Properties of Systems in Equilibrium – Le Châtelier’s Principle

Objectives

- To perturb chemical reactions at equilibrium and observe how they respond.
- To explain these observations using Le Châtelier’s Principle.
- To relate Le Châtelier’s Principle to the concept of coupled reactions.

Background

All chemical reactions eventually reach a state in which the rate of the reaction in the forward direction is equal to the rate of the reaction in the reverse direction. When a reaction reaches this state, it is said to be at chemical equilibrium.

The concentrations of reactants and products at equilibrium are constant as a function of time. Thus, for a homogeneous aqueous system of the form

\[ aA \text{ (aq)} + bB \text{ (aq)} \rightleftharpoons cC \text{ (aq)} + dD \text{ (aq)} \]  

we can express the equilibrium-constant expression for this reaction as,

\[ K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b} \]  

where the values of [A], [B], [C] and [D] correspond to the equilibrium concentrations (or equilibrium positions) of all the aqueous chemical components, and a, b, c and d are their respective stoichiometric coefficients. Note that for a heterogeneous system including pure solids or liquids of the form

\[ aA \text{ (aq)} + bB \text{ (s)} \rightleftharpoons cC \text{ (aq)} + dD \text{ (l)} \]  

that the pure liquids and solids do not appear in the equilibrium-constant expression:

\[ K_c = \frac{[C]^c}{[A]^a} \]  

It has been observed that when a reaction at equilibrium is perturbed by applying a stress, that the reaction will respond by shifting its equilibrium position so as to counteract the effect of the perturbation/stress. In other words, the concentrations of the reactants and products will shift so that the relationship described by Equation (2) is again satisfied. This idea was first proposed by Henri-Louis Le Châtelier and has since been referred to as, “Le Châtelier’s Principle”.

Note that when a reaction makes more products as a response to the perturbation, we call it a right-shift. When a reaction makes more reactants in response to the perturbation, we call it a left-shift. We often designate these respective shifts by drawing right and left arrows below the chemical equation.
For chemical reactions at equilibrium in aqueous solution, the most common types of perturbations include changing the concentration of one of the chemical components, or the total solution volume, or the temperature. The general responses of an aqueous system to these particular perturbations are tabulated below.

<table>
<thead>
<tr>
<th>Perturbation</th>
<th>Effect on Equilibrium Position</th>
<th>Effect on $K_c$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Increase reactant concentration, or, decrease product concentration</td>
<td>Shift to the right</td>
<td>None</td>
</tr>
<tr>
<td>Decrease reactant concentration, or, increase product concentration</td>
<td>Shift to the left</td>
<td>None</td>
</tr>
<tr>
<td>Increase solution volume by addition of solvent</td>
<td>Shift towards more solute particles</td>
<td>None</td>
</tr>
<tr>
<td>Increase temperature of an exothermic reaction</td>
<td>Shift to the left</td>
<td>Decrease</td>
</tr>
<tr>
<td>Decrease temperature of an exothermic reaction</td>
<td>Shift to the right</td>
<td>Increase</td>
</tr>
<tr>
<td>Increase temperature of an endothermic reaction</td>
<td>Shift to the right</td>
<td>Increase</td>
</tr>
<tr>
<td>Decrease temperature of an endothermic reaction</td>
<td>Shift to the left</td>
<td>Decrease</td>
</tr>
<tr>
<td>Addition of an inert substance, catalyst, pure liquid, or pure solid</td>
<td>None</td>
<td>None</td>
</tr>
</tbody>
</table>

Notice that only a temperature change can affect the value of $K_c$; in all other cases the value of $K_c$ remains constant.

In this experiment you will perturb reactions that have reached equilibrium. You will then observe how each reaction responds to that perturbation. In your report you describe these changes in terms of Le Châtelier’s Principle.

**Part A – Acid-Base Equilibrium**

In this part of the experiment you will make use of coupled equilibria to change the equilibrium position of an acid-base reaction.

