AP Chemistry

Textbook: Chemistry: A Molecular Approach 4th ed. Nivaldo Tro.

Course:
Chemistry 130 is a study of fundamental chemistry principles, including atomic structure, chemical bonding, kinetic theory, chemical kinetics, thermodynamics, solutions, electrochemistry, nuclear chemistry and equilibrium. Recommended for pre-professional, engineering and related science and medicine majors. I hope to create a similar college classroom experience to prepare the students for the next steps they will take in college. Students will learn through teacher lead discussions as well as laboratory inquiry experiences throughout the year. The course is structured around the enduring understandings within the six big ideas described in the AP Chemistry Curriculum Framework.

Attendance:
The students are expected to be a class on time ready to learn everyday. As a part of this course there will be a one day per week zero hour lab section starting at 7:00 am. Students are expected to be on time and if there is a prelab exercise it is due upon arrival. If it is not completed the student will not be able to participate in the laboratory activity.

Grading:
The grades will be weighted in this class.

- Homework/Quizzes 15%
- Exams (~ 3 exams/9 weeks) 55%
- Labs 30% (25% of the instructional time will be devoted to laboratory experience)

Students will be performing labs and lab reports in line with expectations of an introductory college level chemistry course and allow students to apply the seven science practices of the AP Framework. They will focus on developing good laboratory technique as well as an inquiry and guided inquiry approach to performing in lab. Students will also gain exposure to the use of technology in the laboratory through the use of Vernier probes and software LoggerPro.

Some labs will be completed on the forms provided, and others will require a formal lab report. The expectations for each lab will be discussed in the prelaboratory discussion.

Laboratory Reports

1. Title
2. Date
3. Purpose
4. Procedure

   A one or two sentence description of the method you are using. Do not include lengthy, detailed directions. A person who understands chemistry should be able read this section and know what you are doing.
5. Data

   Record all your data directly in your lab notebook. Organize data in neat orderly form. Label and use significant figures.
6. Calculations and Graphs

   Give the equation used and show how your values can be applied to the equation. Label everything on the graph.
7. Conclusion

   Make a simple statement concerning what you can conclude from the experiment.
8. Discussion of Theory

   What was demonstrated in this experiment? What do the calculations show? How was the purpose of the experiment fulfilled? Refer back to the purpose of the lab to write this section.
9. Experimental Sources of Error

   What are some specific sources of error, and how do they influence the data? Do they make the values obtained larger or smaller than they should be? Which measurement was the least precise? Instrumental error and human error exist in all experiments, and should not be mentioned as a source of error unless they cause a significant fault. Significant digits and mistakes in calculations are not a valid source of error. If you can calculate a percent error, do so and include in this section.
10. Questions

    Answer any questions included in the lab directions. Answer in such a way that the meaning of the question is obvious from your answer.

Science Practice 1: The student can use representations and models to communicate scientific phenomena and solve scientific problems.
Science Practice 2: The student can use mathematics appropriately.
Science Practice 3: The student can engage in scientific questioning to extend thinking or to guide investigations within the context of the AP course.
Science Practice 4: The student can plan and implement data collection strategies in relation to a particular scientific question.
Science Practice 5: The student can perform data analysis and evaluation of evidence.
Science Practice 6: The student can work with scientific explanations and theories.

Science Practice 7: The student is able to connect and relate knowledge across various scales, concepts and representations in and across domains.

Semester 1

Chapter 1 Matter Measurement and Problem Solving

Atomic Theory and Atomic Structure

1. Modern Chemistry: A Brief Glimpse
   - a. Provide examples of the contributions of chemistry to humanity.
2. Experiment and Explanation
   - a. Describe how chemistry is an experimental science.
   - b. Understand how the scientific method is an approach to performing science.
3. Law of Conservation of Mass
   - a. Explain the law of conservation of mass.
   - b. Apply the law of conservation of mass.
4. Matter: Physical State and Chemical Constitution
   - a. Compare and contrast the three common states of matter: solid, liquid, and gas.
   - b. Describe the classifications of matter: elements, compounds, and mixtures heterogeneous and homogeneous).
   - c. Understand the difference between chemical changes (chemical reactions) and physical changes.
   - d. Distinguish between chemical properties and physical properties.

Physical Measurements

5. Measurement and Significant Figures
   - a. Define and use the terms precision and accuracy when describing measured quantities.
   - b. Learn the rules for determining significant figures in reported measurements.
   - c. Know how to represent numbers using scientific notation.
   - d. Apply the rules of significant figures to reporting calculated values.
   - e. Be able to recognize exact numbers.
   - f. Know when and how to apply the rules for rounding.
   - g. Use significant figures in calculations.
6. SI Units
   - a. Become familiar with the SI (metric) system of units.
   - b. Convert from one temperature scale to another.
7. Derived Units
   - a. Define and provide examples of derived units.
   - b. Calculate the density of a substance.
8. Units and Dimensional Analysis (Factor-Label Method)
   - a. Apply dimensional analysis to solving numerical problems.
   - b. Convert from one metric unit to another metric unit.
   - c. Convert from one metric volume to another metric volume.
   - d. Convert from any unit to another unit.

Chapter 2 Learning Objectives

Atomic Theory and Atomic Structure

1. Atomic Theory of Matter
2. The Structure of the Atom
    a. List the postulates of the atomic theory.
    b. Define element, compound, and chemical reaction in the context of these postulates.
    c. Recognize the atomic symbols of the elements.
    d. Explain the significance of the law of multiple proportions.
    a. Describe Thomson's experiment in which he discovered the electron.
3. Nuclear Structure; Isotopes
4. Atomic Masses
5. Periodic Table of the Elements

    Chemical Substances: Formulas and Names

6. Chemical Formulas; Molecular and Ionic Substances
7. Organic Compounds
8. Naming Simple Compounds
    b. Describe Rutherford's experiment that led to the nuclear model of the atom.
    a. Name and describe the nuclear particles making up the nucleus of the atom.
    b. Define atomic number, mass number, and nuclide.
    c. Write the nuclide symbol for a given nucleus.
    d. Define and provide examples of isotopes of an element.
    e. Write the nuclide symbol of an element.
    a. Define atomic mass unit and atomic mass.

