AACC
STUDENT LEARNING OBJECTIVES FOR CHE 112

Correlated to Chemistry, the Central Science by Brown, LeMay, Bursten, Murphy, Woodward, Stoltzfus 13th ed.

CHAPTER 15 Chemical Equilibrium

REVIEW solution stoichiometry & molarity (Chap 4), enthalpy change (Chap 5), partial pressure (Chap 10), and using the quadratic formula.

NOTE: For section 15.1, read it through equation 15.1 only.

The student will be able to:

1. Explain the concept of equilibrium in terms of distribution of amounts of products and reactants.
2. Explain dynamic equilibrium.
3. Write equilibrium constant expressions for both homogeneous and heterogeneous equilibria.
4. Interrelate values of equilibrium constants, K, for a reaction written in different ways.
5. Apply rule of multiple (simultaneous) equilibria.
6. Relate Kc to K_P.
7. Calculate K for a reaction given equilibrium concentrations.
8. Use the value of K to determine whether the formation of products or reactants is favored.
9. Calculate reaction quotient Q, and use Q value to decide the direction in which the reaction will proceed.
10. Calculate equilibrium concentration of a species, given K and equilibrium concentrations of other reaction species.
11. Calculate equilibrium concentrations of all reaction species given K and initial concentrations (provided equilibrium expression's algebraic order is ≤ 2).
12. Find cases which reduce to simple algebra, or which justify using simplifying approximations (in #11 above). **(OVER)**

13. Explain and apply LeChậtelier’s Principle.

14. Explain the effect on the equilibrium system of changing concentration, pressure, volume, or temperature.

15. Explain why a catalyst does not affect K.

16. Define bold face terms in *Chapter Summary & Key Terms*, p 660.

MEMORIZE:

$K_p = K_c\ (RT)^{\Delta n}$, & know which R value, and that temperature
must be in Kelvin.

CHAPTER 16 Acid-Base Equilibria

(Brown, LeMay, et al., 13th ed.)

REVIEW naming ionic & molecular compounds and acids (Chapter 2); acids and bases & their reactions - including writing net ionic reactions & memorizing the names and formulas of strong acids and strong bases (Chapter 4); Lewis structures; electronegativity; bond strength; oxidation numbers; logarithms.

The student will be able to:

1. List the names and formulas of the strong acids & strong bases.
2. Explain the Arrhenius, the Brønsted-Lowry, & the Lewis models for acids/bases.
3. Describe the properties of ionic oxides in solution. (Note that some compounds reacting with water are in effect strong bases, e.g., Na_2O *reacting with water*.)
4. Explain the concept of autoionization of water.
5. Write the equilibrium reaction for the autoionization of water and write its equilibrium constant expression in terms of K_w.
6. Relate the concepts of acidic, basic, and neutral to $[H^+]$ (or $[H_3O^+]$) and $[OH^-]$.
7. Determine pH given $[H^+]$ or $[OH^-]$, and vice versa.
8. Relate pH, pOH, and pK_W.
9. Calculate the pH (or pOH) of a strong acid or strong base solution of given concentration.
10. Explain the concept of amphiprotic.
11. Write dissociation reactions for weak acids and weak bases, and write their corresponding equilibrium constant expressions.
12. Explain the concept of acid (or base) strength.
13. Interrelate K_a, K_b, K_w, pK_a, pK_b, and pK_w appropriately.
14. Explain the concept of conjugate acid-base pairs, and describe the relation of the relative strength of members of a pair.
15. Explain the concept of simultaneous equilibria with respect to $K_aK_b = K_w$.
16. Use the Brønsted-Lowry model to explain and compare acid and base strengths and position of equilibrium.. [e.g., $HF(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + F^-(aq)$; H_3O^+ stronger acid than HF, equilibrium favors left side of reaction.]

(OVER)

17. Calculate equilibrium concentrations (and pH and pOH and percent ionization) for weak acids and weak bases, given K and initial concentration; calculate quantities for variations of this scenario, e.g. calculate K_a given initial concentration and equilibrium pH.

18. Find cases which justify using simplifying approximations (in # 17 above).

19. Describe and explain the trend in subsequent K_a's for a polyprotic acid. Give the criterion for determining pH by considering only K_{a1}, (i.e., $K_{a1} \geq 10^3 K_{a2}$).

20. Recognize the pH nature of ions.

21. Relate the natures of the cations and anions of a salt solution in order to characterize the pH nature of the solution and to calculate pH of given salt solution of known concentration.

22. Explain the relation of structure & periodicity to acid/base behavior using concepts of 1) bond polarity, 2) bond strength, 3) stability of dissociation product.