Let’s see how such coupled equilibria work. Suppose we have the two reactions described by the chemical equations below:

$$A (aq) \rightleftharpoons B (aq)$$  \hspace{1cm} (5)

$$B (aq) + C (aq) \rightleftharpoons D (aq)$$  \hspace{1cm} (6)

Notice that the species $B (aq)$ is common to both reactions. The presence of this common species couples these two reactions.

We can perturb the equilibrium position of Reaction (6) by the addition of some $C (aq)$. The addition of $C (aq)$ will cause the equilibrium position of Reaction (6) to shift right in accordance with Le Châtelier’s Principle. This right shift in the equilibrium position of Reaction (6) will also result a corresponding decrease in the concentration of $B (aq)$. Because $B (aq)$ is also present in Reaction (5), the decrease in the concentration of $B (aq)$ will in turn result in a right shift in the equilibrium position of Reaction (5). Thus, the addition of $C (aq)$ to Reaction (6) actually results in a right shift in the equilibrium position of Reaction (5) because these two equilibria are
coupled. Many of the other reactions that we observe in this experiment will also involve the use of coupled equilibria.

In Part A of this experiment we will observe the effect of changing various chemical species on an acid-base reaction at equilibrium. The equilibrium system can be written in the general form

$$\text{HA} (aq) \rightleftharpoons \text{H}^+ (aq) + \text{A}^- (aq)$$  (7)

The equilibrium-constant expression for this reaction is

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$  (8)

where we denote the equilibrium constant, $K$, with a subscript $a$ for acid. In your experiment, HA and A–, are the acidic and basic forms of the pH indicator bromothymol blue. Since the two forms are different colors, this will allow you to determine which form is present in the reaction mixture.

Your goal in this part of the experiment is to find a reagent that will shift the equilibrium position of this acid-base equilibria in one direction, and then another reagent that will shift it back. Instead of directly adding acid and base to the system, you will use coupled equilibria to effect these shifts. Specifically, you will effect the equilibrium position of Reaction (7) by using a second chemical reaction that shares a common species with this one. The second equation we shall make use of is the autoionization of water, which can be described by the equation

$$\text{H}_2\text{O} (l) \rightleftharpoons \text{H}^+ (aq) + \text{OH}^- (aq)$$  (9)

The equilibrium constant for this reaction is denoted by $K_w$, where the subscript $w$ stands for water. The value of $K_w$ at 25ºC is given by

$$K_w = [\text{H}^+][\text{OH}^-]$$  (10)

Because Reactions (7) and (9) share a common chemical species (H+), you can use the concept of coupled equilibria to shift the equilibrium in Reaction (7) by increasing or decreasing the concentration of OH– (aq).

**Part B – Solubility Equilibrium**

In this part of the experiment you will test the effects of changing temperature and volume on the solubility of an insoluble salt at equilibrium. Some examples of insoluble salts are AgCl (s), Cu(OH)₂ (s), PbCl₂ (s), and Fe₂S₃ (s), which you should recall are, “insoluble in water,” according to the solubility rules you learned in Chemistry 11. In fact, a very small amount of each of these substances does dissolve in aqueous solution, but the amount is so small that we classify each of these compounds as, “insoluble”.

This type of equilibrium is often called a solubility equilibrium because it is written in the direction of the dissolution of the solid, as shown in the following example:

$$\text{A}_x\text{B}_y (s) \rightleftharpoons x\text{A}^+ (aq) + y\text{B}^- (aq)$$  (11)
The equilibrium-constant expression for Reaction (11) is

\[ K_{sp} = [A^{+}]^x[B^{-}]^y \]  

(12)

where we denote the equilibrium constant, \( K \), with a subscript \( sp \) for solubility product.

