

EXPERIMENT # 4

SHIFT HAPPENS... OR INVESTIGATING CHEMICAL EQUILIBRIA

Objective

To investigate the effect of stress on a chemical system at equilibrium and explain observed phenomena in the light of Le Châtelier's principle.

Introduction

As chemicals are mixed, they often react forming new substances. Consider a situation in which reactants A and B are forming products C and D. The general equation describing the reaction may be written as:



At any given time, all four species, reactants and products, coexist in the system. This means that a chemical reaction does not necessarily proceed until all of the reactants are exhausted or until the only species left in the system are the products. In fact, when substances react they eventually form an equilibrium mixture of reactants and products. At equilibrium the forward reaction which produces C and D proceeds at the same rate as the reverse reaction in which C and D react to produce A and B. Chemical equilibrium then is the state reached by a reaction mixture when the rate of the forward reaction is equal to the rate of the reverse reaction.

Experimental kinetic studies indicate that right after reactants A and B are mixed, the forward reaction is much faster than the reverse reaction. As time passes, however, the rate of the forward reaction decreases while the rate of the reverse reaction increases. Eventually, the two reaction rates become equal and the system reaches equilibrium. For a chemical reaction at a given temperature, the mathematical relationship between the equilibrium concentrations of the reactants and the products is represented by the equilibrium constant, K. The equilibrium constant expression is written in terms of the balanced chemical equation. For the hypothetical reaction in equation 1 the equilibrium constant expression can be written as

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b} \quad \text{Equation 2}$$

where the subscript c on K_c denotes that the concentrations of compounds A, B, C, and D are measured in molarity. The exponents in equation 2 come from the coefficients in the balanced equation (Equation 1).

One important fact to consider when investigating chemical equilibria is that the equilibrium state does not preclude unequal concentrations of species involved. In other words, the concentrations of the reactants and products are rarely equal when the equilibrium has been reached.

Shift Happens.....

According to Le Châtelier's principle, when the equilibrium state of a chemical system is disturbed by changing any factor affecting the equilibrium, the system will change in such a manner so as to counteract the effect of the applied stress.

The equilibrium in a liquid system can be disturbed by:

- a) an increase or decrease in the concentration of one or more of the products or reactants, or

b) an increase or decrease in temperature.

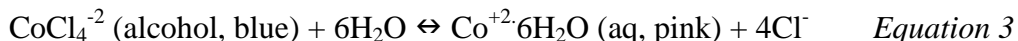
When either the concentrations of the system components or the temperature are changed, the equilibrium will change to produce either more reactants or more products to reach a new equilibrium. This shift in equilibrium can often be monitored by observing the change in the concentration of one or more of the species (e.g. ions) present in the system.

About the experiment

In this experiment, you will investigate the following equilibrium systems:

I. Blue - Pink equilibrium

Hydrated cobalt (II), $\text{Co}^{+2}\cdot 6\text{H}_2\text{O}$, is pink in color. However, if the water of hydration is removed either by prolonged heating of the solid or by dilution of the aqueous ion with alcohol; blue, anhydrous cobalt chloride, CoCl_2 is formed. This is an equilibrium process that is dependent on the relative concentration of water and the temperature of the solution.



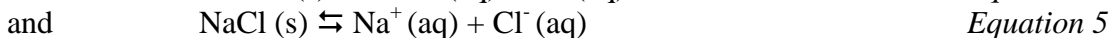
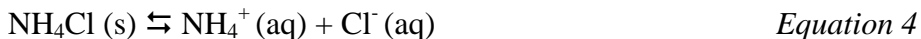
In this experiment you will start with a blue alcoholic solution of anhydrous cobalt (II) chloride. No equilibrium has been established because of the absence of water. By adding water, you will produce the equilibrium described in equation 3. You will then cause a shift in the equilibrium by adding or removing reactants or by changing the temperature. By observing the changes in the color of the solution you can determine the effects these changes on the equilibrium.

The following changes will be made on the solution and their effects will be observed and explained.

- Silver nitrate (AgNO_3) reacts with Cl^- ions forming insoluble AgCl . Addition of AgNO_3 removes a product in equation 3 (Cl^-) causing an equilibrium shift.
- Addition of hydrochloric acid (HCl) adds more Cl^- and causes an equilibrium shift.
- Alternately heating or cooling the reaction mixture shifts the equilibrium and will allow you to determine whether the forward reaction is exothermic or endothermic.

II. Salt dissolution

Precipitation and dissolution of many salts is an equilibrium process. The dissolution equilibrium for the salts ammonium chloride and sodium chloride are represented by the following equations:



In the second part of this experiment you will investigate the effect of a) addition of concentrated hydrochloric acid (a source of Cl^-), or b) heating and cooling, on the equilibria described in equations 4 and 5.

Procedure

Preparation

1. Prepare a hot water bath using a 150-ml beaker with tap water and a hotplate. The water bath should be stabilized by placing a medium sized metal ring around the beaker. Keep the water hot but not boiling. Periodically add more water to the beaker to keep it at least 1/3 full.
2. Prepare an ice bath. Use a 150-ml beaker, ice, and tap water (do not exceed the level of 75 ml). As the ice melts, periodically remove some of the water and add more ice.
3. Prepare four small, clean, and dry test tubes. Label the tubes with consecutive numbers: #1, # 2, #3, and #4.
4. Place 0.5 ml of AgNO_3 solution in a fifth test tube labeled AgNO_3 , and 0.5 ml of HCl in a second test tube labeled HCl. Place one plastic transfer pipet in each solution. Keep the pipets in their appropriate solutions when not using them.
5. Place 5 to 10 ml of distilled water in a 50 mL beaker. Place a medicine dropper in the beaker. Use the dropper to deliver drops of water when needed.

