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Studying Chemical Equilibria and Applying Le Châtelier's Principle Using Microscale Techniques

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Purpose of the Experiment

Determine the effect of a change in a reaction condition on a chemical system at equilibrium. Correlate the observed responses with Le Châtelier's principle.

Background Information

When a chemical reaction proceeds in the direction written until one of the reactants is consumed, we say the reaction goes to completion. For example, if we place a small piece of metallic sodium (Na) in water (H_2O), the reaction shown in Equation 1 proceeds as written until all of the Na has reacted. The single arrow ($\rightarrow$) indicates that the reaction proceeds to completion.

$$2\ Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g) \qquad \text{(Eq. 1)}$$

Some reactions are not so straightforward. For example, when we mix solutions of cerium(IV) ammonium nitrate, $(NH_4)_2Ce(NO_3)_6$, and iron(II) nitrate, $Fe(NO_3)_2$, water-soluble cerium(III) nitrate, $Ce(NO_3)_3$, iron(III) nitrate, $Fe(NO_3)_3$, and ammonium nitrate (NH_4NO_3) are formed, as shown in Equation 2.

$$(NH_4)_2Ce(NO_3)_6(aq) + Fe(NO_3)_2(aq) \rightarrow Ce(NO_3)_3(aq) + Fe(NO_3)_3(aq) + 2\ NH_4NO_3(aq) \qquad \text{(Eq. 2)}$$

The complete ionic equation for this reaction is shown in Equation 3.

$$2\ NH_4^+(aq) + Ce^{4+}(aq) + Fe^{2+}(aq) + 8\ NO_3^-(aq) \rightarrow 2\ NH_4^+(aq) + Ce^{3+}(aq) + Fe^{3+}(aq) + 8\ NO_3^-(aq) \qquad \text{(Eq. 3)}$$

Because equal numbers of ammonium ions (NH_4^+) and nitrate ions (NO_3^-) appear on both sides of the equation, these ions are not directly involved in the reaction and can be eliminated from the equation. We call such ions **spectator ions**. The equation showing only reacting species is called the **net ionic equation**. The net ionic equation for the reaction in Equation 2 is given in Equation 4. This equation represents the **forward reaction** for this chemical system.

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$$Ce^{4+}(aq) + Fe^{2+}(aq) \rightarrow Ce^{3+}(aq) + Fe^{3+}(aq) \quad \text{(Eq. 4)}$$

Before we mix the Ce^{4+} ion and Fe^{2+} ion solutions, neither Ce^{3+} nor Fe^{3+} ions are present in the system. Immediately upon mixing, some of the Ce^{4+} and Fe^{2+} ions react and form Ce^{3+} and Fe^{3+} ions. Then, a few of the Ce^{3+} and Fe^{3+} ions begin to react to form Ce^{4+} and Fe^{2+} ions, as shown in Equation 5. This equation represents the **reverse** or **backward reaction** for the chemical system in Equation 2.

$$Ce^{3+}(aq) + Fe^{3+}(aq) \rightarrow Ce^{4+}(aq) + Fe^{2+}(aq) \quad \text{(Eq. 5)}$$

Shortly after the mixing of the reactants, the rate of the forward reaction (Equation 4) is much greater than the rate of the reverse reaction (Equation 5). As the system continues to change, the rate of the forward reaction decreases, while the rate of the reverse reaction increases. Eventually, the two reaction rates become equal, and the chemical system reaches an **equilibrium**, as shown in Equation 6. The two arrows ($\rightleftharpoons$) indicate that the chemical system is at equilibrium.

$$Ce^{4+}(aq) + Fe^{2+}(aq) \rightleftharpoons Ce^{3+}(aq) + Fe^{3+}(aq) \quad \text{(Eq. 6)}$$

When the system is at equilibrium, Ce^{4+} and Fe^{2+} ions are constantly being converted into Ce^{3+} and Fe^{3+} ions, and vice versa. However, the respective concentrations of the four ions remain constant.

The relationship between the equilibrium concentrations of the reactants and those of the products for a chemical reaction at a given temperature is represented by the **equilibrium constant** (K_{eq}). The magnitude of K_{eq} is dependent on the temperature and on the chemical system involved.

For the general chemical reaction in Equation 7,

$$aA + bB \rightleftharpoons cC + dD \quad \text{(Eq. 7)}$$

the **equilibrium-constant expression** is:

$$K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b} \quad \text{(Eq. 8)}$$

The equilibrium-constant expression is the ratio of the molar concentrations of the products to those of the reactants. Each concentration term in the expression is raised to a power equal to the stoichiometric coefficient for that species in the chemical equation for the reaction. The concentrations of any pure solids or liquids involved in the reaction do not appear in an equilibrium-constant expression, because their concentrations do not vary.

The square brackets in Equation 8 represent the molar concentration (*M*, or moles per liter) of the substance within the brackets.

