CH 223 Spring 2024: **"Le Chatelier's Principle** *(in class)"* ***Lab*** *- Instructions*

Note: **This is the lab for section 01 and H1 of CH 223 only.**

- *If you are taking section W1 of CH 223*, please use this link: http://mhchem.org/q/3b.htm

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Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages Ia-3-2 through Ia-3-12 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Bring the printed copy of the lab with you on Monday, April 8 (section 01) *or* **Wednesday, April 10 (section H1.)** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Three:

Complete the lab work and calculations on your own, then **turn it in (**pages Ia-3-7 through Ia-3-12 *only* to avoid a point penalty**) at the beginning of recitation to the instructor on Monday, April 15 (section 01)** *or* **Wednesday, April 17 (section H1.)** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Le Chatelier's Principle

Chemical systems tend to exist in a state of equilibrium. If this equilibrium is disturbed, the reaction may shift in the forward or reverse direction. If the principles governing the equilibrium system are understood, then predictions can be made as to how to drive the reaction in the forward or reverse directions.

In the following equilibrium system, in which A, B, and C are molecules or ions in solution, their relative concentrations at a given temperature must equal a constant, **K**:

$$a\,A(aq) \rightleftharpoons b\,B(aq) + c\,C(aq) \qquad \text{equilibrium constant} = K = \frac{[B]^b[C]^c}{[A]^a}$$

For K to remain constant, if any concentrations are altered, the other components must also shift in concentration. For example, if the concentration of A is increased, the concentrations of B and C will also increase to keep K at a constant value. The reaction shifts in the forward direction. If the concentration of B or C is increased, the reaction will shift in the reverse direction thereby increasing the concentration of A and lowering the concentrations of both B and C. Furthermore, if the concentration of A is reduced, the reaction will shift in the reverse direction, lowering the concentrations of B and C and subsequently increasing the concentration of A to balance the equilibrium. *If you attempt to change a system in chemical equilibrium, it will react in such a way to counterbalance the change you attempted.*

In addition to changes in concentration, a change in the temperature of the system will affect the equilibrium. If the reaction is exothermic in the forward direction, such that heat is a product, an increase in temperature will shift the reaction in the reverse direction. Therefore, an increase in temperature favors the endothermic reaction.

In solution, there are often more than one equilibria maintained. The most important equilibrium involves the dissociation of water:

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

This equation must always be in equilibrium in aqueous solution. If H^+ ion or OH^- ion are present in another equilibrium solution, the equilibrium can be affected by adding acid or base to the system (parts A and D in the procedure). However, if neither H^+ ion nor OH^- ion is present in the equilibrium equation, the equilibrium will not be influenced by the addition of acids or bases (parts B and C).

PROCEDURE:

In this experiment you will work with several equilibrium systems, varying conditions such as temperature and concentrations thereby forcing shifts in the forward or reverse directions. You will be asked to interpret your observations in terms of the principles presented above.

The waste from parts B and C should be poured in a waste bottle. Parts A and D may be poured down the drain.

PROCEDURE: Part A: Solubility Equilibrium: Finding a value for K_{sp}

Many ionic compounds have limited water solubility, dissolving only partially in water. An example is $PbCl_2$:

$$\mathbf{PbCl_2(s) \rightleftharpoons Pb^{+2}(aq) + 2\ Cl^{-}(aq) \quad K_{sp} = [Pb^{+2}][Cl^{-}]^2}$$

The concentration of $PbCl_2$ does not enter into the equilibrium equation because it is a solid and therefore has a constant effect on the system, independent of its amount. The equilibrium constant for a solubility equilibrium is called the **solubility product constant** and symbolized as $\mathbf{K_{sp}}$. For this equilibrium to exist, there must be some solid $PbCl_2$ present in the system. If there is no solid present, there is no equilibrium.

1. Set up a hot water bath using a 400 mL beaker for step 2 below. To a (room temperature) test tube, add 5.0 mL 0.30 M $Pb(NO_3)_2$. Add 0.30 M HCl in 0.5 mL increments (10 drops) to the $Pb(NO_3)_2$ until a white precipitate forms.

   a. Record the volume of HCl needed and determine the moles of Cl^- present in your solution.
   b. Determine the moles of Pb^{+2} ion present in the final solution.
   c. Calculate K_{sp}. (Note: You must use *diluted* concentrations of ions in mol/L)
   d. Why did the $PbCl_2$ not precipitate immediately on addition of HCl?

