Date Adopted 1965

Dates reviewed 1985

Dates revised 1988, 2001, 2008

Alabama

Department of

Postsecondary Education

Representing Alabama's Public Two-Year College System

Jefferson State Community College

CHM 104

Introduction to General Chemistry

I. CHM 104, Introduction to General Chemistry, 4 Semester Hours

Core Area III, ASCI TSCI (Lec 3 hrs, Lab 2 hrs) (***State guide has 3HR Labs)

II. Course Description

This course is a survey of general chemistry for students who do not intend to major in science or engineering and may not be substituted for CHM 111. Lecture will emphasize the fundamental facts, principles, and theories of general chemistry including math operations, matter and energy, atomic structure, symbols and formulas, nomenclature, the periodic table, bonding concepts, equations, reactions, stoichiometry, gas laws, phases of matter, solutions, pH, chemical equilibrium, and nuclear chemistry.

Laboratory is required.

III. Prerequisites

MTH 098, (***MTH 116?) or equivalent math placement score.

IV. Textbook

Foundations of College Chemistry, Hein and Arena, 12th ed. Wiley
Chemical Education Resources Lab Manual

V. Course Objectives

In the classroom the student will:

A. Understand and apply principles involved in measurement and problem solving.

B. Understand the nature and variety of forms of matter and list the physical properties that characterize each state.
C. Understand the structure of atoms and will apply the periodic laws to predict chemical and physical properties of the elements.
D. Comprehend the nature of compounds, their formation, composition, and nomenclature.
E. Comprehend chemical equations and utilize them in stoichiometric calculations.
F. Understand and apply the principles of gas behavior in ideal situations.
G. Understand the properties of aqueous solutions systems and the theories describing the behavior of acids and bases in aqueous systems.
H. Understand and apply the principles of chemical equilibrium.
I. Understand and apply the principles of radioactivity and nuclear chemistry.

In the laboratory the student will:

A. Develop an understanding of basic laboratory techniques and procedures.
B. Understand basic laboratory safety and will follow all laboratory rules during experimental work.
C. Acquire understanding of the physical and chemical properties of commonly used elements, compounds and mixtures.
D. Be able to make precise measurements and evaluate experimental data through selected qualitative laboratory experiments.
E. Be able to make careful observations, report and interpret experimental data through selected quantitative laboratory experiments.
F. Be able to perform simple calculations from experimental data through selected quantitative laboratory experiments.

VI. Course Outline of Topics

Lecture Topics Stated in Performance Terms

A. The student will understand and apply principles involved in measurement and problem solving.

1. Express any given number in exponential notation form.
2. Express the results of arithmetic operations, of assigned problems, to the proper number of significant figures.
3. Cite, from memory, the basic metric units of mass, length, and volume.
4. Give the numerical equivalent of the metric prefixes deci, centi, milli, micro, deka, hecto, kilo, and mega.
5. Convert any given measurement of mass, length, or volume in American units to metric units and vice versa.
6. Set up and solve assigned problems using the dimensional analysis or factor-label method.
7. Make conversions between Fahrenheit, Celsius, and Kelvin temperatures from assigned problems.
8. Make calculations using the equation $Q=mc\Delta T$, or [heat = (grams of

substance) x (specific heat of substance) x (change in temperature) Δ T].

9. Calculate density, mass, or volume of an object of substance from given data.
10. Calculate the specific gravity when given the density of a substance and vice versa.
11. Define, mass, weight, Metric system, SI, heat, temperature, calories/kilocalorie, Joule, specific heat, density, specific gravity and hydrometer.

B. The student will understand the nature and variety of forms of matter and list the physical properties that characterize each state.

1. Identify the three (3) physical states of matter and list the physical properties that characterize each state.
2. Distinguish between the physical and chemical properties of matter.
3. Classify the given changes undergone by matter as either physical or chemical.
4. Distinguish between a substance and mixture.
5. Distinguish between kinetic and potential energy.
6. State the Law of Conservation of Mass and the Law of Conservation of Energy.
7. Explain why the laws dealing with the conservation of mass and energy may be combined into a single more accurate general statement, the Law of Conservation of Mass and Energy.
8. Classify common materials as elements, compounds, or mixtures.
9. Write the symbols when given the names or write the names when given the symbols of the common elements listed in a given table.
10. State the Law of Definite Composition.
11. Interpret chemical formulas in terms of number of atoms of each element present.
12. Differentiate among atoms, molecules, and ions.
13. List the characteristics of metals and nonmetals.
14. List the seven elements that occur as diatomic molecules.
15. Define element, atom, compound, molecule, ion, metalloid, chemical formula, chemical equation, mixture, metal and nonmetal.

C. The student will understand the structure of atoms and will apply the periodic law to predict chemical and physical properties of the elements.

1. Define nucleus, orbital, atomic number, electron shell, Avogadro's number, noble gas, atomic mass unit, atomic weight, gram-atomic weight and mole.
2. State the major provisions of Dalton's Atomic Theory.
3. Give the names, symbols, charges, and relative masses of the three principal subatomic particles.
4. Describe the atom as conceived by Ernest Rutherford after his alpha particle scattering experiments.
5. Describe the atom as conceived by Niels Bohr.
6. Discuss the contributions to atomic theory made by Dalton, Thomson, Ruth-

erford, Bohr, Chadwick, and Schrodinger.

7. Determine the maximum number of electrons that can exist in a given main energy level;
8. Draw an s orbital and a p-orbital.
9. Give the electron configuration (1s, 2s, 2p ....) for any of the first 56 elements, or identify the element when given the electron configuration.
10. Draw the diagram of any isotope of the first 38 elements, showing the composition of the nucleus and the numbers of electrons in the main energy levels.
11. Give the electron dot structure for any representative element on the periodic table.
12. Name the three isotopes of hydrogen and give the number of protons, neutrons, and electrons in each.
13. List the number of protons, neutrons, and electrons for any element when given the atomic number and atomic weight.
14. Calculate the number of atoms, moles, or grams from appropriate data.
15. Define: periods of elements; groups or families of elements; and transition elements.
16. Describe briefly the contributions of Mendeleev, Meyer, and Moseley to the development of the periodic table.
17. State the periodic law in its modern form.
18. Indicate the location on a periodic table of the metals, the nonmetals, the metalloids, the noble gases, the alkali metals, the alkaline earth metals, the chalcogens and the halogens.
19. Indicate on the periodic table areas in which s, p, d, and f orbitals of electrons are being filled.
20. Describe how atomic radii vary (a) from left to right in a period and (b) from top to bottom in a group.
21. Describe the changes in outer-level electron structure when (a) moving from left to right in a period and (b) going from top to bottom in a group.
22. Predict the formulas of simple binary compounds using the periodic table.
23. Describe the electronic configuration of transition elements.

