

COURSE NUMBER: CH1121

COURSE TITLE: Chemistry

COURSE DESCRIPTION:

This course will develop further the fundamental concepts of chemistry, with emphasis on those relevant to the chemistry of materials and to the processes of polymer chemistry, thermochemistry, chemical reaction rates and equilibrium, electrochemistry, and metals and alloys.

PREREQUISITES: CH1120 – Chemistry

CO-REQUISITES: None

CREDIT VALUE: Four (4)

COURSE HOURS PER WEEK: Three (3)

LAB HOURS PER WEEK: Two (2)

SUGGESTED TEXT: None

LEARNING RESOURCES: To be determined by instructor

MAJOR TOPICS:

- 1.0 Polymer Chemistry
- 2.0 Thermochemistry
- 3.0 Rates and Chemical Equilibrium
- 4.0 Acid/Base and Solubility Equilibrium
- 5.0 Electrochemistry
- 6.0 Metals and Alloys

LEARNING OBJECTIVES:

1.0 Polymer Chemistry

- 1.1 The Bonding of Carbon
 - 1.1.1 Describe the covalent nature of carbon
 - 1.1.2 Describe the breaking of carbon-carbon bonds to form radicals
 - 1.1.3 Define and describe polar covalent bonds using carbon molecules

- 1.2 Hydrocarbons
 - 1.2.1 Define and describe homologous series
 - 1.2.2 Name straight chain alkanes from C₁ to C₁₀.
 - 1.2.3 Describe multiple covalent bonds using carbon compounds
- 1.3 Functional Groups
 - 1.3.1 Given the structural formula, classify according to functional group: part of the molecular framework (alkanes, alkenes, alkynes, aromatics), containing oxygen (alcohols, carboxylic acids, esters, ketones, aldehydes, ethers), containing nitrogen (amines, *amides*)
- 1.4 Simple Reactions of Organic Molecules
 - 1.4.1 Predict the products of the following reactions:
 - 1.4.1.1 Carboxylic acid + NaOH
 - 1.4.1.2 Carboxylic acid + alcohol
 - 1.4.1.3 Amine + HCl
 - 1.4.1.4 Carboxylic acid + amine
- 1.5 Structure of Polymers
 - 1.5.1 Define polymer
 - 1.5.2 Define monomer
 - 1.5.3 Define repeating unit
 - 1.5.4 List a couple of industrial examples of monomers
 - 1.5.5 Describe how the identity of monomers and conditions for polymerization determine the properties of the polymer
 - 1.5.6 Define homopolymer
 - 1.5.7 Define copolymer
 - 1.5.8 Define linear polymer
 - 1.5.9 Define branched polymer
 - 1.5.10 Define network polymer
 - 1.5.11 Define cross-linking
 - 1.5.12 Explain how cross-linking affects the strength of polymers
- 1.6 Types of Polymers
 - 1.6.1 Define and give examples of plastics
 - 1.6.2 Define and give examples of thermoplastic polymers
 - 1.6.3 Define and give examples of thermosetting polymers
 - 1.6.4 Give examples to differentiate between thermoplastic and thermosetting polymers
- 1.7 Bonding in Polymers
 - 1.7.1 List some basic properties of polymers which make them commercially useful
 - 1.7.2 Explain why intermolecular forces contribute useful mechanical properties

- to polymers
- 1.7.3 Explain the importance of dispersion forces to the mechanical properties of polymers
- 1.7.4 Explain the importance of functional groups to the strength of a polymer
- 1.7.5 Explain the relationship between molecular weight and intermolecular forces
- 1.7.6 Explain how intermolecular forces are related to distance of separation
- 1.8 Formation of Polymers
 - 1.8.1 Describe the process of addition polymerization in the formation of polypropylene, polystyrene and PVC
 - 1.8.2 Describe the process of copolymerization in the formation of SB rubber
 - 1.8.3 Describe the process of condensation polymerization using polyester, acetate, rayon and nylon