2. Observing the effect of temperature on the system can determine the sign of enthalpy for this reaction.

   a. Place the test tube in hot water and record your observations.
   b. Place the test tube in cold water and record your observations.
   c. What is the sign of ΔH for the above reaction? Explain.
   d. What happened to the value of K_{sp} in the hot solution? (Increase, decrease, no change) Explain.

3. Add deionized water to the test tube in 0.5 mL increments, stirring well after each addition.

   a. Record the volume of water added when the precipitate just dissolves.
   b. Calculate the molar concentration of Pb^{+2} and Cl^{-1} ions and your reactant quotient (Q).
   c. Explain why the $PbCl_2$ dissolved in water.

Part B: Complex Ion Equilibria

Many metallic ions exist as **complex ions** in solution, combining with other ions or molecules called **ligands**. While complex ions are stable, they may be converted to other complex ions by the addition of ligands that form more stable complexes. Common ligands include OH^-, NH_3, and Cl^-, etc. In this section the cobalt (II) ion can exist as either the **pink** complex ion $Co(H_2O)_6^{+2}(aq)$ or the **blue** complex ion $CoCl_4^{-2}(aq)$ depending on the conditions. The principles of equilibrium can be used to predict which ion will be present.

$$\mathbf{Co(H_2O)_6^{+2}(aq,\ pink) + 4\ Cl^{-}(aq) \rightleftharpoons CoCl_4^{-2}(aq,\ blue) + 6\ H_2O(l)}$$

1. Place a few crystals of $CoCl_2$ - 6 H_2O in a regular test tube. Record the color and determine the Co(II) complex ion present in solution for each of the following. Explain.

a. Add 2 mL (40 drops) of 12 M HCl (CAUTION) and stir to dissolve the crystals.
b. Add 2 mL increments of deionized H_2O to the test tube until no further color change occurs.
c. Place the sample in hot water and cold water and record the color change. What is the sign of ΔH for the above equation? Explain.

Part C: Acid-Base Indicators

Acid-base indicators are chemical substances which change color in solution when $[H^+]$ changes. Methyl violet (HMV) is an example of an acid-base indicator. In solution, HMV dissociates as follows:

$$\textbf{HMV(aq, yellow)} \rightleftharpoons \textbf{H}^+\textbf{(aq)} + \textbf{MV}^-\textbf{(aq, violet)}$$

In solution, HMV has an intense yellow color while MV^- is violet. Therefore, a change in $[H^+]$ will affect the color of the indicator solution.

1. Add 5 mL of deionized H_2O to a test tube. Add a few drops of methyl violet indicator. Record the color.
2. What reagent could be added to shift the equilibrium (change color)? Design and test your hypothesis to demonstrate this equilibrium shift. Explain your results.
3. What reagent could be added to shift the equilibrium back to its original color? Design and test this hypothesis. Explain your results.

Part D: Dissolving Insoluble Solids

Not all solids will dissolve by adding more water or heating the solution. However, in these cases we can employ the principles of equilibrium to dissolve solids. We will use $Zn(OH)_2$ as an example:

$$Zn(OH)_2(s) \rightleftharpoons Zn^{+2}(aq) + 2\ OH^-(aq) \textit{ where } K_{sp} = [Zn^{+2}][OH^-]^2 = 5 \times 10^{-17}$$

The equilibrium constant for this dissociation is quite small, indicating that the reaction does not go very far to the right. Thus, $Zn(OH)_2$ is virtually insoluble in water. In a saturated solution, $[Zn^{+2}] \times [OH^-]^2$ must equal 5×10^{-17}. However, if this product is somehow lowered to less than 5×10^{-17}, then $Zn(OH)_2$ will dissolve until the products equal K_{sp}, where equilibrium will again be obtained. To do this, the concentration of one of the products must be lowered rather drastically. Using a second equilibrium present in solution:

$$H^+(aq) + OH^-(aq) \rightleftharpoons H_2O(l)$$

the addition of an acid to the solution will increase $[H^+]$ and thereby lower $[OH^-]$. This in turn will drive the above reaction to the right, dissolving $Zn(OH)_2$.

Alternatively, we can lower the concentration of Zn^{+2} ion by taking advantage of the fact that zinc ion forms stable **complex ions** with OH^- and NH_3:

$$Zn^{+2}(aq) + 4\ OH^-(aq) \rightleftharpoons Zn(OH)_4^{-2}(aq) \qquad K_1 = 3 \times 10^{15}$$
$$Zn^{+2}(aq) + 4\ NH_3(aq) \rightleftharpoons Zn(NH_3)_4^{+2}(aq) \qquad K_2 = 1 \times 10^{9}$$

In high concentrations of OH^- or NH_3, the above reactions are driven to the right, lowering $[Zn^{+2}]$.