D. The student will comprehend the nature of compounds, their formation, composition, and nomenclature.

1. Define ionization energy, covalent bond, polar covalent bond, polyatomic ion, oxidation, reduction, electronegativity, ionic bond, non-polar covalent bond, coordinate covalent bond, oxidation number, valence electrons and chemical bond.
2. Describe the variation of the ionization energies of the elements with respect to position in the periodic table and with respect to removal of successive electrons.
3. Describe the formation of ions by electron transfer between two elements and the nature of the ionic bond formed.
4. Use the periodic table to predict the formulas of the monatomic ions.

5. Show pictorially with electron dot structures the formation of an ionic compound from atoms.
6. Describe the relative sizes of atoms compared to their ions.
7. Draw electron dot structures for common covalent compounds and polyatomic ions.
8. Explain why ionic bonding results in crystalline compounds while covalent bonding results in molecules.
9. Describe the change in electronegativity in moving across a period and in moving down a family on the periodic table.
10. Predict whether a covalent bond will be polar.
11. Predict whether molecules will be dipoles.
12. Identify which bonds are coordinate covalent in the dot structures of a compound.
13. Classify the bonding in a compound as primarily ionic or primarily covalent.
14. Draw the Lewis structures for simple polyatomic ions.
15. Give the names or formulas of the common ions.
16. Write formulas of compounds which are simple combinations of common ions.
17. Assign oxidation numbers to each element in a compound or ion.
18. Give the name or formula for inorganic binary compounds in which the metal has only one common oxidation state.
19. Give the name or formula for inorganic binary compounds containing metals of variable oxidation state, using either the Stock System or classical nomenclature.
20. Give the name or formula for the following:
    a. inorganic binary compounds that contain two nonmetals;
    b. binary acids;
    c. ternary inorganic acids;
    d. ternary salts;
    e. salts containing more than one kind of positive ion;
    f. inorganic bases.
21. Illustrate, with examples of how each of the following is used in naming inorganic compounds: -ide; -ous; -ic; hypo-; per-; and Roman numerals.
22. Define: formula weight; molecular weight; gram-formula weight; gram-molecular weight empirical formula; and molecular formula.
23. Determine the formula weight or molecular weight of a compound when given the formula.
24. Perform the following calculations:
    a. moles, gram-formula weights, gram-molecular weights, molecules, or grams when given appropriate data;
    b. the percentage composition by weight of a compound when given the formula.
25. Explain the relationship between an empirical formula and a molecular formula.
26. Determine the empirical and/or molecular formula of a compound when given the appropriate information.

E. The student will comprehend chemical equations and utilize them in stoichiometric calculations.

1. Define chemical equation, reactant, balanced equation, decomposition reaction, exothermic reaction, endothermic reaction, word equation, product, combination reaction, single replacement reaction, double replacement or metathesis reaction and heat of reaction.
2. Use the textbook format in setting up chemical equations.
3. Identify and use common symbols in writing chemical equations.
4. Balance simple chemical equations.
5. Interpret a balanced equation in terms of molecules, atoms, grams, or moles of each substance used or produced.
6. Classify reactions as combination, decomposition, single replacement, or double replacement.

F. The student will understand and apply the principles of gas behavior in ideal systems.

1. Define pressure, diffusion, barometer, standard conditions, ideal gas, atmospheric pressure, one atmosphere, molar volume.
2. State the principal assumptions of kinetic molecular theory (KMT).
3. Explain the five (5) given properties of a gas in terms of the KMT.
4. Describe how a gas exerts pressure.
5. Describe how a barometer works.
6. Express one atmosphere in terms of mm of Hg, inches of Hg, torr, and lbs/in^2.
7. State and apply Boyles' law.
8. State and apply Charles' law.
9. Apply the combined gas law to find the volume of a gas when both the temperature and pressure change.
10. Use the molar volume of a gas in conjunction with the combined gas law to solve for molar mass, mass, or volume of a gas.
11. Calculate the density of an ideal gas at STP.
12. State and apply Dalton's law of partial pressures in determining the pressures of component gases in a mixture of gases.
13. State and apply Gay-Lussac's law.
14. State and apply Avogadro's hypothesis.
15. Define an ideal gas.
16. State two valid reasons why real gases may deviate from the behavior predicted for an ideal gas.
17. Recognize the ideal gas equation.

G. The student will understand the properties of aqueous solution systems and the theories describing the behavior of acids and bases in aqueous systems.

1. List the melting point, normal boiling point, heat of fusion, heat of vaporization, specific heat, and density at 4^oC of water.

2. Describe the water molecule with respect to electron dot structure and polarity.
3. Explain the effect of hydrogen bonding on the physical properties of water.
4. Complete and balance equations for acid-base neutralization.
5. Given a list of hydrates, write balanced equations for their decomposition reactions to water and the anhydride.
6. Identify metal oxides as basic anhydrides and write balanced equations for their reactions with water.
7. Identify nonmetal oxides as acid anhydrides and write balanced equations for their reactions with water.
8. Distinguish between peroxides and ordinary oxides.
9. Discuss the occurrences of ozone and its effect on humans.
10. Define solution, solvent, miscible, dilute solution, concentrated solution, unsaturated solution, mass-percent, normality, solute, solubility, immiscible, concentration of a solution, saturated solution, supersaturated solution and molarity.
11. Qualitatively predict the effect of temperature change on the solubility of solids and gases in liquids.
12. Calculate the mass-percent concentration of a solution.
13. Calculate the mass or volume of solute, or mass or volume of solution when given the mass-percent or volume percent concentrations.
14. Calculate the molarity of a solution.
15. Calculate the moles or the mass of solute, or volume of solution when given the molarity and other appropriate data.
16. Relate mass, moles, or gas volume of substances in a chemical reaction when given the chemical equation.
17. Define salt, amphoteric, nonectrolyte, ionization, weak electrolyte, pH, neutralization, spectator ions, hydronium ion, electrolyte, dissociation, strong electrolyte and titration;
18. Give the following definitions of acids and bases:
    a. Arrhenius
    b. Bronsted-Lowry
    c. Lewis
19. Classify common compounds as electrolytes or non-electrolytes.
20. Classify common acids, bases, and salts as strong or weak electrolytes.
21. Relate pH and hydrogen ion concentration.

H. The student will understand and apply the principles of chemical equilibrium, kinetics, and oxidation-reduction.

1. Define equilibrium, chemical equilibrium, chemical kinetics, LeChatelier's Principle, catalyst, common-ion effect, buffer solution and reversible chemical reaction.
2. Describe a reversible reaction.
3. State and explain the qualitative effect of Le Chatelier's principle.
4. Predict how the rate of a chemical reaction is affected by:
    a. changes in concentration of reactants;

b. changes in pressure on gaseous reactants;
c. changes in temperature;
d. the presence of a catalyst.