9. Describe how a mass spectrometer can be used to determine the fractional abundance of the isotopes of an element.
10. Determine the atomic mass of an element from the isotopic masses and fractional abundance.
11. Periodic Table
    a. Identify periods and groups on the periodic table.
    b. Find the main-group and transition elements on the periodic table.
    c. Locate the alkali metal and halogen groups on the periodic table.
    d. Recognize the portions of the periodic table that contain the metals, nonmetals, and metalloids(semimetals).
12. Chemical Representations
    a. Determine when the chemical formula of a compound represents a molecule.
    b. Determine whether a chemical formula is also a molecular formula.
    c. Define ion, cation, and anion.
    d. Classify a compound as ionic or molecular.

e. Define and provide examples for the term formula unit.
f. Specify the charge on all substances, whether ionic and molecular.
g. Write an ionic formula, given the ions.

13. Organic Molecules
    a. List the attributes of molecular substances that make them organic compounds.
    b. Explain what makes a molecule a hydrocarbon.
    c. Recognize some functional groups of organic molecules.
14. Ionic Compounds
    a. Recognize ionic compounds.
    b. Learn the rules for predicting the charges of monatomic ions in ionic compounds.
    c. Apply the rules for naming monatomic ions.
    d. Learn the names and charges of common polyatomic ions.
    e. Name an ionic compound from its formula.
    f. Write the formula of binary compound from its name.
    g. Determine the order of elements in a binary (molecular) compound.
    h. Learn the rules for naming binary molecular compounds.
    i. Name a binary compound from its formula.
    j. Write the formula of a binary compound from its name.
    k. Name a binary molecular compound from its molecular model.
    l. Recognize molecular compounds that are acids.
    m. Determine whether an acid is an oxoacid.
    n. Learn the approach for naming binary acids and oxoacids.
    o. Write the name and formula of an anion from the acid.
    p. Recognize compounds that are hydrates.
    q. Learn the rules for naming hydrates.
    r. Name a hydrate from its formula.
    s. Write the formula of a hydrate from its name.

Chemical Reactions: Equations

9. Writing Chemical Equations
10. Balancing Chemical Equations
    a. Identify reactants and products in a chemical equation.
    b. Write chemical equations using appropriate phase labels, symbols of reaction conditions, and the presence of a catalyst.
    c. Determine if a chemical reaction is balanced.
    d. Master the technique for balancing chemical equations.

Chapter 3 Objectives

Mass and Moles of Substances

1. Molecular Mass and Formula Mass
    a. Define the terms molecular and formula mass of a substance.
    b. Calculate the formula mass from a formula.

c. Calculate the formula mass from molecular models.

2. The Mole Concept
    a. Define the quantity called the mole.
    b. Learn Avogadro's number.
    c. Understand how the molar mass is related to the formula mass of a substance.
    d. Calculate the mass of atoms and molecules.
    e. Perform calculations using the mole.
    f. Convert from moles of substance to grams of substance.
    g. Convert from grams of substance to moles of substance.
    h. Calculate the number of molecules in a given mass of a substance.

Determining Chemical Formulas

3. Mass Percentages from the Formula
    a. Define mass percentage.
    b. Calculate the percentage composition of the elements in a compound.
    c. Calculate the mass of an element in a given mass of compound.
4. Elemental Analysis: Percentages of C, H, and O
    a. Describe how C, H, and O combustion analysis is performed.
    b. Calculate the percentage of C, H, and O from combustion data.
5. Determining Formulas
    a. Define empirical formula.
    b. Determine the empirical formula of a binary compound from the masses of its elements.
    c. Determine the empirical formula from the percentage composition.
    d. Understand the relationship between the molecular mass of a substance and its empirical formula mass.
    e. Determine the molecular formula from the percentage composition and molecular mass.

Stoichiometry: Quantitative Relations in Chemical Reactions

6. Molar Interpretation of a Chemical Equation
    a. Relate the coefficients in a balanced chemical equation to the number of molecules or moles (molar
7. Amounts of Substances in a Chemical Equation
    a. Use the coefficients in a balanced chemical equation to perform calculations.
    b. Relate the quantities of reactant to the quantity of product.
    c. Relation the quantities of two reactants or two products.
8. Limiting Reactant: Theoretical and Percentage Yieldinterpretation).
    a. Understand how a limiting reactant determines how many moles of product are formed during
    b. Calculate with a limiting reactant involving moles.
    c. Calculate with a limiting reactant involving masses.
    d. Define and calculate the theoretical yield of chemical reactions.

e. Determine the percentage yield of a chemical reaction.

Chapter 4 Learning Objectives

Ions in Aqueous Solution

1. Ionic Theory of Solutions and Solubility Rules
    a. Describe how an ionic substance can form ions in aqueous solution.
    b. Explain how an electrolyte makes a solution electrically conductive.
    c. Give examples of substances that are electrolytes.
    d. Define nonelectrolyte and provide an example of a molecular substance that is a nonelectrolyte.
    e. Compare the properties of solutions that contain strong electrolytes and weak electrolytes.
    f. Learn the solubility rules for ionic compounds.
    g. Use the solubility rules.
2. Molecular and Ionic Equations
    a. Write the molecular equation of a chemical reaction.
    b. From the molecular equations for both strong electrolytes and weak electrolytes, determine the
    c. From the complete ionic equation, write the net ionic equation.
    d. Write net ionic equations.

