group 3A and Group 5A elements affect semiconductor properties of Group 4A materials.

Weekly Schedule:

Day	Designation	Activities
Monday	8-1	Lecture, small group and self-directed study
Tuesday	8-2	Laboratory—Aspirin and Oil of Wintergreen Synthesis.
Wednesday	8-3	Laboratory—Aspirin Analysis.

Friday 8-5 Study (1st 46 minutes) and Assessment (2nd 46 minutes)

Week 9 Objectives:

Students will prepare for midterm examination covering chapters 1-10.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	9-1	Lecture, small group and self-directed study
Tuesday	9-2	Lecture, small group and self-directed study
Wednesday	9-3	Midterm Exam (Multiple Choice)
Friday	9-5	Midterm Exam (Constructed Response)

Week 10 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 11):

- Understand the factors that contribute to the large numbers of organic compounds and the wide array of structures.
- Recognize and draw structures of structural isomers and stereoisomers for carbon compounds, including geometric isomers and optical isomers.
- Draw structural formulas and name simple hydrocarbons, including alkanes, alkenes, and alkynes, and aromatic compounds.
- Identify possible isomers that have a given formula.
- Summarize some reactions for the various types of hydrocarbons.
- Describe the physical and chemical properties of hydrocarbons.
- Name and draw structures of alcohol and amines.
- Name and draw structures of carbonyl compounds: aldehydes, ketones, acids, esters, and amides.
- Know the structures and properties of several natural products including carbohydrates, fats and oils, and proteins.
- Write equations for the formation of addition polymers and condensation polymers, and describe the structures.
- Relate properties of the polymers to their structures.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	10-1	Lecture, small group and self-directed study
Tuesday	10-2	Laboratory—Molar Mass of a Volatile Liquid.
Wednesday	10-3	Laboratory—Molar Mass of a Volatile Liquid.
Friday	10-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 11 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 12):

- Describe how pressure measurements are made.
- Use the units of pressure, especially atmosphere (atm) and millimeters of mercury (mm Hg).

- Understand the basis of the gas laws and how to use these laws.
- Understand the ideal gas law and how to use this equation.
- Calculate the molar mass of a compound from knowledge of the pressure of a known quality of gas in a given volume at a known temperature.
- Apply the gas laws to a study of the stoichiometry of reactions.
- Use Dalton's law.
- Apply the kinetic-molecular theory of gas behavior at the molecular level.
- Understand the phenomena of diffusion and effusion and how to use Graham's law.
- Appreciate the fact that gasses usually do not behave as ideal gases. Deviations from ideal behavior are largest at high pressure and low temperature.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	11-1	Lecture, small group and self-directed study
Tuesday	11-2	Laboratory—Vapor Pressure and Enthalpy of H ₂ O.
Wednesday	11-3	Laboratory—Vapor Pressure and Enthalpy of H ₂ O.
Friday	11-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 12 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 13):

- Use the kinetic-molecular theory to define the difference in solids, liquids, and gases.
- Describe the different intermolecular forces in liquids and solids.
- Tell when two molecules can interact through the dipole-dipole attraction or when hydrogen bonding may occur. The latter occurs strongly when H is attracted to O, N, or F.
- Identify instances in which molecules interact only by induced dipoles (dispersion forces).
- Explain how hydrogen bonding affects the properties of water.
- Explain the process of evaporation of a liquid or condensation of its vapor, and use enthalpy of vaporization in calculations.
- Define the concept of equilibrium vapor pressure of a liquid and its relation to the boiling point of a liquid.
- Describe the phenomena of the critical temperature, T_c , and critical pressure, P_c of a substance.
- Describe how intermolecular forces affect (a) the cohesive forces between identical liquid molecules, (b) the energy necessary to break through the surface of a liquid (surface tension), and (c) the resistance of flow, or viscosity, of liquids.
- Characterize different types of solids: metallic (such as copper), ionic such as NaCl and CaF₂), molecular (for example, water and I₂), network (such as diamond), and amorphous (for example, glass and many synthetic polymers).
- Describe the three types of cubic unit cells: primitive or simple cubic (sc); body-centered cubic (bcc); and face centered cubic (fcc).

