Chemical Equilibrium

Le Chatelier's Principle

OVERVIEW

Some chemical reactions are reversible; that is, not only do the reactants react to form products, but the products can in turn reform into the original reactants. When a reversible system reaches a point at which the rate of the forward reaction equals the rate of the reverse reaction, the system is said to be at equilibrium. At equilibrium, no observable changes in the system can be noted. It is important to understand that at equilibrium all reaction participants are present – all reactant particles, as well as all product particles. This does not mean, however, that all are present in equal amounts.

Le Chatalier's Principle tells us that if a system at equilibrium is subjected to a stress, the system will shift in order to minimize the effect of that stress. In this laboratory exercise you will examine how various stresses cause equilibrium systems to shift. These particular equilibrium systems undergo color changes as the equilibrium shifts between the reactants and products, allowing you to see which side of the reaction becomes favored.

PURPOSE

- To observe the effect of various stresses (ion concentration; temperature) on equilibrium systems.

SAFETY

1. Use extreme caution when handling the acidified cobalt(II) chloride hexahydrate, $CoCl_2 \cdot 6\ H_2O$. HCl, used to make the acidified solution, is highly corrosive. Report any spills immediately to the teacher. Spills on the skin should be flushed with cold water.
2. Potassium chromate, K_2CrO_4 is a hazardous substance. Use with care.
3. Use caution when using the hot water bath.

THE REACTIONS

Part 1. Chromate – Dichromate Equilibrium

$$2\ CrO_4^{2-}{}_{(aq)} + 2\ H_3O^+{}_{(aq)} \rightleftharpoons Cr_2O_7^-{}_{(aq)} + 3H_2O_{(l)}$$

yellow *orange*

Part 2. Iron(III) – Thiocyanate Ion Complex

$$Fe^{3+}{}_{(aq)} + SCN^-{}_{(aq)} \rightleftharpoons Fe(SCN)^{2+}{}_{(aq)}$$

light brown *red*

Part 3. Cobalt(II) Chloride Complex; Effect of Temperature

$$[CoCl_4]^{2-} + 6\ H_2O \rightleftharpoons [Co(H_2O)_6]^{2+} + 4\ Cl^-$$

blue *pink*

EQUIPMENT AND MATERIALS

Amounts needed per group are approximate

Part 1. Chromate – Dichromate Equilibrium

- 0.1 *M* K_2CrO_4, 5 mL per group
- 0.1 *M* $K_2Cr_2O_7$, 5 mL per group
- 1 *M* HCl, dropper bottle
- 1 *M* NaOH, dropper bottle
- 2 test tubes in a test tube rack

Part 2. Iron(III) – Thiocyanate Ion Complex

- 0.1 *M* $FeCl_3$, 10 mL per group
- 0.1 *M* KSCN, 10 mL per group
- 0.1 *M* KCl, 5 mL per group
- 10 mL graduated cylinder
- 250 mL or larger beaker
- distilled water, approx 100 mL
- 4 test tubes in a test tube rack
- dropper pipette

Part 3. Cobalt(II) Chloride Complex; Effect of Temperature

- 0.2 *M* acidified $CoCl_2 \cdot 6\ H_2O$, 15 mL per group
- hot water bath (approx. 90° C)
- ice water bath
- 3 test tubes in a test tube rack

PROCEDURE

Part 1. Chromate – Dichromate Equilibrium

1. Fill a test tube approximately half-full with potassium chromate, K_2CrO_4, (Tube 1).

2. Fill another test tube approximately half-full with dichromate, K_2CrO_4 (Tube 2).

3. To Test Tube 1 add several drops of HCl. HCl is an acid; adding HCl increases the concentration of H_3O^+ ions in the equilibrium system. Note the color change.

4. After recording the color change in Test Tube 1, add several drops of NaOH. NaOH is a base; adding a base decreases the concentration of H_3O^+ ions in the equilibrium system. Record the color change.