Types of Chemical Reactions

3. Precipitation Reactions
    a. Recognize precipitation (exchange) reactions.
    b. Write molecular, complete ionic, and net ionic equations for precipitation reactions.
    c. Decide whether a precipitation reaction will occur.
    d. Determine the product of a precipitation reaction.
4. Acid–Base Reactions
    a. Understand how an acid–base indicator is used to determine whether a solution is acidic or basic.
    b. Define Arrhenius acid and Arrhenius base.
    c. Write the chemical equation of an Arrhenius base in aqueous solution.
    d. Define Brønsted–Lowry acid and Brønsted–Lowry base.
    e. Write the chemical equation of a Brønsted–Lowry base in aqueous solution
    f. Write the chemical equation of an acid in aqueous solution using the hydronium ion.
    g. Learn the common strong acids and strong bases.
    h. Distinguish between a strong acid and a weak acid and the solutions they form.
    i. Distinguish between a strong base and a weak base and the solutions they form.
    j. Classify acids and bases as strong or weak.
    k. Recognize neutralization reactions.
    l. Write an equation for a neutralization reaction.
    m. Write the reaction equations for a polyprotic acid in aqueous solution

n. Recognize acid–base reactions that lead to gas formation.
o. Write an equation for a reaction with gas formation.

5. Oxidation–Reduction Reactions
    a. Define oxidation–reduction reaction.
    b. Learn the oxidation-number rules.
    c. Assign oxidation numbers.
    chemical reaction and how much excess reactant remains.
    complete ionic equation.
    d. Write the half-reactions of an oxidation–reduction reaction.
    e. Determine the species undergoing oxidation and reduction.
    f. Recognize combination reactions, decomposition reactions, displacement reactions, and combustion
    g. Use the activity series to predict when displacement reactions will occur.
6. Balancing Simple Oxidation–Reduction Equations
    a. Balance simple oxidation–reduction reactions by the half-reaction method

Working with Solutions

7. Molar Concentration
    a. Define molarity or molar concentration of a solution.
    b. Calculate the molarity from mass and volume.
    c. Use molarity as a conversion factor.
8. Diluting Solutions
    a. Describe what happens to the concentration of a solution when it is diluted.
    b. Perform calculations associated with dilution.
    c. Describe the process for diluting a solution.

Quantitative Analysis

9. Gravimetric Analysis
    a. Determine the amount of a species by gravimetric analysis.
10. Volumetric Analysis
    a. Calculate the volume of reactant solution needed to perform a titration.
    b. Understand how to perform a titration.
    c. Calculate the quantity of substance in a titrated solution.

Chapter 5-Learning Objectives

1. Gas Pressure and Its Measurement
    a. Define pressure and its units.
    b. Convert units of pressure.
    2. Empirical Gas Laws
    a. Express Boyle’s law in words and as an equation.
    b. Use Boyle’s law.
    c. Express Charles’s law in words and as an equation.

d. Use Charles’s law.
e. Express the combined gas law as an equation.
f. State Avogadro’s law.
g. Define standard temperature and pressure (STP).

3. The Ideal Gas Law
    a. State what makes a gas an ideal gas.
    b. Learn the ideal gas law equation.
    c. Derive the empirical gas laws from the ideal gas law.
    d. Use the ideal gas law.
    e. Calculate gas density.
    f. Determine the molecular mass of a vapor.
    g. Use an equation to calculate gas density.
4. Stoichiometry Problems Involving Gas Volumes
    a. Solving stoichiometry problems involving gas volumes.
5. Gas Mixtures; Law of Partial Pressures reactions.
    a. Learn the equation for Dalton’s law of partial pressures.
    b. Define the mole fraction of a gas.
    c. Calculate the partial pressure and the mole fraction of a gas in a mixture.
    d. Describe how gases are collected over water and how to determine the vapor pressure of water.
    e. Calculate the amount of gas collected over water.
6. Kinetic Theory of An Ideal Gas
    a. List the five postulates of the kinetic theory.

b. Provide a qualitative description of the gas laws based on the kinetic theory.

7. Molecular Speeds; Diffusion and Effusion
    a. Describe how the root-mean square (rms) molecular speed of gas molecules varies with temperature.
    b. Describe the molecular-speed distribution of gas molecules of different temperatures.
    c. Calculate the rms speed of a molecule.
    d. Define effusion and diffusion.
    e. Describe how individual gas molecules move while undergoing diffusion.
    f. Calculate the ratio of effusion rates of gases.
8. Real Gases
    a. Explain how and why a real gas is different from an ideal gas.

Chapter 6-Thermochemistry

Understanding Heats of Reaction

1. Energy and Its Units
    a. Define energy, kinetic energy, and internal energy.
    b. Define the SI unit of energy (joule) as well as the common unit of energy (calorie).
    c. Calculate the kinetic energy of a moving object.
    d. State the law of conservation of energy.

2. Heat of Reaction
    a. Define a thermodynamic system and its surroundings.
    b. Define heat and heat of reaction.
    c. Distinguish between an exothermic process and an endothermic process.
3. Enthalpy and Enthalpy Changes
    a. Define enthalpy and enthalpy of reaction.
    b. Explain how the terms enthalpy of reaction and heat of reaction are related.
    c. Explain how enthalpy and internal energy are related.
4. Thermochemical Equations
    a. Define a thermochemical equation.
    b. Write a thermochemical equation given pertinent information.
    c. Learn the two rules for manipulating (reversing and multiplying) thermochemical equations.
    d. Manipulate a thermochemical equation using these rules.
5. Applying Stoichiometry to Heats of Reaction
    a. Calculate the heat absorbed or evolved from a reaction given its enthalpy of reaction and the mass of a reactant or product.
6. Measuring Heats of Reaction
    a. Define heat capacity and specific heat.
    b. Relate the heat absorbed or evolved to the specific heat, mass, and temperature change.
    c. Perform calculations using the relationship between heat and specific heat.
    d. Define a calorimeter.
    e. Calculate the enthalpy of reaction from calorimetric data (temperature change and heat capacity).

