CHEM 142 Exam 2 Study Guide – Spring 2012

I. Chapter 13: Kinetics (Integrated Rate Laws and Half-Life only)
Key concepts:
1) Integrated rate laws
2) Half-life

Problem-solving: A copy of Table 13.2 Rate Law Summary Table will be provided during the exam.
1) Work on Example 13.3 and Practice 13.3 (p. 574), Example 13.4 and Practice 13.4 (p. 575), Example 13.5 and Practice 13.5 (p. 575), and Example 13.6 and Practice 13.6 (p. 579).

II. Chapter 14: Chemical Equilibria
Concepts and *skills* that you need to know/possess:
1) Chemical equilibrium
   - Definition
     - ➢ A condition where the rates of forward and reverse reaction are equal
     - ➢ No net change in concentration of reactants and products
     - ➢ Does not mean concentrations of reactants and products are equal
   - Explain the difference between reaction rates (kinetics) and equilibrium (thermodynamics) – see lecture notes
   - Dynamic process (Review Question 2, p. 650)
     - ➢ Relate to reversibility of chemical reactions

2) Equilibrium constant expression, K (or K_{eq} or K_c; Review Questions 21-24, 33-34, p. 651-652)
   - Write equilibrium constant expressions for:
     - ➢ Gaseous reactants and products (Use K_P instead of K)
     - ➢ Heterogeneous reactions (don't include solids (s) and liquids(l))
     - ➢ Other chemical reactions, such as those involving species in aqueous (aq) solutions
   - Meaning of large K (>1), small K (< 1) and K = 1 in terms of direction of equilibrium and relative abundance (or concentration) of reactants and products (Review Questions 23-24)

3) LeChâtelier's principle (Review Questions 61, 63, 65, 67-70, p. 654)
   - State LeChâtelier's principle in terms of disturbance to a system in equilibrium
   - Explain the effect of each of the following changes to an equilibrium system:
     - ➢ Concentration of reactants or products
     - ➢ Temperature (treat "heat" as reactant or product depending on whether the reaction is exothermic or endothermic)
     - ➢ Pressure (only affects gases; relate to change in volume)

4) Calculating K given the equilibrium concentrations of all species (Review Questions 35-38 (but only for the lowest temperature in 37-38), p. 652)

5) Calculating K from initial concentrations and at least one equilibrium concentration (Example 14.6 and Review Questions 41-44, pp. 652-653)

6) Calculating equilibrium concentrations when K is given and the equilibrium concentrations of all but one species (Example 14.8, p. 632 and Review Questions 51, 53 & 55, p. 653)

III. Chapter 15: Acid-Base Equilibria

A. Terminologies and Concepts
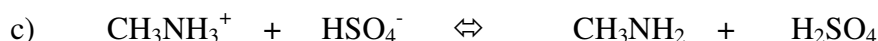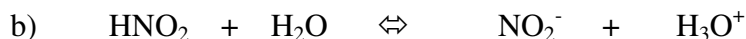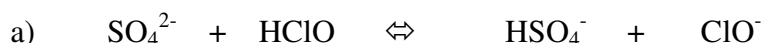 1. Acid reflux or GURD
 2. Identify several general properties of <u>acids</u>
 3. Identify several general properties of <u>bases</u>
 4. Binary acids - define and give examples
 5. Oxyacids - define and give examples
 6. Neutralization
 7. Arhhenius definitions - acids vs. bases; give examples
 8. Bronsted-Lowry definitions - acids vs. bases; give examples
 9. Amphoteric substances - define and give examples
10. Salts - define and give examples
11. Conjugate acid-base pairs - define and give examples

B. Equations
Get used to equations in the handout provided to you

C. Problem-Solving
 1. For each of the following reactions, indicate the Bronsted-Lowry acids and bases, as well as the conjugate acid-base pairs. (For simplicity, the subscript (aq) has been omitted, but all species are aqueous.)

   a) $SO_4^{2-} + HClO \Leftrightarrow HSO_4^- + ClO^-$

   b) $HNO_2 + H_2O \Leftrightarrow NO_2^- + H_3O^+$

   c) $CH_3NH_3^+ + HSO_4^- \Leftrightarrow CH_3NH_2 + H_2SO_4$

 2. According to the Bronsted-Lowry theory, which of the following would you expect to act as an acid? Which as a base? Which is ampotheric?

   a) NH_4^+ b) CN^- c) O^{2-} d) HPO_3^{2-} e) CHO_2^-
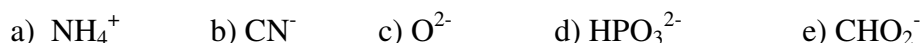
 3. Complete the table below by writing the formula of the conjugate acid or conjugate base of each of the following species. You must provide the correct ionic charge, if applicable.

Conjugate acid	Conjugate base	Conjugate acid	Conjugate base
H_2CO_3			PO_4^{-3}
	Br^-	NH_4^+	
	O^{2-}		$CH_3CO_2^-$
$HBrO_2$		H_2SO_4	