2.0 Thermochemistry

- 2.1 Energy
 - 2.1.1 Define energy
 - 2.1.2 Define various forms of energy: radiant energy, thermal energy, chemical energy
 - 2.1.3 State the Law of Conservation of Energy
- 2.2 Heat Capacity
 - 2.2.1 Define heat
 - 2.2.2 Define thermochemistry
 - 2.2.3 Define system and surroundings
 - 2.2.4 Describe an open system and a closed system
 - 2.2.5 Define and give examples of exothermic reactions and endothermic reactions
- 2.3 Enthalpy
 - 2.3.1 Define enthalpy and enthalpy of reaction, ΔH
 - 2.3.2 Describe the relationship between the sign ΔH (+/-) and the processes (physical and chemical changes) occurring
 - 2.3.3 Use thermochemical equations to predict quantities of energy consumed or produced in chemical and physical processes
- 2.4 Calorimetry
 - 2.4.1 Define calorimetry
 - 2.4.2 Define specific heat
 - 2.4.3 Define heat capacity
 - 2.4.4 Calculate, using the relationship $q = mc\Delta T$, either q , m , c or ΔT , given appropriate data for the other variables (for processes involving changes of state without change of phase)

- 2.4.5 Differentiate between a constant volume bomb calorimeter and a constant pressure (coffee cup) calorimeter
- 2.4.6 Determine heat changes for processes occurring in coffee cup calorimeters
- 2.4.7 Calculate enthalpy changes, ΔH , for processes occurring in coffee cup calorimeters
- 2.5 Energy Diagrams and Hess's Law
 - 2.5.1 Construct enthalpy (energy) diagrams to represent changes for which the appropriate thermochemical data is supplied
 - 2.5.2 Label appropriately, using the terms exothermic, endothermic, heat of reaction, reactants, products, and the +/- sign convention, a set of changes for which the differences in enthalpy between the initial and final states are given, either in the form of data or enthalpy diagrams
 - 2.5.3 Distinguish between the magnitude of enthalpy changes involved in changes of phase, and chemical reactions
 - 2.5.4 Use Hess's Law to calculate heats of reactions given a set of thermochemical data

3.0 Rates and Chemical Equilibrium

- 3.1 Rates of Reaction
 - 3.1.1 Concept of Reaction Rate
 - 3.1.1.1 Define reaction rate as the measure of how fast a reactant is used up or how fast a product is formed
 - 3.1.1.2 State that the rate of a reaction decreases over time
 - 3.1.1.3 Calculate the average rate of a reaction
 - 3.1.2 Factors Influencing Reaction Rate
 - 3.1.2.1 State, that for a solid, the rate of reaction is dependent on its degree of subdivision
 - 3.1.2.2 State that for a solid, the rate of reaction in solution increases with concentration of reactants
 - 3.1.2.3 State that the rate of reaction of gaseous reactants increases with increased pressure
 - 3.1.2.4 State that the rate of reaction generally increases considerably with increasing temperature
 - 3.1.2.5 List and describe the factors (temperature, particle size, quantity of reactants, retarders and accelerators) which influence the rate of concrete formation
 - 3.1.3 Collision Theory
 - 3.1.3.1 Discuss using collision theory how the rate of a reaction changes with changes in reactant concentrations and with temperature
 - 3.1.4 Activation Energy and Transition States