Using Heats of Reaction
7. Hess's Law
    a. State Hess's law of heat summation.
    b. Apply Hess's law to obtain the enthalpy change for one reaction from the enthalpy changes of a number
8. Standard Enthalpies of Formation
    a. Define standard state and reference form.
    b. Define standard enthalpy of formation.
    c. Calculate the heat of a phase transition using standard enthalpies of formation for the different phases.
    d. Calculate the heat (enthalpy) of reaction from the standard enthalpies of formation of the substances in

Chapter 7-Quantum Theory of the Atom

Light Waves, Photons, and the Bohr Theory
1. The Wave Nature of Light

a. Define the wavelength and frequency of a wave.
b. Relate the wavelength, frequency, and speed of light.
c. Describe the different regions of the electromagnetic spectrum.

2. Quantum Effects and Photons
   a. State Planck's quantization of vibrational energy.
   b. Define Planck's constant and photon.
   c. Describe the photoelectric effect.
   d. Calculate the energy of a photon from its frequency or wavelength.
3. The Bohr Theory of the Hydrogen Atom
   a. State the postulates of Bohr's theory of the hydrogen atom.
   b. Relate the energy of a photon to the associated energy levels of an atom.
   c. Determine the wavelength or frequency of a hydrogen atom transition.
   d. Describe the difference between emission and absorption of light by an atom.

Quantum Mechanics and Quantum Numbers

4. Quantum Mechanics
   a. State the de Broglie relation.
   b. Calculate the wavelength of a moving particle.
   c. Define quantum mechanics.
   d. State Heisenberg's uncertainty principle.
   e. Relate the wave function for an electron to the probability of finding the electron at a location in space.
5. Quantum Numbers and Atomic Orbitals
   a. Define atomic orbital.
   b. Define each of the quantum numbers for an atomic orbital.
   c. State the rules for the allowed values for each quantum number.
   d. Apply the rules for quantum numbers.
   e. Describe the shapes of s, p, and d orbitals.

Chapter 8-Electron Configurations and Periodicity

Electronic Structure of Atoms

1. Electron Spin and the Pauli Exclusion Principle
   a. Define electron configuration and orbital diagram.
   b. State the Pauli exclusion principle.
   c. Apply the Pauli exclusion principle.
2. Building-Up Principle and the Periodic Table
   a. Define building-up principle.
   b. Define noble-gas core, pseudo-noble-gas core, and valence electron.
   c. Define main-group element and (d-block and f-block) transition element.
3. Writing Electron Configurations Using the Periodic Table
   a. Determine the configuration of an atom using the building-up principle.

b. Determine the configuration of an atom using the period and group number.

4. Orbital Diagrams of Atoms; Hund's Rule
   a. State Hund's rule.
   b. Apply Hund's rule.
   c. Define paramagnetic substance and diamagnetic substance.

Periodicity of the Elements

5. Mendeleev's Predictions from the Periodic Table
   a. State the periodic law.
   b. State the general periodic trends in size of atomic radii.
   c. Define effective nuclear charge.
   d. Determine relative atomic sizes from periodic trends.
   e. State the general periodic trends in ionization energy.
   f. Define first ionization energy.
   g. Determine relative ionization energies from periodic trends.
6. Some Periodic Properties
   a. Describe how Mendeleev predicted the properties of undiscovered elements.
7. Periodicity in the Main-Group Elements
   a. Define electron affinity.
   b. State the broad general trend in electron affinity across any period.
   a. Define basic oxide, acidic oxide, and amphoteric oxide.
   b. State the main group corresponding to an alkali metal, an alkaline earth metal, a chalcogen, a halogen, and a noble gas.
   c. Describe the change in metallic/nonmetallic character (or reactivities) in going through any main group of elements.

Chapter 9-Ionic and Covalent Bonding

Ionic Bonds

1. Describing Ionic Bonds
   a. Define ionic bond.
   b. Explain the Lewis electron-dot symbol of an atom.
   c. Describe the energetics of ionic bonding.
   d. Define lattice energy.
   e. Describe the Born–Haber cycle to obtain a lattice energy from thermodynamic data.
   f. Describe some general properties of ionic substances.
2. Electron Configurations of Ions
   a. State the three categories of monatomic ions of main-group elements.
   b. Write the electron configuration and Lewis symbol for a main-group ion.
   c. Note the polyatomic ions given earlier in Table 9.2.
   d. Note the formation of +2 and +3 transition-metal ions.
   e. Write the electron configurations of transition-metal ions
3. Ionic Radii

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a. Define ionic radius.
b. Define isoelectronic ions.
c. Use periodic trends to obtain relative ionic radii.

Covalent Bonds

4. Describing Covalent Bonds
    a. Describe the formation of a covalent bond between two atoms.
    b. Define Lewis electron-dot formula.
    c. Define bonding pair and lone (nonbonding) pair of electrons.
    d. Define coordinate covalent bond.
    e. State the octet rule.
    f. Define single, double, and triple bond.
5. Polar Covalent Bonds; Electronegativity
    a. Define polar covalent bond.
    b. Define electronegativity.
    c. State the general periodic trends in electronegativity.
    d. Use electronegativity to obtain relative bond polarity.
6. Writing Lewis Electron-Dot Formulas
    a. Write Lewis formulas having single bonds.
    b. Write Lewis formulas having multiple bonds.
    c. Write Lewis formulas for ionic species
7. Delocalized Bonding: Resonance
    a. Define localized bonding.
    b. Define resonance description.
    c. Write resonance forms.
8. Exceptions to the Octet Rule
    a. Write Lewis formulas (exceptions to the octet rule).
    b. Note exceptions to the octet rule in Group IIA and Group IIIA.
9. Formal Charge and Lewis Formulas
    a. Define formal charge.
    b. State the rules for obtaining the formal charge.
    c. State two rules useful in writing Lewis formulas.
    d. Use formal charges to determine the best Lewis formula.
10. Bond Length and Bond Order
    a. Define bond length (bond distance).
    b. Define covalent radii.
    c. Define bond order.
    d. Explain how bond order and bond length are related.
    e. Bond Energy
    f. Define bond energy.

g. Estimate ΔH from bond energies.