23. Define bold face terms in *Chapter Summary & Key Terms*, pp. 713-714.

MEMORIZE:

$K_W = [H^+] [OH^-] = 1.0 \times 10^{-14}$ (numerical value true at 25^oC)

$pK_w = pH + pOH = 14.00$ (numerical value true at 25^oC)

$pH = -\log_{10} [H^+]$

$(pX = -\log_{10} X)$

In calculations, $[H^+] = [H_3O^+]$

$K_w = K_a \cdot K_b$ (for conjugate pair)

CHAPTER 17 Additional Aspects of Aqueous Equilibrium

(Brown, LeMay, et al., 13th ed.)

REVIEW precipitation reactions, solubility rules (guidelines), & writing net ionic reactions (Chapter 4).

The student will be able to:

1. Explain the Common Ion Effect with respect to weak acid or weak base equilibria.

2. Calculate pH and concentration of ions after a common ion is added to weak acid or weak base.

3. Define and describe buffers, prepare a buffer for a given pH, calculate pH of a given buffer using both equilibrium considerations and the Henderson-Hasselbalch equation.

4. Calculate pH of a given buffer after acid or base is added using the two-step method: 1) stoichiometry, 2) K_a or K_b equilibrium.

5. Recognize which types of acid-base reactions go to completion.

6. Calculate pH at any given point of a titration involving either: a) strong acid and strong base, b) strong acid and weak base, or c) weak acid and strong base.

7. Explain the salient features of titration curves.

8. Explain how indicators work and how they are chosen.

9. Explain factors affecting ionic solubility.

10. Write equilibrium expressions (K_{sp}) for slightly soluble ionic compounds.

11. Use K_{sp} to calculate concentration of ions and approximate solubility of the compound, and to qualitatively judge relative solubility.

12. Calculate Q value and compare it with K_{sp} to predict whether a precipitate forms.

13. Explain the affect of a common ion on the solubility of a slightly soluble ionic compound and calculate concentrations of all ions. (Common Ion Effect).

14. Describe the coordination sphere of a complex ion using terms like central atom, ligand, coordination number, and Lewis base.

15. Write reactions and equilibrium expressions for formation of complex ions in terms of K_f, the formation constant.

16. Use K_f to calculate concentrations of reaction species.

(OVER)

17. Use concepts of multiple equilibria (e.g., K_{sp} & K_f) to determine reaction outcomes.

18. Explain the classical method of qualitative analysis for inorganic ions.

19. Define bold face terms in *Chapter Summary & Key Terms*, p. 765.

BE ABLE TO USE: Henderson-Hasselbalch Equation

$$pH = pK_a + \log_{10}([\text{conj base}]/[\text{conj acid}])$$

CHAPTER 23 Transition Metals and Coordination Chemistry

(Brown, LeMay, et al., 13th ed.)

REVIEW Wave nature of light and visible spectrum; d orbitals; electron configurations (Chapter 6).

1. Describe the characteristics of a central atom which promote complex ion formation.

2. Distinguish among monodentate and polydentate ligands.

3. Use (do not need to memorize) the rules for nomenclature.

4. Describe the geometry of a given complex ion (linear, tetrahedral, square planar, octahedral).

5. Recognize isomerism: [a) constitutional—linkage and coordination sphere]
   b) stereo—cis/trans (geometric).

6. Use the Crystal Field Theory to describe the electron configurations of complex ions for octahedral complexes.

7. Describe how Crystal Field Theory explains color and magnetic properties.

8. Classify a complex as low-spin or high-spin.

9. Use the spectrochemical series (do not need to memorize) to predict colors of complexes.

10. Define bold face terms in *Chapter Summary & Key Terms*, pp. 1030-1031.

EQUATIONS TO MEMORIZE:

$c = \lambda\nu$
$E = h\nu = hc / \lambda$

CHAPTER 19 Thermodynamics and Equilibrium

(Brown, LeMay, et al., 13th ed.)

REVIEW Chapter 5 “Thermochemistry”.