The equilibrium-constant expression for the reaction in Equation 6 is given in Equation 9.

$$K_{eq} = \frac{[Ce^{3+}][Fe^{3+}]}{[Ce^{4+}][Fe^{2+}]} \quad \text{(Eq. 9)}$$

Equation 9 indicates that, at a given temperature, K_{eq} for the reaction in Equation 6 is equal to the molar concentration of Ce^{3+} ion times the molar concentration of Fe^{3+} ion, divided by the molar concentration of Ce^{4+} ion times the molar concentration of Fe^{2+} ion. An equilibrium mixture, at a specified temperature, for Equation 6 can contain any combination of molar concentrations of Ce^{4+}, Ce^{3+}, Fe^{3+}, and Fe^{2+} ions, but the ratio of the product of the molar concentrations of the products to that of the concentrations of the reactants will always equal K_{eq}.

The equilibrium-constant expression and molar concentrations of reactants and products can be used to predict the direction of a chemical reaction. We can substitute the initial concentrations of the components of a chemical system into the equilibrium-constant expression for that system at a given temperature. Then, we can calculate the **reaction quotient** (*Q*) for that system. A chemical system is at equilibrium when the equilibrium constant equals the reaction quotient ($K_{eq} = Q$).

Under non-equilibrium conditions, *Q* does not equal K_{eq}. For instance, if a system at equilibrium is disturbed by a change in the concentration of one of the components, *Q* will no longer equal K_{eq}. If *Q* is less than K_{eq} ($Q < K_{eq}$), the relative amount of reactants is too large compared to that of products to satisfy the equilibrium constant. In such a case, the forward reaction proceeds more rapidly than the reverse reaction, until $Q = K_{eq}$. We say that the reaction proceeds from left to right.

On the other hand, if *Q* is greater than K_{eq} ($Q > K_{eq}$), the relative amount of products is too large compared to that of reactants to satisfy the equilibrium constant. In such a case, the reverse reaction proceeds more rapidly than the forward reaction, until $Q = K_{eq}$. We say that the reaction proceeds from right to left.

One qualitative approach to predicting the effect of a change on a system at equilibrium is based on **Le Châtelier's principle**. According to this principle, when any factor governing an equilibrium is changed, the chemical system responds to partially counteract or reduce the effect of the change, resulting in the establishment of a new equilibrium state. An **equilibrium state** is a set of concentrations that make *Q* equal the equilibrium constant for that chemical system at that

temperature. The individual concentrations of the components in the new equilibrium state will differ from those in the previous equilibrium state. However, K_{eq} or Q for the chemical system do not change.

We can disturb a system at equilibrium in several ways. For example, we can increase or decrease the concentration of one of the components. If we add a Ce^{4+} ion solution to the system represented by Equation 6, the system will no longer be at equilibrium ($Q \neq K_{eq}$). At the moment we add the solution, the magnitude of the denominator in the Q calculation is too large with respect to the magnitude of the numerator for the system to remain at equilibrium. Hence, Q will be less than K_{eq}, and the forward reaction will temporarily proceed faster than the reverse reaction, until the system reaches a new equilibrium state. When the rate of the forward reaction once again equals that of the reverse reaction, another new equilibrium state is reached. However, K_{eq} does not change, even though the individual equilibrium concentrations of the components have changed.

Similarly, if we remove some of the Fe^{3+} ion from the equilibrium mixture described by Equation 6, we will temporarily disturb the system. At the moment we remove the Fe^{3+} ion, the magnitude of the numerator in the Q calculation is too small with respect to the magnitude of the denominator for the system to remain at equilibrium. Hence, Q will be less than K_{eq}, and the forward reaction will temporarily proceed faster than the reverse reaction, until the system reaches a new equilibrium state. When the rate of the forward reaction once again equals that of the reverse reaction, the system has reached a new equilibrium state. Again, K_{eq} does not change, even though the individual equilibrium concentrations of the components have changed.

Most changes in chemical systems are accompanied by the transfer of heat. When we write a chemical equation for a reaction, we can include heat either as a reactant or a product, whichever is appropriate. A reaction in which heat is released is an **exothermic reaction**, and one in which heat is absorbed is an **endothermic reaction**. Another way to disturb a chemical system at equilibrium is to heat or cool the equilibrium mixture. Although heat does not appear in the equilibrium-constant expression, we can predict the effect of heating or cooling a system at equilibrium. To do so, we consider heat either as a reactant or product and apply Le Châtelier's principle.

For example, heat is absorbed during the dissolution of solid NH_4NO_3 in water, which forms ammonium ions (NH_4^+) and nitrate ions (NO_3^-), as shown in Equation 10. Because the forward reaction is endothermic, we can include heat as a reactant in the equation.

$$\text{heat} + NH_4NO_3(s) \rightleftharpoons NH_4^+(aq) + NO_3^-(aq) \qquad \text{(Eq. 10)}$$

If we heat the system at equilibrium described by Equation 10, the effect is as though we had increased the concentration of one of the reactants. Hence, the chemical system will respond by forming more NH_4^+ and NO_3^- ions until the system has reached a new equilibrium state. In such cases, we say that the equilibrium shifts to the right, meaning that more products form from the reactants. If we cool this new equilibrium mixture, the equilibrium will shift left, producing more reactant and heat. When we heat or cool a system at equilibrium, K_{eq} will change, because it is temperature-dependent.