- Understand the relation between the unit cell for an ionic compound and its formula.
- Define and use the enthalpy of fusion.
- Identify the different points (triple point, normal boiling point, freezing point) and regions (solid,, liquid, vapor) of a phase diagram, and use the diagram to evaluate the vapor pressure of a liquid or the relative densities of a liquid and solid.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	12-1	Lecture, small group and self-directed study
Tuesday	12-2	Laboratory—Empirical Formula of Silver Oxide.
Wednesday	12-3	Laboratory—Empirical Formula of Silver Oxide.
Friday	12-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 13 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 14):

- Define the terms solution, solute, solvent, and colligative properties.
- Use the following concentration units: molality, mole fraction, weight percent and parts per million.
- Describe the process of dissolving a solute into a solvent, including the energy changes that might occur.
- Understand the distinctions between saturated, unsaturated, and supersaturated solutions.
- Define and illustrate the terms miscible and immiscible.
- Relate lattice energy and enthalpy of hydration to the enthalpy of solution for an ionic solute.
- Describe the effect of pressure and temperature on the solubility of a solute.
- Use Henry's law to calculate the solubility of a gas in a solvent.
- Apply Le Chatelier's principle to the change of solubility in gases with pressure and temperature changes.
- Calculate the mole fraction of a solute or solvent (X_{solv}) and the effect of a solute on solvent vapor pressure (P_{solv}) using Raoult's law.
- Calculate the boiling point elevation or freezing point depression caused by a solute in a solvent.
- Use colligative properties to determine the molar mass of a solute.
- Characterize the effect of ionic solutes on colligative properties.
- Use the van't Hoff factor, i , in calculations involving colligative properties.
- Calculate the osmotic pressure (Π) for solutions, and use the equation defining osmotic pressure to determine the molar mass of a solute.
- Recognize the difference between a homogenous solution, a suspension, and a colloid (or colloidal dispersion).
- Recognize hydrophobic and hydrophilic colloids.
- Describe the action of a surfactant.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	13-1	Lecture, small group and self-directed study
Tuesday	13-2	Laboratory—The Structure of Crystals.
Wednesday	13-3	Laboratory—The Structure of Crystals.
Friday	13-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 14 Objectives:

Students will prepare for an examination covering chapters 11-14.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	14-1	Lecture, small group and self-directed study
Tuesday	14-2	Lecture, small group and self-directed study
Wednesday	14-3	CH 11-14 Exam (Multiple Choice)
Friday	14-5	CH 11-14 (Constructed Response)

Week 15 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 15):

- Explain the concept of reaction rate.
- Describe the average and instantaneous rate of a reaction from experimental information.
- Describe the various conditions that affect reaction rate (that is, reactant concentrations, temperature, presence of catalyst, and the state of the reactants).
- Define the various parts of a rate equation (the rate constant and order of the reaction) and understand their significance
- Derive a rate equation from experimental information.
- Describe and use the relationships between reactant concentration and time for zero-order, first-order, and second-order reactions.
- Apply graphical methods for determining reaction order and the rate constant from experimental data.
- Use the concept of half life ($t_{1/2}$), especially for first-order reactions.
- Describe the collision theory of reaction rates.
- Relate activation energy (E_a) to the rate and thermodynamics of a reaction using reaction coordinate diagrams.
- Use the collision theory to describe the effect of reactant concentration on reaction rate.
- Describe the functioning of a catalyst and its effect on the activation energy and mechanism of a reaction.
- Define homogenous and heterogeneous catalysts.
- Understand the effect of molecular orientation on reaction rate.
- Describe the effect of temperature on reaction rate using the collision theory of reaction rates and the Arrhenius equation.

- Use the appropriate equations to calculate the activation energy from experimental data.
- Understand the concept of reaction mechanism (the sequence of bond-making and bond-breaking steps that occurs during the conversion of reactants to products) and the relation of the mechanisms to the overall, stoichiometric equation for a reaction.
- Describe the elementary steps of a mechanism and give their molecularity.
- Recognize the rate-determining step in a mechanism and identify any reaction intermediates.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	15-1	Lecture, small group and self-directed study
Tuesday	15-2	Laboratory—Kinetics—Oxidation of Iodine by Bromate.
Wednesday	15-3	Laboratory—Kinetics—Oxidation of Iodine by Bromate.
Friday	15-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 16 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 16):

