Lab 3. Le Châtelier's Principle

Prelab Assignment

Before coming to lab:

- Follow the guidelines in the "Lab Notebook Policy and Format for Lab Reports" handout to complete in your lab notebook the following sections of the report for this lab exercise: Title, Introduction, Materials/Methods and Data Tables. An outline or flow chart of the procedure is appropriate for the Materials/Methods section. Ensure that the table of contents of your lab notebook is current.
- Read the lab thoroughly and answer the pre-lab questions that appear at the end of this lab exercise. Background information for this lab can be found in chapter 17 of your textbook (*Silberberg 5th ed*).

Purpose

In this experiment you will observe shifts in three equilibrium systems, and learn to explain the observed changes in terms of molecular/ionic interactions and LeChâtlier's Principle.

Introduction

LeChâtlier's Principle states that when a stress is placed on a system at equilibrium the system will react to relieve the stress. The stress might be an increase or decrease in the concentration of a reactant or product, a change in pressure if a gas is present in the reaction, or a change in temperature. The system reacts by shifting in the direction of either products or reactants, whichever would counteract the effect of the stress. Often this shift produces observable results. For example, the color of the solution may change or a precipitate may form or dissolve.
To understand how changes in concentration might shift a chemical reaction at equilibrium consider the following generic equation:

$$\mathbf{A + B \rightleftharpoons C + D}$$

How might the equilibrium be shifted in the direction of the products? There are two possibilities.

1. **Add More Reactants to the System**. As additional reactants are added (the "stress"), the reaction system attempts to remove them by shifting to the right ("relief of the stress'), forming more products. This is fairly intuitive in that it seems reasonable that if you add more reactants, you will get more products since the addition of reactants increases the reactant concentration which increases the frequency of collisions between reactant particles thus increasing the rate of the forward reaction.

2. **Remove Products from the System**. This is less intuitive. Removal of a product (the "stress") results in a lower concentration of products and therefore fewer collisions between reactant particles—hence the reverse reaction decreases in rate relative to the forward reaction. This results in the equilibrium shifting towards product formation and "relief of the stress." If we can devise a way to remove the products from the reaction system, the equilibrium will shift to the right to compensate. How can this be done? There are several possibilities.

   If one of the products were a gas, simply letting the gas escape will remove the product and shift the equilibrium to the right. Remember what happens when you leave a soda bottle open. The CO_2 can escape, shifting the equilibrium, below, to the right, and the soda goes flat. $\mathbf{CO_2\ (aq) \rightleftharpoons CO_2\ (g)}$

If one of the products were a base, for example OH^-, we could remove OH^- from the system by adding an acid. In this instance H^+ ions react with the OH^- to form water, removing the product OH^-. The equilibrium will then shift to the right to compensate. If one of the products were an acid, then adding a base would have a similar effect.

In short, anything you can do to remove products from the reaction system will shift the equilibrium to the right. The reverse is also true; if reactants are removed, the equilibrium will shift to the left.

3. ***Changing the temperature*** of a reaction can also be thought of as adding a reactant or product. Recall that energy can be written into an equation, as follows:

   Endothermic reaction: $\textbf{heat} + \textbf{A} + \textbf{B} \rightleftharpoons \textbf{C} + \textbf{D}$,

   Exothermic reaction: $\textbf{A} + \textbf{B} \rightleftharpoons \textbf{C} + \textbf{D} + \textbf{heat}$

   Changing the temperature will have the opposite effect on exothermic and endothermic reactions. In the case of an endothermic reaction, heating the system can be thought of as adding additional reactant (heat). The system reacts by shifting to the right, using up the heat. Conversely, cooling the system effectively removes reactant (heat), and the system reacts by shifting to the left.

   For exothermic reactions the reverse is true. Adding heat puts a stress on the product side of the equilibrium and the system relieves the stress by shifting to the left. On the other hand, cooling causes a shift to the right.

In this laboratory exercise you will investigate three different equilibrium systems, and will see how they react to changes in concentration and temperature. The three systems are:

Part1. Equation 1. $\mathbf{Ca^{2+}_{(aq)} + 2\ OH^-_{(aq)} \rightleftharpoons Ca(OH)_{2\ (s)}}$

Calcium hydroxide is a fairly insoluble salt, that is, the equilibrium is strongly shifted to the right. In this part of the investigation you will examine how the solubility of $Ca(OH)_2$ can be increased, *i.e.*, how can the equilibrium be shifted to the left?