Now let’s consider the process of precipitation. In a typical precipitation reaction two aqueous salt solutions are mixed together resulting in the production an insoluble salt. Notice that this process corresponds to a left shift of Reaction (11), and so Equation (12) can also be used to examine the conditions required for the precipitation of a solid to occur. We can denote the product \([A^{+}]^x[B^{-}]^y\) under any conditions (not necessarily at equilibrium) as,

\[ Q_{sp} = [A^{+}]^x[B^{-}]^y \]  

(13)

where \( Q_{sp} \) is called the solubility product reaction quotient. Note that, upon mixing two solutions, one containing \( A^{+} \) and the other containing \( B^{-} \), if \( Q_{sp} < K_{sp} \) the system is not at equilibrium, but since no solid \( A_B \) is present the reaction cannot shift to the right and therefore no reaction will be observed. On the other hand if \( Q_{sp} > K_{sp} \) the solution contains an excess of aqueous species, and so Reaction (11) will shift left, forming the solid precipitate \( A_B \) until the system reaches a state of equilibrium where \( Q_{sp} = K_{sp} \). Thus, we can use the values of \( Q_{sp} \) and \( K_{sp} \) to predict the conditions under which a precipitation reaction will occur.

In Part B of this experiment we will study the solubility equilibrium of \( \text{PbCl}_2 \) (s). We will observe how its solubility equilibrium is effected by changes in solution volume (or quantity of solvent) and temperature. We will express these changes in terms of their respective values of \( Q_{sp} \) and \( K_{sp} \).

**Part C – Complex Ion Equilibrium**

Certain metal ions, most often transition metals, exist in solution as complex ions in combination with other ions or molecules, called ligands. Common ligands include \( \text{H}_2\text{O}, \text{NH}_3, \text{Cl}^- \) and \( \text{OH}^- \). Many of these complex ions exhibit vibrant colors in solution. For example, the \( \text{Co(H}_2\text{O)}_6^{2+} \) \( (aq) \) complex ion is pink and the \( \text{CoCl}_4^{2-} \) \( (aq) \) complex ion is blue.

In Part C of this experiment you will study the following complex ion formation reaction:

\[
\text{Co(H}_2\text{O)}_6^{2+} \ (aq) \ + \ 4 \ \text{Cl}^- \ (aq) \ \rightleftharpoons \ \text{CoCl}_4^{2-} \ (aq) \ + \ 6 \ \text{H}_2\text{O} \ (l) \ 
\]  

(14)

The equilibrium-constant expression for Reaction (15) is

\[
K_f = \frac{[\text{CoCl}_4^{2-}]}{[\text{Co(H}_2\text{O)}_6^{2+}][\text{Cl}^-]^4} \ 
\]  

(15)

where we denote the equilibrium constant, \( K \), with a subscript \( f \) for complex ion formation.

Your goal in Part C of this experiment is to observe how the Reaction (14) shifts from its equilibrium position as the result of various perturbations.
Part D – Dissolving Insoluble Solids

In this part of the experiment you will use coupled equilibria to affect the solubility equilibrium of the insoluble salt, Zn(OH)\(_2\) (s), whose solubility equilibrium reaction can be described by the equation

\[
\text{Zn(OH)}_2 (s) \rightleftharpoons \text{Zn}^{2+} (aq) + 2 \text{OH}^- (aq) \quad K_{sp} = 5 \times 10^{-17} \text{ M}^3 \tag{16}
\]

Notice that \(K_{sp} \ll 1\) for this reaction, confirming that Zn(OH)\(_2\) (s) is indeed an essentially insoluble salt in aqueous solution.

Now consider the reactions described by the following chemical equations, each of which shares a common species with the Reaction (16):

\[
\text{H}_2\text{O} (l) \rightleftharpoons \text{H}^+ (aq) + \text{OH}^- (aq) \quad K_w = 1 \times 10^{-14} \text{ M}^2 \tag{17}
\]

\[
\text{Zn}^{2+} (aq) + 4 \text{OH}^- (aq) \rightleftharpoons \text{Zn(OH)}_4^{2-} (aq) \quad K_f = 3 \times 10^{15} \text{ M}^{-4} \tag{18}
\]

\[
\text{Zn}^{2+} (aq) + 4 \text{NH}_3 (aq) \rightleftharpoons \text{Zn(NH}_3)_4^{2+} (aq) \quad K_f = 1 \times 10^9 \text{ M}^{-4} \tag{19}
\]

Because Reactions (17), (18), and (19) each share a common species with Reaction (16) they can be coupled together. Thus, in Part D of this experiment you will observe the effect of coupling each of these equilibria on the solubility of Zn(OH)\(_2\) (s).
Procedure

Equipment

From lab/locker: Six large test tubes, test tube rack, stirring rod, scoopula, small 10-mL graduated cylinder, large 100-mL graduated cylinder, 400-mL beaker, wire gauze, stand with ring clamp, and Bunsen burner.