- Understand the nature and characteristics of the state of equilibrium: (a) chemical reactions and reversible and (b) equilibria and dynamic.
- Write the reaction quotient, Q , for a chemical reaction. When the system is at equilibrium, the (products/reactants) quotient is called the equilibrium constant expression and has a constant value called the equilibrium constant, which is symbolized by K .
- Recognize that the concentrations of solids, pure liquids, and solvents (such as water) are not included in the equilibrium constant expression.
- Recognize that the large value of K ($K \gg 1$) means the reaction is product-favored, and the product concentrations are greater than the reactant concentrations at equilibrium. A small value of K ($K \ll 1$) indicates a reactant-favored reaction in which the product concentrations are smaller than the reactant concentrations at equilibrium.
- Appreciate the fact that equilibrium concentrations may be expressed in terms of reactant and product concentrations (expressed in moles per liter), and K is sometimes designated as K_c . Alternatively, concentrations of gases may be represented by partial pressures, and K for such cases is designated as K_p .
- Use the reaction quotient (Q) to decide if a reaction is at equilibrium ($Q = K$), or if there will be a net conversion of reactants to products ($Q < K$) or products to reactants ($Q > K$) to attain equilibrium.
- Use equilibrium constants to calculate to calculate the concentration of reactant or product at equilibrium.
- Know how K changes as different stoichiometric coefficients are used in a balanced equation, or if the equation is reversed, or if several equations are added to give a new, net equation.

- Know how to predict, using Le Chatelier's Principle, the effect of the disturbance on a chemical equilibrium: a change in temperature, a change in concentration, or a change in volume or pressure for a reaction involving gases.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	16-1	Lecture, small group and self-directed study
Tuesday	16-2	Laboratory— K_{sp} of Calcium Hydroxide.
Wednesday	16-3	Laboratory— K_{sp} of Calcium Hydroxide.
Friday	16-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 17 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 17):

- Define and use the Brønsted concept of acids and bases.
- Recognize common monoprotic and polyprotic acids and bases and write balanced equations for their ionization in water.
- Appreciate when a substance can be amphoteric.
- Recognize the Brønsted acid and base in a reaction and identify the conjugate partner of each.
- Understand the concept of water autoionization and its role in Brønsted acid-base chemistry. Use the water ionization constant, K_w .
- Use the pH concept.
- Identify common strong acids and bases.
- Recognize some common weak acids, including neutral molecules (such as acetic acid), cations (such as NH_4^+ or hydrated metal ions such as $\text{Fe}(\text{H}_2\text{O})_6^{2+}$) and anions (such as HCO_3^-).
- Write equilibrium constant expressions for weak acids and bases.
- Calculate $\text{p}K_a$ from K_a (or K_a from $\text{p}K_a$) and understand how $\text{p}K_a$ is correlated with acid strength.
- Understand the relation of K_a for a weak acid to K_b for its conjugate base.
- Write acid-base reactions and decide whether they are product- or reactant- favored.
- Recognize the types of acid-base reactions and describe their result.
- Calculate the equilibrium constant for a weak acid (K_a) or weak base (K_b) from experimental information (such as pH, $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$).
- Use the equilibrium constant and other information to calculate the pH of a solution of weak acid or weak base.
- Describe the acid-base properties of salts and calculate the pH of a solution of a salt of a weak acid or of a weak base.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	17-1	Lecture, small group and self-directed study
Tuesday	17-2	Laboratory—Dissociation Constant Determination of Weak Acids.
Wednesday	17-3	Laboratory—Dissociation Constant Determination of Weak Acids.

Friday 17-5 Study (1st 46 minutes) and Assessment (2nd 46 minutes)

Week 18 Objectives:

Students will prepare for semester 1 final examination covering chapters 1-17.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	18-1	Lecture, small group and self-directed study
Tuesday	18-2	Lecture, small group and self-directed study
Wednesday	18-3	Final Exam (Multiple Choice)
Friday	18-5	Final Exam (Constructed Response)

Semester 2

Week 19 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 18):

- Predict the effect of the addition of a common ion on the pH of the solution of a weak acid or base.
- Describe the functioning of buffer solutions.
- Use the Henderson-Hasselbalch equation to calculate the pH of a buffer solution of given composition.
- Describe how a buffer solution of a given pH can be prepared.
- Calculate the pH of a buffer solution before and after adding excess acid or base.
- Predict the pH of an acid-base reaction at its equivalence point.