- 3.1.4.1 Define activation energy as the difference between the initial energy of the reactants and the transition state
 - 3.1.4.2 Define transition state as the energy maximum at which the activated complex is formed
 - 3.1.4.3 Explain that activation energy, and orientation of molecules at collision, determine whether or not a collision of molecules will result in a chemical reaction
 - 3.1.4.4 Sketch a fully labeled energy diagram from data provided
- 3.1.5 Catalysis
 - 3.1.5.1 Define a catalyst as a substance that influences the rate of a reaction whilst itself remaining chemically unchanged at the end of the reaction
 - 3.1.5.2 Explain that a catalyst works by lowering the activation energy of a reaction
 - 3.1.5.3 State examples of catalysts of industrial and biological importance
- 3.1.6 Rate Law
 - 3.1.6.1 Define rate law (rate equation) as a representation of the dependence of the rate of reaction on the concentration of the species
 - 3.1.6.2 Define rate constant
 - 3.1.6.3 Define order and reaction order (0,1,2)
 - 3.1.6.4 Write the rate law given the order of the reactants
 - 3.1.6.5 Calculate the rate of reaction given the rate law, k , and concentration of reactants using orders of 0,1,2
- 3.1.7 Half Life
 - 3.1.7.1 Define half-life
 - 3.1.7.2 Calculate the amount of reactant left in a reaction up to 4 half-lives
- 3.2 Chemical Equilibrium
 - 3.2.1 Reversible Reactions
 - 3.2.1.1 Explain the term reversible reaction
 - 3.2.2 Equilibrium Constant
 - 3.2.2.1 Write an equation for the equilibrium constant, K_c , given the equation for the reaction concerned, using the symbol \rightleftharpoons
 - 3.2.2.2 Calculate K_c for a given reaction given equilibrium concentrations of the reactants and products
 - 3.2.2.3 Given the value of K^c , calculate the value of K_c for the reverse reaction
 - 3.2.2.4 Write the equilibrium equation, K_c , for heterogeneous equilibria

- involving pure solids and liquids
- 3.2.2.5 Write the equilibrium equation, K_p , for a typical all-gaseous reaction
- 3.2.2.6 Using the value of K_c for an equilibrium reaction, indicate whether reactants or products are favored
- 3.2.3 Le Chatelier's Principle
 - 3.2.3.1 Predict the effect of alteration in initial concentrations and pressures on the composition of the equilibrium mixture using Le Chatelier's Principle
- 3.2.4 Effect of Temperature on Equilibrium
 - 3.2.4.1 State that K_c and K_p vary with temperature
 - 3.2.4.2 Predict, using Le Chatelier's Principle, the effect of a change in temperature on equilibrium composition, given the sign of ΔH for the reaction concerned
 - 3.2.4.3 Explain, in terms of its equal effects on the rates of forward and reverse reactions, that a catalyst does not affect the composition of an equilibrium mixture

4.0 Acid/Base and Solubility Equilibrium

- 4.1 Definition
 - 4.1.1 Define Bronsted-Lowry acid and base
 - 4.1.2 Write the conjugate acid of a given base or a conjugate base of a given acid
 - 4.1.3 For the equation of a weak acid equilibrium system, identify the conjugate acid/base pairs
- 4.2 pH and Indicators
 - 4.2.1 Describe the self-ionization of water in terms of:

$$\text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq)$$
 - 4.2.2 Describe pH as a numerical scale and recognize what values describe acids and bases as well as the neutral point
 - 4.2.3 Define titration and perform simple acid/base calculations
 - 4.2.4 Describe the action of acids and bases on the color of common indicators
 - 4.2.5 State the equation for and the value of the ionic product of water, K_w , and hence deduce the values of $[\text{H}^+_{(aq)}]$ and $[\text{OH}^-_{(aq)}]$ in pure water
 - 4.2.6 Discuss the range of values for $[\text{H}^+_{(aq)}]$ and $[\text{OH}^-_{(aq)}]$ found in aqueous solutions of acids and bases
 - 4.2.7 State, that at a constant temperature, K_w remains constant in aqueous solutions of acids and bases
 - 4.2.8 Calculate values of pH and pOH from the relationships $\text{pH} = -\log[\text{H}^+_{(aq)}]$ and $\text{pOH} = -\log[\text{OH}^-_{(aq)}]$