In this experiment, you will study several systems at equilibrium, using microscale techniques. You will begin by studying the dissociation of a weak acid, hydrogen sulfate ion (HSO_4^-), in water. The net ionic equation for this reaction is shown in Equation 11.

$$HSO_4^-(aq, \text{colorless}) + H_2O(l, \text{colorless}) \rightleftharpoons H_3O^+(aq, \text{colorless}) + SO_4^{2-}(aq, \text{colorless}) \qquad \text{(Eq. 11)}$$

You will add Thymol Blue to the reaction mixture in order to detect changes in H_3O^+ ion concentration. In this system, Thymol Blue turns increasingly red as H_3O^+ ion concentration increases, and increasingly orange-yellow as H_3O^+ ion concentration decreases. You will observe the effects on the equilibrium when you add sodium sulfate solution (Na_2SO_4) and solid sodium hydrogen sulfate ($NaHSO_4$). Recall that both Na_2SO_4 and $NaHSO_4$ are strong electrolytes and essentially completely dissociate into their respective ions in solution.

In the second part of the experiment, you will study the dissociation equilibrium of a slightly soluble salt, magnesium hydroxide, $Mg(OH)_2$. You will prepare the equilibrium mixture, shown in Equation 12, by mixing magnesium chloride ($MgCl_2$) and NaOH solutions.

$$MgCl_2(aq) + 2\ NaOH(aq) \rightleftharpoons Mg(OH)_2(s, \text{white}) + 2\ NaCl(aq) \qquad \text{(Eq. 12)}$$

The net ionic equation for this reaction is given in Equation 13a, but for convenience, we will consider this reaction written as shown in Equation 13b.

$$Mg^{2+}(aq) + 2\ OH^-(aq) \rightleftharpoons Mg(OH)_2(s) \qquad \text{(Eq. 13a)}$$

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2\ OH^-(aq) \qquad \text{(Eq. 13b)}$$

You will observe the effect of adding hydrochloric acid (HCl), a source of H_3O^+ ion, to the equilibrium

shown in Equation 13b. The resulting reactions are shown in Equations 14 and 15.

$$HCl(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq) \quad \text{(Eq. 14)}$$

$$H_3O^+(aq) + OH^-(aq) \rightleftharpoons 2\,H_2O(l) \quad \text{(Eq. 15)}$$

Next, you will observe the effect of adding the tetrasodium salt of ethylenediamine tetraacetic acid (Na_4EDTA) to another sample of the equilibrium mixture. In solutions such as this one, containing relatively high concentrations of OH^- ion, $EDTA^{4-}$ ion forms strong complexes with Mg^{2+} ion, effectively removing the Mg^{2+} ion from the system.

You will also determine how the equilibrium shown in Equation 13b is affected by heating and cooling. Before doing this, you will add phenolphthalein to the equilibrium mixture. Phenolphthalein is an indicator that is colorless in solutions, except in the presence of excess OH^- ion, in which case it turns a bright pink.

Finally, you will study the equilibrium between tetrachlorocobalt(II) ion, $[CoCl_4]^{2-}$, and hexaaquacobalt(II) ion, $[Co(H_2O)_6]^{2+}$, as shown in Equation 16.

$$[CoCl_4]^{2-}(\text{alc, blue}) + 6\,H_2O(l, \text{colorless}) \rightleftharpoons [Co(H_2O)_6]^{2+}(\text{aq, pink}) + 4\,Cl^-(\text{aq, colorless}) \quad \text{(Eq. 16)}$$

By observing color changes, you will determine the effects on this equilibrium of adding H_2O, ethyl alcohol (C_2H_6O, or ethanol), silver nitrate ($AgNO_3$), or HCl. Recall that ethyl alcohol combines with H_2O to form relatively strong hydrogen bonds, and that silver ion (Ag^+) reacts with Cl^- ion to form a white precipitate of silver chloride (AgCl), as shown in Equation 17.

$$Ag^+(aq) + Cl^-(aq) \rightleftharpoons AgCl(s, \text{white}) \quad \text{(Eq. 17)}$$

Finally, again by observing color changes, you will determine how the system at equilibrium described in Equation 16 is affected by heating and cooling.

Procedure

Chemical Alert

0.1*M* cobalt(II) chloride—toxic and irritant
concentrated hydrochloric acid—toxic and corrosive
95% ethyl alcohol—flammable
0.5*M* Na_4EDTA—irritant
1*M* magnesium chloride—irritant
0.1*M* silver nitrate—toxic and oxidant
sodium hydrogen sulfate monohydrate—corrosive and moisture sensitive
0.1*M* sodium hydrogen sulfate—corrosive
0.5*M* sodium hydroxide—toxic and corrosive
1*M* sodium sulfate—irritant

Caution: Wear departmentally approved eye protection while doing this experiment.

Note: If Beral pipets are not being used for this experiment, your laboratory instructor will describe and demonstrate the use of the available pipets.

Note: The numbers appearing in parentheses indicate the specific lines on your Data Sheet on which the indicated data should be entered.

