

Expanded Course Objectives: CH116 General Chemistry II
Based on Chemistry 7th Edition, E. Chang
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Chapter 11

Intermolecular Forces and Liquids and Solids

This chapter introduces the student to the concepts that are used to explain the properties of liquids and solids. Upon completion of this chapter, the student should be able to:

1. Characterize the properties of gases, liquids and solids in terms of density, compressibility and motion of molecules.
2. Distinguish between intermolecular and intramolecular forces.
3. Identify and give examples of the following forces:
 - ion – ion
 - ion – dipole
 - ion – induced dipole
 - dipole – dipole
 - dipole – induced dipole
 - induced dipole – induced dipole
 - van der Waals
 - dispersion
 - hydrogen bonding
4. Relate polarizability and intermolecular forces.
5. Suggest why H₂O, HF and NH₃ do not follow the expected trend as shown by the plot in Figure 11.7
6. Use the concepts of intermolecular forces to explain surface tension, capillary action, cohesion, adhesion and viscosity.
7. Describe the structure of water and relate it to why ice is less dense than water.
8. Sketch density vs. temperature curve for water and relate how this plot has major significance for aquatic life.
9. Distinguish between crystalline and amorphous solids and give examples of each.
10. Identify coordination numbers for atoms in simple cubic, face-centered cubic and body-centered cubic structures.
11. Determine the number of atoms contained in a unit cell for simple cubic, face-centered cubic and body-centered cubic structures.
12. Use geometric concepts to relate a , the edge length of a unit cell, to r , the radius of atoms, in simple cubic, face-centered cubic and body-centered cubic structures.
13. Calculate what fraction of a unit cell is occupied by atoms, and show that the face-centered cubic structure has the closest packing for simple cubic, face-centered cubic and body-centered cubic structures.
14. Perform calculations involving crystal structure, density, atomic radius and the number of atoms per unit cell.
15. Describe how x-ray diffraction is used to determine geometric parameters of solids.
16. Characterize ionic, covalent, molecular and metallic crystals including general properties and examples of each.

17. Discuss the following:
 - evaporation
 - condensation
 - molar heat of vaporization
 - molar heat of fusion
 - boiling point
 - melting point
 - supercooling
 - sublimation

- deposition
 - molar heat of sublimation
 - phase changes
 - critical temperature and pressure
18. Use Clausius – Clapeyron equation to solve for vapor pressure at a specific temperature given associated data.
 19. Sketch a typical heating curve and identify various aspects of it.
 20. Use phase diagrams to identify what phase(s) is/are present given specific conditions.

Chapter 12

Physical Properties of Solutions

This chapter brings together several concepts dealing with solutions. The primary emphasis in this chapter is aqueous solutions; however, solid solutions are of major importance especially in the manufacturing of metal alloys such as steel. Upon completion of this chapter, the student should be able to:

1. Use the terms saturated, unsaturated and supersaturated to describe solutions.
2. Distinguish between crystallization and precipitation.
3. Predict relative solubilities given the net dipole moment of solute and solvent.
4. Define, determine and inter-convert between each of the following:
 - molarity
 - percent by mass
 - mole fraction
 - molality
5. Suggest a shortcoming of molarity and explain why molality is a preferred concentration unit under certain conditions.
6. Use the concept of fractional crystallization to show how dissolved solids can be separated.
7. Describe how thermal pollution may effect the oxygen content in lakes or streams.
8. State Henry's law and use it to determine the solubility of gases in liquids.
9. Rationalize why two common materials (NH_3 or CO_2) when dissolved in water do not follow Henry's law.
10. Define colligative properties and give four examples (vapor – pressure lowering, freezing – point lowering, boiling – point elevation and osmotic pressure).
11. Use Raoult's law to find vapor pressures or concentrations of solutions.
12. Describe the apparatus used in fractional distillation.
13. Predict the plot of pressure versus mole fraction for an ideal solution.
14. Rationalize the possibility of either positive or negative deviations from Raoult's law by non-ideal solutions.
15. Perform calculations involving boiling-point elevation, freezing-point depression, K_f , K_b and molality.
16. Use the concepts of osmotic pressure to describe the processes of osmosis and reverse osmosis.
17. Describe the following terms:
 - semi-permeable membrane
 - isotonic
 - hypertonic
 - hypotonic
 - crenation
18. Use the concepts of colligative properties to determine molar mass.
19. Define the van't Hoff factor and demonstrate how it is incorporated into the colligative property equations.
20. Give examples of common types of colloids and describe the dispersing medium and dispersed phase for each.