5. Discuss the common ion effect on a system at equilibrium.
6. Explain how a buffer solution is able to counteract the addition of small amounts of either H^+ or OH^- ions.
7. Draw the relative energy diagram of a reaction in terms of activation energy, exothermic or endothermic reaction, and the effect of a catalyst.
8. Define oxidation and reduction.
9. Assign oxidation numbers to elements in chemical compounds and ions.
10. Describe the difference between electrolytic and voltaic cells.
11. Describe the relationship among elements in an activity series of metals.

I. The student will understand and apply the principles of radioactivity in nuclear chemistry.

1. Define radioactivity, beta particle, trans-uranium elements, alpha particle and gamma ray;
2. Give the major contribution of the following to the historical development of nuclear chemistry:
   a. Henri Becquerel
   b. Wilhelm Roentgen
   c. Marie and Pierre Curie
   d. Ernest Rutherford
3. List the characteristics that distinguish alpha, beta, and gamma rays from the standpoint of man, charge relative velocities and penetrating power.
4. Use the periodic table and identify the trans-uranium elements.

Laboratory Topics

A. The student will develop an understanding of basic laboratory techniques and procedures.

1. Properly operate the Bunsen burner.
2. Operate a single pan balance.

B. The student will understand basic laboratory safety and will follow all laboratory rules during experimental work.

1. Follow basic laboratory safety rules as set forth by the department and the instructor.
2. Locate laboratory safety and first aid equipment.

C. The student will acquire understanding of the physical and chemical properties of commonly used elements, compounds, and mixtures.

1. Distinguish between physical and chemical properties of substances.

2. Determine physical properties such as density, volume, mass, etc.
3. Make specific and accurate observations of materials and reactions as to color, odor, energy changes, gas evolution, precipitation, etc.
4. Identify evidence of chemical changes.

D. The student will be able to make precision measurements and evaluation of experimental data through selected quantitative laboratory experiments.

1. Use a meter stick to measure length of any object in cm, mm, and meters.
2. Read centigrade thermometers and convert to Kelvin and Fahrenheit.
3. Read the volume contained in any graduated cylinder to within 0.5 ml.
4. Use a laboratory balance to determine the mass of any object to within 0.01 g.

E. The student will be able to make careful observations, report and interpret experimental results through selected qualitative laboratory experiments.

1. Interpret evidence of solubility and miscibility.
2. Collect a precipitate by filtration.
3. Predict the formation of precipitates based on principles of solubility.
4. Make accurate observations of state, color, and odor of elements, compounds, and mixtures.
5. Distinguish between elements, compounds, and mixtures.
6. Record evidence of chemical change occurring in a reaction.
7. Determine the relative activities of two metals in a single replacement reaction.
8. Arrange a group of metals from most active to least based upon observations of a series of single replacement reactions.

F. The student will be able to perform simple calculations from experimental data through selected quantitative laboratory experiments.

1. Calculate the densities of selected solids and water.
2. Calculate the percent of water in selected unknown hydrated salts.
3. Calculate the empirical formula for strontium iodide salt or other compound.
4. Calculate the weight percent of potassium dichromate in a saturated solution.

VII. Evaluation and Assessment

The student will have demonstrated attainment of the general course objectives if he accumulates a minimum of 70 percent of the points possible.

Grades will be composed of tests, lab work, a comprehensive final exam, and may include other assignments. Lecture will count for 75 – 80% and the laboratory component will count for 20-25% of the student's grade. A minimum of three lecture exams and a comprehensive final exam will be given. In lab a minimum of one exam and a final exam will be given.

Grades will be given based upon A = 90 – 100%, B = 80 – 89%, C = 70 – 79%, D = 60 – 69%, and F = below 60%.

VIII. Attendance

Students are expected to attend all classes for which they are registered. Students who are unable to attend class regularly, regardless of the reason or circumstance, should withdraw from that class before poor attendance interferes with the student's ability to achieve the objectives required in the course. Withdrawal from class can affect eligibility for federal financial aid.

IX. Statement on Discrimination/Harassment

The College and the Alabama State Board of Education are committed to providing both employment and educational environments free of harassment or discrimination related to an individual's race, color, gender, religion, national origin, age, or disability. Such harassment is a violation of State Board of Education policy. Any practice or behavior that constitutes harassment or discrimination will not be tolerated.

X. Americans with Disabilities

The Rehabilitation Act of 1973 (Section 504) and the Americans with Disabilities Act of 1990 state that qualified students with disabilities who meet the essential functions and academic requirements are entitled to reasonable accommodations. It is the student's responsibility to provide appropriate disability documentation to the College. The ADA Accommodations office is located in FSC 300 (205-856-7731).

Date Adopted: 1965, 1967
Dates Revised: 1967, 1998, 2001, 2008

Representing Alabama's Public Two-Year College System
Alabama Department of Postsecondary Education

Jefferson State Community College

CHM 105

Introduction to Organic Chemistry

I. CHM 105, Introduction to Organic Chemistry, 4 Semester Hours

Core Area III, ASCI TSCI (Lec 3 hrs, Lab 2 hrs) (***State guide has 3HR Labs)

II. Course Description

This is a survey course of organic chemistry and biochemistry for students who do not intend to major in science or engineering. Topics will include basic nomenclature, classification of organic compounds, typical organic reactions, reactions involved in life processes, function of biomolecules, and the handling and disposal of organic compounds. Laboratory is required.

III. Prerequisite

CHM 104 or CHM 111

IV. Textbook

Fundamentals of Organic Chemistry, 6th Ed. McMurray
Safety-Scale Laboratory Experiments for Chemistry for Today, 5th Ed. Seager and Slabaugh

V. Course Objectives

At the end of the course the student will be able to:

A. Understand the relationship of the hybridization of the carbon atom to the various types of carbon bonds.

B. Apply the concepts of carbon bonding to distinguish and identify types of organic compounds.
C. Apply the concepts of the versatility of carbon bonds in naming organic compounds using the IUPAC System of nomenclature and in drawing the structural formulas.
D. Apply the concepts of the wide versatility of carbon compounds in understanding the preparation and the chemical reactions of these compounds.
E. Understand the relationship of the structure of organic molecules to the physical properties.
F. Possess an appreciation of the wide economic implications of organic compounds by knowing the source and use of these varied structures.
G. Apply the concepts and knowledge of simple organic molecules to understanding and recognizing the more complicated polymers.
H. Understand steriochemistry and its relationship to rotation of light and the implications in studying biomolecules such as enzymes.
I. Discuss basic structure and function of carbohydrates, lipids, amino acids, proteins and nucleic acids.

VI. Course Outline of Topics

Lecture Topics Stated in Performance Terms

A. The student will be able to understand the relationship of the hybridization of the carbon atom to the various types of carbon bonds.

1. Describe the sp^3, sp^2, and sp hybridization of the carbon atom using the electronic configuration and orbital sketches.
2. Relate each of the types of hybridization to single, double, and triple bonds.
3. Distinguish between a sigma and a pi bond.
4. Distinguish between a molecular formula, structural formula and condensed structural formula.

