Le Chatelier's Principle in a Cobalt Complex

A reversible chemical reaction is subjected to stress using various means, and the effects are observed and discussed.

Overall Goal

Demonstration of Le Chatelier's principle through various stresses on a reversible chemical reaction involving a hydrated Cobalt complex.

Background Information

To understand the demonstration, students should previously be familiar with the following topics:

- Solid working knowledge of chemical equations
- Exposure to idea of equilibrium
- Exposure to concept of hydrous form
- Solid knowledge of factors affecting rate of reaction
- Knowledge of stoichiometry and limiting reagents
- (Optional) Exposure to equilibrium constants

Before the demonstration, students should be able to:

- Be able to place heat on the proper side of a chemical equation for endo/exothermic equations
- Be able to interpret chemical equations
- Be able to correctly identify state of matter in a chemical equation

This demonstration should follow the discussion, qualification and quantification of chemical reactions. Ideally, it should follow a discussion of equilibrium in a chemical reaction.

Equipment

- Graduated cylinder
- 2 beakers
- Suitable gloves for acid handling
- 1 glass stirring rod
- Pipette or dropper
- Overhead projector
- Test tube
- Test tube holder or ring stand and clamp
- Bunsen burner
- Matches
- Ice bath in suitable container, insulated preferred

Reagents

- 20 mL of 0.2 M Cobalt chloride solution: 26 g of $CoCl_2$ per liter of water.
- ~20 mL of 0.1 M Silver nitrate solution: 1.7g of $AgNO_3$ per 100 mL of water.
- ~40 mL of 6M hydrochloric acid
- ~20 mL of Distilled water

Preparation

- Prepare the stock solutions.
- Measure out stock solutions.
- Set up overhead projector and place two beakers on it. Make sure it is focused and the projected beakers are clearly visible.
- Place 20 mL Cobalt chloride solution in 1 beaker. Pink color should be observable through projection.

Procedure

1) Turn on overhead, project setup.
2) Shift of equilibrium to the left: Slowly, add 40 mL of concentrated HCl dropwise into Cobalt solution in beaker, stopping when deeper blue color forms. Note the formation of blue color.
3) Divide the resulting blue solution in half between the two beakers.
4) Shift of equilibrium to the right: Add 20 mL of distilled water dropwise to one of the beakers, until color changes away from blue. Note the color change to pink.
5) Light burner using matches.
6) Pour remaining blue solution from step 3 from beaker into test tube.
7) Shift of equilibrium to the right: Cool test tube in ice bath until solution changes color. Note pink color.
8) Shift of equilibrium to the left: Using test tube holder or ring stand, heat test tube until solution changes color. Note blue color.
9) Slowly pour heated solution back into the empty beaker, to avoid breakage. Let cool slightly.
10) Shift of equilibrium to the right: add silver nitrate solution dropwise to the blue solution until a precipitate forms. Note the formation of a pink color.

Hazards 6M HCl is very corrosive! Exercise caution around solution and vapors. Always make sure to add acid to less concentrated solutions, rather than adding less concentrated solutions to the acid. Always clean up any spills!

Disposal Cobalt Chloride and Silver Nitrate must be disposed of as hazardous chemicals.

Tips and Suggestions

- Raw solution concentrations are not as important as having clearly visible colors. Let your judgment guide you after you practiced the demonstration.
- Changing the temperature is not as dramatic as the effect of changing the concentrations by adding water, HCl, or Silver Nitrate. Consider skipping the temperature effects, if pressed for time.
- Make sure to have students PREDICT what will happen before each equilibrium shift. Have them write down their predictions. Alternately, consider taking a poll. Or do both!

Explanation of Why the Demonstration Works

We begin by dissolving the Cobalt Chloride in water, forming the pink hydrous form. The system's equilibrium is then stressed by adding substances that affect concentrations or other parameters.

$$[CoCl_4]^{2-}_{(aq)} + 6H_2O_{(l)} \qquad [Co(H_2O)_6]^{2+}_{(aq)} + 4Cl^{-}_{(aq)}$$

(blue) (pink)

The reason this equilibrium is easy to study is because (1) it doesn't lie too far to either one side or the other and (2) the reactants are of a different color (pink) than the products (blue). Thus, when reactants predominate, the solution looks pink and when products predominate, the solution looks blue.

Adding water would shift equilibrium to favor the formation of the hydrous form. Adding HCl adds Cl^- ions, shifting the equilibrium toward production of the blue complex. Adding Silver Nitrate removes Cl^- ions due to the formation of Silver Chloride, thus shifting the equilibrium toward production of the pink complex to restore the consumed Cl^- ions.

To understand the effect of changing the temperature, it helps to remember that the process has an energy component:

$$[CoCl_4]^{2-}_{(aq)} + 6H_2O_{(l)} \qquad [Co(H_2O)_6]^{2+}_{(aq)} + 4Cl^-_{(aq)} + \textit{Heat}$$

(blue) (pink)

So, the addition of heat would favor the production of the blue complex, whereas removing heat would favor the production of the pink complex to restore the lost energy.

As far as the formation of various colors, Cobalt atoms in salts are positive ions with a 2+ charge. They attract negative particles such as chloride ions (Cl) and the oxygen end of water molecules. When most of the negative species around the cobalt ion are water molecules, the ion absorbs light so that it appears pink. When the water molecules bonded to the Cobalt atoms are gone (evaporated or chemically stripped off), the negative chloride sticks to the positive cobalt ions, and the cobalt appears blue. The water and the chloride are different and cause the electrons in the cobalt to absorb different energies. These different energies result in different colors absorbed by the cobalt.

Questions for Students

- What might be other ways of shifting the equilibrium? (Depending on how many steps you have demonstrated so far.)
- How many times do you think the equilibrium can be shifted before it stops working? Why?
- Can the shift in equilibrium you observed upon heating and cooling be explained in terms of Le Chatelier's Principle if we think of HEAT as a reactant? Which side would heat go on?
- Is this an endo or exothermic reaction? Does it depend on which way the reaction is proceeding?

Additional Suggestions

- The pink Cobalt solution can be applied to fiber materials like fabric, paper or filter material. Let this dry and use it as a moisture detector. As a follow-up activity, have students make humidity detectors and place them around the classroom or around school.
- Use the original pink cobalt solution to write a message on a piece of filter paper. Leave it out to dry for the class or use a hairdryer to "magically" turn the letters blue.
- "The Magic Handkerchief"
  Make a different cobalt solution using 26g Hydrous $CoCl_2$ in 100 mL distilled water. Soak a white handkerchief or small white towel in this solution. Let dry or use a hair dryer to help. Store it in an airtight bag to keep it from picking up humidity from the air. With classroom theatrics in mind,

use the “blue” handkerchiefs or towels to dry just washed hands and see them magically turn color where your wet hands touched them.

modified for Chemical Demonstrations by Jesse Fitzgerald 2008