EXPERIMENT 12

The Common Ion Effect

INTRODUCTION

In this experiment we will measure the common ion effect on the solubility of the slightly soluble salt potassium hydrogen tartrate, \( \text{KHC}_{4}\text{H}_{4}\text{O}_{6} \). To demonstrate the effect of a common ion on solubility, an example of Le Chatelier’s Principle, we will compare the solubility of this salt in pure water, and then the solubility of the salt in a solution containing KCl.

From the previous experiment, we know that a slightly soluble salt in water is in equilibrium with its ions. For example, AgCl (s) in water is represents the following equilibrium reaction:

\[
\text{AgCl} (s) \rightleftharpoons \text{Ag}^+ (aq) + \text{Cl}^- (aq)
\]

The equilibrium constant for this process is called the solubility product constant, \( K_{sp} \), and is represented by the following expression:

\[
K_{sp} = [\text{Ag}^+] [\text{Cl}^-]
\]

The value of the solubility constant depends only on temperature for a given salt and is completely independent of the quantity of the solid, undissolved salt that remains. The solubility product constant for AgCl in pure water at 25.0°C is \( 1.6 \times 10^{-10} \). If solid AgCl is dissolved in pure water to form a saturated solution, then:

\[
K_{sp} = [\text{Ag}^+] [\text{Cl}^-] = 1.6 \times 10^{-10}
\]

Mathematically we can write \( K_{sp} \) expressions using molar solubility, \( x \), in place of the individual ion concentrations. If we define \( x \) as the molar solubility of AgCl, then \( x \) moles of this salt dissolve in one liter of solution. In a saturated solution of pure AgCl, \( [\text{Ag}^+] = [\text{Cl}^-] = x \)

\[
\text{AgCl} (s) \rightleftharpoons \text{Ag}^+ (aq) + \text{Cl}^- (aq)
\]

\[
\begin{array}{c|cc}
& 0 & 0 \\
+ & +x & +x \\
\hline
& x & x \\
\end{array}
\]

initial molarities
change in molarities
equilibrium molarities

and

\[
K_{sp} = x^2 = 1.6 \times 10^{-10}
\]

And solving for the solubility, \( x \), gives

\[
x = \sqrt{1.6 \times 10^{-10}} = 1.3 \times 10^{-5} \text{ M}
\]

This means that, at 25.0°C, \( 1.3 \times 10^{-5} \) moles of AgCl will dissolve in one liter of solution to produce a saturated solution.
In the previous discussion, we considered the case in which the salt AgCl was dissolved in pure water. If we dissolved AgCl in a dilute solution containing either Ag⁺ or Cl⁻ ions (from another source), we have to consider the effect of the already present ions on the solubility of AgCl. Suppose we dissolve AgCl in a 0.10 M solution of sodium chloride. On a molecular level, there is no distinction among Cl⁻ ions in solution, whether they were produced from the solid AgCl dissolving or from another salt. At constant temperature, the solubility product constant must be satisfied for AgCl regardless of what else is present in the solution. However, the concentration of Cl⁻(aq) in the solution is already 0.10 M before AgCl was added to the solution. When substituting into the expression for $K_{sp}$ we must consider the already present 0.10 M chloride ions.

$$K_{sp} = [x][x + 0.10] = 1.6 \times 10^{-10}$$

In this case, $x$ is the molar solubility of AgCl in the presence of 0.10 M NaCl. The concentration of Ag⁺ in this solution is $x$, but [Cl⁻] = [0.10 + $x$].

$$\text{AgCl(s)} \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)$$

<table>
<thead>
<tr>
<th></th>
<th>C/Initial</th>
<th>C/Change</th>
<th>C/Equilibrium</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0.10</td>
<td>+$x$</td>
<td>0.10 + $x$</td>
</tr>
</tbody>
</table>

Since $K_{sp}$ is very small, we can make the approximation $(0.10 + x) \approx 0.10$ and easily solve for [Ag⁺] as shown below:

$$K_{sp} = [x][0.10] = 1.6 \times 10^{-10}$$

$$x = 1.6 \times 10^{-9} M$$

Since the only source of Ag⁺ ion is from the dissolving of AgCl(s), then $1.6 \times 10^{-9}$ mol/L of AgCl(s) must be dissolved to reach saturation in the presence of the common ion, Cl⁻(aq). This is approximately 10,000 times less AgCl(s) than dissolved in one liter of pure water. The presence of the common ion already in the solvent has decreased the solubility of AgCl(s) by five orders of magnitude.
In this experiment we will use the salt potassium hydrogen tartrate, KHC\(_4\)H\(_4\)O\(_6\) (s). Potassium hydrogen tartrate is a slightly soluble salt and dissolves in water with the following equilibrium:

\[
\text{KHC}_4\text{H}_4\text{O}_6 (s) \rightleftharpoons \text{K}^+ (aq) + \text{HC}_4\text{H}_4\text{O}_6^- (aq)
\]

The solubility product constant expression for the salt is given by:

\[
K_{sp} = [\text{K}^+][\text{HT}^-]
\]

where HT\(^-\) represents HC\(_4\)H\(_4\)O\(_6^-\).