Part 2. Equation 2. $\mathbf{Cu^{2+}_{(aq)} + 2\ OH^-_{(aq)} \rightleftharpoons Cu(OH)_{2\ (s)}}$

Equation 3. $\underset{\text{Blue}}{\mathbf{Cu^{2+}_{(aq)}}} \mathbf{+ 6\ NH_{3\ (aq)} \rightleftharpoons} \underset{\text{Deep Blue}}{\mathbf{Cu(NH_3)_6^{2+}{}_{(aq)}}}$

The copper(II) ion forms a precipitate with hydroxide ions, as indicated in equation 2, above. The copper(II) ion also reacts with ammonia, NH_3, to form a complex ion, as indicated in equation 3, above. In part 2 of today's lab activity you will learn how the solubility of $Cu(OH)_2$ is affected by the presence of ammonia.

Part 3. Equation 4. $\underset{\text{Pink}}{\mathbf{Co(H_2O)_6^{2+}{}_{(aq)}}} \mathbf{+ 4\ Cl^-_{(aq)} \rightleftharpoons} \underset{\text{Blue}}{\mathbf{CoCl_4^{2-}{}_{(aq)}}} \mathbf{+ 6\ H_2O_{(l)}}$

In this system you will test the effects of adding both H_2O and Cl^- to the reaction. You will also observe the effect of temperature on the equilibrium, in order to determine whether the reaction is exothermic or endothermic.

Procedure (work in teams of two)

Note: Throughout the procedure are several "***thought" questions***, which are ***highlighted in bold***. These are designed to stimulate your thinking about the equilibrium systems as you work with each. Take time to answer each question as you proceed, being sure to write your answers in your lab notebook—***letter your responses in your notebook as the questions are lettered below***. You will find that your answers will then provide a basis for your discussion and analysis.

*** Be sure that you have your goggles on before beginning any work in the lab. *** Then clean and dry several smaller test tubes and arrange them in a test–tube rack.

Part I. $Ca^{2+}_{(aq)} + 2\ OH^{-}_{(aq)} \rightleftharpoons Ca(OH)_{2\ (s)}$

1. Label a test tube #1, and add a few mL of NaOH to it.

2. Then add $CaCl_2$ dropwise until a noticeable precipitate of $Ca(OH)_2$ is present.
   - **a. What happens to the OH^- ions? Are they all "in" the precipitate, or do some remain in solution? How can we prove where they are?**
   - ⇒ A test is to use the indicator phenolphthalein, which can be used to detect hydroxide ions in solution. To learn how the indicator works do the following:
     - Add a few mL of deionized water to a second test tube.
     - Add a few drops of phenolphthalein.
     - Then add a few drops of NaOH (aq) to the test tube.
   - **b. What happened?** Now add a few drops of the indicator to the test tube #1 with the $Ca(OH)_2$.
   - **c. What happened? What does this demonstrate?**
   - **d. What is the role of the Na+ ions in this experiment? The Cl^- ions?**

3. Add <u>6 *M* HCl</u> (NOT 12 *M*!!!) dropwise to the first test tube (#1) with the $Ca(OH)_2$ until a change is observed. Record your results.
   - **e. How does the equilibrium respond to the addition of HCl? What happens to the precipitate?**
   - **f. What happens to the OH^- ions?**
   - **g. Which OH^- ions are reacting in this test, the ones in the precipitate, the ones in the solution or both? How do you know? What is the evidence?**
   - **h. What happens at the molecular/ionic level as the HCl is added? What is the role of the H^+? The Cl^-? List all species that are present in the test tube at the end of the experiment.**

