Exploring Equilibrium
It Works Both Ways

Introduction

The word equilibrium has two roots: *æqui,* meaning equal, and *libra,* meaning weight or balance. Our physical sense of equilibrium—in the motion of a seesaw or the swing of a pendulum—suggests an equal balance of opposing forces. How does this physical sense of equilibrium translate to chemical equilibrium? Let's explore the nature and consequences of equilibrium in chemical reactions.

Concepts

- Reversible reactions
- Chemical equilibrium
- Complex-ion reaction
- Acid–base indicators

Background

Physical changes, such as melting ice or dissolving sugar, are often introduced by noting that these processes can be easily reversed. Some common examples of chemical change, such as burning wood or spoiling food, generally cannot be reversed. A closer look at chemical change, however, reveals that many chemical reactions are also reversible.

Consider the following example of a reversible chemical reaction. At high pressures and in the presence of a catalyst, nitrogen and hydrogen react to form ammonia. If the temperature is high enough, however, ammonia decomposes to reform its constituent elements. The reaction can go both ways! This reversible reaction is represented symbolically using double arrows (Equation 1).

$$N_2(g) + 3H_2(g) \rightleftarrows 2NH_3(g) \qquad \textit{Equation 1}$$

What happens when nitrogen and hydrogen are allowed to react? In a closed system, the concentrations of nitrogen and hydrogen will decrease and the concentration of ammonia will steadily increase as the reaction proceeds in the forward direction. Soon, however, the concentration of ammonia will be large enough that the reverse reaction will begin to take place at a significant rate as well. Eventually, as the reaction occurs in both the forward and the reverse directions, the number of ammonia molecules being formed will become equal to the number of ammonia molecules being consumed. At this point, no further changes will be observed in the overall concentrations of nitrogen, hydrogen, and ammonia. This is the point of chemical equilibrium.

Chemical equilibrium is defined as the state where the rate of the forward reaction equals the rate of the reverse reaction and the concentrations of reactants and products remain constant with time. Note that this definition describes a dynamic picture of equilibrium. The reactions continue, but there is an equal balance of opposing reaction rates.

What happens when the equilibrium is disturbed? Any factor that changes the rate of the forward or the reverse reaction will change the amounts of reactants and products that are present at equilibrium. Reaction conditions that are known to affect the rates of chemical reactions include the concentrations of reactants and the temperature. In this experiment, we will investigate how changes in reaction conditions affect the amounts of reactants and products present at equilibrium.

Experiment Overview

The purpose of this experiment is to explore the nature of chemical equilibrium and to identify conditions that affect the position of equilibrium. Two different reversible reactions will be studied.

Reaction of iron(III) nitrate with potassium thiocyanate will be used to study *complex-ion equilibrium*. Iron(III) ions react with thiocyanate ions to form $FeSCN^{2+}$ complex ions (Equation 2). The effects of changing the concentrations of reactants and of changing the reaction temperature will be investigated.

$$Fe^{3+}(aq) + SCN^-(aq) \rightleftarrows FeSCN^{2+}(aq) \qquad \textit{Equation 2}$$

The properties of an indicator will be used to study *acid–base equilibrium*. An indicator is a dye that can gain or lose hydrogen ions to form substances that have different colors. Equation 3 summarizes the reversible reaction of the indicator bromcresol green (HIn). HIn represents an uncharged indicator molecule and In^- an indicator anion formed after the molecule has lost a hydrogen ion. The color of the indicator in the presence of either excess H^+ or OH^- ions (see Equation 4) will show how changing the concentration of a product affects the equilibrium shown in Equation 3.

$$\underset{\textit{Color A}}{HIn(aq)} \rightleftarrows H^+(aq) + \underset{\textit{Color B}}{In^-(aq)} \qquad \textit{Equation 3}$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \qquad \textit{Equation 4}$$

Pre-Lab Questions

1. True or False: At equilibrium, no more reactants are transformed into products. If false, rewrite the statement so that it correctly describes the nature of chemical equilibrium.
2. True or False: At equilibrium, the concentrations of reactants and products are equal. If false, rewrite the statement so that it correctly describes the nature of chemical equilibrium.
3. Paper coated with cobalt chloride is sold commercially as moisture-sensitive test strips to estimate relative humidity levels between 20 and 80 percent in air. The following reversible reaction takes place with water:

$$\underset{\textit{Blue}}{CoCl_2(s)} + H_2O(g) \rightleftarrows \underset{\textit{Pink}}{CoCl_2{\cdot}6H_2O(s)}$$

(a) What color do you think the paper will be when the humidity is low (20%)? What color will it be when the humidity is high (80%)?