B. The student will be able to apply the concepts of carbon bonding to distinguish and identify types of organic compounds.

1. Using structural formulas, the student will be able to
2. Name and give the general formula for at least nine classes of organic compounds.
3. Identify an alkyl group and an alkyl halide.
4. Define an isomer and give three isomers when given the molecular formula.
5. Distinguish between cis and trans isomers using structural formulas.
6. Distinguish between a primary, secondary, and tertiary alcohol.
7. Name four classes of carboxylic acids and give an example of each.
8. Identify glycerol.

9. Identify a triglyceride.
10. Distinguish between primary, secondary, and tertiary amines.
11. Distinguish between fats and oils (structural formula not needed here).
12. Distinguish between a soap and syndet.

C. The student will be able to apply the concepts of the versatility of carbon bonds in naming organic compounds using the IUPAC System of nomenclature and in drawing the structural formulas.

1. Name any one of the alkanes, alkenes, alkynes, cycloalkanes, aromatic compounds, alcohols, aldehydes, ketones, ethers, carboxylic acids, esters, amines and amides when given the structural formula.
2. Draw the structural formulas for any of the groups listed above when given the name.

D. The student will apply the concepts of the wide versatility of carbon compounds in understanding the preparation and the chemical reactions of these compounds.

1. Halogenation, dehydrogenation and combustion of alkanes.
2. Addition and oxidation of alkenes and alkynes.
3. Oxidation of primary, secondary, and tertiary alcohols.
4. Addition reactions that illustrate an understanding of “Markovnikov's Rule.”
5. Intermolecular and intramolecular dehydration of alcohols.
6. Esterfication of alcohols.
7. Oxidation and reduction of aldehydes and ketones.
8. Tollens' test and the Benedict's test used to distinguish between aldehydes and ketones.
9. Aldol condensation of aldehydes and ketones.
10. Preparation of the Grignard reagent.
11. Grignard reaction that will produce a given alcohol.
12. Preparation of carboxylic acids by oxidation of alcohols, hydrolysis of esters and fats, oxidation of aromatic hydrocarbons, and hydrolysis of nitriles.
13. Formation of esters, amides, and acid chlorides using carboxylic acid.
14. Preparation of esters.
15. Preparation of a quaternary ammonium salt.
16. Illustration of the basic properties of amines.

E. The student will be able to understand the relationship of the structure of organic molecules to the physical properties.

1. Relate the physical state of the hydrocarbons to the length of the carbon chain.
2. List the physical properties of aromatic hydrocarbons.

3. Relate the solubility and boiling point of the alcohols, ethers, and carboxylic acids to their chemical structure.
4. List two physical properties of formaldehyde, acetaldehyde, paraldehyde, and benzaldehyde.
5. List the acidic properties of carbocylics.
6. List four physical properties of esters.
7. Give the major physical difference in fats and oils and state the reason for this difference.
8. Describe the cleaning action of soap.
9. List two physical properties of amines.

F. To gain and appreciation of the economic implications of organic compounds by knowing the source and use of these varied structures.

1. List the major source and uses of the hydrocarbons, alcohols, ethers, aldehydes, ketones, esters, carboxylic acids, amines and amino acids.
2. Discuss the physiology of alcohol as a food, drug, and poison.

G. The student will be able to apply the concepts and knowledge of simple organic molecules to understanding and recognizing the more complicated polymers.

1. Define the terms polymer, monomer and polymerization.
2. Name and describe both physically and chemically the two general types of polymers.
3. List six general groups of synthetic polymers.
4. Give the three basic steps in the formation of an addition polymer.
5. Define a copolymer.
6. Identify specific polymers such as styrene, polyurethane and polyvinyl chloride.

H. The student will be able to understand stereochemistry and its relation to rotation of light and the implications in studying biomolecules such as enzymes.

1. Define stereoisomerism and chirality.
2. List two types of stereoisomers.
3. Describe plane-polarized light
4. Describe optical activity and specific rotation.
5. Use sequence rules for specifying configuration of optical isomers.
6. Describe the difference between enantiomers, diasteriomers, meso compounds, and racemic mixtures.
7. Describe and give examples of chirality in nature and, particularly, for enzymes.

I. Discuss basic structure and function of carbohydrates, lipids, amino acids, proteins and nucleic acids.

1. Discuss basic structure and function of carbohydrates, lipids, amino acids, proteins and nucleic acids.
2. Demonstrate an understanding of the classification system for carbohydrates.
3. Draw glucose as both a Fischer structure and a Haworth structure.
4. List four disaccharides and their alternate names.
5. Describe the monosaccharides combinations that compose the four disaccharides, the enzymes needed to cleave the disaccharides, and the types of linkages used to form the disaccharides.
6. Describe the difference between the 2 common polysaccharides.
7. Describe the general structure of an α-amino acid and give three specific examples.
8. Show how α-amino acids form peptide linkages.
9. Define "isoelectric point".
10. List the different classifications of proteins.
11. Define "lipids".

Laboratory Topics

A. Safety Procedures; rules and regulations
B. Ball and Stick Models of Compounds Representing Various Functional Groups
C. Melting point determination
D. Preparation of Aspirin
E. Purification by Recrystallization
F. Characteristic Reactions of Alkenes
G. Characteristic Reactions of Alcohols
H. Characteristic Reactions of Aldehydes
I. Characteristic Reactions of Ketones
J. Characteristic Reactions of Amines
K. Preparation of Esters
L. Optical Activity and Polarimetry
M. Thin Layer Chromatography of Analgesics

VII. Evaluation and Assessment

The student will have demonstrated attainment of the general course objectives if he accumulates a minimum of 70 percent of the points possible.

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Jefferson State Community College

CHM 111

College Chemistry I

I. CHM 111, College Chemistry I, 4 Semester Hours

Core Area III, ASCI TSCI (Lec 3 hrs, Lab 2 hrs) (***State guide has 3HR Labs)

II. Course Description

This is the first course in a two-semester sequence designed for the science or engineering major who is expected to have a strong background in mathematics. Topics in this course include measurement, nomenclature, stoichiometry, atomic structure, equations and reactions, basic concepts of thermochemistry, chemical and physical properties, bonding, molecular structure, gas laws, kinetic-molecular theory, liquids and solids, solutions, and colloids. Lab is required.