1. To each of three small test tubes, add about 2 mL 0.1 M $Zn(NO_3)_2$. Add one drop 6M NaOH to each test tube and stir. Record your observations.

   a. To the first test tube add 6 M HCl drop by drop. Record and explain your observations.
   b. To the second test tube add 6 M NaOH drop by drop. Stir well and explain your observations.
   c. To the third test tube add 6 M NH_3 (or NH_4OH) drop by drop. Stir well and explain your observations.

2. Repeat the above procedure using 2 mL of 0.1 M $Mg(NO_3)_2$ instead of $Zn(NO_3)_2$ in three separate test tubes. Add one drop of 6 M NaOH to each test tube and stir. Record your observations.

   a. Use the same volume of HCl as you used with the $Zn(NO_3)_2$, above, and stir well. Record your observations.
   b. Use the same volume of 6 M NaOH as you used with the $Zn(NO_3)_2$, above, and stir well. Record your observations.
   c. Use the same volume of 6 M NH_3 (or NH_4OH) as you used with the $Zn(NO_3)_2$, above, and stir well. Record your observations.

3. Explain the following in your report:

   a. What insoluble hydroxide formed when one drop of NaOH was added to the $Zn(NO_3)_2$? To the $Mg(NO_3)_2$?
   b. Explain why some of the reactions were similar between the two metal ions and why some were different.
   c. Why does an insoluble hydroxide tend to dissolve in acidic solution?
   d. Does Mg^{+2} ion appear to form complex ions with OH^- and NH_3? Explain.

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Le Chatelier's Principle

YOUR NAME: ______________________ ***LAB PARTNER(s):*** ________________

Purpose: To explore Le Chatelier's Principle through four different chemical scenarios. Follow the instructions and learn about Le Chatelier's Principle!

Goal #1: **Part A: Solubility Equilibrium: Finding a value for K_{sp}**

1. Set up a hot water bath using a 400 mL beaker for step 2 below. To a (room temperature) test tube, add **5.0 mL 0.30 M $Pb(NO_3)_2$.** Add **0.30 M HCl in 0.5 mL increments (10 drops)** to the $Pb(NO_3)_2$ until a white precipitate forms.

How many mL of $Pb(NO_3)_2$ were used? _________________ mL

How many mL of HCl did you use? _________________ mL

Total volume of $Pb(NO_3)_2$ and HCl = _________ mL *(add up two previous volumes)*

Calculate the moles of Pb^{2+} added to the solution: __________________ moles

Calculate the moles of Cl^{-1} added to the solution: __________________ moles

Calculate the concentration (M) of Pb^{2+} at equilibrium: __________________ M (*use total volume!)*

Calculate the concentration (M) of Cl^{-1} at equilibrium: __________________ M (*use total volume!)*

Calculate the value of $K_{sp} = [Pb^{2+}][Cl^{-1}]^2$ K_{sp} = ________________________

Remember: $[Pb^{2+}] = Pb(NO_3)_2$ (M) and $[Cl^{-1}]$ = HCl (M). **Show work!**

Goal #1: **Part A** *continued*

2. Observing the effect of temperature on the system can determine the sign of enthalpy for this reaction.

* Place the test tube in hot water and record your observations. What happened to the solid $PbCl_2$ as it was heated up?

* Place the test tube in cold water and record your observations. What happened to the $PbCl_2$ solution as it was cooled down?

* Remembering that K_{sp} is defined here as: **$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\ Cl^{-1}(aq)$**, is this reaction **endothermic** or **exothermic**? Explain.

3. Add deionized water to the test tube in 0.5 mL increments, stirring well after each addition

How much water did you add to the solution to make the $PbCl_2(s)$ disappear? _________ mL

Calculate the value of $Q = [Pb^{2+}][Cl^{-1}]^2$ Q = ________________________ **Show work!**

Hint: use moles of $Pb(NO_3)_2$ and HCl from the previous page, divide each by the total volume (which includes the water added in the previous question) to find $[Pb^{2+}]$ and $[Cl^{-1}]$, then calculate Q.

Goal #2: **Part B: Complex Ion Equilibria**

1. Place a few crystals of $CoCl_2$ - 6 H_2O in a regular test tube. Record the color and determine the Co(II) complex ion ("pink" or "blue" will be fine) present in solution for each of the following.

* Add 2 mL (40 drops) of 12 M HCl (CAUTION) and stir to dissolve the crystals.