I. Preparing Water Baths

Caution: Take extreme care when using a hot-water bath. Avoid touching the hot beaker and the hot plate. Do not let the water boil.

1. Prepare a hot-water bath. Half fill a 150-mL beaker with tap water. Heat the beaker and water on a hot plate, adjusting the temperature control from time to time, so that the water will be hot by the time you are ready to do Step 11. Because you will use the hot-water bath throughout the experiment, check it periodically and replenish the water if the level becomes less than a quarter full.

2. Prepare an ice-water bath. Half fill a 150-mL beaker with ice. Add tap water until the ice-water combination fills half the beaker. Because you will use the ice-water bath throughout the experiment, check it periodically and add ice if all of the ice has dissolved.

II. Labeling Equipment

Note: Equipment can be labeled either by using a lead pencil on the etched spots on the glassware or by applying a gummed label or tape to the equipment and writing on the label or tape.

Your laboratory instructor will give you additional directions for labeling your well plate, test tubes, and Beral pipets.

3. The columns and rows of your 24-well plate may be identified with letters and numbers, as shown in Figure 1. If they are not identified, label them as shown in Figure 1. You can then use the combination of a letter and a number to quickly identify any specific well. Thus, Well *A*-1 refers to the well in the first column (1) of the first row (*A*).

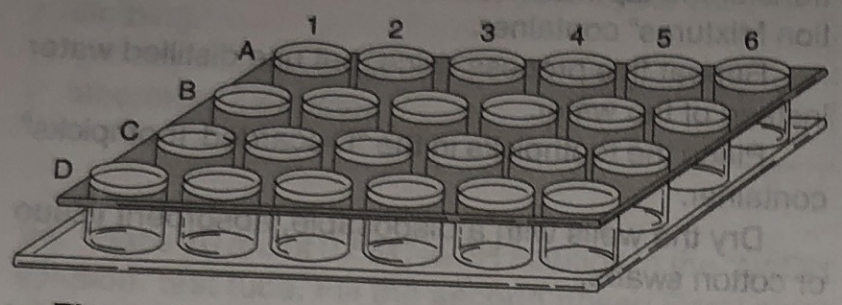

Figure 1 *Labeling a 24-well plate*

4. Label five clean, dry 15 × 125-mm test tubes as follows:

"$NaHSO_4$ solution" "Na_2SO_4 solution"
"$MgCl_2$ solution" "NaOH solution"
"$CoCl_2$ solution"

III. Studying the Equilibrium of a Weak Acid, HSO_4^- Ion, in Water

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq) \quad \text{(Eq. 11)}$$

Caution: Solid $NaHSO_4$ is a corrosive, moisture-sensitive substance, and $NaHSO_4$ solution is also corrosive. Prevent contact with your eyes, skin, or clothing.

If any solution is spilled, immediately notify your laboratory instructor.

5. In the test tube labeled "$NaHSO_4$ solution," place 5 mL of $NaHSO_4$ solution. Stand the test tube in a test tube support rack or a 400-mL beaker.

In the test tube labeled "Na_2SO_4 solution," place 5 mL of Na_2SO_4 solution. Stand this test tube in the rack or beaker.

6. Fill a clean, dry Beral pipet with $NaHSO_4$ solution. Add 10 drops of the solution to Wells *A*-1, *A*-2, and *A*-3. Add 20 drops of the solution to Wells *A*-4 and *A*-5. Place the pipet in the "$NaHSO_4$" test tube.

Note: To prevent contamination of solutions, you should mix them using a clean toothpick or small glass stirring rod. Take care to avoid scratching the bottom of the wells.

Another technique is to grasp the base of the well plate and gently move it in small circles on the laboratory bench. Take care to prevent the solutions from swirling so fast that they splash out of the wells.

A third approach is to draw the solution into a Beral pipet and ***gently*** expel it back into the well.

Your laboratory instructor will inform you of the mixing technique you are to use.

7. Add 1 drop of Thymol Blue indicator solution from a dropper bottle to the $NaHSO_4$ solution in each well. Thoroughly stir each mixture with a clean toothpick or small glass stirring rod. Leave the toothpicks in the wells. Observe the color of the mixture in each well. Record this color on your Data Sheet (1).

Note: In Steps 8 and 9, use the color of the solution in Well *A*-2 as a standard against which to compare the colors of the solutions in Wells *A*-1 and *A*-3.

8. Fill a clean, dry Beral pipet with Na_2SO_4 solution. Add one drop of Na_2SO_4 solution to the mixture in Well *A*-1. Place the pipet in the "Na_2SO_4" test tube.

Thoroughly mix the solution with the toothpick in the well. Compare the color of the mixture with that of the standard in Well *A*-2.

Continue to add drops of Na_2SO_4 solution to Well *A*-1 until you see a color change, stirring after each addition. Record your observations on your Data Sheet (2).

9. Using a micro spatula, add a crystal of $NaHSO_4$ to Well *A*-3. Stir the mixture to dissolve the crystal.

Continue adding individual crystals of $NaHSO_4$ until you see a color change. Record your observations on your Data Sheet (3).

