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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:

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

IV. Textbook

General Chemistry. Ebbing & Gammon. 10th edition. Brooks/Cole.

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 corre- sponding 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

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).