21. Describe the Tyndall effect.

Identify hydrophilic and hydrophobic colloids and describe the cleansing action of soap.

Chapter 13

Chemical Kinetics

This chapter introduces the student to the concept of rates of chemical reactions. Upon completion of this chapter your students should be able to:

- 1) Distinguish between average and instantaneous rates of chemical reactions.
- 2) Define rate constant.
- 3) Use the concepts of stoichiometry to write reaction rate expressions in terms of the disappearance of reactants and the appearance of products.
- 4) Sketch the rate of reaction versus concentration of reactant for zero and first order reactions.
- 5) Use rate data to determine rate laws and rate constants.
- 6) Show mathematically that

$$\ln [A] = -kt + \ln [A]_0$$

for first order reactions.

- 7) Determine the time required for the concentration of a reactant to change a desired amount given the initial concentration and the rate constant for a first order reaction.
- 8) Show that half-life is independent of initial concentration of the reactant in a first order reaction.
- 9) Use the concept of half-life to determine concentration of reactants over time.
- 10) Show mathematically that

$$t_{1/2} = \frac{1}{k[A]_0}$$

for second order reactions.

- 11) Describe the Collision Theory of Chemical Kinetics using the terms activation energy, activated complex (transition state), potential energy profiles, endothermic and exothermic reactions.
- 12) Use Arrhenius equation to determine the activation energy of a reaction.
- 13) Show that the sum of elementary steps is the overall reaction for a reaction mechanism and that intermediates appear in the reaction mechanism but not in the overall reaction.
- 14) Define molecularity of unimolecular, bimolecular, and termolecular reactions.
- 15) Relate the importance of the rate-determining step in determination of reaction mechanisms.
- 16) Suggest how the use of isotopes can be used to experimentally determine reaction mechanisms.
- 17) Describe what a catalysis does, how it effects activation energy and the difference between homogeneous and heterogeneous catalysis.
- 18) Relate the importance of enzymes as biological catalysts.

Chapter 14

Chemical Equilibrium

This is the first of several chapters that deals with the concepts of chemical equilibrium. The concept of chemical equilibrium seems to be a topic that is difficult for many students. Only after several different applications of the topic do students finally grasp the subject. Upon completion of this chapter your students should be able to:

- 1) Describe chemical equilibrium using the terms forward and reverse reactions and dynamic process.
- 2) Write the equilibrium constant in terms of the equilibrium concentration of products and reactants and their respective stoichiometric coefficients for both homogenous and heterogeneous equilibria.
- 3) Derive the relationship between K_p and K_c .
- 4) Determine equilibrium constant given equilibrium concentration data.
- 5) Show that if a reaction can be expressed as the sum of two or more reactions, the equilibrium constant for the overall reaction is given by the product of the equilibrium constants of the individual reactions.

- 6) Relate equilibrium constant to rate constants from chemical kinetics.
- 7) Describe the relationship between reaction quotient and equilibrium constant and predict the direction a reaction will proceed to reach equilibrium.
- 8) Use the concepts of equilibrium to determine concentration of all species in a solution.
- 9) Use Le Ch&gttelier's Principle to describe how changing concentration, volume, pressure, or temperature will shift the reaction so that an equilibrium will be maintained.
- 10) Describe the effect of a catalyst has on equilibrium concentrations.