The student will be able to:

1. Explain the concept of entropy.
2. Explain the relation of probability to entropy.
3. Explain the relation of the distributions of translational, vibrational, and rotational energy to entropy.
4. Explain factors affecting entropy.
5. Calculate entropy change for a reaction.
6. Correlate algebraic sign of entropy change for a process with the relevant factors affecting entropy.
7. Explain the three laws of thermodynamics.
8. Calculate the (Gibbs) Free Energy change (ΔG) for a reaction using $\Delta G = \Delta H - T\Delta S$, and the equation $\Delta G^0_{rxn} = \Sigma n \cdot \Delta G_f^0$ (products) $- \Sigma m \cdot \Delta G_f^0$ (reactants), when appropriate.
9. Predict spontaneity of a reaction based on the value of ΔG_{rxn}.
10. Define the standard free energy change, ΔG^0 and standard free energy change of formation, ΔG_f^0.
11. Explain the relation between free energy and work.
12. Relate reaction spontaneity to the algebraic signs of ΔH_{rxn} and ΔS_{rxn}, and to temperature.
13. Calculate the temperature at which $\Delta G^0_{rxn} = 0$ (when there is one).
14. Calculate and interpret ΔG for a reaction involving non-standard conditions (i.e., variation of temperature, pressure, and concentration).
15. Use the relation between ΔG_{rxn} and the equilibrium constant.
16. Define bold face terms in *Chapter Summary & Key Terms*, p. 844.

(OVER)

EQUATIONS TO MEMORIZE:

$\Delta H^0_{rxn} = \Sigma n \cdot \Delta H_f^0$ (products) $-$ $\Sigma m \cdot \Delta H_f^0$ (reactants)

$\Delta S^0_{rxn} = \Sigma n \cdot S^0$ (products) $-$ $\Sigma m \cdot S^0$ (reactants)

$\Delta G^0_{rxn} = \Sigma n \cdot \Delta G_f^0$ (products) $-$ $\Sigma m \cdot \Delta G_f^0$ (reactants), and know this equation is only good at temp at which table is compiled.
(where n & m refer to balancing coefficients)

$\Delta G^0 = \Delta H^0 - T\Delta S^0$

$\Delta G = W_{max}$ (done by system—if spontaneous)

$\Delta G = W_{min}$ (done on system—if nonspontaneous)

EQUATIONS to be able to use:

$\Delta G = \Delta G^0 + RT \ln Q$	Know which R value and that temperature must be in Kelvin.
$\Delta G^0 = RT \ln K$	Know which R value and that temperature must be in Kelvin.

CHAPTER 20 Electrochemistry

(Brown, LeMay, et al., 13th ed.)

REVIEW Oxidation-Reduction Reactions (Chapter 4)

The student will be able to:

1. Explain and describe examples of oxidation and reduction.
2. Determine oxidation numbers, including from Lewis Structures.
3. Use oxidation numbers to analyze oxidation-reduction (redox) reactions.
4. Separate a redox reaction into its half reactions.
5. Balance redox reactions.
6. Use electrical units: ampere, Coulomb, and volt.
7. Draw and explain the operation of voltaic (galvanic) cells.
8. Calculate cell voltages using the Table of Standard Reduction Potentials.
9. Qualitatively state the relation between cell voltage and spontaneity of reaction.
10. Determine relative strength of oxidizing agents and reducing agents using Standard Reduction Potentials.
11. Use the Nernst Equation to calculate the effect of concentration on voltage.
12. Calculate equilibrium constant from standard cell voltage.
13. Explain the chemistry of at least one commercial voltaic cell (battery).
14. Describe corrosion in terms of the electrochemical principles involved.
15. Explain corrosion-prevention techniques.
16. Draw and explain the operation of electrolytic cells.
17. Predict reaction products when a solution is electrolyzed.
18. Calculate quantities relating to electrolysis.
19. Define bold face terms in *Chapter Summary & Key Terms*, pp. 897-898.

(OVER)

EQUATION TO MEMORIZE:

$$E^0 = E^0_{ox} + E^0_{red} = E^0_{red}(\text{cathode}) - E^0_{red}(\text{anode})$$

EQUATIONS TO BE ABLE TO USE:

$$\Delta G = -nFE$$

$$E = E^0 - [RT/(nF)]\ 1n\ Q = E^0 - [2.30RT/(nF)]\ \log Q$$

$$E^0 = [RT/(nF)]\ 1n\ K = [2.30RT/(nF)]\ \log K$$

and know which R value, and that temp must be in Kelvin.

CHAPTER 14 Chemical Kinetics

(Brown, LeMay, et al., 13th ed.)

REVIEW logarithms *thoroughly*.