From stockroom: Bucket of ice

Chemicals

Bromothymol blue, a 6 M strong acid, a 6 M strong base, 0.3 M Pb(NO₃)₂ (aq), 0.3 M HCl (aq), CoCl₂•6H₂O (s), 12 M HCl (aq), 0.1 M Zn(NO₃)₂ (aq), 0.1 M Mg(NO₃)₂ (aq), 6 M NaOH (aq), 6 M HCl (aq), 6 M NH₃ (aq) (often labeled as NH₄OH), and deionized water

Safety and Waste Disposal

1. 12-M HCl is extremely caustic and great care must be taken to avoid contact with eyes or skin. The bottle should be kept in a plastic tray and not removed from the fume hood. Should any of this solution enter your eyes rinse immediately in the emergency eyewash. Should any of this solution come in contact with your skin rinse with copious amounts of water and apply saturated sodium bicarbonate to the affected area from the stock bottle located on the sink.

2. The solutions you will use in Part B contain lead. Be certain that all of these lead-containing solutions are disposed of in the proper waste container and rinse your hands following this procedure.

3. Many of the chemicals used in this lab are hazardous to the environment. All waste must be disposed of in the hazardous-waste container in the fume hood. Rinse all glassware directly into the waste container twice using a small squirt bottle to be certain all hazardous waste ends up in the waste container.

General Procedural Notes

The amounts of reagents used in this experiment are approximate only. If you are unsure how to estimate a milliliter, then measure out about one milliliter of water using your graduated cylinder, transfer this amount to a large test tube, and then use this approximate volume as a reference throughout the experiment.

All glassware needs to be rinsed at least once with deionized water but not dried because all reagent volumes are approximate and all solutions used for the experiments are aqueous.
Part A: Acid-Base Equilibrium

In this part of the experiment, you will find a reagent that will shift the acid-base equilibrium given by Reaction (7) in one direction and then a second reagent that will cause the equilibrium position to shift back in the opposite direction.

Reagents needed for this part are: deionized water, bromothymol blue solution, a 6 M strong acid and a 6 M strong base.

1. Add approximately 5 mL of deionized water to a large test tube. Add 3 drops of the bromothymol blue indicator solution. Report the color of your solution on your data sheet.

2. Your solution from Step 1 currently contains one form of bromothymol blue (see background). Now predict which of the two 6 M reagents you obtained, the strong acid or the strong base, will cause a color change in your solution by making the bromothymol blue indicator shift to its other form. Add the 6 M reagent of your choice drop-by-drop and if your solution changes color, write the color of the solution and formula of the reagent on your data sheet. If the addition of your reagent does not result in a color change, try other reagents until you are successful.

3. Since equilibrium systems are reversible, it is possible to shift a reaction left or right repeatedly by changing the conditions. Now find another 6 M reagent that will cause your solution from Step 2 to revert back to its original color. Add the 6 M reagent of your choice drop-by-drop and if your solution changes color, write the formula of the reagent on the data sheet. If the addition of your second reagent does not result in a color change, try other reagents until you are successful.

Part B: Solubility Equilibrium

In this part of the experiment you will test the effects of changing the solution volume and temperature on the solubility equilibrium of lead(II) chloride in aqueous solution.

Reagents needed for this part are: 0.3 M Pb(NO₃)₂ (aq), 0.3 M HCl (aq), and deionized water.

1. Set up a hot-water bath by filling a 400-mL beaker about half-full with tap water and heating the water using a Bunsen burner while you work on the following steps. The hot-water bath will be used in Step 6 (and also in Part C).

2. Measure out 5.0 mL of the 0.3 M Pb(NO₃)₂ solution into a large test tube using a 10-mL graduated cylinder.

3. Thoroughly rinse the graduated cylinder with deionized water, then measure out 5.0 mL of the 0.3 M HCl solution using the 10-mL graduated cylinder.