Chapter 15

Acids and Bases

This is the first of two chapters dealing with acids and bases. Upon completion of this chapter your students should be able to:

- 1) Compare and contrast Arrhenius, Brønsted, and Lewis acids and bases.
- 2) Describe what is meant by conjugate acid-base pairs and give several examples.
- 3) Use K_w to determine $[H^+]$ and $[OH^-]$ of solutions.
- 4) Discuss the pH scale and calculate pH and pOH given either $[H^+]$ or $[OH^-]$.
- 5) Define strong and weak acids and bases and give several examples of each.
- 6) Relate properties of conjugate acid-base pairs.
- 7) Determine K_a from experimental data.
- 8) Calculate pH, $[H^+]$, weak acid concentration, and conjugate base concentration given K_a and the initial concentration of the weak acid using the quadratic equation or the method of successive approximation as needed.
- 9) Calculate percent ionization for a weak acid.
- 10) Calculate pH, $[OH^-]$, weak base concentration, and conjugate acid concentration given K_b and the initial concentration of the weak base using the quadratic equation or the method of successive approximation as needed.
- 11) Show the relationship between K_a , K_b , and K_w .
- 12) Calculate concentrations of all species present at equilibrium for diprotic and polyprotic acids.
- 13) Relate molecular structure and the strength of acids.
- 14) Predict the relative strengths of oxoacids.
- 15) Describe salt hydrolysis and explain how some salts produce neutral solutions, some acidic solutions and others basic solutions.
- 16) Calculate the pH of salt solutions and determine the percent hydrolysis.
- 17) Describe acid-base properties of oxides and hydroxides.
- 18) Give several examples of Lewis acid-base reactions.

Chapter 16

Acid-Base Equilibria and Solubility Equilibria

Chapter 16 is the third chapter of the three-chapter sequence which deals with equilibrium concepts. Upon completion of this chapter, your student should be able to:

1. Describe the common ion effect as a special case of Le Ch&>telier's principle.
2. Derive the Henderson-Hasselbalch equation.
3. Use the Henderson-Hasselbalch equation to determine the pH of a solution containing a weak acid (weak base) and its conjugate base (conjugate acid).
4. Describe what a buffer solution is and its importance in chemical and biological systems.
5. Calculate the pH of a buffer solution using the Henderson-Hasselbalch equation.
6. Calculate the pH of a buffer solution after the addition of H^+ or OH^- .
7. Describe how to prepare a buffer of a desired pH.
8. Predict the pH profile of a strong acid-strong base titration and calculate the pH at any stage of the titration.

- Predict the pH profile of a strong acid-weak base (or strong base-weak acid) titration and calculate the pH at any stage.
- Distinguish between end point and equivalence point of a titration.
- Describe common acid-base indicators and suggest the correct method of selection for a specific titration.
- Use the concepts of equilibrium to relate ion product, Q , with K_{sp} to predict if a solution is unsaturated, saturated or supersaturated.
- Distinguish between solubility product, molar solubility and solubility.
- Use the concept of fractional precipitation to predict concentration of insoluble ions in a solution.
- Calculate the solubility of an in soluble ion when a common ion is present.
- Describe how changing pH can effect solubility.
- Use the concept of equilibrium, formation constant (K_f) and complex ion formation to predict ion concentration in solutions.
- Describe how solubility product principle is used in qualitative analysis.

Chapter 18

Entropy, Free Energy, and Equilibrium

This chapter covers the concepts of entropy and free energy as part of thermodynamics and relates these concepts to equilibrium. Upon completion of this chapter, your students should be able to:

- State the three laws of thermodynamics.
- Provide several examples of spontaneous processes.
- Give several examples of endothermic spontaneous processes.
- Define entropy using the terms disorder or randomness.
- Justify why entropy of one mole of steam is greater then the entropy of one mole of water.
- Predict the sign on the change in entropy (ΔS) for common processes.
- Describe what is meant by $\Delta S_{universe}$, ΔS_{system} , and $\Delta S_{surroundings}$.
- Relate $\Delta S_{universe}$ for spontaneous processes and for processes at equilibrium.
- Use thermodynamic tables to determine $\Delta S_{reaction}$.
- Mathematically relate $\Delta S_{surroundings}$ to ΔH_{system} and the temperature.
- Express Gibbs free energy in terms of H, T and S.
- Predict if a reaction is spontaneous, spontaneous in the reverse direction or at equilibrium from the sign on ΔG .
- Use thermodynamic tables to determine ΔG of a reaction.
- Rationalize the direction of a spontaneous reaction given ΔH , ΔS and T.
- Calculate ΔS for phase changes given ΔH_{fusion} or $\Delta H_{vaporization}$ and the melting and boiling point temperatures.
- Mathematically describe ΔG in terms of ΔG° , the gas constant, temperature and the reaction quotient.
- Relate ΔG° and K, equilibrium constant.
- Calculate K, equilibrium constant, using data from thermodynamic tables.
- Describe the concept of coupled reactions and discuss the importance of this concept in biological systems.

Chapter 19

Electrochemistry

This chapter uses the concepts developed in thermodynamics to describe the interconversion of chemical energy to electrical energy. Upon completion of this chapter, your students should be able to:

- Describe the concept of redox reactions using such terms as reduction, oxidation, reducing agents, oxidizing agents and oxidation numbers.
- Balance redox equations in both acidic and basic solutions.
- Describe an electrochemical cell using such terms as oxidation, reduction, galvanic cell, electrolytic cell, anode, cathode, half-cell reactions, salt bridge, cell voltage and emf.
- Use standard cell diagrams to describe an electrochemical cell.
- Use standard reduction potentials to predict the emf of a cell.
- Use standard reduction potentials to justify the activity series for metals and the activity series for halogens as

described in section 4.4 of this textbook.

7. Define what SHE means and relate its significance to the standard reduction potential table.
8. Predict the outcome of reactions based on standard reduction potentials and $E^{\circ}_{\text{overall}}$.
9. Relate what a faraday is in terms of coulombs and number of electrons.
10. Mathematically relate ΔG° , K and E°_{cell} to each other and to reactions that are spontaneous, at equilibrium or are non-spontaneous.
11. State the Nernst equation and describe how it is derived.
12. Use the Nernst equation to predict if a reaction is spontaneous, at equilibrium or non-spontaneous when the concentrations are not at standard conditions.
13. Describe how emf may be generated by concentration cells.
14. Describe dry cell, mercury, lead storage and solid-state lithium batteries.
15. Explain how measurement of the density of the electrolyte of a lead storage battery is an indication of charge on the battery.
16. Compare fuel cell efficiency to combustion in the generation of electricity.
17. Describe the hydrogen – oxygen fuel cell to include E°_{cell} and the overall chemical reaction.
18. Describe corrosion and give several examples.
19. Indicate several methods to reduce corrosion including the use of inhibitors, coatings, passivation and cathodic protection.
20. Predict the expected results from the electrolysis of molten sodium chloride, water and aqueous sodium chloride.
21. Use the term “overvoltage” to justify what reaction occurs at the anode for the electrolysis of aqueous sodium chloride.
22. Use the concepts of electrolysis to predict the quantitative results of processes involving the passage of a known current for a designated time.

Chapter 23

Nuclear Chemistry

The importance of the interaction of electron clouds dominates most of the chemistry that we study; yet nuclear reactions (the atomic bomb) changed the world forever. Upon completion of this chapter, your students should be able to:

1. Compare chemical reactions to nuclear reactions.
2. State the charge, mass and symbol for each of the following:
 - proton
 - neutron
 - electron or beta particle
 - positron
 - alpha particle
3. Balance nuclear equations using the concepts of conservation of mass and conservation of atomic numbers.
4. Explain how nuclei having a neutron-to-proton ratio that is too high results in beta-particle emission and nuclei having a neutron-to-proton ratio that is too low results in either positron emission or electron capture.
5. Use the concepts of nuclear binding energy, mass defect and Einstein’s mass-energy equivalence relationship to calculate nuclear binding energy per nucleon.
6. Describe a radioactive decay series using the terms parent and daughter.
7. Use half-life and first order kinetics to find number of radioactive atoms left after a given length of time.
8. Use first order kinetics and disintegrations per time to calculate half-life.