Part II. $Cu^{2+}_{(aq)} + 2\ OH^{-}_{(aq)} \rightleftharpoons Cu(OH)_{2\ (s)}$

Deep Blue

$Cu^{2+}_{(aq)} + 6\ NH_{3\ (aq)} \rightleftharpoons Cu(NH_3)_6^{2+}{}_{(aq)}$

1. Label a test tube #3, and then add a few mL of the Cu^{+2} solution to it.
2. Add 6 M NaOH dropwise until a precipitate of $Cu(OH)_2$ is observed.
   - **i. What happens to the Cu^{+2} ions? Are they all "in" the precipitate, or do some remain in solution? How do you know? What is the evidence?**
3. Add 6 *M* NH_3 dropwise (Use Caution!! Are those goggles still on!) until a change is observed. You will probably need to stir or shake your test tube. Record your observations.
   - **j. What happens at the molecular/ionic level as NH_3 is added? How is the position of the equilibrium affected? Include a careful explanation of your reasoning.**
4. Add 6 *M* HCl dropwise until a change is observed. Stir or shake as you add the HCl. Record all observations.
   - **k. What is happening at the molecular/ionic level as HCl is added?**
   - **l. How does the equilibrium respond to the addition of HCl?**

Pink Blue

Part III. $Co(H_2O)_6^{2+}(aq) + 4\ Cl^{-}(aq) \rightleftharpoons CoCl_4^{2-}(aq) + 6\ H_2O\ (l)$

1. Do this step and step 2 in the hood. Label a test tube #4, and add a few mL of the pink $Co(H_2O)_6^{+2}$ solution to a test tube. Add 12 *M* HCl dropwise (Use Caution!! Goggles still on!) until a color change is observed. Record your observations.
   - **m. What is happening at the molecular/ionic level as HCl is added? How does the equilibrium respond to the addition of HCl?**
   - **n. What is the role of the H^+ ions, and of the Cl^- ions?**
2. Add deionized water dropwise until a color change is observed. Record your observations.
   - **o. How does the equilibrium react when water is added?**
   - **p. Is the water just diluting the solution or is it reacting chemically with the solute? How do you know? What is the evidence?**
3. Use a hot plate to heat a 150 mL beaker of water to boiling. Label a test tube #5, and half fill it with a new portion of the pink $Co(H_2O)_6^{2+}$ solution and add 12 M HCl dropwise until a color between pink and blue. Gently heat the test tube by placing it in the boiling water bath. Observe and record any color changes. Cool the solution in a beaker of tap water and observe any color changes.
   - **q. What happens to the position of the equilibrium as heat is added to, or removed from the system? Does this prove that the forward reaction is endothermic or exothermic? Explain.**

Analysis

Your answers to the questions above will constitute the majority of your analysis. Be sure to explain all observed changes in terms of Le Chatelier's Principle—what was the stress induced upon the equilibrium and why the stress resulted in a shift to one side or the other. You should also include a summary of the principles demonstrated/learned in this experiment.

Acknowledgement: This lab is adapted from similar labs produced by my Chemistry colleagues at GRCC.

Lab 3. Le Châtelier's Principle
Prelab Questions

Name________________________________

Team Number _____

<u>Instructions</u>: Complete the following questions and hand in at the start of your lab period or when instructed by your instructor. Show your work with units and correct significant figures for all questions that involve a calculation.

1. Assume that the following reaction is in chemical equilibrium:

$$\mathbf{N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g) + heat}$$

Use Le Châtelier's Principle to explain the effect each of the following changes will have upon the system—will the equilibrium shift toward the product or reactant side? Why?

a. If more hydrogen is added to the system the equilibrium will shift to the…..
(circle one and explain below) **i.** Right **ii.** Left **iii.** Remains unchanged

b. If ammonia is removed from the system the equilibrium will shift to the…..
(circle one and explain below) **i.** Right **ii.** Left **iii.** Remains unchanged

c. If nitrogen is removed from the system the equilibrium will shift to the…..
(circle one and explain below) **i.** Right **ii.** Left **iii.** Remains unchanged

d. If the temperature is raised the equilibrium will shift to the…..
(circle one and explain below) **i.** Right **ii.** Left **iii.** Remains unchanged

e. If the pressure of the system is decreased by doubling the total volume the equilibrium will shift to the: (circle one and explain below) **i.** Right **ii.** Left **iii.** Remains unchanged

2. How would you expect the equilibrium in part 1 to be affected if calcium chloride solution were added? What visible clues would be noticed if a change occurred? ***Explain***.

3. How would you expect the equilibrium in part 3 to be affected if silver nitrate, $AgNO_3$, solution were added? What visible clues would be noticed if a change occurred? ***Explain***.

4. How would you expect the equilibrium in part 3 to be affected if sodium iodide powder were added? What visible clues would be noticed if a change occurred? ***Explain.***