Acid	Base	pH at Equivalence Point
Strong	Strong	=7 (neutral)
Strong	Weak	<7 (acidic)
Weak	Strong	>7 (basic)
Weak	Weak	Depends on <i>K</i> values of conjugate base and acid

- Calculate the pH at the equivalence point in the reaction of a strong acid with a strong base or weak base, or in the reaction of a strong base with a weak acid.
- Understand the differences between the titration curves for a strong acid/strong base titration versus a weak acid/strong base titration in which one of the substances is weak.
- Describe how an indicator functions in an acid-base titration.
- Write the equilibrium constant expression- the solubility product constant, K_{sp} - for any insoluble salt.
- Calculate K_{sp} from experimental data.
- Estimate the solubility of a salt from the value of K_{sp} .
- Calculate the solubility of a salt in the presence of a common ion.
- Understand the effect of basic anions on the solubility of a salt.
- Decide if a precipitate will form when the ion concentrations are known.

- Calculate the ion concentrations that are required to begin the precipitation of an insoluble salt.
- Understand that the formation of a complex ion can increase the solubility of an insoluble salt.
- Use K values to devise a method of separating ions in solution from one another.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	19-1	Lecture, small group and self-directed study
Tuesday	19-2	Laboratory—Freezing Point Depression of Naphthalene.
Wednesday	19-3	Laboratory—Freezing Point Depression of Naphthalene.
Friday	19-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 20 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 19):

- Understand that entropy is a measure of matter dispersal or disorder.
- Recognize that entropy can be determined experimentally as the heat change of a reversal process.
- Know how to calculate entropy changes.
- Identify common processes that are entropy-favored.
- Use entropy and enthalpy changes to predict whether a reaction is spontaneous.
- Recognize how temperature influences whether a reaction is spontaneous.
- Understand the connection between enthalpy and entropy changes and the Gibbs free energy change for a process.
- Calculate the change in free energy at standard conditions for a reaction from the enthalpy and entropy changes or from the standard free energy of formation of reactants and products (ΔG°_f).
- Know how free energy changes with temperature.
- Describe the relationship between the free energy change, product favorability and equilibrium constants. Calculate K from $\Delta G^\circ_{\text{rxn}}$.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	20-1	Lecture, small group and self-directed study
Tuesday	20-2	Laboratory— K_{eq} Determination of FeSCN^{2+}
Wednesday	20-3	Laboratory— K_{eq} Determination of FeSCN^{2+}
Friday	20-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 21 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 20):

- Balance equations for oxidation-reduction reactions in acidic or basic solutions using the half-reaction approach.

- In a voltaic cell identify the half-reactions occurring at the anode and cathode, the polarity of the electrodes. The direction of electron flow in the external connection, and the direction of ion flow in the salt bridge.
- Know the chemistry and recognize the advantages and disadvantages of dry cells, alkaline batteries, mercury batteries, lead storage batteries, and Ni-Cad batteries.
- Understand how fuel cells work.
- Understand the process by which standard reduction potentials are determined and identify standard conditions as applied to electrochemistry.
- Describe the standard hydrogen electrode ($E^\circ = 0.00 \text{ V}$) and explain how it is used to determine standard potentials of half-reactions.
- Know how to use a Table of Standard Reduction Potentials to rank the strengths of oxidizing and reducing agents, to predict whether redox reactions are product-favored or reactant-favored.
- Use the Nernst equation to calculate the cell potential under nonstandard conditions.
- Explain how cell voltage relates to the concentration of ions and explain how this relationship allows the determination of pH and other ion concentrations.
- Describe the chemical processes occurring in an electrolysis. Recognize the factors that determine which substances are oxidized and reduced at the electrodes.
- Relate the amount of a substance oxidized or reduced to the amount of the current and the time of the current flows.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	21-1	Lecture, small group and self-directed study
Tuesday	21-2	Laboratory—Activity Series of Metals
Wednesday	21-3	Laboratory—Activity Series of Metals
Friday	21-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 22 Objectives:

Students will understand the following items conceptually and with relation to problem solving (Chapter 23):

- Identify α , β , and γ radiation, the three major types of radiation from radioactive elements in nature.
- Write balanced equations for nuclear reactions.
- Predict whether a radioactive isotope will decay by α or β emission, or by positron emission or electron capture.
- Understand the nature and origin of γ radiation.
- Calculate the binding energy and binding energy per nucleon for a particular isotope.
- Recognize the significance of a graph of binding energy per nucleon versus mass number and understand how it relates to nuclear stability.
- Understand and use mathematical equations and characterize rates of radioactive decay.
- Use the half-life to estimate the time required for an isotope to decay to a particular activity.