Chapter 10-Molecular Geometry and Directional Bonding

1. The Valence-Shell Electron-Pair Repulsion (VSEPR) Model
    a. Define molecular geometry.
    b. Define valence-shell electron-pair repulsion model.
    c. Note the difference between the arrangement of electron pairs about a central atom and molecular geometry.
    d. Note the four steps in the prediction of geometry by the VSEPR model.
    e. Predict the molecular geometry for a molecule with two, three, or four electron pairs.
    f. Note that a lone pair tends to require more space than a corresponding bonding pair and that a multiple bond requires more space than a single bond.
    g. Predict the molecular geometry for a molecule with five or six electron pairs.
2. Dipole Moment and Molecular Geometry
    a. Define dipole moment.
    b. Explain the relationship between the dipole moment and molecular geometry.
    c. Note that the polarity of a molecule can affect certain properties, such as boiling point.
3. Valence Bond Theory
    a. Define valence bond theory.
    b. State the two conditions needed for bond formation according to valence bond theory.
    c. Define hybrid orbitals.
    d. State the five steps in describing bonding, following the valence bond theory.
    e. Apply valence bond theory to a molecule with two, three, or four electron pairs.
    f. Apply valence bond theory to a molecule with five or six electron pairs.
4. Description of Multiple Bonding
    a. Define an s (sigma) bond.
    b. Define a p (pi) bond.
    c. Apply valence bond theory (multiple bonding).
    d. Explain geometric, or cis-trans, isomers in terms of the p-bond description of a double bond.
5. Molecular Orbitals and Delocalized Bonding
    a. Describe the delocalized bonding in molecules such as O3.

Chapter 11-States of Matter: Liquid and Solid

1. Comparison of Gases, Liquids, and Solids
    a. Recall the definitions of gas, liquid, and solid given in Section 1.4.
    b. Compare a gas, a liquid, and a solid using a kinetic molecular theory description.
    c. Recall the ideal gas law and the van der Waals equation for gases (there are no similar simple equations for liquids and solids).

Changes of State

2. Phase Transitions
    a. Define change of state (phase transition)
    b. Define melting, freezing, vaporization, sublimation, and condensation.
    c. Define vapor pressure.
    d. Describe the process of reaching a dynamic equilibrium that involves the vaporization of a liquid and condensation of its vapor.
    e. Describe the process of boiling.
    f. Define freezing point and melting point.
    g. Define heat (enthalpy) of fusion and heat (enthalpy) of vaporization.
    h. Calculate the heat required for a phase change of a given mass of substance.
    i. Describe the general dependence of the vapor pressure (ln P) on the temperature (T).
    j. State the Clausius–Clapeyron equation (the two-point form).
    k. Calculate the vapor pressure at a given temperature.
    l. Calculate the heat of vaporization from vapor pressure.

3. Phase Diagrams
    a. Define phase diagram.
    b. Describe the melting-point curve and the vapor-pressure curves (for the liquid and the solid) in a phase diagram.
    c. Define triple point.
    d. Define critical temperature and critical pressure.
    e. Relate the conditions for the liquefaction of a gas to its critical temperature.

Liquid State

4. Properties of Liquids; Surface Tension and Viscosity
    a. Define intermolecular forces.
    b. Define dipole–dipole force.
    c. Describe the alignment of polar molecules in a substance.
    d. Define London (dispersion) forces.
    e. Note that London forces tend to increase with molecular mass.

5. Intermolecular Forces; Explaining Liquid Properties
    a. Define surface tension.
    b. Describe the phenomenon of capillary rise.
    c. Define viscosity.

5. Intermolecular Forces; Explaining Liquid Properties (con’t)
    a. Relate the properties of liquids to the intermolecular forces involved.
    b. Define hydrogen bonding.
    c. Identify the intermolecular forces in a substance.
    d. Determine relative vapor pressures on the basis of intermolecular attractions.

Solid State

6. Classification of Solids by Type of Attraction of Units
    a. Define molecular solid, metallic solid, ionic solid, and covalent network solid.
    b. Identify types of solids.
    c. Relate the melting point of a solid to its structure.
    d. Determine relative melting points based on types of solids.
    e. Relate the hardness and electrical conductivity of a solid to its structure.
7. Crystalline Solids; Crystal Lattices and Unit Cells
    a. Define crystalline solid and amorphous solid.
    b. Define crystal lattice and unit cell of a crystal lattice.
    c. Define simple cubic unit cell, body-centered cubic unit cell, and face-centered cubic unit cell.
    d. Determine the number of atoms in a unit cell.
    e. Describe the two kinds of crystal defects.
10. Determining Crystal Structure by X-Ray Diffraction
    a. Describe how constructive and destructive interference give rise to a diffraction pattern.
    b. Note that diffraction of x rays from a crystal gives information about the positions of atoms in the crystal.

Chapter 12-Solutions

Solution Formation

1. Types of Solutions
    a. Define solute and solvent.
    b. Define miscible fluid.
    c. Provide examples of gaseous solutions, liquid solutions, and solid solutions.
2. Solubility and the Solution Process
    a. List the conditions that must be present to have a saturated solution, to have an unsaturated solution, and to have a supersaturated solution.
    b. Describe the factors that make one substance soluble in another.
    c. Determine when a molecular solution will form when substances are mixed.
    d. Learn which conditions must be met to create an ionic solution.
3. Effects of Temperature and Pressure on Solubility
    a. State the general trends for the solubility of gases and solids with temperature.
    b. Explain how the solubility of a gas changes with temperature.
    c. Apply Henry’s law.