10. Fill an empty 150-mL beaker with tap water. Rinse the Beral pipets you used in Steps 6 and 8 by filling each with tap water and expelling the water into the container labeled "Discarded HSO_4^- Reaction Mixtures." Repeat the rinse, discarding the water into the drain. Rinse each pipet twice with distilled or deionized water. Expel as much water as possible from each.

Note: Be cautious when squeezing Beral pipet bulbs. Any drops of solution in the pipet tip or stem can spurt out when the pipet is squeezed or inverted.

Caution: Beral pipets are heat sensitive. Do not touch the hot plate. The plastic will melt. Do not leave the pipets in hot water for prolonged periods.

11. Draw the mixture in Well *A*-4 into one of the rinsed Beral pipets. Carefully invert the pipet, allowing the solution to fill the pipet bulb. Hold the inverted pipet tip and snap the stem directly below the tip with your forefinger, as shown in Figure 2. Continue doing this until any solution in the stem drains into the bulb. Place this pipet, bulb down, in the hot-water bath.

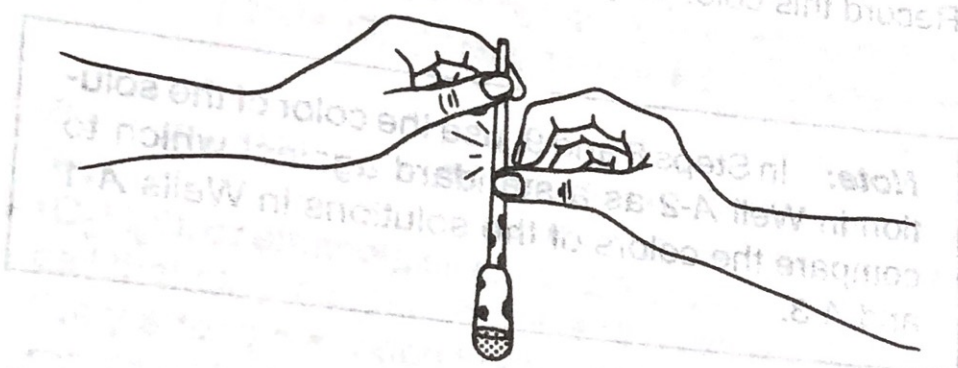

Figure 2 *Draining liquid from the stem into the bulb of a Beral pipet*

12. Draw the mixture in Well *A*-5 into the second rinsed pipet, following the procedure in Step 11. Place this pipet, bulb down, in the ice-water bath.

13. Occasionally swirl the pipet and contents in the hot-water bath to speed the heating process. Observe the color of the mixture in the pipet. Record any color change on your Data Sheet (4).

14. Occasionally swirl the pipet and contents in the ice-water bath to speed the cooling process. Observe the color of the mixture in the pipet. Record any color change on Data Sheet (5).

15. Exchange the pipets in the two baths.

16. Observe the color of the mixtures in the two pipets. Record any color changes on your Data Sheet (6, 7).

17. Transfer the mixtures in Step 16 to the "Discarded HSO_4^- Reaction Mixtures" container. Use one of the Beral pipets to transfer the contents of Wells *A*-1, *A*-2, and *A*-3 into the same container.

18. Fill Wells *A*-1 through *A*-5 with tap water. Thoroughly stir the water in each well. Use a Beral pipet to transfer the tap water to the "Discarded HSO_4^- Reaction Mixtures" container.

Repeat this process twice, but use distilled water instead of tap water.

Place the toothpicks in the "Discarded Toothpicks" container.

Dry the wells with a disposable, absorbent tissue or cotton swabs.

19. Discharge any solution remaining in the Beral pipets into the "Discarded HSO_4^- Reaction Mixtures" container.

20. Fill each of the Beral pipets with tap water, and carefully expel the water into the container mentioned in Step 19.

Repeat this procedure twice with distilled water. Remove as much water from the pipet as possible.

21. Dispose of the contents of your "$NaHSO_4$ solution" and "Na_2SO_4 solution" test tubes in the container labeled "Discarded $NaHSO_4^-$ Reaction Mixtures." Rinse the test tubes twice with tap water, pouring the rinse water into the same container. Rinse the test tubes twice with distilled water. Pour this rinse water into the drain.

Note: Your laboratory instructor will inform you whether you are to complete Data Sheet items 8–14 at this time, or if you are to immediately begin work on Part IV.

IV. Studying the Equilibrium of a Slightly Soluble Salt, $Mg(OH)_2$

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2\ OH^-(aq) \qquad \text{(Eq. 13b)}$$

22. Place 10 mL of 1.0*M* $MgCl_2$ solution in the "$MgCl_2$ solution" test tube. Fill one of the Beral pipets you rinsed in Step 20 with the $MgCl_2$ solution. Add 20 drops of the $MgCl_2$ solution to Wells *B*-1, *B*-2, *B*-3, and *B*-4. Place the pipet in the "$MgCl_2$ solution" test tube.

