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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

Introductory Chemistry, An Active Learning Approach. Mark S. Cracolice . 5th edition. Brooks/Cole.

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).