LeChatelier's Principle

Influence of Concentration & Temperature on Reaction Equilibrium

In this experiment, the influence of concentration and temperature changes on the equilibrium of the following reaction will be observed.

$$\underset{\text{blue}}{Cu(H_2O)_4^{2+}} + 2Cl^- \leftrightarrow \underset{\text{green}}{Cu(H_2O)_2Cl_2} + 2H_2O$$

Cupric ion (Cu^{2+}) in water solution forms covalent bonds with 4 water molecules to form blue colored $Cu(H_2O)^{2+}$, called the **tetraaquacopper (ll) complex ion.** Adding chloride ion to a cupric ion solution causes the forward reaction to occur, consuming part of the added chloride ions to form green colored $Cu(H_2O)_2(Cl)_2$, called the **diaquadichlorocopper (ll) complex,** a soluble neutral complex. If only a little chloride is added, the equilibrium will be to the left side of the reaction where blue $Cu(H_2O)_4^{2+}$ is in greatest concentration, and the solution will appear blue. As more and more chloride ion is added, the equilibrium shifts towards the right side and the increasing concentration of $Cu(H_2O)_2(Cl)_2$ causes the solution color to become increasingly greener. If enough Cl^- is added, the equilibrium of the reaction moves far enough to the right to make the solution appear completely green.

This reaction is not thermoneutral. By observing color changes that occur upon heating and cooling equilibrium solutions, you should be able to deduce which direction the reaction is exothermic and which endothermic.

Procedure

A. Effect of Concentration

1. Obtain 30 mL of 0.5M $CuSO_4$ in a 50 mL Erlenmeyer flask and 5g solid NaCl in a plastic weighing dish.

2. Add the NaCl to the $CuSO_4$ solution, in approximately 1 g increments, using a spatula. Swirl the solution after each 1 g addition until the solid is dissolved. Observe the progressive color change of the solution. Record your observations in a data table.
   Do not add more than 5 grams NaCl.

3. Pour approximately 10 mL of the final solution into a test tube and set it aside for Part B, Step 1.

4. Using a 10 mL graduated cylinder, slowly add 10 mL H_2O, in increments of approximately 1 mL, to the remaining solution that was not set aside. Observe the progressive color change. Record your observations in a data table.

B. Effect of Temperature

1. Place the test tube containing 10 mL of solution set aside in Part A, Step 3, in an ice bath for 5 min. Observe the color change that occurs after a few minutes. Remove it from he dry ice and allow it to warm to room temperature. Again, observe any color change. Record your observations in a data table.

2. Obtain 30 mL 0.5M $CuSO_4$ solution in a 50 mL Erlenmeyer flask and add 2g NaCl. Swirl the solution until the solid is dissolved.

3. Heat the solution gently until a color change is observed. This might take about 5 minutes. Record your observations in your data table.
   Do not boil the solution.

4. Allow the solution to cool back to room temperature. Observe any color change and record your observations on the data sheet.

Analysis and Questions:

1. What effect does adding NaCl have on the equilibrium? Why?

2. What effect does adding water have on the equilibrium? Why?

3. Explain why the data in Procedure A supports or contradicts LeChatelier's Principle.

4. Which reaction is the exothermic one?

5. Which reaction is the endothermic one?

6. Explain your reasoning to support your answers to questions 4 & 5.

7. Sometimes chemists want a given reaction to run to completion without establishing an equilibrium. How might Le Chatelier's Principle be used to keep a reaction from reaching equilibrium?