The student will be able to:

1. Express reaction rates and equivalent rates.
2. Calculate average rates and instantaneous rates.
3. Determine the Rate Law for a reaction from initial rate studies (i.e., rate vs. concentration).
4. Determine order of reaction using the integrated rate equations for zero, first, and second order.
5. Use integrated rate equations (for zero, first, and second) to calculate rate constants and concentration—time relations.
6. Calculate the rate constant, given the half-life (for zero, first, and second order processes.)
7. Explain the theoretical basis for the Arrhenius Equation in terms of collision frequency, effective collisions, activation energy, and transition states.
8. Calculate activation energy using the Arrhenius Equation.
9. Calculate rate constant—temperature relations using the Arrhenius Equation.
10. Explain catalysis and relate it to activation energy.
11. Analyze reaction mechanisms for consistency with a given rate law.
12. Explain reaction mechanisms using concepts of rate determining, molecularity, and intermediates.
13. Define bold face terms in *Chapter Summary & Key Terms*, p. 614.

(OVER)

EQUATIONS: Be able to use:

$[A] = -kt + [A]_o$ (zero order)

* $1n\,[A] = -kt + 1n\,[A]_o$ & $1n\,[A]/1n\,[A]_o = -kt$ (first order)

$1/[A] = kt + 1/[A]_o$ (second order)

$t_{1/2} = [A]_o/2k$ (zero order)

$t_{1/2} = 0.693/k$ (first order)

$t_{1/2} = 1/(k[A]_o)$ (second order)

Note: Must be able to recognize order for each equation above. They will not be labeled.

* $1n\,k_1 - 1n\,k_2 = 1n\,(k_1/k_2) = (E_a/R)(1/T_2 - 1/T_1)$
Know which R value is used and that temperature
must be in Kelvin.

* Given one form of equation, be able to obtain the other.

CHAPTER 21 Nuclear Chemistry

(Brown, LeMay, et al., 13th ed.)

REVIEW subatomic particles, mass number, atomic number, isotopes, atomic mass units (Chapter 2).

The student will be able to:

1. Explain why there is natural radioactivity.

2. Describe types of radioactivity.

3. Write balanced nuclear equations.

4. Describe guidelines which indicate nuclear stability.

5. Describe modes of detection of radiation, and units of decay rate.

6. Rank relative penetrating power of types of radiation.

7. Describe ionizing radiation with respect to biological effects.

8. Recognize that nuclear decay processes obey the first-order reaction rate law.

9. Calculate activity, ratio of nuclei, or time of decay using the first-order integrated rate equation.

10. Describe applications of radioactive nuclei, including performing calculations relating to radioactive dating.

11. Describe transmutation.

12. Use the Einstein Equation ($E=mc^2$) to convert change in mass to change in energy (and vice versa).

13. Calculate Binding Energies.

14. Relate Binding Energy to fusion and fission process; explain these processes.

15. Describe a nuclear fission reactor.

16. Define bold face terms in *Chapter Summary & Key Terms*, pp. 944-945.

EQUATION TO MEMORIZE: $E = mc^2$ or $\Delta E = c^2 \Delta m$

EQUATIONS TO BE ABLE TO USE:

$1n\ N = 1n\ N_0 - kt$ & $1n\ (N/N_0) = -kt$

$t_{1/2} = 0.693/k$

CHAPTER 24 The Chemistry of Life: Organic and Biological Chemistry

(Brown, LeMay, et al., 13th ed.)

REVIEW Lewis structures & resonance (Chap 9); electronegativity (Chap 8); sp, sp^2, sp^3 hybrid orbitals (Chap10)

The student will be able to:

1. Explain the basis of the organization of organic chemistry in terms of structural theory.

2. Categorize hydrocarbons as alkanes, alkenes, alkynes, cyclic alkanes, and aromatics.

3. Describe the characteristics of alkanes, alkenes, and alkynes using hybrid orbitals, geometry, electron density, and intermolecular forces.

4. Memorize and use the IUPAC system to name alkanes, alkenes, and alkynes.

5. Interconvert condensed formulas and structural formulas.

6. Define and recognize constitutional isomers and stereoisomers.

7. Determine structures of isomers for a given molecular formula.

8. Recognize the following families by their characteristic functional groups: alcohols, ethers, aldehydes, ketones, carboxylic acids, esters, amines, amides, and amino acids.

9. Define bold face terms in *Chapter Summary & Key Terms*, through Organic Functional Groups, p. 1082.

MEMORIZING reactions is not required.

CHAPTER 18 Chemistry of the Environment

(Brown, LeMay, et al., 13th ed.)

The student will be able to:

1. Define lithosphere, hydrosphere, and atmosphere.
2. Describe qualitatively the composition of the various layers of the atmosphere.
3. Explain what is meant by a pollutant.
4. Describe and give examples of how a given chemical can be harmless in one context, and harmful in another.
5. Use course concepts, such as kinetics, equilibria, electromagnetic radiation and thermodynamics, to approach environmental problems, such as acid rain, green house gases, and stratospheric ozone depletion.