Colligative Properties

4. Ways of Expressing Concentration
    a. Define colligative property.

b. Define molarity.
c. Define mass percentage of solute.
d. Perform calculations with the mass percentage of a solute.
e. Define molality.
f. Calculate the molality of a solute.
g. Define mole fraction.
h. Calculate the mole fractions of components.
i. Convert molality to mole fraction.
j. Convert mole fraction to molality.
k. Convert molality to molarity.
l. Convert molarity to molality

5. Vapor Pressure of a Solution
    a. Explain vapor-pressure lowering of a solvent.
    b. State Raoult's law.
    c. Calculate vapor-pressure lowering.
    d. Describe an ideal solution.
6. Boiling-Point Elevation and Freezing-Point Depression
    a. Define boiling-point elevation and freezing-point depression.
    b. Calculate boiling-point elevation and freezing-point depression.
    c. Calculate the molecular mass of a solute from molality.
    d. Calculate the molecular mass from freezing-point depression.
7. Osmosis
    a. Describe a system where osmosis will take place.
    b. Calculate osmotic pressure.
8. Colligative Properties of Ionic Solutions
    a. Determine the colligative properties of ionic solutions.

Colloid Formation

9. Colloids
    a. Define colloid.
    b. Explain the Tyndall effect.
    c. Give examples of hydrophilic colloids and hydrophobic colloids.
    e. Describe coagulation.
    f. Explain how micelles can form an association colloid.

Semester 2

Chapter 13-Rates of Reaction

1.Definition of a Reaction Rate
    a. Define reaction rate.
    b. Explain instantaneous rate and average rate of a reaction.

c. Explain how the different ways of expressing reaction rates are related.
d. Calculate average reaction rate.

2. Experimental Determination of Rate
    a. Describe how reaction rates may be experimentally determined.
3. Dependence of Rate on Concentration
    a. Define and provide examples of a rate law, rate constant, and reaction order.
    b. Determine the order of reaction from the rate law.
    c. Determine the rate law from initial rates.
4. Change of Concentration with Time
    a. Learn the integrated rate laws for first-order, second-order, and zero-order reactions.
    b. Use an integrated rate law.
    c. Define half-life of a reaction.
    d. Learn the half-life equations for first-order, second-order, and zero-order reactions.
    e. Relate the half-life of a reaction to the rate constant.
    f. Plot kinetic data to determine the order of a reaction.
5. Temperature and Rate; Collision and Transition-State Theories
    a. State the postulates of collision theory.
    b. Explain activation energy (Ea).
    c. Describe how temperature, activation energy, and molecular orientation influence reaction rates.
    d. State the transition-state theory.
    e. Define activated complex.
    f. Describe and interpret potential-energy curves for endothermic and exothermic reactions.
6. Arrhenius Equation
    a. Use the Arrhenius equation.

Reaction Mechanisms

7. Elementary Reactions
    a. Define elementary reaction, reaction mechanism, and reaction intermediate. Determine the rate law from initial rates.
    b. Write the overall chemical equation from a mechanism.
    c. Define molecularity.
    d. Give examples of unimolecular, bimolecular, and termolecular reactions.
    e. Determine the molecularity of an elementary reaction.
    f. Write the rate equation for an elementary reaction.
8. The Rate Law and the Mechanism
    a. Explain the rate-determining step of a mechanism.
    b. Determine the rate law from a mechanism with an initial slow step.
    c. Determine the rate law from a mechanism with an initial fast, equilibrium step.
9. Catalysis
    a. Describe how a catalyst influences the rate of a reaction.
    b. Indicate how a catalyst changes the potential-energy curve of a reaction.

c. Define homogeneous catalysis and heterogeneous catalysis.
d. Explain enzyme catalysis.

Chapter 14-Chemical Equilibrium

Describing Chemical Equilibrium

1. Chemical Equilibrium—A Dynamic Equilibrium
    a. Define dynamic equilibrium and chemical equilibrium.
    b. Apply stoichiometry to an equilibrium mixture.
2. The Equilibrium Constant
    a. Define equilibrium-constant expression and equilibrium constant.
    b. State the law of mass action.
    c. Write equilibrium-constant expressions.
    d. Describe the kinetics argument for the approach to chemical equilibrium.
    e. Obtain an equilibrium constant from reaction composition.
2. The Equilibrium Constant (cont)
    f. Describe the equilibrium constant Kp; indicate how Kp and Kc are related.State the law of mass action.
    g. Obtain Kc for a reaction that can be written as a sum of other reactions of known Kc values.
3. Heterogeneous Equilibria; Solvents in Homogeneous Equilibria
    a. Define homogeneous equilibrium and heterogeneous equilibrium.
    b. Write Kc for a reaction with pure solids or liquids.

Using the Equilibrium Constant
4. Qualitatively Interpreting the Equilibrium Constant
    a. Give a qualitative interpretation of the equilibrium constant based on its value.
5. Predicting the Direction of Reaction
    a. Define reaction quotient, Q.
    b. Describe the direction of reaction after comparing Q with Kc.
    c. Use the reaction quotient.
6. Calculating Equilibrium Concentrations
    a. Obtain one equilibrium concentration given the others.
    b. Solve an equilibrium problem (involving a linear equation in x).
    c. Solve an equilibrium problem (involving a quadratic equation in x).

Le Châtelier’s Principle
7. Removing Products or Adding Reactants
    a. State Le Châtelier’s principle.
    b. State what happens to an equilibrium when a reactant or product is added or removed.
    c. Apply Le Châtelier’s principle when a concentration is altered.

