What Happens to an Equilibrium System when It Is Disturbed?

According to your text, Le Chatelier's Principle describes the effect that applying various types of stress will have on the position of equilibrium: whether it will shift to increase or decrease the concentration(s) of products in the equilibrium system. These stresses include variations in such factors as concentrations of reactants or products, temperature of the system, and, for reactions involving gases, the pressure.

Some investigations are done with systems in a water solution. Here, unless gases are involved in the reaction, the volume of the system is generally defined by the volume of the solution and pressure is of little or no consequence. This sort of system permits us to simplify Le Chatelier's Principle to read:

> For any system at equilibrium in solution: If anything is added to the system, it will try to consume whatever was added. If anything is removed from the system, it will try to replace whatever was removed.

Note that the word, "anything" refers to energy (heat) as well as to any of the reactants or products shown in the reaction equation.

The purpose of this experiment is to let you observe for yourself what Le Chatelier's Principle means. Your investigation will deal with two complex ions, both containing cobalt(II): they are $Co(H_2O)_6^{2+}$ and $CoCl_4^{2-}$. The procedure is short and should require only about 25 minutes to complete. When you have finished and cleaned your work area, return to your desk for the post-lab discussion to discuss what you saw and what it signifies in terms of the reaction system being investigated.

Materials

50 mL beaker (1)	thin-stem pipets
24-well plate (1)	10 mL ethanol
$CoCl_2{\cdot}6H_2O(s)$	0.1 M $AgNO_3$
$CaCl_2(s)$	12 M hydrochloric acid
deionized water	hot plate
spatula (1)	ice bath

PROCEDURE

Caution: Put on your goggles and apron now!!

1. Pour 10 mL of ethanol into a 50 mL beaker, using the measurements on the side of the beaker.
2. Place several pieces of the solid cobalt(II) chloride in one of the wells in your 24-well plate. Note both its color and the formula for the compound, as shown on the label of the stock bottle.
3. Add 4 or 5 crystals of the cobalt II chloride solid to the ethanol in the beaker until a blue solution results. Add more crystals, if necessary.
4. Using a thin-stem pipet, transfer one-fifth of the blue solution to four of the wells in the 24-well plate. Be sure to leave a small amount of the solution in the beaker.
5. To one of the wells from step #4, add 5 drops of deionized water, one drop at a time. Record your observations after each drop. Repeat this step in two more of the wells so that all three of them exhibit the same color.
6. Take your 24-well plate to the fume hood. Use the thin-stem pipet provided in the acid bottle of 12 M hydrochloric acid and carefully add one drop at a time until you have added five drops to the first well from step #5.

   WARNING: Hydrochloric acid is caustic and corrosive. Avoid contact and immediately rinse all spills with copious amounts of water.
7. To the second well from step #5, add 2 small lumps of solid calcium chloride.
8. To the third well from step #5, add 10 drops of 0.1 M silver nitrate ($AgNO_3$). **Caution: Silver nitrate will stain your skin and clothing!**
9. Retain the solution in the fourth well to use for comparison purposes.
10. To the remaining solution in the beaker, add just enough deionized water to get a purple color that is about half-way between the blue and pink shades. Place the beaker on a hot plate and warm the beaker until a color change occurs.
11. Chill the beaker in an ice bath to see if the color change in step #10 is reversible.

QUESTIONS

1. The net-ionic equation for the equilibrium reaction you have been investigating is:

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

pink blue

On the reagent bottle, the formula for solid cobalt(II) chloride is $CoCl_2{\bullet}6H_2O$. What name do we give to compounds which have water molecules bound to their structure?

2. Which cobalt complex (see equation) was favored by addition of water in step #5? Use Le Chatelier's Principle to explain the color change.
3. Which cobalt complex was favored in both steps #6 and #7? What ion is common to both of the reagents that caused the color changes? Use Le Chatelier's Principle to explain why the color changes occurred in each case.
4. Silver chloride, AgCl, is a white solid. For the equilibrium:

$$AgCl(s) \rightleftarrows Ag^+ + Cl^-$$

The solubility product is K_{sp}= 1.77 X 10^{-10}. At equilibrium, would you expect to have mostly silver and chloride ions in solution, or mostly solid silver chloride? Explain.

5. What color was the solid you formed in step #8? What must it have been? What color did the liquid in the well turn? Which complex of cobalt was favored? Explain. Use Le Chatelier's Principle to explain why the liquid in the well underwent the color change that you observed.
6. Which cobalt complex was favored by addition of heat in step #10? Rewrite the equation for the reaction, including the energy term directly in the equation. The value of ΔH for the process is +50 kJ/mol. Use Le Chatelier's Principle and the equation that you just wrote to explain the color changes that resulted from the heating and cooling.