To experimentally measure the common ion effect, we will dissolve enough potassium hydrogen tartrate in (1) pure water to make a saturated solution of KHT, and (2) a dilute solution of KCl (aq) to make a saturated solution of KHT. Since the hydrogen tartrate ion, HT\(^-\), behaves as a weak acid its concentration can be determined in both solutions by acid-base titration using NaOH as the titrant in the following equation:

\[
\text{OH}^- (aq) + \text{HT}^- (aq) \rightarrow \text{H}_2\text{O} (l) + \text{T}^2^- (aq)
\]

Knowing the concentration of the base used in the titration and the volume of base needed to reach the equivalence point, we can determine the number of moles of base needed in the titration. Since we are focusing on the equivalence point, the number of moles of base needed must be equal to the number of moles of the acid HT\(^-\) in the titration sample. Dividing the number of moles of HT\(^-\) by the volume of HT\(^-\) solution used in the titration, we can determine the concentration of the HT\(^-\).

Since the only source of HT\(^-\) ion in both solutions is from the dissolving of KHT (s), the molarity of the HT\(^-\) will equals the molar solubility, \(x\), of KHT in both solutions. Therefore, the first solution will allow you to determine the molar solubility of KHT in pure water, and the second solution will allow you to determine the molar solubility of KHT in the solution with the common ion, K\(^+\).
PROCEDURE

1. Students at odd-numbered lab stations will perform Parts A, C, and D. Student at even-numbered lab stations will perform Parts B, C, and E. Students at odd-numbered and even-numbered lab stations that are next to each other will share their data. Each student will then individually perform all of the calculations in the experiment. Except for the laboratory handout, remove all books, purses, and such items from the laboratory bench top, and placed them in the storage area by the front door. For laboratory experiments you should be wearing closed-toe shoes. Tie back long hair, and do not wear long, dangling jewelry or clothes with loose and baggy sleeves. Open you lab locker. Put on your safety goggles, your lab coat, and gloves.

2. Obtain one 50-mL buret (tolerance ±0.03 mL) and one 25-mL volumetric pipet (tolerance ±0.03 mL) from the cart.

PART A – PREPARING THE PURE KHT SOLUTION

3. Using a milligram balance, weigh out approximately 0.8 grams of KHT in a 250 mL Erlenmeyer flask. This mass does not need to be accurate, and does not need to be recorded in your Data Table.

4. Add 100 mL of deionized water to the flask, and label it Pure KHT.

5. Add your magnetic stir bar to this solution, place the flask on a stirring plate, and stir for 10 minutes using the stir bar.

PART B – PREPARING THE KHT IN KCl SOLUTION

6. Using a milligram balance, weigh accurately about 0.08 grams of KCl into a clean, dry 50-mL beaker. Record the mass of KCl to the nearest milligram.

7. Add about 30 mL of deionized water to the beaker to dissolve the potassium chloride and transfer quantitatively into your 100-mL volumetric flask with a tolerance of ±0.05 mL.

8. Bring the volume of the solution to the hair-line mark with deionized water. Invert several times to mix and transfer the contents to a clean and dry 250-mL Erlenmeyer flask.

9. Calculate the concentration of KCl in this solution and label the flask KHT in KCl, along with its concentration.

10. To the flask add about 0.8 g of potassium hydrogen tartrate, but this mass does not need to be accurate, and does not need to be recorded in your Data Table. Add your magnetic stir bar to this solution, place the flask on a stirring plate, and stir for 10 minutes using the stir bar.
PART C – PREPARING THE BURET FOR TITRATION

11. Obtain about 75 mL of 0.05 M sodium hydroxide in a clean beaker, and from its container record its actual molarity in your Data Table. Condition your buret by rinsing it three times with 5 mL portions of the sodium hydroxide solution, holding the buret in a horizontal position with the stopcock closed and rolling it to make sure that the sodium hydroxide solution wets the entire inside surface. Drain the sodium hydroxide solution through the buret tip into a waste beaker. *All chemicals can be washed down the sink with water.* Add the remaining sodium hydroxide to the buret and be sure that the tip of the buret is filled with sodium hydroxide and there are no air bubbles in the buret tip. Clamp the base-filled buret to the stand above your stir plate.

PART D – TITRATION OF THE PURE KHT SOLUTION

12. Obtain a clean, dry 250-mL Erlenmeyer flask and a circle of medium grade filter paper (do not wet!). Fold the filter paper as shown below, and place it into a *dry* glass funnel, supported over the 250-mL Erlenmeyer flask. Filter out the solid from the pure KHT solution through the filter the solid out of the the pure KHT solution by pouring it through the filter paper in the funnel, and into the Erlenmeyer flask. Label the flask and stopper it.