5. To Test Tube 2 add several drops of NaOH until a color change is observed.

6. After recording the color change, add several drops of HCl to Test Tube 2. Again note the change in color.

Part 2. Iron(III) – Thiocyanate Ion Comlex

1. Pour 5 mL of 0.1 *M* $FeCl_3$ into the beaker.

2. After rinsing the graduated cylinder, measure 5 mL of 0.1 *M* KSCN. Add to the beaker containing the $FeCl_3$. Note the color change.

3. Add enough distilled water to the beaker to dilute the solution to a light brown color. Pour some into a test tube to check the color.

4. Pour about 10 mL of this solution into each of the four numbered test tubes. The first test tube will serve as a control.

5. To Test Tube 2 add several drops of $FeCl_3$ until a color change is observed. Adding more $FeCl_3$ increases the concentration of Fe^{3+} in solution. Record the color change.

6. To Test Tube 3 add several drops of KSCN until a color change is observed. Adding more KSCN increases the concentration of SCN^- in solution. Record the color change.

7. To Test Tube 4 add several drops of KCl until a color change is observed. Adding KCl causes the concentration of Fe^{3+} to decrease because the Fe^{3+} and Cl^- react to form $FeCl_4^-$.

Part 3. Cobalt(II) Chloride Complex; Effect of Temperature

1. Fill three test tubes approximately half-full with the acidified $CoCl_2 \cdot 6\ H_2O$ solution.

2. Test tube 1 will serve as the control. Keep this test tube at room temperature. Record the initial colour of the solution.

3. Place the second test tubes in the hot water bath. After a few minutes a colour change will occur. Record the colour.

4. Place the third test tube in the cold water bath. Record any colour change.

5. Reverse tubes 2 and 3. Observe any colour changes that occur.

RESULTS

Part 1. Chromate – Dichromate Equilibrium

Solution		
K_2CrO_4	initial colour	
	HCl added	
	NaOH added	
$K_2Cr_2O_7$	initial colour	
	NaOH added	
	HCl added	

Part 2. Iron(III) – Thiocyanate Ion Complex

Test Tube	**Stress Applied**	**Initial Color**	**Final Color**
1	Control		--
2	Fe^{3+} added		
3	SCN^- added		
4	Cl^- added: decreases $[Fe^{3+}]$		

Part 3. Cobalt(II) Chloride Complex; Effect of Temperature

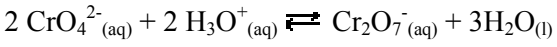
Temperature	**Solution Colour**
room temperature	
hot water bath	
cold water bath	

CONCLUSIONS AND QUESTIONS

Part 1. Chromate – Dichromate Equilibrium

1. Use Le Chatelier's Principle to explain the color changes observed in both test tubes with the addition of both HCl and NaOH.

   Write your answer as a clear, descriptive paragraph.



$$2\ CrO_4^{2-}{}_{(aq)} + 2\ H_3O^+{}_{(aq)} \rightleftharpoons Cr_2O_7^-{}_{(aq)} + 3H_2O_{(l)}$$

yellow *orange*

   Adding HCl, an acid, increases the concentration of H_3O^+ ions; adding NaOH, a base, decreases the concentration of H_3O^+ ions.

Part 2. Iron(III) – Thiocyanate Ion Comlex

2. Use Le Chatelier's Principle to explain the color changes observed in Test Tubes 2 – 3 upon the addition of $FeCl_3$, KSCN, and KCl.

   Write your answer as a clear, descriptive paragraph.

$$Fe^{3+}{}_{(aq)} + SCN^-{}_{(aq)} \rightleftharpoons Fe(SCN)^{2+}{}_{(aq)}$$

light brown *red*

Part 3. Cobalt(II) Chloride Complex; Effect of Temperature

$$[CoCl_4]^{2-} + 6\ H_2O \rightleftharpoons [Co(H_2O)_6]^{2+} + 4\ Cl^-$$

blue *pink*

3. Based on the colour changes observed in the hot water and cold water baths, determine whether the forward reaction is endothermic or exothermic.

   Rewrite the equation with a simple energy term (" + heat") included on the appropriate side of the equation. You may find it easier if you begin by using only the terms "heat", "pink", and "blue" in your equation.