4. Add about 1 mL of the 0.3 M HCl solution to the Pb(NO₃)₂ solution in the large test tube. Stir the mixture gently using your stirring-rod and record your observations on your data sheet.

5. Continue to add the 0.3 M HCl solution to the Pb(NO₃)₂ solution in the large test tube in roughly 1 mL increments until you begin to see white PbCl₂ solid appear in your test tube. To confirm that the solid is present, let the test tube sit on the bench for about 3 minutes,
which will allow all solid to fall to the bottom and make it easier to see. Record the total volume of 0.3 M HCl added to produce the solid on your data sheet.

6. Put the test tube containing the solid into your hot-water bath from Step 1. Stir the contents of the test tube gently for a few seconds using the glass stirring-rod and record your observations on your data sheet. Continue heating and stirring until a change is observed.

7. Make a cold-water bath by filling a 400-mL beaker half full with tap water and ice. Cool the test tube down by placing it into this cold-water bath. Observe what happens and record your observations. Save this cold-water bath for use in Part C.

8. Dispose of the remainder of the 0.3 M HCl solution in your 10-mL graduated cylinder by pouring it into the appropriate waste container and then rinse the cylinder using deionized water. Then measure 5.0 mL of deionized water using this graduated cylinder.

9. Add about 1 mL of deionized water to the large test tube containing the PbCl₂ solid. Stir the mixture and observe what happens. Continue to add the deionized water in 1 mL increments until the white PbCl₂ solid just dissolves (or disappears). Record the volume of water needed to do this on your data sheet.

10. Pour the contents of the large test tube into your large graduated cylinder and measure the total volume of the solution. Record this volume on your data sheet.

11. Dispose of all solutions used in this part of the experiment in the proper waste container. Rinse all glassware twice using a squirt bottle of deionized water to be certain all of the lead solution has been removed and transferred to the waste container. Save the hot- and cold-water baths for Part C.

Part C: Complex Ion Equilibrium

In this part of the experiment you will test the effects of changing the volume and temperature on the complex ion equilibrium between Co(H₂O)₆²⁺ (aq) and CoCl₄²⁻ (aq) as in Reaction (14).

Reagents needed for this part are: CoCl₂•6H₂O (s), 12 M HCl (aq) (do not remove from fume hood), and deionized water.

1. Reheat your hot-water bath from part B of this laboratory to a near boil. The hot-water bath will be used in Step 4.

2. Put a few small crystals (an amount that fits at the tip of the scoopula) of CoCl₂•6H₂O (s) in a large test tube. Working in the fume hood, carefully add 2 mL of 12 M HCl (aq) (this acid is extremely caustic; see cautions in the safety section of this experiment). Stir to dissolve the crystals. Record the color of this solution.

3. Using a 10-mL graduated cylinder, add deionized water to the solution in your test tube in approximately 2 mL increments, stirring after each addition until no further color change occurs. Record the new color.

4. Place the test tube into the hot-water bath from Step 1 and record any color change.

5. Cool the solution down in the cold-water bath and record any color change.
Part D: Dissolving Insoluble Solids

In this part of the experiment you will further examine how one reaction can affect the behavior of another reaction when the two share one or more common chemical species.

Reagents needed for this part are: 0.1 M Zn(NO₃)₂ (aq), 0.1 M Mg(NO₃)₂ (aq), 6 M NaOH (aq), 6 M HCl (aq), and 6 M NH₃ (aq) (often labeled as NH₄OH).

1. Label three large test tubes A, B, and C. Add about 2 mL of 0.1 M Zn(NO₃)₂ solution to each test tube. Add one drop of 6 M NaOH solution to each test tube. Stir each solution and record your observations.

2. Into test tube A, add at least 20 drops of 6 M HCl (aq) drop-by-drop while stirring. Record all observations.

3. Into test tube B, add at least 20 drops of 6 M NaOH (aq) drop-by-drop while stirring. Record all observations.

4. Into test tube C, add at least 20 drops of 6 M NH₃ (aq) drop-by-drop while stirring. Record all observations.

5. Label three additional large test tubes D, E, and F. Add about 2 mL of 0.1 M Mg(NO₃)₂ solution to each test tube. Add one drop of 6 M NaOH solution to each test tube. Stir each solution and record your observations.