(b) The test strips come with a color chart to estimate intermediate humidity levels. Predict the intermediate color that might be observed when the humidity is about 50%.

Materials

Bromcresol green indicator, 0.04%, 1 mL
Iron(III) nitrate solution, $Fe(NO_3)_3$, 0.1 M, 4 mL
Hydrochloric acid, HCl, 0.1 M, 2 mL
Potassium thiocyanate solution, KSCN, 0.1 M, 4 mL
Sodium hydroxide solution, NaOH, 0.1 M, 2 mL
Sodium phosphate (monobasic) solution, NaH_2PO_4, 0.1 M, 1 mL
Water, distilled or deionized
Wash bottle
Beaker, 50-mL
Beakers, 250- or 400-mL, 2
Beral-type pipets, graduated, 8
Hot plate
Ice
Labeling or marking pen
Stirring rod
Test tubes, small, 6
Test tube rack
Thermometer

Safety Precautions

Potassium thiocyanate is toxic by ingestion. Dilute hydrochloric acid and sodium hydroxide solutions are skin and eye irritants. Iron(III) nitrate solution is also a possible skin and body tissue irritant; it will stain clothes and skin. Avoid contact of all chemicals with eyes and skin. Clean up all chemical spills immediately. Wear chemical splash goggles and chemical-resistant gloves and apron. Wash hands thoroughly with soap and water before leaving the laboratory.

Procedure

Part A. Complex-Ion Equilibrium of Iron(III) and Thiocyanate Ion

1. Fill two beakers (250- or 400-mL) half-full with tap water. Add ice to one beaker to prepare an ice-water bath (0–5 °C) for use in step 8. Heat the second beaker on a hot plate to prepare a hot water bath (70–80 °C) for use in step 9. Do not boil the water.
2. Observe and record the initial colors of the $Fe(NO_3)_3$ and KSCN solutions.
3. Prepare a stock solution of $FeSCN^{2+}$: In a clean 50-mL beaker, measure 40 mL of distilled water. Using separate Beral-type pipets for each solution, add 1 mL of 0.1 M $Fe(NO_3)_3$ and 2 mL of 0. 1 M KSCN. Mix thoroughly with a stirring rod.
4. Label six clean test tubes 1–6. Using a graduated, Beral-type pipet, add 1 mL of the $FeSCN^{2+}$ stock solution to each test tube 1–6.
5. Add 10 drops of distilled water to test tube 1. Gently swirl the test tube to mix the solution and record the color of the solution in the data table. Test tube 1 will be used as the control solution for comparison purposes in steps 6–10.
6. Add 10 drops of 0.1 M $Fe(NO_3)_3$ to test tube 2. Gently swirl the test tube to mix the solution and compare the color of the resulting solution to the control in test tube 1. Record the color comparison in the data table.
7. Add 10 drops of 0.1 M KSCN to test tube 3. Gently swirl the test tube to mix the solution and compare the color of the resulting solution to the control in test tube 1. Record the color comparison in the data table.

8. Add 10 drops of distilled water to test tube 4 and place the sample in an ice-water bath. After 3–5 minutes, remove the test tube from the ice bath and compare the color of the solution to the control in test tube 1. Record the color comparison in the data table.

9. Add 10 drops of distilled water to test tube 5 and place the sample in a hot water bath at 70–80 °C. After 2–3 minutes, remove the tube from the hot water bath and compare the color of the solution to the control in test tube 1. Record the color comparison in the data table.

10. To test tube 6, add 10 drops of 0.1 M NaH_2PO_4. Record the color and appearance of the solution in the data table.

11. Wash the contents of the test tubes down the drain with excess water and rinse with distilled water.

Part B. Acid–Base Equilibrium of Bromcresol Green

12. Obtain 2 mL of distilled water in a clean test tube and add 5 drops of 0.04% bromcresol green. Swirl gently and record the color of the solution in the data table.

13. Add 3 drops of 0.1 M HCl solution to the test tube. Swirl gently and record the new color of the solution in the data table.