III. Prerequisite: MTH 112 (Precalculus Algebra) or equivalent math placement score.

IV. Textbook

Chemistry and Chemical Reactivity, Kotz and Treichel, 7th Ed.
Modular Lab Notebook, Chemical Education Resources

V. Course Competencies

In the classroom the student will:

A. Understand the basic mathematical principles involved in chemical calculations and will have a thorough understanding of the metric system of measurement.
B. Understand the classification of matter into various groups based upon similarity of chemical and physical properties.
C. Comprehend and interpret chemical symbols, formulas, names, chemical equations, and calculations that apply, stressing stoichiometry.
D. Understand the electronic arrangement of the atom in terms of the quantum theory, and will be able to use the periodic table to link electronic configuration to the properties of the element.
E. Understand chemical periodicity.
F. Understand the nature of covalent bonding, ionic bonding, and the main concepts of the three covalent bonding theories (VB, VSEPR, MO).
G. Comprehend and apply the principles of gas behavior in ideal as well as real gas systems.
H. Understand the characteristics of the solid and liquid states of matter and phase diagrams.
I. Understand terms used in solution chemistry.
J. Understand the nature of aqueous solution systems and apply the principles of solubility, colligative properties and concentration in problem solving.
K. Understand and apply chemical principles of acids and bases.

In the laboratory the student will:

A. Develop an understanding of basic laboratory techniques and procedures.
B. Understand basic laboratory safety and will follow all laboratory rules during experimental work.
C. Acquire understanding of the physical and chemical properties of commonly used elements, compounds and mixtures.
D. Be able to make precise measurements and evaluate experimental data through selected qualitative laboratory experiments.
E. Be able to make careful observations, report and interpret experimental data through selected quantitative laboratory experiments.
F. Be able to perform simple calculations from experimental data through selected quantitative laboratory experiments.

VI. Course Outline of Topics

Lecture Topics Stated in Performance Terms

The student will be required to demonstrate that he has attained each general course competency by performing the objectives listed under each competency.

A. Understand the basic mathematical principles involved in chemical calculations, and will have thorough understanding of the metric system of measurement.

1. Express any given number in exponential notation, and use exponential notation in calculations.
2. Apply the rules of significant figures to calculations based upon experimental measurements.
3. Cite from memory the meaning of all metric prefixes listed in a reference table.
4. Cite from memory at least one conversion factor from reference table relating metric and English units of (a) mass, (b) length, and (c) volume.
5. Use dimensional analysis to convert within the metric system and between metric and non-metric units.
6. Carry out calculations relating density, specific gravity, mass, and volume to one another from assigned problems.
7. Convert a specified temperature in degrees Celsius, Fahrenheit or Kelvin to the other two scales.
8. Carry out calculations relating heat capacity or specific heat to the heat transfer that accompanies temperature changes.

B. Understand the classification of matter into various groups based upon similarity of chemical and physical properties.

1. State the Law of Conservation of Matter, the Law of Conservation of Energy, and the Law of Matter and Energy.
2. Express these Laws in words other than those used in the text and give an example of each.
3. Describe and distinguish among the general properties of gases, liquids, and solids.
4. Define and give examples of each: a chemical change, a physical change, a chemical property, a physical property, an intensive property, and extensive property, an endothermic change and exothermic change.
5. Define, distinguish among, and give an example of: a pure substance, a compound, a mixture, an atom, an element and molecule.
6. Distinguish between: mass and weight; density and specific gravity, accuracy

and precision, heat and temperature.

C. Comprehend and interpret chemical symbols, formulae, names, chemical equations, and calculations that apply, stressing stoichiometry.

1. Memorize the names and symbols of the elements given in a reference table.
2. Interpret a chemical formula in terms of the type and number of atoms present.
3. Given an ionic or molecular formula, determine the formula weight.
4. Relate the numbers of particles (atoms, molecules, or ions) and the mass in grams of a sample of matter.
5. Given the formula of a substance, relate the number of moles and the mass in grams of the sample.
6. Write and interpret the formulas for some common substances.
7. Given the formula of a compound, calculate the percentages by mass of the elements.
8. Determine the empirical formula of compound, given the mass percentages of the elements or the analytical data from which these can be calculated.
9. Determine the molecular formula of a compound, given the simplest formula and at least an approximated molecular mass.
10. Use the Periodic Table to obtain the charges of ions formed by the main-group elements.
11. Write the formula for an ionic compound given either the formulas of the ions or the name of the compound.
12. Given the formula for a compound, give its name.
13. Describe some experimental methods of determining percent composition.
14. Write and balance chemical equations and interpret the various symbols used in chemical equations to represent the condition of the reaction system.
15. Relate the number of moles of any two substances taking part in a reaction.
16. Relate the masses of any two substances taking part in a reaction.

17. Given or having calculated two of the three quantities, concentration, number of moles of solute, volume of solution, determine the other quantity.

18. Given the balanced equation for a reaction involving species in solution, relate the volumes or concentrations of two reactant species.

19. Describe water solution reactions involving precipitation, acid-base, and oxidation-reduction.

20. Given the number of moles or masses of all reactants, determine which is the limiting reagent and calculate the theoretical yield of any product.

21. Calculate the percent yield, given the actual or theoretical yields.

D. Understand the electronic arrangement of the atom in terms of quantum theory, and will be able to use the periodic table to link electronic configuration to the properties of the elements.

1. State and describe the postulates of quantum theory and compare quantum and classical theories.

2. Relate the wavelength and frequency of a spectral line to the energy of photons and to the change in energy of an atom.

3. Discuss the contributions to the atomic theory made by Dalton, Thomson, Rutherford, Bohr, Chadwick, deBroglie and Schrodinger.

4. State and apply the Aufbau principle.

5. Determine the number of electrons that may be accommodated by any given principal energy level or sublevel.

6. Given the atomic number of an element, write the electron configuration.

7. Given the electron configuration, state and apply Hund's rule and draw orbital diagram of the atom.

8. Describe the four quantum numbers, and the rules for assigning them.

9. Apply the rules and assign them to each of the various electrons in an atom.

10. State and apply Pauli's Exclusion principle.

11. Relate electronic configurations to the periodic table and to periodicity.

12. Using the periodic table, predict the relative values of ionization energy,

electronegativity, atomic radius, and ionic radius.

E. Understand chemical periodicity.

1. Describe what is meant by periodicity.
2. Relate the electronic configuration of an atom to the position of the element in the periodic table.
3. Categorize elements according to alkali metals, representative elements, chalcogens, alkaline earth metals, halogens, d-transition elements, noble gases, lanthanide series, actinide series, inner transition elements and f-transition elements.
4. Summarize horizontal and vertical trends in the periodic table for each of the following properties: atomic radii, electron affinity, ionization energy, ionic radii and electronegativity.
5. State the contrasting physical and chemical properties of metals and non-metals.
6. Know where the metals, nonmetals, and metalloids are located in the periodic table.

F. Understand the nature of covalent bonding, ionic bonding and the main concepts of the three covalent bonding theories (VB, VSEPR, MO).

1. Distinguish between the major aspects of ionic and covalent bonding.
2. Identify some of the major differences in physical properties of ionic and covalent compounds.
3. Predict whether bonding between a given pair of elements or in a given compound would be covalent or ionic.
4. Predict the formulas of binary ionic compounds from the positions of the constituent.
5. Given a periodic table, predict and compare bond polarity.
6. Draw Lewis structures for molecules and polyatomic ions.
7. Given or having written the Lewis structure of a molecule or ion, predict its geometry.