- Describe how artificial nuclear reactions are carried out.
- Describe nuclear-chain reactions, nuclear fission, and nuclear fusion.
- Describe the units used to measure intensity and understand how they pertain to health issues.
- Describe some uses of radioisotopes.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	22-1	Lecture, small group and self-directed study
Tuesday	22-2	Laboratory—Electrochemical Cells
Wednesday	22-3	Laboratory—Electrochemical Cells
Friday	22-5	Study (1 st 46 minutes) and Assessment (2 nd 46 minutes)

Week 23 Objectives:

Students will prepare for examination covering chapters 1-20, and 23.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	23-1	Lecture, small group and self-directed study
Tuesday	23-2	Lecture, small group and self-directed study
Wednesday	23-3	Exam (1999 AP Multiple Choice)
Friday	23-5	Exam (1999 AP Constructed Response)

Week 24 Objectives:

Students will prepare for qualitative analysis of cations practicum.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	24-1	Laboratory—Qualitative Analysis of Cations.
Tuesday	24-2	Laboratory—Qualitative Analysis of Cations.
Wednesday	24-3	Laboratory—Qualitative Analysis of Cations
Friday	24-5	Qualitative Analysis of Cations Practicum.

Week 25 Objectives:

Students will prepare for qualitative analysis of anions practicum.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	25-1	Laboratory—Qualitative Analysis of Anions.
Tuesday	25-2	Laboratory—Qualitative Analysis of Anions.
Wednesday	25-3	Laboratory—Qualitative Analysis of Anions
Friday	25-5	Qualitative Analysis of Anions Practicum.

Week 26 Objectives:

Students will prepare for the AP examination.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	26-1	Lecture, small group and self-directed study
Tuesday	26-2	Lecture, small group and self-directed study
Wednesday	26-3	Exam (2002 AP Multiple Choice)
Friday	26-5	Exam (2002 AP Constructed Response)

Week 27 Objectives:

Students will examine the processes of chromatography.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	27-1	Lecture, small group and self-directed study
Tuesday	27-2	Spinach Extract Chromatography.
Wednesday	27-3	Sep-Pak C18 Column Chromatography
Friday	27-5	Sep-Pak C18 Column Chromatography

Week 28

Students will prepare for the AP examination.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	28-1	Exam (REA1 AP Multiple Choice)
Tuesday	28-2	Exam (2006 AP Constructed Response)
Wednesday	28-3	Review Exam (REA1 AP Multiple Choice)
Friday	28-5	Review Exam (2006 AP Constructed Response)

Week 29

Students will prepare for the AP examination.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	29-1	Exam (REA1 AP Multiple Choice)
Tuesday	29-2	Exam (2006 AP Constructed Response)
Wednesday	29-3	Review Exam (REA1 AP Multiple Choice)
Friday	29-5	Review Exam (2006 AP Constructed Response)

Week 30

Students will prepare for and take the SAT Chemistry examination.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	30-1	Practice SAT Chemistry Subject Test 1
Tuesday	30-2	Review Exam (Practice SAT Chemistry Subject Test 2)
Wednesday	30-3	Practice SAT Chemistry Subject Test 1
Friday	30-5	Review Exam (Practice SAT Chemistry Subject Test 2)

Saturday

Take SAT Chemistry Exam

Week 31

Students will prepare for the AP examination.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	31-1	Exam (FT5 #2 AP Multiple Choice)
Tuesday	31-2	Exam (FT5 #2 AP Constructed Response)
Wednesday	31-3	Review Exam (FT5 #2 AP Multiple Choice)
Friday	31-5	Review Exam (FT5 #2 AP Constructed Response)

Week 32

Students will prepare for the AP examination.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	32-1	Exam (FT5 #1 AP Multiple Choice)
Tuesday	32-2	Exam (FT5 #1 AP Constructed Response)
Wednesday	32-3	Review Exam (FT5 #1 AP Multiple Choice)
Friday	32-5	Review Exam (FT5 #1 AP Constructed Response)

Week 33

Students will prepare for and take the AP examination.

Weekly Schedule:

<u>Day</u>	<u>Designation</u>	<u>Activities</u>
Monday	32-1	Lab Review
Tuesday	32-2	Take AP Exam