14. Add 0.1 M NaOH dropwise to the solution until the original color is restored. Shake gently and record the number of drops of NaOH added and the color of the solution in the data table.

15. Continue adding 0.1 M NaOH dropwise until a total of 5 drops of NaOH have been added in steps 14 and 15 combined.

Can the process be reversed to obtain a color that is intermediate between that in steps 13 and 14?

16. Add 0.1 M HCl again dropwise very slowly until the solution reaches a "transition" color midway between the two colors recorded above (steps 13 and 14). Swirl gently between drops to avoid overshooting the transition color. Record the number of drops of HCl required and the color in the data table. *Note:* It may be necessary to add half a drop at a time.

17. Wash the contents of the test tube down the drain with excess water and rinse with distilled water.

Name: ______________________________

Class/Lab Period: ______________________________

Exploring Equilibrium

Data Table

Part A. Complex-Ion Equilibrium of Iron(III) and Thiocyanate Ions		
Color of $Fe(NO_3)_3$ Solution		**Color of KSCN Solution**
Test tube 1	Color of control solution (step 5)	
Test tube 2	Color after addition of $Fe(NO_3)_3$ (step 6)	
Test tube 3	Color after addition of KSCN (step 7)	
Test tube 4	Color of solution after cooling (step 8)	
Test tube 5	Color of solution after heating (step 9)	
Test tube 6	Color after addition of NaH_2PO_4 (step 10)	

Part B. Acid–Base Equilibrium of Bromcresol Green	
Initial color of indicator solution (step 12)	
Color after addition of HCl (step 13)	
Color after addition of NaOH (step 14)	
Number of drops of NaOH added (step 14)	
Amount of HCl required to obtain "transition" color (step 16)	
Transition color (step 16)	

Post-Lab Questions *(Use a separate sheet of paper to answer the following questions.)*

1. Write the chemical equation for the reversible reaction of iron(III) ions with thiocyanate ions in Part A. Label this Equation A. Use the information in the data table to write the color of each reactant and product underneath its formula.

2. How did the color of the solution in Part A change when additional reactant—either $Fe(NO_3)_3$ in step 6 or KSCN in step 7—was added? *Explain the observed color changes:* Adding more reactant to an equilibrium mixture of reactants and products increases the rate of the (forward/reverse) reaction and thus (increases/decreases) the amount of product.

3. How do the results obtained in steps 6 and 7 demonstrate that both reactants and products must be present at equilibrium?

4. How did the color of the solution in Part A change when it was cooled (step 8) or heated (step 9)? How do these results demonstrate that the reaction shown in Equation A does

indeed occur in both the forward and reverse directions?

5. In step 10, $H_2PO_4^-$ ions combined with iron(III) ions and removed them from solution. How did the color of the solution in Part A change when NaH_2PO_4 was added? *Explain the observed color change:* Removing one of the reactants from an equilibrium mixture of reactants and products decreases the rate of the (forward/reverse) reaction and thus (increases/decreases) the amount of product.

6. After observing the effect of NaH_2PO_4 on the equilibrium mixture in step 10, a student doubted that both Fe^{3+} and SCN^- ions were still present in solution. Suggest additional experiments that could be done to prove that both reactants are still present at this point.

7. Write the chemical equation for the reversible reaction of bromcresol green with water in Part B. Label this Equation B. *Hint:* Refer to Equation 3 in the *Background* section.

8. Use the color changes observed for the indicator before and after adding HCl (steps 12 and 13) to predict the colors of the HIn and In^- forms of bromcresol green. Write the colors of HIn and In^- underneath their formulas in Equation B. Explain your reasoning. *Hint:* Adding HCl increases the concentration of H^+ ions. Which reaction, forward or reverse, would that increase?

9. *Explain the observed color change:* Adding more product to an equilibrium mixture of reactants and products increases the rate of the (forward/reverse) reaction and thus (increases/decreases) the amount of product.

10. In step 14, hydroxide ions reacted with and removed H^+ ions from solution (see Equation 4 in the *Background* section). What color change was observed when NaOH was added? *Explain the observed color change:* Removing one of the products from an equilibrium mixture of reactants and products decreases the rate of the (forward/reverse) reaction and thus (increases/decreases) the amount of product.

11. What form(s) of the indicator were most likely present when the transition color was observed in step 16? How does this observation provide visual proof that not all reactions "go to completion?"