8. Predict molecular polarity from Lewis structures.
9. Predict molecular geometry from orbital hybridization.
10. Predict the kind and number of sigma and pi bonds in a molecular species.
11. Write molecular orbital diagrams for simple diatomic species.

G. Comprehend and apply the principles of gas behavior in ideal as well as real gas systems.

1. Describe and apply Boyle's, Charles', Gay-Lussac's and Avogadro's law.
2. Apply the ideal gas law to predict the effect of a change in conditions upon a variable such as volume.
3. Apply gas laws to calculate the density of a gas at a given temperature and pressure.
4. Use the ideal gas law to calculate the molecular mass of a gas, knowing the mass of a given volume or the density at a known pressure and temperature.
5. Relate volumes of gases involved in chemical reactions from information obtained from chemical equations.
6. Apply Dalton's law of partial pressures of gases in mixtures.
7. List the assumptions of kinetic-molecular theory and describe gas behavior in terms of the theory.
8. Describe and apply Graham's law to relate molecular masses, rates of effusion, and times of effusion of gases.
9. Describe how real gases deviate from the assumptions of the ideal gas law, and indicate the conditions where these deviations are most significant for most gases.

H. Understand the characteristics of the solid and liquid states of matter and phase diagrams.

1. Determine vapor pressure of liquids in given temperatures and pressures.
2. Predict and describe the various intermolecular forces present in molecular substances.

3. Classify a given substance as ionic, nonpolar, polar, macromolecular, or metallic.

4. List the general physical properties associated with each of the five categories of substances listed above.

5. Write equations for the thermal decomposition of carbonates, hydroxides, and hydrates.

6. Determine enthalpy change associated with a given phase change.

7. Interpret phase diagrams and apply them to predict phase changes associated with changes in temperature and pressure.

I. Understand terms used in solution chemistry.

1. Describe and distinguish among solvent, solute, solution, dispersion and colloid.

2. Give examples of various kinds of solutions involving different combinations of solids, liquids, and gases as dispersing medium and dispersed substances

3. Describe the relative effects on solubility of the following kinds of interactions: solute-solute; solvent-solvent; solvent-solute.

4. Describe and illustrate the mechanism of dissolution of ionic solids and polar covalent substances in water.

5. State the effects of exo- or endothermicity and of an increase in disorder on the spontaneity of the dissolution process.

6. Distinguish among unsaturated, saturated, and supersaturated solutions.

7. Distinguish between exothermic and endothermic dissolution process.

J. Understand the nature of aqueous solution system and apply the principles of solubility, colligative properties and concentration in problem solving.

1. Given the formula for a substance, predict whether it will be an electrolyte or a non-elecrolyte in aqueous solution.

2. Predict the relative solubilities of different solutes in water.

3. Predict the effect on solubility of a change in temperature or pressure.

4. Given a molecular, write the corresponding total and net ionic equations.

5. Using the solubility rules, predict whether two soluble ionic compounds will react to form a precipitate; if they will, write the net ionic equation.

6. Utilizing an equation for a precipitation reaction, relate the mole relationship of reactants and products.

7. List the colligative properties of solutions and state on what factors they depend.

8. Use the colligative properties to determine the molecular weights of solutes.

9. Applying Raoult's law, calculate the vapor pressure of a solution.

10. Applying Henry's law, given data for one set of conditions, determine the concentration of a gas in solution at a second set of conditions.

11. Perform calculations related to molarity, normality, mole fraction, and mass percent;

12. Perform calculations related to molarity and colligative properties.

L. Understand and apply chemical principles of acids and bases.

1.Describe the similarities and differences among the Arrhenius, Bronsted-Lowry, and Lewis theories of acids and bases.

2.Classify any given species in a reaction as an acid or base, according to acid-base pairs.

3. Describe the nature of the proton in aqueous solution, with special attention to the hydronium ion, H_3O^+.

4. Name the common strong acids and bases.

5. Write equations for the dissociation of strong acids or strong bases.

6. Write equations for the ionization equilibria of weak acids and weak bases in aqueous solutions.

7. Write equations for the reactions of acids and bases, describe the solutions that result as acidic, basic or neutral.

8. Describe the ionization of a poly-protic acid in aqueous solutions.

9. Predict the relative strengths of acids and bases from a given set of molecular structures.

10. Use titration data for an acid-base reaction to determine: the concentration of an acid or a base in aqueous solutions, and molecular mass of an acid or a base.

11. Select an acid-base indicator appropriate for a given acid-base titration.

12. Given the following $[H^+]$, $[OH^-]$, pH, or pOH calculate any of the others.

Laboratory Topics

A. The student will develop an understanding of basic laboratory techniques and procedures.

1. Properly operate the Bunsen burner.
2. Operate a single pan balance.

B. The student will understand basic laboratory safety and will follow all laboratory rules during experimental work.

1. Follow basic laboratory safety rules as set forth by the department and the instructor.
2. Locate laboratory safety and first aid equipment.

C. The student will acquire understanding of the physical and chemical properties of commonly used elements, compounds, and mixtures.

1. Distinguish between physical and chemical properties of substances.
2. Determine physical properties such as density, volume, mass, etc.
3. Make specific and accurate observations of materials and reactions as to color, odor, energy changes, gas evolution, precipitation, etc.
4. Identify evidence of chemical changes.

D. The student will be able to make precise measurements and evaluate experimental data through selected quantitative laboratory experiments.

1. Use a meter stick to measure length of any object in cm, mm, and meters.
2. Read centigrade thermometers and convert to Kelvin and Fahrenheit.
3. Read the volume contained in any graduated cylinder to within 0.5 ml.
4. Use a laboratory balance to determine the mass of any object to within 0.01 g.

E. The student will be able to make careful observations, report and interpret experimental results through selected qualitative laboratory experiments.

1. Interpret evidence of solubility and miscibility.
2. Collect a precipitate by filtration.
3. Predict the formation of precipitates based on principles of solubility.
4. Make accurate observations of state, color, and odor of elements, compounds, and mixtures.
5. Distinguish between elements, compounds, and mixtures.
6. Record evidence of chemical change occurring in a reaction.
7. Determine the relative activities of two metals in a single replacement reaction.
8. Arrange a group of metals from most active to least based upon observations of a series of single replacement reactions.

F. The student will be able to perform simple calculations from experimental data through selected quantitative laboratory experiments.

1. Calculate the densities of selected solids and water.
2. Calculate the percent of water in selected unknown hydrated salts.
3. Calculate the empirical formula for strontium iodide salt or other compound.
4. Determine the concentration of a basic solution when titrated with acid of a known concentration.
5. Determine the molar mass of an impurity due to freezing point depression.
6. Apply the rules of significant figures, rounding off, exponential notation, and instrument precision to the numerical results of measurements and calculations.
7. Measure volumes of liquids correctly with burettes, graduated cylinders and pipettes.
8. Apply titration techniques to the standardization of solutions.
9. Use pH paper, universal indicator, and special indicators in determinations of pH of solutions.
10. Use a spectrophotometer to measure percent transmission or absorbance of a

solution, and use the solutions of known concentration to prepare a standard curve.

