Name____Date______________

37 Chemical Equilibrium and Le Chatelier's Principle

PRE-LAB DISCUSSION

In most of the chemical reactions you have studied thus far, at least one of the reactants has been "used up." The point at which a reactant is used up marks the end of the reaction, and the reaction is said to have "gone to completion." Under ordinary circumstances, the product(s) of such reactions are not able to react to re-form the original reactants. Thus, these are "one-way" reactions. They proceed in one direction only.

Many other chemical reactions do not go to completion. Rather, the products of these reactions remain in contact with each other and react to re-form the original reactants. Such reactants are said to be *reversible.* In a reversible reaction, the forward and reverse reactions proceed at the same time. When the *rates* of the two reactions are equal, a state of **chemical equilibrium** is said to exist. Under such conditions, both the forward and reverse reactions continue with *no net change* in the quantities of either products or reactants.

A state of equilibrium is affected by concentration and temperature and, if gases are involved, by pressure. If a system at equilibrium is subjected to a change in one or more of these factors, a stress is placed on the system. According to Le Chatelier's principle, when a stress is placed on a system at equilibrium, the equilibrium will shift in the direction that tends to relieve the stress. Equilibrium will be re-established at a different point, that is, with different concentrations of reactants and products.

In this experiment, we will study two equilibrium systems. The equilibrium equation for the reversible reaction of the first system is:

$$\underset{\text{(light brown)}}{Fe^{3+} + SCN^-} \rightleftharpoons \underset{\text{(red)}}{Fe(SCN)^{2+}}$$

The addition of any substance to the system that increases the concentration of Fe^{3+} or SCN^- will favor the forward reaction. This will cause the equilibrium to shift to the right. The addition of any substance that decreases the concentration of these ions will have the opposite effect.

The equilibrium equation for the second system is:

$$\underset{\text{(yellow)}}{2CrO_4^{2-}} + 2H_3O^+ \rightleftharpoons \underset{\text{(orange)}}{Cr_2O_7^{2-}} + 3H_2O$$

The addition of an acid to this system increases the H_3O^+ concentration and causes the equilibrium to shift to the right. The addition of any substance that causes a decrease in H_3O^+ concentration will have the opposite effect.

By studying these two systems, you should achieve a better understanding of equilibrium systems and their response to stress.

PURPOSE

Study equilibrium systems and their responses to stress as described by Le Chatelier's principle.

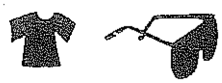
EQUIPMENT

beaker, 100-mL
graduated cylinder, 10-mL
test tubes, 13x100 (5)
test tube rack
dropper pipet
marking pencil
safety glasses

MATERIALS

0.1 M $FeCl_3$
0.1 M KSCN
0.1 M KCl
distilled water
0.1 M K_2CrO_4
0.1 M $K_2Cr_2O_7$
1.0 M HCl
1.0 M NaOH

PROCEDURE



Part I

1. Using a marking pencil, number 4 test tubes 1 through 4, respectively, and stand the tubes in a test tube rack.
2. Measure out 5 mL of 0.1 M $FeCl_3$ and pour it into a 100-mL beaker. Add 5 mL of 0.1 M KSCN to the same beaker. Dilute the contents of the beaker with distilled water until the solution is a light reddish-orange color. Divide the solution equally among the four numbered test tubes. Set test tube #1 at one end of the rack to be used for color comparison.
3. Using a dropper pipet, add 0.1 M $FeCl_3$ drop-by-drop to the solution in test tube #2 until a color change occurs. Record your observations in the data section. Rinse the pipet.
4. Repeat step 3, but instead of $FeCl_3$, add the following solutions drop-by-drop to the test tube indicated. Rinse the pipet after each use.

   test tube # 3   0.1 M KSCN
   test tube # 4   0.1 M KCl

   Record your observations in the data section.
5. Discard the solutions. Wash and rinse the test tubes and invert them in the rack to drain.

Part II

6. Using a marking pencil, number 4 test tubes 5 through 8, respectively. Stand the tubes in a rack.
7. Measure out 10 mL of 0.1 M K_2CrO_4. Pour 5 mL each into test tube #5 and test tube #6. Rinse the graduated cylinder and measure out 10 mL of 0.1 M $K_2Cr_2O_7$. Divide this equally between test tube #7 and test tube #8.

8. Using a dropper pipet, add 1.0 M HCl drop-by-drop to test tube #5 until the color changes. Record your observations.

9. Repeat step 8 with test tube #6. As soon as the color changes, rinse the pipet and use it to add 1.0 M NaOH drop-by-drop to the solution until the color changes again. Record your observations.

CAUTION: HCl and NaOH are highly corrosive substances. Solutions of these chemicals can cause serious injury if they come in contact with skin or eyes. Handle these solutions with extreme care.

10. Using the pipet, add 1.0 *M* NaOH to test tube #7 until the color changes. Record your observations.
11. Repeat step 10 with test tube #8. As soon as the color changes, rinse the pipet and use it to add 1.0 *M* HCl to the solution until the color changes again. Record your observations.

OBSERVATIONS AND DATA

Part I

Color

test tube #2 ____________________________

test tube #3 ____________________________

test tube #4 ____________________________

Part II

Color Change

test tube #5 ____________________________

test tube #6 ____________________________

test tube #7 ____________________________

test tube #8 ____________________________

CONCLUSIONS AND QUESTIONS

1. Write equilibrium equations for the reversible reactions that take place in Part I and Part II.
2. Using Le Chatelier's principle, explain how the addition of $FeCl_3$ to the solution in test tube #2 (step 3) affected the equilibrium that existed in the solution. Give similar explanations for the addition of each of the other substances (step 4).
3. Using the equilibrium equation for the reaction and Le Chatelier's principle, explain the color changes noted in Part II.