23. Add 1 drop of phenolphthalein solution from a dropper bottle to the solution in Wells *B*-3 and *B*-4.

Thoroughly stir each mixture with a separate toothpick. Leave the toothpicks in the wells.

> ***Caution:*** A 0.5*M* sodium hydroxide solution (NaOH) is corrosive, toxic, and can cause skin irritation. Prevent contact with your eyes, skin, or clothing.
>
> If any solution is spilled, immediately notify your laboratory instructor.

24. Place 10 mL of 0.5*M* NaOH solution in the "NaOH solution" test tube. Fill the second rinsed Beral pipet with NaOH solution. Add 10 drops of NaOH solution to the $MgCl_2$ solution in Wells *B*-1, *B*-2, *B*-3, and *B*-4. Thoroughly stir the mixtures with toothpicks and leave the toothpicks in the wells. Record the appearance and color of the mixtures in Wells *B*-1, *B*-2, *B*-3, and *B*-4 and the presence or absence of a precipitate on your Data Sheet (15, 18).

> ***Caution:*** Concentrated HCl is a very toxic, corrosive substance. Prevent eyes, skin, and clothing contact. Avoid inhaling the vapors or ingesting the substance.
>
> If you spill any acid, immediately notify your laboratory instructor.

25. Add ***one*** drop of concentrated HCl from the dropper bottle to the mixture in Well *B*-1. Stir the mixture. Continue to add HCl dropwise until you see a color change. Record your observations on your Data Sheet (16).

26. Add ***one*** drop of Na_4EDTA solution from the dropper bottle to the mixture in Well *B*-2. Stir the mixture. Continue to add Na_4EDTA solution dropwise until you see a color change. Record your observations on your Data Sheet (17).

27. Draw the solution and as much of the precipitate as possible from Well *B*-3 into the bulb of the "$MgCl_2$ solution" pipet. Carefully expel the solution back into the well in order to mix it.

Following the procedure in Step 11, draw the mixed solution and precipitate in Well *B*-3 into the pipet bulb and invert the pipet. Place the pipet, bulb down, in your hot-water bath.

28. Repeat Step 27, drawing the solution and precipitate from Well *B*-4 into a rinsed Beral pipet. Place this pipet, bulb down, in your ice-water bath.

29. Allow each pipet to remain in its water bath for 3 or 4 min, allowing the temperature of the mixture in the pipet to approach that of the water bath. Swirl the pipets and contents occasionally to speed the heating and cooling processes. Observe the mixtures and record your observations of each mixture on your Data Sheet (19, 20).

30. Exchange the two pipets in the water baths. Allow the pipets to remain in their new respective water baths for 3 or 4 min. Observe the mixtures carefully and record your observations on your Data Sheet (21, 22).

31. Expel the contents of the pipets used in Step 30 into the container labeled "Discarded $Mg(OH)_2$ Reaction Mixtures". Use one of your Beral pipets to transfer the contents of Wells *B*-1 and *B*-2 into the same container.

32. Fill Wells *B*-1, *B*-2, *B*-3, and *B*-4 with tap water. Thoroughly stir the water. Use a Beral pipet to transfer the tap water into the container mentioned in Step 31.

Repeat this process twice, but use distilled water instead of tap water. Place the toothpicks in the "Discarded Toothpicks" container. Dry the wells with a disposable, absorbent tissue or cotton swabs.

33. Discharge any solution remaining in the Beral pipets into the "Discarded $Mg(OH)_2$ Reaction Mixtures" container.

Fill your Beral pipets with tap water and carefully expel the water into the same container.

Repeat this procedure twice with distilled water. Remove as much water from each pipet as possible.

34. Dispose of the contents of your "$MgCl_2$ solution" test tube in the "Discarded $Mg(OH)_2$ Reaction Mixtures" container. Rinse the test tube twice with tap water, pouring the rinse water into the same container. Rinse the test tube twice with distilled water. Pour this rinse water into the drain.

> ***Note:*** Your laboratory instructor will inform you whether you are to complete Data Sheet items 23–29 at this time, or if you are to immediately begin work on Part V.

V. Studying the Equilibrium of the Hexaaquacobalt(II) Complex Ion

$$[CoCl_4]^{2-}(alc, blue) + 6\ H_2O(l, colorless) \rightleftharpoons [Co(H_2O)_6]^{2+}(aq, pink) + 4\ Cl^-(aq, colorless) \quad \text{(Eq. 16)}$$

> ***Note:*** In Steps 35, 39, and 41, the Beral pipet used to transfer the $CoCl_2$ solution must be ***absolutely clean and dry*** in order to obtain acceptable results.

35. Place 10 mL of 0.10*M* cobalt(II) chloride solution ($CoCl_2$) in the clean, ***dry*** "$CoCl_2$ solution" test tube. Using a clean, ***dry*** Beral pipet, add 25 drops of the $CoCl_2$ solution to Well *C*-1. Record the initial color of the mixture on your Data Sheet (30).

36. Using a clean, rinsed Beral pipet, add ***one*** drop of distilled or deionized water to the mixture in Well *C*-1. Thoroughly stir the mixture with a clean, dry toothpick, and leave the toothpick in the well. Continue to add water dropwise, stirring the mixture after each drop, ***just*** until you see a color change. Record the color change on your Data Sheet (31).