- 4.2.9 Define acid rain
- 4.2.10 Describe the formation of acid rain
- 4.2.11 Describe the effects of acid rain on buildings and lakes
- 4.3 Strong/Weak Acids and Bases
 - 4.3.1 Identify common strong and weak acids and bases
 - 4.3.2 Differentiate between strong and weak acids and bases with regards to the degree of ionization
 - 4.3.3 Write the equilibrium equation, K_a , for a typical acid/base reaction
 - 4.3.4 Predict the relative strengths of weak acids from K_a values
- 4.4 Buffers
 - 4.4.1 Define buffer
 - 4.4.2 Describe the composition of a simple buffer
 - 4.4.3 Explain qualitatively the response of a buffer to the addition of small quantities of acid or base
 - 4.4.4 State examples of common buffering systems
- 4.5 Solubility Equilibrium
 - 4.5.1 Define K_{sp}
 - 4.5.2 Write the equilibrium equation, K_{sp} , for a typical solution of a solid dissolved in a liquid
 - 4.5.3 Calculate the molar solubility of a substance given K_{sp}

5.0 Electrochemistry

- 5.1 Balance Redox Reactions
 - 5.1.1 Write oxidation and reduction half reactions for redox reactions
 - 5.1.2 Balance redox equations using half-reaction equations/Ion-Electron method in acid and basic medium
- 5.2 Electrochemical Cells
 - 5.2.1 How Electrochemical Cells Work
 - 5.2.1.1 Recognize that spontaneous redox reactions are a source of electrical energy
 - 5.2.1.2 Draw a simple electrochemical cell with reference to zinc in aqueous copper sulphate using a porous partition or a salt bridge
 - 5.2.1.3 Define anode and cathode
 - 5.2.1.4 Describe the external flow of electrons from anode to cathode
 - 5.2.1.5 State that oxidation occurs at the anode and reduction occurs at the cathode
 - 5.2.1.6 Write the oxidation and reduction half-reactions
 - 5.2.1.7 Write the overall reaction for the cell
 - 5.2.1.8 Describe the use of a salt bridge or porous material (cup, plates, etc.)

- 5.2.1.9 Describe the flow of ions through the salt bridge
- 5.2.2 Standard Reduction Potentials
 - 5.2.2.1 Describe the construction of electrochemical cells using gaseous reactants and platinum electrodes
 - 5.2.2.2 Draw cell diagrams for electrochemical cells using gaseous reactants and platinum electrodes
 - 5.2.2.3 State that any standard half-cell potential can be determined by combining it with a hydrogen electrode
 - 5.2.2.4 Define standard reduction potential
 - 5.2.2.5 State that standard reduction potentials and standard oxidation potentials are equal in magnitude but opposite in sign
 - 5.2.2.6 Use a table of reduction potentials to calculate E°_{cell}
 - 5.2.2.7 Describe the construction of the hydrogen electrode and how it works
- 5.2.3 Batteries
 - 5.2.3.1 Distinguish between primary and secondary batteries
 - 5.2.3.2 Describe the features of:
 - 5.2.3.2.1 Dry cell batteries
 - 5.2.3.2.2 Alkaline batteries
 - 5.2.3.2.3 Lead-storage batteries
 - 5.2.3.2.4 Nickel-cadmium batteries
 - 5.2.3.2.5 Fuel cell batteries
 - 5.2.3.2.6 Lithium batteries
- 5.3 Electrolysis
 - 5.3.1 Electrolysis of Aqueous Solutions
 - 5.3.1.1 Sketch a simple electrolytic cell
 - 5.3.1.2 Compare and contrast electrochemical and electrolytic cells
 - 5.3.2 Predicting Electrode Reactions
 - 5.3.2.1 Predict products of electrolysis with inert electrodes
 - 5.3.3 Electroplating
 - 5.3.3.1 Describe electroplating and give examples
 - 5.3.4 Faraday's Laws
 - 5.3.4.1 Use Faraday's Law to relate time, mass and amount of work done
- 5.4 Industrial Applications
 - 5.4.1 Metallurgy
 - 5.4.1.1 State that metals occur in nature as minerals
 - 5.4.1.2 State that mineral deposits from which metals can be produced economically are called ores