VII. Evaluation and Assessment

The student will have demonstrated attainment of the general course objectives if he accumulates a minimum of 70 percent of the points possible.

Grades will be composed of tests, lab work, a comprehensive final exam, and may include other assignments. Lecture will count for 75 – 80% and the laboratory component will count for 20-25% of the student's grade. A minimum of three lecture exams and a comprehensive final exam will be given. In lab a minimum of one exam and a final exam will be given.

Grades will be given based upon the traditional scale:

A = 90 – 100%, B = 80 – 89%, C = 70 – 79%, D = 60 –69%, and F = below 60%.

VIII. Attendance

Students are expected to attend all classes for which they are registered. Students who are unable to attend class regularly, regardless of the reason or circumstance, should withdraw from that class before poor attendance interferes with the student's ability to achieve the objectives required in the course. Withdrawal from class can affect eligibility for federal financial aid.

IX. Statement on Discrimination/Harassment

The College and the Alabama State Board of Education are committed to providing both employment and educational environments free of harassment or discrimination related to an individual's race, color, gender, religion, national origin, age, or disability. Such harassment is a violation of State Board of Education policy. Any practice or behavior that constitutes harassment or discrimination will not be tolerated.

X. Americans with Disabilities

The Rehabilitation Act of 1973 (Section 504) and the Americans with Disabilities Act of 1990 state that qualified students with disabilities who meet the essential functions and academic requirements are entitled to reasonable accommodations. It is the student's responsibility to provide appropriate disability documentation to the College. The ADA Accommodations office is located in FSC 300 (205-856-7731).

Date Adopted: 1965
Date Reviewed: 1985

Date Revised: 1999, 2001, 2003, 2008

Representing Alabama's Public Two-Year College System

Alabama Department of Postsecondary Education

Jefferson State Community College

CHM 112

College Chemistry II

I. CHM 112, College Chemistry II, 4 Semester Hours

Core Area III, ASCI TSCI (Lec 3 hrs, Lab 2 hrs) (***State guide has 3HR Labs)

II. Course Description

This is the second course in a two-semester sequence designed primarily for the science or engineering major who is expected to have a strong background in mathematics. Topics in this course include chemical kinetics, chemical equilibria, acids and bases, ionic equilibria of weak electrolytes, solubility product principle, chemical thermodynamics, electrochemistry, oxidation-reduction, nuclear chemistry, and selected topics in organic chemistry, biochemistry, atmospheric chemistry, and descriptive chemistry, including the metals, nonmetals, semi-metals, coordination compounds, transition compounds, and post-transition compounds. Laboratory is required.

III. Prerequisite:

CHM 111 or equivalent course and MTH 112 (Precalculus Algebra) or equivalent math placement score.

IV. Textbook

Chemistry and Chemical Reactivity, Kotz and Treichel, 7th Ed.

V. Course Competencies

At the end of the course the student will be able to:

A. Understand and apply the principles of chemical thermodynamics.
B. Understand and apply the principles of chemical kinetics.
C. Comprehend the nature of equilibrium systems.
D. Understand the properties of electro-chemical systems involving oxidation-reduction reactions.
E. Apply the concepts on ionic equilibrium and solubility in solving problems.
F. Apply the concepts of acid-base neutralization reactions.
G. Apply the concepts of co-ordination compounds to better understand the complexes formed in metal ions.
H. Apply the concepts of nuclear chemistry.

VI. Course Outline of Topics

Lecture Topics Stated in Performance Terms

The student will be required to demonstrate that he has attained each general course competency by performing the objectives listed under each competency.

A. The student will understand and apply the principles of chemical thermodynamics.

1. Define the following terms:

system	surroundings
state of a system	state function
standard conditions	endothermic process
spontaneous process	non-spontaneous process

2. State the First Law of Thermodynamics both in words and in mathematical form, then summarize its implications with respect to reaction spontaneity;
3. Explain the relationship between energy change and enthalpy change;
4. Discuss what is meant by the standard enthalpy change of a reaction;
5. Perform the calculations of Hess's Law to determine enthalpy changes for given reactions;
6. State the Second Law of Thermodynamics and summarize its implications with respect to reaction spontaneity;
7. Predict ΔS for many kinds of common changes, both chemical and physical;
8. Explain how ΔH and $T\Delta S$ are related to spontaneity of a reaction;
9. Discuss the meaning of "Gibbs free energy change" for a reaction and relate

it to enthalpy and entropy changes;

10. Use the standard Gibbs free energy changes for a reaction as an indicator of spontaneity;
11. Given the standard molar entropies of reactants and products, calculate the standard entropy change for a reaction;
12. Given the enthalpy change and the standard entropy change for a reaction, calculate the standard free energy change at 298K and at any other temperature;
13. Given the enthalpy change and the standard entropy change for a reaction, calculate the temperature at which equilibrium will exist at l atm;
14. Quantitatively relate the standard free energy change and the E cell for a given reaction at 298K;
15. Quantitatively relate the standard free energy change and the K_c for a reaction in an aqueous system;
16. Apply the laws of thermochemistry to calculations involving standard entropy change, standard free energy change, and enthalpy change.

B. The student will understand and apply the principles of chemical kinetics.

1. Determine the order of a reaction, given the rate as a function of concentration of reactants;
2. Determine the order of a reaction, given the concentration of a reactant as a function of time;
3. Given the order of a reaction, write a rate expression for the reaction, and calculate the rate constant given the rate at a known concentration;
4. Use rate equations to determine original concentrations and the rate constants;
5. Use the rate equations to determine the time required for concentration of reactant to drop to a particular value, given the rate constant and the original concentration; (Also be able to determine the initial concentration given the concentration at some particular time.)
6. Given either half-life or rate constant for a first order reaction, calculate the other quantity;
7. Describe and assess energy diagrams showing energy of activation and enthalpy change and describe the effect of catalysis;
8. List and describe the three factors which affect rates of reaction according to collision theory;
9. Use the Arrhenius equation to obtain the rate constant at T_2 given its value at T_1 and the energy of activation;
10. Use the Arrhenius equation to obtain the activation energy given rate constants at two different temperatures.

