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SAULT COLLEGE OF APPLIED ARTS & TECHNOLOGY

SAULT STE. MARIE, ONTARIO

COURSE OUTLINE

Course Title: PRINCIPLES OF CHEMISTRY II THEORY & LAB

Code No.: CHM 218-5

Program: WATER RESOURCES II & PULP & PAPER II

Semester: II

Date: MAY, 1988

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New: ______ Revision: x

APPROVED: ______________________ August 2/88
Chairperson Date

CALENDAR DESCRIPTION

PRINCIPLES OF CHEMISTRY II — Course Name

CHM 218-5 — Course Number

PHILISOPHY/GOALS:

CHM 218-5 is a continuation of CHM 104-4 from semester 1. The major emphasis in the lab is on Quantitative Analysis; the student is expected to analyze a variety of samples and arrive at satisfactory results. The theory component of the course deals with the following concepts: solution chemistry, chemical calculations, K_{sp}, K_{eq}, K_a, K_b, K_w, acid-base chemistry, pH, H^+, pOH, OH^-, % ionization of weak acids and bases.

CHM 218 serves as a prerequisite for CHM 230-33 (Water Chemistry) and Pulp and Paper - PPE 220-4 (Pulp Testing II).

TEXTBOOK(S):

Ebbing, Darrell D., General Chemistry, 2nd ed. Houghton Mifflin Co., 1984.

LAB MANUAL:

Lab Experiments for CHM 218 - Sault College, Heggart & Korrey.

METHOD OF GRADING:

Grades:
90% > - A+
80% > - A
70% > - B
60% > - C

EVALUATION:

The final grade is arrived at by totalling the theory marks (50%) and the lab marks (50%).

The lab mark is the sum of all marks awarded for the analysis plus the written report for each of the five experiments. The analysis is graded on accuracy and precision. The report is graded on format, content, and neatness.

The theory mark is the sum of all tests, assignments, mid-term and final examinations.

EVALUATION - Continued

Term Test/Quizzes/Assignments/Final Exam	50 marks
Lab Work	50 marks
	100 marks

Assignments are due on the date specified. Late assignments will not be accepted so it is critical that you submit as much of the assignment as possible on the due date. Lab reports are due one week from completion of the lab. Late labs will be downgraded 10% per week.

ATTENDANCE:

Your grade will be greatly affected by attendance at scheduled classes and labs. 85% is required at all theory classes while 100% is needed for all labs. Serious illness (doctor's care) is the only valid excuse.

EXEMPTION:

The theory grade is the sum of all test and assignments. Tests will include all work up to the time of each test. All students having 70% or more on term work and mid-term exam are exempt from the final exam which will cover the whole course and counts 50% of the theory grade. The final exam will be held during the exam week at the end of the semester.

UNIT I: SOLUTIONS

1 Types of Solutions
Gaseous Solutions
Liquid Solutions
Solid Solutions

2 Ways of Expressing Concentration
Mass
Percentage of Solute
Conversion of Concentration Units

3 Equivalents and Normality

IONS IN SOLUTION; IONIC EQUATIONS

Continuation of solution chemistry problems.

4-1 Electrolytes
A note about the Hydrogen Ion
Introduction to Chemical Equilibrium
Strong and Weak Electrolytes

4-2 Ionic Equations

4-3 Types of Reactions

METATHESIS REACTIONS

4-4 Solubility and Precipitation
Solubility Rules
Precipitation Reactions

5 Reactions of Acids, Bases and Salts
Neutralization
Reactions of Salts
Formation of a Gas

6 Introduction to Oxidation-Reduction Reactions
Terminology
Understanding oxidation-Reduction Equations

7-1 Properties of Water
Hydrogen Bonding and the Physical
Properties of Water
Chemical Properties of Water

UNIT II: CHEMICAL EQUILIBRIUM

DESCRIBING CHEMICAL EQUILIBRIUM

1 Chemical Equilibrium - A Dynamic Equilibrium

2 The Equilibrium Constant
Definition of the Equilibrium Constant K_c
Obtaining Equilibrium Constants for Reactions
The Equilibrium Constant K_p