- 5.4.1.3 Explain the industrial process of roasting using zinc, nickel or iron as an example
- 5.4.1.4 Explain the industrial process of reduction of an ore to a free metal using the reduction of Fe_2O_3 to Fe
- 5.4.1.5 Describe the Industrial Hall Process of producing aluminum from bauxite ($\text{Al}_2\text{O}_3 \cdot x\text{H}_2\text{O}$)
- 5.4.1.6 Describe the production of copper from copper ore

6.0 Metals and Alloys

6.1 Types of Solids

- 6.1.1 Define crystalline solid
- 6.1.2 Define amorphous solid
- 6.1.3 State that solids can be either crystalline or amorphous
- 6.1.4 State that crystalline solids can be categorized as ionic, molecular, covalent network or metallic
- 6.1.5 Differentiate between the properties of ionic, molecular, covalent network or metallic solids

6.2 Network Solids

- 6.2.1 Recognize and sketch common covalent giant lattice structures such as diamond, graphite and silica
- 6.2.2 Differentiate between the properties of diamond and graphite based on their structure
- 6.2.3 Differentiate between crystalline forms of SiO_2 and amorphous forms of SiO_2 (quartz and glass) in terms of their properties
- 6.2.4 Describe how the color of glass can be influenced by the addition of certain metal oxides
- 6.2.5 Describe how the strength of glass can be influenced by the addition of certain chemical substances
- 6.2.6 Define superconductor
- 6.2.7 Describe how network solids can be used as superconductors

6.3 Metallic Bonding

- 6.3.1 “Sea of Electrons” Model
 - 6.3.1.1 Explain the bonding of metals in terms of the “sea of electrons” model
- 6.3.2 Alloys
 - 6.3.2.1 Explain the difference between interstitial and substitutional alloys
 - 6.3.2.2 Explain why alloys are usually harder and stronger than pure metals

6.4 Iron and Steel

- 6.4.1 The Nature of Iron
 - 6.4.1.1 State the abundance of iron in nature
 - 6.4.1.2 Explain why pure iron is of limited industrial use
- 6.4.2 Corrosion of Iron
 - 6.4.2.1 State that metals corrode because they are easily oxidized by the atmosphere
 - 6.4.2.2 Describe the process of iron corrosion in terms of anode and cathode half-reactions, and the overall reaction
 - 6.4.2.3 State the factors that promote metal corrosion
- 6.4.3 Inhibiting Corrosion of Iron
 - 6.4.3.1 Describe methods of inhibiting corrosion with reference to metal coatings, alloys, and cathodic protection
- 6.4.4 The Composition of Steel
 - 6.4.4.1 Describe the chemical composition of steel as an alloy of iron and carbon
 - 6.4.4.2 Differentiate between carbon steel and alloy steel
 - 6.4.4.3 Describe and differentiate between chromium stainless steel and nickel stainless steel
 - 6.4.4.4 Describe the industrial uses of steel
 - 6.4.4.5 Define galvanized steel
- 6.5 Nonferrous Metals
 - 6.5.1 Aluminum
 - 6.5.1.1 Explain the properties of aluminum that make it a good engineering material
 - 6.5.1.2 Describe the industrial uses of aluminum alloys
 - 6.5.2 Copper, Brass, Bronze
 - 6.5.2.1 Explain the advantages of copper as an industrial material
 - 6.5.2.2 Explain why copper for electrical applications must be pure
 - 6.5.2.3 Explain the advantages of using alloys of copper (bronze, brass) in industrial materials
 - 6.5.3 Lead, Tin, Zinc
 - 6.5.3.1 Describe the uses of lead, tin and zinc alloys

EVALUATION:

In-class Quizzes:	40%
Laboratories:	10%
Final Exam:	50%

**** What would be considered as a “valid documented reason” will be at the discretion of the campus administrator in consultation with the faculty responsible for the two Chemistry courses CH1120 and CH1121.**

REVISION NUMBER: 2 **DATE REVISED:** November 2010