C. The student will comprehend the nature of equilibrium systems.

1. Given a balanced equation for a reaction involving gases, write the corresponding expression for K_C;

2. Interpret the magnitude of K_C in relation to the extent of forward and reverse reactions.

3. Given initial concentrations of all species on a reaction and the value of K for this reaction, calculate equilibrium concentrations of all species;

4. Distinguish between the reaction quotient, Q, and the equilibrium constant, K_C;

5. Use Q to determine whether or not a given system is at equilibrium, and if not, how it must proceed to approach equilibrium;

6. For a given equation, calculate the numerical values of K_{eq} knowing the K_C equilibrium concentrations of all species;

7. For a given equation, calculate the numerical value of K_C knowing the original concentrations of all species and the equilibrium concentration of one species;

8. Write the equilibrium constant expression for K_C and K_P and calculate their values for a given reaction;

9. Given the value of K_C , predict the direction in which a chemical system will move to reach equilibrium;

10. Given the value of K_C, predict the equilibrium concentrations of one species, knowing the concentrations of all the other species at equilibrium;

11. Given the value of K_C , predict the equilibrium concentrations of all species, given their initial concentrations;

12. Use Le Chatelier's Principle to predict the direction in which a system at equilibrium will shift (if it does) when stresses of the following kinds are applied:

    a. change in pressure
    b. change in volume
    c. change in temperature
    d. change in amounts of reactants or products present
    e. addition of a catalyst

13. Given concentrations or partial pressures of all species in a system at equilibrium, to which a stress is then applied, determine the concentrations or partial pressures of all species after equilibrium is re- established;

14. Given the equilibrium constant for a reaction at a particular temperature, calculate the standard free-energy change, ΔG, at that temperature and vice versa;

15. Given the standard enthalpy change, ΔH , and the equilibrium constant at a particular temperature, calculate the equilibrium constant at a different temperature.

D. The student will understand the properties of electrochemical systems involving oxidation-reduction reactions.

1. Determine the oxidation number of each atom in a molecule or an ion when given the molecular or ionic formula;
2. Define oxidation, oxidizing agent, electrolytic cell, anode, reduction, reducing agent, voltaic cell and cathode;
3. Balance molecular and net ionic equations for redox reactions using the ion-electron, (half-reaction) method;
4. Label the oxidizing and the reducing agents and the species being oxidized and reduced in a balanced oxidation-reduction reaction;
5. Given the balanced equation for a redox reaction and titration data for the reaction, calculate the concentration of one of the reactant species;
6. Utilize standard voltages to decide whether or not a given redox reaction will occur at standard concentration and pressure at 298K;
7. Apply the Nernst equation to the determination of electrode potentials and the cell potentials under nonstandard conditions;
8. For a given redox reaction, write the expression for the Nernst equation and use the equation to calculate the voltage E of a cell, given E^O , and the concentrations of all other species;
9. Summarize the relationship of the values of E^O , ΔG and the equilibrium constant, K to reaction spontaneity and equilibrium, and when given the value for one of these parameters, be able to determine the value of the other two.

E. The student will be able to apply the concepts of ionic equilibrium and solubility in solving problems.

1. The solubility product constant (K_{sp}) given the concentration in moles/liter, grams/liter, or pH;

2. The solubility of a compound in moles/liter and grams/liter given K_{sp};

3. The concentration of ions necessary to start precipitation and the ions remaining in solution after precipitation given the K_{sp} and the fact that precipitation will occur when the molar concentration exceeds the K_{sp};

4. To determine if precipitation will occur in simultaneous equilibrium given the K_{sp}'s and the concentration of solutions;

5. The H_3O^+ and OH^- ion concentration, the pH and pOH in strong acids and bases, given the concentration of the solution using the K_w of water;

6. The ionization constant of a weak monoprotic acid (K_a) or a weak base (K_b) given the percent ionization and the concentration of the acid or base;

7. The K_a and K_b given the pH or pOH;

8. The concentration of all the species in a weak monoprotic acid or weak base, given the K_a or K_b and using the quadratic formula, if necessary;

9. The percent ionization of a weak monoprotic acid or weak base, given the concentration and K_a or K_b;

10. The H_3O^+ ion concentration and the pH of a solution given the concentration of the salt and the acid using the Henderson-Hasselbach equation - illustrating the common ion effect in calculations;

11. The H_3O^+ ion and pH change in a buffer solution to which a strong acid or base has been added, given the concentration of the acid and salt in the buffer solution and the strong acid or base added;

12. The H_3O^+ ion concentration in the preparation of a buffer solution, given the volume and concentration of salt and acid used and the K_a;

13. The concentrations of all species in a poly-protic acid ,given the K_a for all steps;

14. The hydrolysis constant K_b of a salt ,given the K_a or K_b;

15. The $[OH^-]$, pH and the percent hydrolysis of a salt of the strong base and weak acid ,given the concentration of the salt and K_a;

16. The $[H_3O^+]$, the pH, and the percent hydrolysis of a salt of a strong acid and weak base, given the concentration of the salt and K_b;

17. The pH and the percent hydrolysis, given the concentration of the metal salt, the hydrolytic constant, and using the equation for the hydrolysis of the salt in calculations;

18. State three ways to dissolve a precipitates, giving equations for the three methods.

F. The student will apply the concepts of acid-base neutralization reactions to make

and interpret titration curve.

1. Plot titration curves and determine the equivalence point of acid/base titrations;
2. Choose the proper pH range of indicators to be used in titrations of acids and bases using the titration curves for: strong acid/strong base; weak acid/strong base; strong acid/weak base; weak acids/strong bases.

G. The student will be able to apply the concepts of coordination compounds to better understand the complexes formed by metal ions.

1. Define the following terms in coordination chemistry:

   a. coordination chemistry
   b. Ligands
   c. Coordination number
   d. Donor atoms
   e. Bidentate
   f. Coordination sphere

2. Name coordination compounds using IUPAC rules of nomenclature for coordination compounds.

3. Name and identify the four types of structural isomers.
4. Name two types of stereoisomers.
5. Indicate the type of hybridization for specific metal ions with a coordination number of six, using the electron configuration of the ion.
6. Indicate whether it is an inner or outer "d" orbital after the hybridization of the specific ion.
7. State the proposition of the Crystal Field Theory.
8. Give two names for the site of "d" orbitals after the split.
9. Indicate whether a complex is high spin or low spin using field strength of the ligand and the electronic configuration of the "d" orbitals to make the determination.
10. Calculate the Crystal Field Stabilization Energy (CSFE) and relate the CSFE to the stability of the complex.

H. The student will be able to apply the concepts of nuclear chemistry.

1. Characterize the three major types of radiation observed in natural radioactive decay;
2. Write a balanced equation for a nuclear reaction;

3. Decide whether a particular radioactive isotope will decay by alpha, beta, positron emission or by electron capture;
4. Calculate the binding energy for a particular isotope and understand the relationship between binding energy and nuclear stability;
5. Perform kinetic calculations involving half-life and the time required for an isotope to decay to a particular activity;
6. Describe nuclear fission and nuclear fusion;
7. Relate some uses of radioisotopes.

Laboratory Topics

A. Semimicro Qualitative Analysis of Cations
B. Preparation and Investigation of Voltaic Cells

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