When HCl was added to the crystals, what color was observed? (circle one) **Pink** **Blue**

* Add 2 mL increments of deionized H_2O to the test tube until no further color change occurs.

When water was added to the solution, what color was observed? (circle one) **Pink** **Blue**

* Place the sample in hot water and cold water and record the color change.

Explain what happened when the sample was added to hot water and cold water:

Is this reaction **endothermic** *(positive ΔH)* or **exothermic** *(negative ΔH)*? Explain.

Goal #3: **Part C: Acid-Base Indicators**

1. Add 5 mL of deionized H_2O to a test tube. Add a few drops of methyl violet indicator.

When drops of methyl violet are added to water, what color is observed? **Yellow** **Violet**
(circle one)

What chemical (reagent) can be used to change the color of the solution? Add the chemical and state what chemical you used. (*hint:* this is an *acid-base* indicator.)

What chemical (reagent) can be used to change the color of the solution back to the original color? Add the chemical to observe the change; state what chemical you used. Explain why adding this reagent worked.

Goal #4: **Part D: Dissolving Insoluble Solids - the Zinc Test**

1. To each of three small test tubes, add about 2 mL 0.1 M $Zn(NO_3)_2$. Add one drop 6M NaOH to each test tube and stir.

When one drop of NaOH was added to each of the three Zn-containing test tubes, what was observed? Explain why this happened and identity the white solid.

* To the first test tube add 6 M HCl drop by drop. Record and explain your observations.

Explain what happened when HCl was added to the first white solid-containing test tube. Do acids appear to dissolve insoluble hydroxides?

* To the second test tube add 6 M NaOH drop by drop. Stir well and explain your observations.

Explain what happens when excess NaOH is added to the second white solid-containing test tube. Do you think that the complex ion (**$Zn(OH)_4^{-2}$(aq)**) could have formed? Explain.

* To the third test tube add 6 M NH_3 (or NH_4OH) drop by drop. Stir well and explain your observations.

Explain what happens when NH_3 is added to the third white solid-containing test tube. Do you think that the complex ion (**$Zn(NH_3)_4^{+2}$(aq)**) could have formed? Explain.

***Goal #4:* Part D: Dissolving Insoluble Solids - the Magnesium Test**

2. Repeat the above procedure using 2 mL of 0.1 M $Mg(NO_3)_2$ instead of $Zn(NO_3)_2$ in three separate test tubes. Add one drop of 6 M NaOH to each test tube and stir.

When one drop of NaOH is added to each of the three Mg-containing test tubes, what is observed? Explain why this happened and identity the white solid.

* Use the same volume of HCl as you used with the $Zn(NO_3)_2$, above, and stir well. Record your observations.

Explain what happened when HCl was added to the first white solid-containing test tube. Do acids appear to dissolve insoluble hydroxides?

* Use the same volume of 6 M NaOH as you used with the $Zn(NO_3)_2$, above, and stir well. Record your observations.

Explain what happened when excess NaOH was added to the second white solid-containing test tube. Do you think that a complex ion could have formed? Explain.

* Use the same volume of 6 M NH_3 (or NH_4OH) as you used with the $Zn(NO_3)_2$, above, and stir well. Record your observations.

Explain what happened when NH_3 was added to the third white solid-containing test tube. Do you think that a complex ion could have formed? Explain.

In terms of electron orbitals, what differences can be found between zinc and magnesium?

Do complex ions appear to form easier with main group metals or transition metals? Why?

Postlab Questions:

1. Methyl orange (HMO) is a common acid-base indicator. In solution it ionizes according to the equation:
   HMO(aq, red) $\rightleftharpoons$ H^+(aq) + MO^- (aq, yellow)

   Chat adds 5 mL of deionized water to a test tube and adds 5 drops of methyl orange. He is captivated by the lovely yellow color of the solution.

   a. If Chat adds 6 M HCl to his solution, what color will he observe? Explain.

   b. If Chat next adds 6 M NaOH to the solution, what color will be observed? Explain.

2. Zinc hydroxide is relatively insoluble in water. The reaction is represented as:
   $Zn(OH)_2$(s) $\rightleftharpoons$ Zn^{+2}(aq) + 2 OH^-(aq)

   a. What is the equilibrium expression for the above reaction?

   b. For $Zn(OH)_2$ to be soluble in solution, [Zn^{+2}] or [OH^-] must be relatively small. Explain.

   c. Would you expect $Zn(OH)_2$ to dissolve if acid were added to the solution? Explain.