3 Heterogeneous Equilibria

USING AN EQUILIBRIUM CONSTANT

4 Qualitatively Interpreting an Equilibrium Constant

5 Predicting the Direction of Reaction

6 Calculating Equilibrium Concentrations

CHANGING THE REACTION CONDITIONS AND THE APPLICATION OF LeCHATELIER'S PRINCIPLE

7 Adding a Catalyst

8 Removing or Adding Reactants or Products

9 Changing the Pressure and Temperature
Effect of Pressure Change
Effect of Temperature Change
Choosing the Optimum Conditions for Reaction

UNIT III: ACID-BASE CONCEPTS

1 Arrhenius Concept of Acids and Bases

2 Self-Ionization of Water

3 The pH of a Solution

UNIT III - Continued

4 Bronsted-Lowry Concept of Acids and Bases

5 Relative Strengths of Acids and Bases

6 Molecular Structure and Acid Strength

7 Acid-Base Properties of Salt Solutions

8 Lewis Concept of Acids and Bases

UNIT IV: ACID-BASE EQUILIBRIA

SOLUTIONS OF A WEAK ACID OR BASE OR SALT

1 Acid Ionization Equilibria
Experimental Determination of K_a
Calculations from K_a

2 Polyprotic Acids

3 Base Ionization Equilibria

4 Hydrolysis

SOLUTIONS OF A WEAK ACID OR BASE WITH ANOTHER SOLUTE

5 Common-Ion Effect

6 Buffers 611

7 Acid-Base Titration Curves
Titration of a Strong Acid by a Strong Base
Titration of a Weak Acid by a Strong Base
Titration of a Weak Base by a Strong Acid

UNIT V: SOLUBILITY AND COMPLEX-ION EQUILIBRIA

SOLUBILITY EQUILIBRIA

1 The Solubility Product Constant

2 Solubility and the Common-Ion Effect

3 Precipitation Calculations
Criterion for Precipitation
Completeness of Precipitation
Fractional Precipitation

LABORATORY WORK

The student will complete five of the experiments designated for this course in the allotted time. The following experiments are required: (#1 to 4 and 3 or 5)

1. Titration of Acids and Bases - standardization of NaOH, and determination of unknown KHP.
2. Gravimetric Cl^- - Cl^- in a known (NH_4Cl) plus Cl^- in an unknown.
3. Volumetric Cl^- - Cl^- in a known (NaCl) and in unknown (use same unknown as Exp. #2)
4. Gravimetric Ni - use organic precipitant DMG.
5. Determination of sulphate in a known (Na_2SO_4) and an unknown.
6. Hardness of water - Volumetric determination using ED7A.

In addition to the above the student will be able to subject his results to statistical analysis and determine:

1. Precision
2. Relative error
3. Average deviation
4. Standard deviation
5. Whether a result should be excluded by the 2.5d rule, 4.0d rule and by the Q test.

CHM 218-5 OBJECTIVES

Unit I: "Reactions in Aqueous Solution"
Ref: Ebbing - General Chemistry

Upon completion of this unit, the student should be able to:

1. Using solubility rules, decide whether two soluble ionic compounds will or will not react to form a precipitate. If they will, write the net ionic equation.

2. Write the moleculous equation, and then the net ionic equation for the neutralization of an acid and a base.

3. Given a reaction between a carbonate, sulphide, or sulphite, and a strong acid, or the reaction between an ammonium compound and a strong base, write the molecular and ionic equations.

4. Write a skeleton equation given as balanced oxidation-reduction equation. Label the oxidizing and reducing agents, the oxidized and reduced species, and the oxidation and reduction parts of the equation. Comment on the reaction by referring to the commonly observed oxidation states.

5. Given an oxidation-reduction equation (an unbalanced or a skeleton equation), complete and balance it by the half-reaction method and/or the oxidation number method.

6. Explain how bond polarity in water molecules contributes to hydrogen bonding between molecules.

Unit II: Chemical Equilibrium
Ref: Chapter 16 - Ebbing - General Chemistry

After completion of this unit, the student should be able to :

1. Given the starting amounts of reactants and the amount of one substance at equilibrium, find the equilibrium composition (Example 16.1).

