Chemistry CP

Name:

Investigating Chemical Equilibrium

Section:

Prelab

Be sure to attach this page to your lab report!

1. What is a spectator ion?

2. What will you observe if both ions added are spectator ions?

3. Only eleven solutions are tested. What is the purpose of the twelfth well on the microplate?

4. What macroscopic change indicates an increase in concentration of $FeSCN^{2+}$?

5. How do spectator ions affect the equilibrium concentrations?

6. Give the symbols for the ions present in a solution of sodium sulfate.

7. In your own words, describe the purpose of this experiment.

8. On a separate sheet of paper, construct a data table to record the following information: the name and formula of each reagent, the ions contained in each reagent, the well number (such as A-1 or D-5) of each reagent, and the color change (lighter, darker, or no change).

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When a reaction occurs in a closed system, the products of the reaction may react with each other and reform the original reactants in a reverse reaction. When the rates of the forward and reverse reactions are equal, the system is said to be at *equilibrium.* At equilibrium, the macroscopic properties of the system such as color, pressure, and concentration remain constant even though reactions continue to take place. However, *LeChatelier's principle* states that any change in the conditions of the system causes a stress that shifts the equilibrium.

In this experiment, aqueous solutions of iron (III) chloride and potassium thiocyanate will be mixed. Iron (III) ions, Fe^{3+}, react with colorless SCN^- ions to for the brick-red complex ion, $FeSCN^{2+}$. The equation is as follows:

$$Fe^{3+}(aq) + 3\ Cl^-(aq) + K^+(aq) + SCN^-(aq) \Leftrightarrow FeSCN^{2+}(aq) + K^+(aq) + 3Cl^-(aq)$$

The potassium and chloride ions are unchanged in the process. Because these ions do not participate in the reaction, they are called "spectator ions." The net equilibrium equation is as follows:

$Fe^{3+}(aq)$	+	$SCN^-(aq)$	$\Leftrightarrow$	$FeSCN^{2+}(aq)$
Colorless		colorless		brick-red

At equilibrium, the color of the $FeSCN^{2+}$ ion is constant because the rate of formation of $FeSCN^{2+}$ is equal to the rate of its decomposition. You will disturb the equilibrium by adding pairs of ions to this system and observe any changes in color. At least one of the ions added in each pair is a spectator ion. If both added ions are spectator ions, there is no change in the equilibrium concentrations and therefore no change in color. If the color becomes darker, the shift in equilibrium must have caused an increase in the concentration of $FeSCN^{2+}$. If the color becomes lighter, the concentration of $FeSCN^{2+}$ must have decreased while the concentrations of Fe^{3+} and SCN^- increased.

Hypothesis

Predict which of the solutions tested in this experiment will shift the equilibrium to the left, and which will shift the equilibrium to the right.

Objectives

- Analyze color changes and relate them to shifts in equilibrium.
- Infer molecular events, which account for the observed changes.
- Present and explain the observed changes using LeChatelier's principle.
- Formulate a model to account for shifts in equilibrium and predict the results of other changes.

Materials

96-well microplate
50 mL beaker
25 mL graduated cylinder
0.1 M iron (III) chloride
0.1 M potassium thiocyanate
0.1 M sodium thiocyanate
0.1 M sodium chloride
0.1 M iron (III) nitrate
0.1 M ammonium nitrate
0.1 M sodium carbonate
0.1 M potassium bromide
0.1 M ammonium thiocyanate
0.1 M sodium hydroxide
0.1 M calcium iodide
Distilled water

Procedure

1. Put on your laboratory apron and safety goggles.
2. To the 50 mL beaker, add 3 drops of iron (III) chloride and 3 drops of potassium thiocyanate. Record your observations in your data table.
3. Add enough distilled water to the beaker so that the red color is less intense. Color changes due to shifts in equilibrium will be more noticeable in the more dilute solution.
4. Using a micropipette, add a few drops of the diluted solution to each of 12 wells of the microplate. Be sure to leave two or three empty wells between each filled well.
5. Choose one of the wells near the center of the plate to be your control. Using a micropipette, add one drop of water to this well; you will compare all other color changes to this well.
6. Place a sheet of white paper under the microplate. This will help you see the color changes more easily.
7. Add one drop of the first solution listed on your data table and record the result as lighter, darker, or no change.
8. Repeat step 7 for each of the remaining solutions.
9. Dispose of the solutions as directed by your teacher.
10. Before leaving the laboratory, clean up all materials and wash your hands thoroughly.

Analyze and Apply Questions

1. List the formulas for all of the spectator *ions*.
2. Explain how you recognized the spectator ions.
3. List the formulas of all ions that increased the concentration of $FeSCN^{2+}$.
4. List the formulas of all ions that decreased the concentration of $FeSCN^{2+}$.
5. Predict the color effect and the change in concentration when one drop of each of the following solutions is added. (Note: this is NOT your data table!)

Solution	Color Change	Concentration of $FeSCN^{2+}$
Potassium carbonate		
Calcium thiocyanate		
Sodium bromide		
Potassium hydroxide		
Potassium nitrate		
Iron (III) bromide		
Ammonium bromide		
Iron (III) iodide		
Calcium hydroxide		
Ammonium carbonate		

6. Describe the submicroscopic changes that take place when one drop of a solution of ammonium thiocyanate is added to the solution in equilibrium. (Be as detailed as possible.)
7. State LeChatelier's principle.
8. Apply LeChatelier's principle to question 6.
9. Imagine that a lethal amount of the equilibrium mixture was swallowed. If $FeSCN^{2+}$ is very poisonous but Fe^{3+} and SCN^- are only slightly toxic, which solutions used in your experiment can be used as antidotes?

Remember to write an appropriate conclusion. Restate your findings and discuss possible sources of error.