> ***Caution:*** Concentrated HCl is a very toxic, corrosive substance. Prevent eyes, skin, and clothing contact. Avoid inhaling the vapors or ingesting the substance.
>
> If you spill any acid, immediately notify your laboratory instructor.

37. Add ***one*** drop of concentrated HCl from the dropping bottle to the mixture in Well *C*-1. Stir the mixture. Continue to add HCl dropwise, stirring after each drop, until you observe a color change. Record your observations on your Data Sheet (32).

> ***Caution:*** A 0.1*M* $AgNO_3$ solution is toxic and can cause skin burns or discoloration. Prevent eyes, skin, or clothing contact. If any of the solution touches your skin, immediately wash with copious amounts of water, and then notify your laboratory instructor.

38. Add ***one*** drop of $AgNO_3$ solution from the dropper bottle to the mixture in Well *C*-1. Stir the mixture. Continue to add $AgNO_3$ solution dropwise until you observe a color change. Record your observations on your Data Sheet (33).

39. Using a second clean, ***dry*** Beral pipet, half fill Well *C*-2 with $CoCl_2$ solution. Add ***one*** drop of distilled water to the mixture. Stir the mixture. Continue to add water dropwise ***just*** until the solution has changed color. Record the color change on your Data Sheet (34).

> ***Caution:*** Absolute ethyl alcohol is flammable and should not be used near open flames.

40. Add ***one*** drop of absolute ethyl alcohol from the dropper bottle to the mixture in Well *C*-2. Stir the mixture. Continue adding ethyl alcohol dropwise until you observe a color change. Record your observations on your Data Sheet (35).

41. Using a clean, ***dry*** Beral pipet, half fill Well *C*-3 with $CoCl_2$ solution. Add ***one*** drop of distilled water to the mixture. Stir the mixture. Continue to add water dropwise ***just*** until the solution has changed color. Record the color change on your Data Sheet (36).

> ***Caution:*** Any liquid in the stem of a Beral pipet will spurt out of the tip when the pipet bulb is warmed. Do not allow this liquid to touch your skin, eyes, clothing, or the hot plate. The alcohol solvent is flammable.

42. Draw all of the mixture in Well *C*-2 into a clean, ***dry*** Beral pipet. Invert the pipet so that the solution fills the pipet bulb. Follow the procedure in Step 11 to drain any solution in the pipet stem into the bulb. Place the pipet, bulb down, in your hot-water bath. Swirl the pipet and contents occasionally to speed the heating process. Allow the pipet to remain in the hot-water bath until you observe a color change. Record your observations on your Data Sheet (37).

43. Holding the pipet stem, carefully remove the pipet from the hot-water bath. Turn off your hot plate.

Place the pipet, bulb down, in the ice-water bath. Swirl the pipet and contents occasionally to speed the cooling process. Allow the pipet to remain in the ice-water bath until you observe a color change. Record your observations on your Data Sheet (38).

44. Use one of your Beral pipets to transfer the mixtures in Wells *C*-1 and *C*-3 into the "Discarded $[Co(H_2O)_6]^{2+}$ Reaction Mixtures" container.

45. Wash, rinse, and dry Wells *C*-1, *C*-2, and *C*-3, following the procedure in Step 32. Discard the toothpicks in the "Discarded Toothpicks" container.

46. Discharge any solutions remaining in your Beral pipets into the "Discarded $[Co(H_2O)_6]^{2+}$ Reaction Mixtures" container.

47. Fill your Beral pipets with tap water and carefully expel the water into the container mentioned in Step 46. Repeat this procedure twice with distilled water. Remove as much water from each pipet as possible.

Do not discard the Beral pipets.

48. Dispose of the contents of your "$CoCl_2$ solution" test tube in the "$[Co(H_2O)_6]^{2+}$ Reaction Mixture" container. Rinse the test tube twice with tap water, pouring the rinse water into the same container. Rinse the test tube once with distilled water. Pour the rinse water into the drain.

Note: Your laboratory instructor will inform you whether you are to complete Data Sheet items 39–49 before you leave the laboratory.

Caution: Wash your hands thoroughly with soap or detergent before leaving the laboratory.

name _____ section _____ date _____

Data Sheet

III. Studying the Equilibrium of a Weak Acid, HSO_4^- Ion, in Water

(1) What is the initial color of the mixtures in Wells *A*-1, *A*-2, *A*-3, *A*-4, and *A*-5?

clear

(2) What color change do you observe when you add Na_2SO_4 solution to Well *A*-1?

Redish Pink

(3) What color change do you observe as you add solid $NaHSO_4$ to Well *A*-3?

Purple

(4) What color change do you observe when you heat the mixture from Well *A*-4 in the hot-water bath?

light pink

(5) What color change do you observe when you cool the pipet containing the mixture from Well *A*-5 in the ice-water bath?

Neon pink

(6) What color change do you observe when you cool the pipet containing the heated mixture from Well *A*-4 in the ice-water bath?