2. Given the chemical equation, write the equilibrium-constant expression (Examples 16.2 and 16.4).

3. Given the equilibrium composition, find K_c (Example 16.3).

4. Given the concentrations of substances in a reaction mixture, predict the direction of reaction (Example 16.5).

5. Given K_c and all concentrations of substances but one in an equilibrium mixture, calculate the concentration of this one substance (Example 16.6).

6. Given the starting composition and K_c of a reaction mixture, calculate the equilibrium composition (Examples 16.7 and 16.8).

7. Given a reaction, use LeChatelier's principle to decide the effect of adding or removing a substance (Example 16.9), changing the pressure (Example 16.10), or changing the temperature (Example 16.11).

Unit III: Acid-Base Concepts
Ref: Chapter 17 - Ebbing

After completing this unit, the student should be able to:

1. Given the concentration of hydroxide ion (or concentration of strong base), calculate the hydrogen-ion concentration (Example 17.1).

2. Given the hydrogen-ion concentration (or concentration of strong acid), calculate the pH (Example 17.2); given the pH, calculate the hydrogen-ion concentration (Example 17.3).

3. Given a proton-transfer reaction, label the Brønsted-Lowry acids and bases, and name the conjugate acid-base pairs (Example 17.4).

4. Given a Brønstead-Lowry acid-base reaction and the relative strengths of acids (or bases), decide whether reactants or products are favored at equilibrium (Example 17.5).

5. Decide whether an aqueous solution of a given salt will be acidic, basic, or neutral (Example 17.6).

6. Given a reaction involving the donation of an electron pair, identify the Lewis acid and Lewis base (Example 17.7).

Unit IV: Acid-Base Equilibria
Ref: Chapter 18 - Ebbing

After completion of this chapter, the student should be able to:

1. Given the molarity and pH of a solution of a weak acid, calculate the acid ionization constant, K_a (Example 18.1). Given K_a, calculate the hydrogen-ion concentration and pH of a solution of a weak acid of known molarity (Examples 18.2 and 18.3).

2. Given the molarity and pH of a solution of a weak base, calculate the base ionization constant, K_b (similar to Example 18.1). Given K_b, calculate the hydrogen-ion concentration and pH of a solution of a weak base of known molarity (Example 18.5).

3. Calculate the K_a for a cation or the K_b for an anion from the ionization constant of the conjugate base or acid (Example 18.6).

4. Given the concentrations of weak acid and strong acid in a solution, calculate the degree of ionization and concentration of the anion of the weak acid (Example 18.8).

5. Given the K_a and the concentrations of weak acid and its salt in a solution, calculate the pH (Example 18.9). Given the K_b and the concentrations of weak base and its salt in a solution, calculate the pH (Example 18.10).

6. Calculate the pH of a given volume of buffer solution (given the concentrations of conjugate acid and base in the buffer) to which a specified amount of strong acid or base is added (Example 18.11).

7. Calculate the pH during the titration of a strong acid and strong base, given the volumes and concentrations of the acid and base (Example 18.12).

Unit V: Solubility Equilibria
Ref: Chapter 19 - Ebbing

After completion of this unit, the student should be able to:

1. Write the solubility product expression for a given ionic compound (Example 19.1).

2. Given the solubility of a slightly soluble ionic compound, calculate K_{sp} (Examples 19.2 and 19.3). Given K_{sp}, calculate the solubility of an ionic compound (Example 19.4).

3. Given the solubility product constant, calculate the molar solubility of a slightly soluble ionic compound in a solution containing a common ion (Examples 19.5 and 19.6).

4. Given the concentrations of ions originally in solution, determine if a precipitate is expected to form (Example 19.7). Determine if a precipitate is expected to form when two solutions of known volume and molarity are mixed (Example 19.8). For both problems, you will need the solubility product constant.

5. Calculate the concentration and percentage of an ion remaining after the corresponding ionic compound precipitates from a solution of known concentrations of ions (Example 19.9). K_{sp} is required.