6. Into test tube D, add at least 20 drops of 6 M HCl (aq) drop-by-drop while stirring. Record all observations.

7. Into test tube E, add at least 20 drops of 6 M NaOH (aq) drop-by-drop while stirring. Record all observations.

8. Into test tube F, add at least 20 drops of 6 M NH₃ (aq) drop-by-drop while stirring. Record all observations.

Part A – Equilibrium and an Acid-Base Indicator

Equilibrium system: \( \text{HA (aq)} \rightleftharpoons \text{H}^+ \ (aq) + \text{A}^- \ (aq) \)

Observations

Record your results upon completing each of the following steps:

<table>
<thead>
<tr>
<th>Step 1</th>
<th>Color of bromothymol blue in distilled water</th>
</tr>
</thead>
<tbody>
<tr>
<td>Step 2</td>
<td>Name of reagent “A” causing color change when added</td>
</tr>
<tr>
<td>Step 3</td>
<td>Name of reagent “B” causing a return to original color</td>
</tr>
</tbody>
</table>

Analysis

- Complete the following:
  - The acidic form of the bromothymol blue indicator, \( \text{HA (aq)} \), is _____________ in color.
  - The basic form of the bromothymol blue indicator, \( \text{A}^- \ (aq) \), is _____________ in color.

- Explain why reagent A (in Step 2) caused the color change observed.

- Explain why reagent B (in Step 3) caused the color change observed.
Part B – Solubility Equilibrium and $K_{sp}$

Equilibrium system:  \[ \text{PbCl}_2 (s) \rightleftharpoons \text{Pb}^{2+} (aq) + 2 \text{Cl}^- (aq) \]

Observations

<table>
<thead>
<tr>
<th>Step</th>
<th>Observations</th>
</tr>
</thead>
<tbody>
<tr>
<td>4</td>
<td>Observations upon addition of just 1.0 mL of HCl to the Pb(NO$_3$)$_2$ solution</td>
</tr>
<tr>
<td>5</td>
<td>Total volume of HCl required for noticeable precipitation mL</td>
</tr>
<tr>
<td>6</td>
<td>Observations upon placing the test tube with precipitate in hot water</td>
</tr>
<tr>
<td>7</td>
<td>Observations upon placing the test tube with precipitate in cold water</td>
</tr>
<tr>
<td>9</td>
<td>Volume of water added to just dissolve PbCl$_2$ precipitate mL</td>
</tr>
<tr>
<td>10</td>
<td>Total solution volume when finished mL</td>
</tr>
</tbody>
</table>

Analysis

- Why didn’t any solid PbCl$_2$ precipitate out immediately upon addition of 1 mL of HCl (aq) in Step 4? What condition must be met by [Pb$^{2+}$] and [Cl$^-$] if solid PbCl$_2$ is to form?

- Consider your observation in hot water in Step 6:
  - In which direction did the equilibrium shift? ____________________________
  - Did the value of $K_{sp}$ get smaller or larger? ____________________________
  - Is the dissolution of PbCl$_2$ (s) exothermic or endothermic? ____________________________
    Explain below.
• Explain why the solid PbCl$_2$ dissolved when water was added to it in Step 9. What was the effect of this water on [Pb$^{2+}$], [Cl$^-$], and $Q_{sp}$? In which direction would such a change drive the equilibrium system?

• The point at which the PbCl$_2$ precipitate just dissolves in Step 9 can be used to determine the value of $K_{sp}$ for this equilibrium system, where $K_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]^2$. Calculate [Pb$^{2+}$] and [Cl$^-$] in the final solution (consider the “dilution effect”). Then use these equilibrium concentrations to determine the value of $K_{sp}$ for this system. Show all work below.