8. Changing the Pressure and Temperature
    a. Describe the effect of a pressure change on chemical equilibrium.
    b. Apply Le Châtelier’s principle when the pressure is altered.
    c. Describe the effect of a temperature change on chemical equilibrium.
    d. Apply Le Châtelier’s principle when the temperature is altered.
    e. Describe how the optimum conditions for a reaction are chosen.
9. Effect of a Catalyst
    a. Define catalyst.
    b. Compare the effect of a catalyst on rate of reaction with its effect on equilibrium.
    c. Describe how a catalyst can affect the product formed.

Chapter 15 Learning Objectives

Acid Base Concepts

1. Arrhenius Concept of Acids and Base
    a. Define acid and base according to the Arrhenius concept.
2. Brønsted–Lowry Concept of Acids and Bases
    a. Define acid and base according to the Brønsted–Lowry concept.
    b. Define the term conjugate acid–base pair.
    c. Identify acid and base species.
    d. Define amphiprotic species.
3. Lewis Concept of Acids and Bases
    a. Define Lewis acid and Lewis base.
    b. Identify Lewis acid and Lewis base species.

Acid and Base Strengths

4. Relative Strengths of Acids and Bases
    a. Understand the relationship between the strength of an acid and that of its conjugate base.
    b. Decide whether reactants or products are favored in an acid–base reaction
5. Molecular Structure and Acid Strength
    a. Note the two factors that determine relative acid strengths.
    b. Understand the periodic trends in the strengths of the binary acids HX.
    c. Understand the rules for determining the relative strengths of oxoacids.
    d. Understand the relative acid strengths of a polyprotic acid and its anions.

Self-Ionization of Water and pH

6. Self-Ionization of Water
    a. Define self-ionization (or autoionization).
    b. Define the ion-product constant for water.
7. Solutions of a Strong Acid or Base

a. Calculate the concentrations of H3O

8. The pH of a Solution
    a. Define pH.
    b. Calculate the pH from the hydronium-ion concentration.
    c. Calculate the hydronium-ion concentration from the pH.
    d. Describe the determination of pH by a pH meter and by acid–base indicators.

Chapter 16- Learning Objectives

Solutions of a Weak Acid or Base

1. Acid-Ionization Equilibria
    a. Write the chemical equation for a weak acid undergoing acid ionization in aqueous solution.
    b. Define acid-ionization constant and degree of ionization.
    \+ and OHin solutions of a strong acid or base
    c. Determine Ka from the solution pH.
    d. Calculate concentrations of species in a weak acid solution using Ka (approximation method).
1. Acid-Ionization Equilibria (cont)
    a. State the assumption that allows for using approximations when solving problems.
    b. Calculate concentrations of species in a weak acid solution using Ka (quadratic formula).
2. Polyprotic Acids
    a. State the general trend in the ionization constants of a polyprotic acid.
    b. Calculate concentrations of species in a solution of a diprotic acid.
3. Base-Ionization Equilibria
    a. Write the chemical equation for a weak base undergoing ionization in aqueous solution.
    b. Define base-ionization constant.
    c. Calculate concentrations of species in a weak base solution using Kb.
4. Acid–Base Properties of Salt Solutions
    a. Write the hydrolysis reaction of an ion to form an acidic solution.
    b. Write the hydrolysis reaction of an ion to form a basic solution.
    c. Predict whether a salt solution is acidic, basic, or neutral.
    d. Obtain Ka from Kb or Kb from Ka.
    e. Calculating concentrations of species in a salt solution.

Solutions of a Weak Acid or Base with Another Solute

5. Common-Ion Effect
    a. Explain the common-ion effect.
    b. Calculate the common-ion effect on acid ionization (effect of a strong acid).
    c. Calculate the common-ion effect on acid ionization (effect of a conjugate base).
6. Buffers

    a. Define buffer and buffer capacity.
    b. Describe the pH change of a buffer solution with the addition of acid or base.
    c. Calculate the pH of a buffer from given volumes of solution.
    d. Calculate the pH of a buffer when a strong acid or a strong base is added.
    e. Learn the Henderson–Hasselbalch equation.
    f. State when the Henderson–Hasselbalch equation can be applied

7. Acid–Base Titration Curves
    a. Define equivalence point.
    b. Describe the curve for the titration of a strong acid by a strong base.
    c. Calculate the pH of a solution of a strong acid and a strong base.
    d. Describe the curve for the titration of a weak acid by a strong base.
    e. Calculate the pH at the equivalence point in the titration of a weak acid by a strong base.

7. Acid–Base Titration Curves (cont)
    f. Describe the curve for the titration of a weak base by a strong acid.
    g. Calculate the pH of a solution at several points of a titration of a weak base by a strong acid.

Chapter 18-Thermodynamics and Equilibrium

1. First Law of Thermodynamics; Enthalpy
    a. Define internal energy, state function, work, and first law of thermodynamics.
    b. Explain why the work done by the system as a result of expansion or contraction during a chemical reaction is -

1. First Law of Thermodynamics; Enthalpy (cont)
    c. Relate the change of internal energy, ΔU, and heat of reaction, q.
    d. Define enthalpy, H.
    e. Show how heat of reaction at constant pressure, qp, equals the change of enthalpy, ΔH.

Spontaneous Processes and Entropy

2. Entropy and the Second Law of Thermodynamics
    a. Define spontaneous process.
    b. Define entropy.
    PΔV.
    c. Relate entropy to disorder in a molecular system (energy dispersal).
    d. State the second law of thermodynamics in terms of system plus surroundings.

2. Entropy and the Second Law of Thermodynamics (cont)
    e. State the second law of thermodynamics in terms of the system only.
    f. Calculate the entropy change for a phase transition.
    g. Describe how ΔH - TΔS functions as a criterion of a spontaneous reaction.