Neon purple

(7) What color change do you observe when you heat the cooled mixture from Well *A*-5 in the hot-water bath?

light pink

(8) Write the net ionic equation for the equilibrium involving HSO_4^- ion in this experiment.

HSO_4^- (aq colorless) + H_2O (l, colorless) $\rightleftharpoons$ H_3O^+ (aq) colorless) + SO_4^{2-} (aq colorless)

(9) Predict the direction in which the equilibrium should shift when you add Na_2SO_4 solution to the mixture in Well *A*-1. Briefly explain.

(10) Briefly explain how your observations do or do not support your prediction in (9).

(11) Predict the direction in which the equilibrium should shift when you add solid $NaHSO_4$ to the mixture in Well *A*-3. Briefly explain. addition of solid $NaHSO_4$ would shift the equilibrium towards right hand side, that is more SO_4^{2-} would be formed by using excess of $NaHSO_4$ added to the solution. This is according to the Le Chatelier's principle which states that any

(12) Briefly explain how your observations do or do not support your prediction in (11).

(13) Based upon your observations of heating and cooling the mixtures in Wells *A*-4 and *A*-5, is the forward reaction for the HSO_4^- ion equilibrium endothermic or exothermic? Briefly explain.

(14) Write the net ionic equation for the HSO_4^- ion equilibrium again, but include heat in the equation this time.

IV. Studying the Equilibrium of a Slightly Soluble Salt, $Mg(OH)_2$

(15) What is the initial color and appearance of the mixtures in Wells *B*-1 and *B*-2? If a precipitate is present, note that fact, as well as the color of the precipitate.

clear

(16) What color change do you observe after adding concentrated HCl to Well *B*-1?

no change

(17) What color change do you observe after adding Na_4EDTA solution to Well *B*-2?

no change

(18) What is the initial color and appearance of the mixtures in Wells *B*-3 and *B*-4? If a precipitate is present, note that fact and the color of the precipitate.

Neon pink

(19) What color change do you observe when you heat the mixture from Well *B*-3?

light pink

(20) What color change do you observe when you cool the pipet containing the heated mixture from Well *B*-4?

dark pink

(21) What color change do you observe when you cool the pipet containing the heated mixture from Well *B*-3?

became darker

(22) What color change do you observe when you heat the pipet containing the cooled mixture from Well *B*-4?

became lighter

(23) Write the net ionic equation for the equilibrium involving $Mg(OH)_2$ in this experiment.

(24) Predict the direction in which the equilibrium should shift when you add concentrated HCl to the mixture in Well *B*-1. Briefly explain.

(25) Briefly explain how your observations do or do not support your prediction in (24).

(26) Predict the direction in which the equilibrium should shift when you add Na_4EDTA solution to the mixture in Well *B*-2. Briefly explain.

(27) Briefly explain how your observations do or do not support your prediction in (26).

(28) Based on your observations of heating the mixture from Well *B*-3, is the forward reaction endothermic or exothermic? Briefly explain.

(29) Write the net ionic equation for this equilibrium again, but include heat in the equation this time.

name _____ section _____ date _____

V. Studying the Equilibrium of the Hexaaquacobalt(II) Complex Ion

(30) What was the initial color of the mixture in Well *C*-1?

blue

(31) What color change do you observe when you add H_2O to the mixture in Well *C*-1?

light pink

(32) What color change do you observe when you add concentrated HCl to the mixture in Well *C*-1?

light blue

(33) What color change do you observe when you add 0.1*M* $AgNO_3$ solution to the mixture in Well *C*-1?

light pink

(34) What color change do you observe when you add H_2O to the mixture in Well *C*-2?

light pink

(35) What color change do you observe when you add ethyl alcohol to the mixture in Well *C*-2?

light pink

(36) What color change do you observe when you add H_2O to the mixture in Well *C*-3?

light pink

(37) What color change do you observe when you heat the mixture from Well *C*-2?

no change

(38) What color change do you observe when you cool the warm mixture from Well *C*-2?

no change

(39) Write the net ionic equation for the equilibrium involving the hexaaquacobalt(II) ion in this experiment.

(40) Predict the direction in which the equilibrium should shift when you add H_2O to the mixture in Well *C*-1. Briefly explain.

(41) Briefly explain how your observations do or do not support your prediction in (40).

(42) Predict the direction in which the equilibrium should shift when you add concentrated HCl to the mixture in Well *C*-1. Briefly explain.

(43) Briefly explain how your observations do or do not support your prediction in (42).

(44) Predict the direction in which the equilibrium should shift when you add $AgNO_3$ solution to the mixture in Well *C*-1. Briefly explain.

(45) Briefly explain how your observations do o
do not support your prediction in (44).

(46) Predict the direction in which the equilibrium should shift when you add ethyl alcohol to the mixture in Well *C*-2. Briefly explain.

(47) Briefly explain how your observations do or do not support your prediction in (46).

(48) Based on your observations of heating the mixture from Well *C*-2, is the forward reaction endothermic or exothermic? Briefly explain.

(49) Write the net ionic equation for this equilibrium again, but include heat in the equation this time.