**Part C – Complex Ion Equilibria**

Equilibrium system:  

\[
\text{Co(H}_2\text{O)}_{6}^{2+} (aq) + 4 \text{Cl}^- (aq) \rightleftharpoons \text{CoCl}_4^{2-} (aq) + 6 \text{H}_2\text{O} (l) 
\]

**Observations**

<table>
<thead>
<tr>
<th>Step</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>Step 1</td>
<td>Color of solution in 12 M HCl</td>
</tr>
<tr>
<td>Step 3</td>
<td>Color of solution upon addition of water</td>
</tr>
<tr>
<td>Step 4</td>
<td>Color of solution in hot water</td>
</tr>
<tr>
<td>Step 5</td>
<td>Color of solution in cold water</td>
</tr>
</tbody>
</table>
Analysis

• What form of the complex ion, $\text{Co(H}_2\text{O)}_6^{2+} (aq)$ or $\text{CoCl}_4^{2-} (aq)$, is found predominately in:
  
  the 12 M HCl (aq) __________________________
  
  the diluted solution __________________________
  
  the heated solution __________________________

• Explain why you obtained the observed color in 12 M HCl (aq) in Step 2.

• Explain the observed color change that occurred when water was added to the solution in Step 3. Consider how water affects the ion concentrations in this system, and $Q$.

• Consider your observation in hot water in Step 4.
  
  – In which direction did the equilibrium shift? __________________________
  
  – Did the value of $K$ get smaller or larger? __________________________
  
  – Is the reaction (as written) exothermic or endothermic? __________________________
  
  Explain.
Part D – Dissolving Insoluble Solids

Equilibrium system: \( \text{Zn(OH)}_2 (s) \rightleftharpoons \text{Zn}^{2+} (aq) + 2 \text{OH}^- (aq) \quad K_{sp} << 1 \)

Observations

<table>
<thead>
<tr>
<th>Step 1</th>
<th>Adding 1 drop of NaOH (aq) to Zn(NO_3)_2 (aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Step 2</td>
<td>Tube A: Effect when HCl (aq) is added</td>
</tr>
<tr>
<td>Step 3</td>
<td>Tube B: Effect when NaOH (aq) is added</td>
</tr>
<tr>
<td>Step 4</td>
<td>Tube C: Effect when NH_3 (aq) is added</td>
</tr>
</tbody>
</table>

Analysis

- Explain your observation upon addition of HCl (aq) to the precipitate in terms of Le Châtelier’s Principle and the reaction equations for the various equilibria that are occurring. Consider the effect of HCl on [OH\(^-\)].

- Explain your observations upon addition of NaOH (aq) to the precipitate in terms of Le Châtelier’s Principle and the reaction equations for the various equilibria that are occurring. Consider that Zn\(^{2+}\) forms stable complex ions with OH\(^-\) if supplied in large concentrations.

- Explain your observations upon addition of NH_3 (aq) to the precipitate in terms of Le Châtelier’s Principle and the reaction equations for the various equilibria that are occurring. Consider that Zn\(^{2+}\) forms stable complex ions with NH_3.
Equilibrium system:  \[ \text{Mg(OH)}_2(s) \rightleftharpoons \text{Mg}^{2+}(aq) + 2 \text{OH}^-(aq) \quad K_{sp} << 1 \]

**Observations**

<table>
<thead>
<tr>
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</tr>
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<tbody>
<tr>
<td>Step 1</td>
<td>Adding 1 drop of NaOH (aq) to Mg(NO₃)₂ (aq)</td>
</tr>
<tr>
<td>Step 2</td>
<td>Tube A: Effect when HCl (aq) is added</td>
</tr>
<tr>
<td>Step 3</td>
<td>Tube B: Effect when NaOH (aq) is added</td>
</tr>
<tr>
<td>Step 4</td>
<td>Tube C: Effect when NH₃ (aq) is added</td>
</tr>
</tbody>
</table>

**Analysis**

- Explain your observation upon addition of HCl (aq) to the precipitate in terms of Le Châtelier’s Principle and the reaction equations for the various equilibria that are occurring. Consider the effect of HCl on [OH⁻].

- Based on your observations in Steps 3 and 4 do you think that Mg²⁺ forms stable complex ions? Explain your reasoning.