I. Blue-Pink equilibrium

Take notes concerning the initial color of the solution and record all changes including color and turbidity (precipitate formation) after each step is performed! Add reagents one drop at a time, mixing after each drop by swirling or tapping the test tube!

Using a pipet, add 25 drops of blue cobalt (II) chloride alcoholic solution to test tubes # 1 and # 2.

Test tube #1

Experiment 1A. Using the prepared medicine dropper, add distilled water, just until you observe a color change. Do not add more water than necessary or you will have to repeat the experiment. Record your observations.

Experiment 1B. Add drops of concentrated HCl to the mixture from experiment 1-A just until the color changes. Record your observations.

Experiment 1C. Add drops of AgNO_3 solution to the mixture from experiment 1-B, just until the color changes. Record your observations.

Discard all solutions in the designated waste container. Wash the tubes, rinse with distilled water, and dry.

Test tube #2

Experiment 2A. Using the prepared medicine dropper, add drops of distilled water just until you observe a color change. Record your observations.

Experiment 2B. Place the test tube from experiment 2-A in the hot water bath. Leave the test tube in the bath, swirling occasionally, until you see a color change. Record your observations.

Experiment 2C. Remove the test tube from experiment 2-B from the hot bath and place it in the ice bath. Leave it there, swirling occasionally, until the color changes. Record your observations

Discard all solutions into a designated waste container. Wash the test tubes, rinse with distilled water, and dry.

II. Salt Dissolution

There are no color changes involved in this experiment. However, you should observe and record any other changes.

Test tubes #3 & #4

Place 20 drops of saturated ammonium chloride in tube #3, and twenty drops of saturated sodium chloride solution in test tube #4.

Experiment 3A. Add concentrated hydrochloric acid to tube #3 one drop at a time. Mix after each drop; continue until a change is observed. Record your observations.

Experiment 3B. Place the test tube from experiment 3A in a hot water bath. Leave the test tube in the bath until a change is observed. Record your observations.

Experiment 3C. Remove the test tube from the hot water bath and place it in the prepared ice bath until a change is observed. Record your observations.

Experiments 4A-C. Repeat the procedures followed with experiments 3A-C with test tube #4. Record your observations.

CLEANUP

To wash the pipets, draw tap water into them, then discard water to a designated waste container. Repeat three times. Next, rinse the pipets with distilled water (2 times) using identical technique, but discard the water into the sink. Return the pipets to the cart.

Report

In the discussion, write the equilibrium equations for the reactions studied in this lab (equations 3, 4, and 5). Determine whether the reaction is exothermic or endothermic and include heat (Q) as a product or reactant. Remember that there is only one equilibrium equation for a given system and the added reagents are not part of that equation. For each stress applied (reagent added, heat added or removed) identify the stress and explain the effect of the stress in detail, according to Le Chatelier's principle. Include which way the equilibrium shifted and why. Remember that the observations you made and recorded are not explanations; rather they are results needing to be explained.

Chemistry 1225 Experiment 4 **Data Sheet** Name _____

Experiment 1A observations:

Experiment 1B observations:

Experiment 1C observations:

Experiment 2A observations:

Experiment 2B observations:

Experiment 2C observations:

Experiment 3A observations:

Experiment 3B observations:

Experiment 3C observations:

Experiment 4A observations:

Experiment 4B observations:

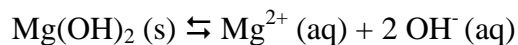
Experiment 4C observations:

Discussion:

Chemistry 1225 Experiment 4 **Prelab**

Name _____

1. A student obtained a test tube with a suspension of white, slightly soluble magnesium hydroxide in water. This system was at equilibrium as represented by the following equation:



Write the equilibrium constant expression for this reaction.

2. According to Le Châtelier's principle, what will happen to the equilibrium in question 1 if nitric acid, HNO_3 , is added? Hint, will HNO_3 react with any of the products or reactants? Explain your answer.

3. According to Le Châtelier's principle, what would happen to the equilibrium in question 1 if soluble magnesium chloride was added to the solution? Explain your answer.

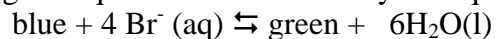
4. The student heated the solution described in question 1 and found that more of the Mg(OH)_2 dissolved. Is the reaction endothermic or exothermic? Explain your answer.

Chemistry 1225 Experiment 4 **Postlab** Name _____

1. The alcoholic, blue solution from part I of your experiment is commonly used in weather-forecasting devices. Based on your observations in the lab, explain how this reaction can indicate coming rain.

2. It is common knowledge that most ionic salts are more soluble in warm water than in cold water. Do the solubilities of NH_4Cl and NaCl fit this pattern? Explain why or why not based on your lab observations.

3. A student is studying the equilibrium described by the equation:



a) The system is at equilibrium and has a blue color. The student added several crystals of sodium bromide (NaBr) to this blue equilibrium solution. Predict happened in the reaction mixture and explain why.

b) The student placed the green equilibrium solution in an ice bath. After a short time, the solution turned blue. What does it tell you about the thermochemical character of the forward reaction? Explain your answer.