

Chemical Equilibrium & Le Chatelier's Principle

Purpose:

To explore the effects of concentration and temperature on a well-known reaction, and to estimate the equilibrium constant for that reaction at 25 °C.

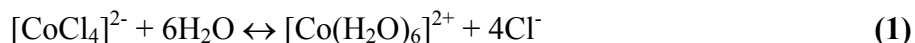
References:

1. Zumdahl, S. S. *Chemistry*; 2nd ed.; D.C Heath: Lexington, MA, 1989; Chapters 13, 15 and 20.
2. Shakhshiri, B. Z. *Chemical Demonstrations*; Volume 1; University of Wisconsin: Madison, 1985; pp. 280-285.
3. Dr. Thomas Janik, State University of New York College at Fredonia

Introduction and Background:

By this point we have hopefully covered equilibrium. It should be apparent that a major portion of the course deals with equilibrium. Remember everything is in some sort of state of equilibrium.

A particularly interesting chemical reaction from this perspective involved the interconversion of two Co(II) complexes:



A useful characteristic of this system is that the cobalt-containing reactant and product are visually distinct. $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ is pale pink in color, while $[\text{CoCl}_4]^{2-}$ is a bright, deep blue.

Le Chatelier's principle is a simple concept of great power for predicting the behavior of equilibrium systems. In this experiment you and your laboratory partner will explore this principle.

Pre-Lab Questions:

1. Given their coordination numbers, predict the geometries of these cobalt complexes. Can you think why Co(II) might bind simultaneously to six water molecules, but only four chloride ions?
2. Write the mass action expression for the above reaction; assume that H₂O is the solvent.
3. A certain system has $[\text{CoCl}_4]^{2-} = [\text{Co}(\text{H}_2\text{O})_6]^{2+}$. Derive an expression relating K_{eq} to $[\text{Cl}^-]$ for that system.

Procedure:

1. Obtain about 25 mL of ~0.3 M Co(II) solution. (A more precise concentration will be provided.) Note the color. Pour 5 mL of this solution into a test tube. In a second test tube mix about 0.5 mL of the Co(II) solution with 4.5 mL of 12 M HCl. [CAUTION: 12 M HCl is extremely corrosive!] What happens? Record and explain the results. Set these aside, but retain them for color reference. *What is the approximate Cl⁻/Co(II) molar ratio in the mixture?*
2. Prepare and intermediate (red-purple) solution as follows. Add 8.0 mL of 12 M HCl to a test tube contained 4.0 mL of the stock Co(II) solution, and mix thoroughly with a stirring rod. Assuming that your solution is blue, add more Co(II) in 0.5 mL increments, mixing each time, until the desired color is reached. Pour the resulting mixture into five test tubes, with roughly equal amounts in each. Each of those tubes contains a chemical system at equilibrium. You will disturb this equilibrium in four different ways; the fifth tube is for color comparison.
3. Heat the contents of one tube to about 75°C using a water bath. Cool a second in ice. What are the results? Record and rationalize.
4. While (3) is in progress, use a medicine dropper to add absolute ethanol to a third tube, while stirring. Does a color change occur? Explain what has happened, and why. [Hint: like water, the ethanol molecule is capable of intermolecular hydrogen bonding.]
5. To the fourth portion add about 1 mL of AgNO₃ solution. [CAREFUL: AgNO₃ stains hands and clothing, and is toxic when taken internally!] A two-phase system should result. Stir and then let stand for a few minutes. Describe and explain your observation. [Hint: Is AgCl soluble in water?]
6. Develop and carry out a procedure (based on #2, above) to estimate the equilibrium constant for equation (1) at room temperature. Don't expect to achieve an accuracy of more than two significant figures.

Report:

1. Write up a lab report. (Directions shown in class.)
2. Calculate K_{eq} for equation (1). Give an estimate of the uncertainty, and defend that estimate.
3. Answer the following questions.
 - a. Is the reaction described by equation (1) exo- or endothermic as written? How do you know?
 - b. Suggest one or more ways in increase the accuracy of your determination of K_{eq} .