3. Standard Entropies and the Third Law of Thermodynamics

a. State the third law of thermodynamics.
b. Define standard entropy (absolute entropy).
c. State the situations in which the entropy usually increases.
d. Predict the sign of the entropy change of a reaction.
e. Express the standard change of entropy of a reaction in terms of standard entropies of products and reactants.
f. Calculate ΔS

Free-Energy Concept

4. Free Energy and Spontaneity
    a. Define free energy, G.
    b. Define the standard free-energy change.
    c. Calculate ΔG
    d. Define the standard free energy of formation, ΔG
    e. Calculate ΔG
    f. State the rules for using ΔG
    g. Interpret the sign of ΔG
5. Interpretation of Free Energy
    a. Relate the free-energy change to maximum useful work.
    b. Describe how the free energy changes during a chemical reaction.

Free Energy and Equilibrium Constants

6. Relating ΔG
    a. Define the thermodynamic equilibrium constant, K.
    b. Write the expression for a thermodynamic equilibrium constant.
    c. Indicate how the free-energy change of a reaction and the reaction quotient are related.
    d. Relate the standard free-energy change to the thermodynamic equilibrium constant.
    e. Calculate K from the standard free-energy change (molecular equation).
    f. Calculate K from the standard free-energy change (net ionic equation).
7. Change of Free Energy with Temperature
    a. Describe how ΔG is a criterion for spontaneity
    b. Describe how the spontaneity or nonspontaneity of a reaction is related to each of the four possible combinations
    c. Calculate ΔG from standard free energies of formation.

Chapter 19-Electrochemsitry

1. Balancing Oxidation–Reduction Reactions in Acidic and Basic Solutions for a reaction.
    a. Learn the steps for balancing oxidation–reduction reactions using the half-reaction method.
    b. Balance equations by the half-reaction method (acidic solutions).

  c. Learn the additional steps for balancing oxidation–reduction reactions in basic solution using the half-reaction method.
  d. Balance equations using the half-reaction method (basic solution).
2. Construction of Voltaic Cells
  a. Define electrochemical cell, voltaic (galvanic cell), electrolytic cell, and half-cell.
  b. Describe the function of the salt bridge in a voltaic cell.
  c. State the reaction that occurs at the anode and the cathode in an electrochemical cell.
  d. Define cell reaction.
  e. Sketch and label a voltaic cell.
3. Notation for Voltaic Cells
  a. Write the cell reaction from the cell notation.
4. Cell Potential
  a. Define cell potential and volt.
  b. Calculate the quantity of work from a given amount of cell reactant.
5. Standard Cell Potentials and Standard Electrode Potentials
  a. Explain how the electrode potential of a cell is an intensive property.
  b. Define standard cell potential and standard electrode potential.
  c. Interpret the table of standard reduction potentials.
  d. Determine the relative strengths of oxidizing and reducing agents.
  e. Determine the direction of spontaneity from electrode potentials.
  f. Calculate cell potentials from standard potentials.
6. Equilibrium Constants from Cell Potentials
  a. Calculate the free-energy change from electrode potentials.
  b. Calculate the cell potential from free-energy change.
  c. Calculate the equilibrium constant from cell potential.
7. Dependence of Cell Potential on Concentration
  a. Calculate the cell potential for nonstandard conditions.
  b. Describe how pH can be determined using a glass electrode.
8. Some Commercial Voltaic Cells
  a. Describe the construction and reactions of a zinc–carbon dry cell, a lithium–iodine battery, a lead storage cell, and a nickel–cadmium cell.
  b. Explain the operation of a proton-exchange membrane fuel cell.
  c. Explain the electrochemical process of the rusting of iron.
  d. Define cathodic protection.
9. Electrolysis of Molten Salts
  a. Define electrolysis.
10. Aqueous Electrolysis
  a. Learn the half-reactions for water undergoing oxidation and reduction.
  b. Predict the half-reactions in an aqueous electrolysis.
11. Stoichiometry of Electrolysis
  a. Calculate the amount of charge from the amount of product in an electrolysis.
  b. Calculate the amount of product from the amount of charge in an electrolysis.

Chapter 20-Nuclear Chemistry

1. Radioactivity
    a. Define radioactive decay and nuclear bombardment reaction.
    b. Learn the nuclear symbols for positron, gamma photon, electron, neutron, and proton.
    c. Write a nuclear equation
    d. Deduce a product or reactant in a nuclear equation.
    e. Describe the shell model of the nucleus.
    f. Explain the band of stability.
    g. Predict the relative stabilities of nuclides
    h. List the six types of radioactive decay.
    i. Predict the type of radioactive decay.
    j. Define radioactive decay series.
2. Nuclear Bombardment Reactions
    a. Define transmutation.
    b. Use the notation for a bombardment reaction.
    c. Locate the transuranium elements on the periodic table.
    d. Determine the product nucleus in a nuclear bombardment reaction.
3. Radiation and Matter: Detection and Biological Effects
    a. State the purpose of a Geiger counter and a scintillation counter.
    b. Define activity of a radioactive source and curie (Ci).
    c. State the relationship between a rad and a rem.
4. Rate of Radioactive Decay
    a. Define radioactive decay constant.
    b. Calculate the decay constant from activity.
    c. Define half-life.
    d. Draw a typical half-life decay curve of a radioactive element.
    e. Calculate the half-life from the decay constant.
    f. Calculate the decay constant and activity from the half-life.
    g. Determine the fraction of nuclei remaining after a specified time.
    h. Apply the carbon-14 dating method.
5. Applications of Radioactive Isotopes
    a. State the ways in which radioactive isotopes are used for chemical analysis.
    b. Describe how isotopes are used for medical therapy and diagnosis.
6. Mass–Energy Calculations
    a. Calculate the energy changes for a nuclear reaction.
    b. Define nuclear binding energy and mass defect.
    c. Compare and contrast nuclear fission and nuclear fusion.
7. Nuclear Fission and Nuclear Fusion
    a. Explain how a controlled chain reaction is applied in a nuclear fission reactor using a critical mass of fissionable material.
    b. Write the reaction of the nuclear fusion of deuterium and tritium.