13. Using the 25-mL volumetric pipet, remove a small portion of the filtered Pure KTH solution to condition the pipet, and discard the washings. Use the conditioned 25-mL volumetric pipet to transfer 25 mL of filtered pure KHT solution into a clean 250-mL Erlenmeyer flask. Add 2 drops of phenolphthalein indicator to the flask. Clean and dry the magnetic stir bar, and carefully slide it into the Erlenmeyer flask. Center the Erlenmeyer flask on the stirring plate, and adjust your buret so the tip is slightly inside the mouth of the flask. Read the buret and record it as the *Initial Buret Reading*. Do this by holding a card with a thick black line behind the buret and reading the bottom of the meniscus. Have another student verify the reading. Turn on the stirring plate, slowly increasing the speed until that the stirring bar creates a vortex in the liquid, but does not collide with the sides of the vessel.

14. Titrate the sample of Pure KHT solution to the light pink phenolphthalein end point. Read the buret and record it as the *Final Buret Reading*.

15. *Complete two titrations in which you obtain a pale pink endpoint. Report in your Data Table the two calculated molarities of the hydrogen tartrate. Precise work will give you two molarities that are nearly identical. The expected precision limit for volumetric analysis is 1%. This means that the two molarities should agree within 1% of each other.*

16. Calculate the molarity of the acid HT⁻ in the solution, and then the molar solubility of KHT in pure water.

17. Calculate the $K_{sp}$ of KHT from its molar solubility in pure water.
PART E – TITRATION OF THE KHT SOLUTION IN KCl

18. Obtain a clean, dry 250-mL Erlenmeyer flask and filter the KHT in KCl solution using a dry glass funnel and medium grade filter paper (do not wet!) into the flask. Label the flask and stopper it.

19. Using the 25-mL volumetric pipet, remove a small portion of the filtered KTH in KCl solution to condition the pipet, and discard the washings. Use the conditioned volumetric pipet to transfer 25 mL of filtered KHT in KCl solution into a clean 250-mL Erlenmeyer flask. Add 2 drops of phenolphthalein indicator to the flask. Clean and dry the magnetic stirrer, and carefully slide it into the Erlenmeyer flask. Place the Erlenmeyer flask on the stirring plate, and adjust your buret so the tip is slightly inside the mouth of the flask. Read the buret and record it as the Initial Buret Reading. Do this by holding a card with a thick black line behind the buret and reading the bottom of the meniscus. Have another student verify the reading. Turn on the stirring plate, slowly increasing the speed to a setting of 6.

20. Titrate the sample of KHT in KCl solution to the light pink phenolphthalein end point. Read the buret and record it as the Final Buret Reading.

21. Complete two titrations in which you obtain a pale pink endpoint. Report in your Data Table the two calculated molarities of the hydrogen tartrate. Precise work will give you two molarities that are nearly identical. The expected precision limit for volumetric analysis is 1%. This means that the two molarities should agree within 1% of each other.

22. Calculate the molarity of the acid HT- in the KCl solution, and then the molar solubility of KHT in the KCl solution.

23. Calculate the $K_{sp}$ of KHT from its molar solubility in the KCl solution, along with the potassium ion concentration in the solution.

24. At the end of the experiment, all solutions can be rinsed down the sink. Do not let your magnetic stir bar go down the sink! Clean and dry the magnetic stir bar before returning it to your lab locker. Rinse the pipet and buret thoroughly, including their tips, several times with tap water, then three times with deionized water, and then dry off the outsides. Return these items to the cart. Return all clamps to their proper drawers in the lab room, and place the stirring plate back in the fume hood.

25. Clean and wipe dry your laboratory work area and all apparatus. When you have completed your lab report have the instructor inspect your working area. Once your working area has been checked your lab report can then be turned in to the instructor.
## EXPERIMENT 12 LAB REPORT

Name: ________________________________    Student Lab Score: _____________
Date/Lab Start Time: ____________________    Lab Station Number: ______________

### DATA TABLE

#### PART B

<table>
<thead>
<tr>
<th>Description</th>
<th>Unit</th>
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</thead>
<tbody>
<tr>
<td>Mass of KCl for Standard KCl Solution</td>
<td>g</td>
</tr>
<tr>
<td>Volume of Standard KCl Solution</td>
<td>mL</td>
</tr>
<tr>
<td>1  Molarity of Standard KCl Solution</td>
<td>M</td>
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#### PART D

<table>
<thead>
<tr>
<th>Description</th>
<th>TRIAL 1</th>
<th>TRIAL 2</th>
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<tbody>
<tr>
<td>Volume of Pure KHT Solution</td>
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<tr>
<td>Molarity of NaOH Solution</td>
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<td>M</td>
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<tr>
<td>Initial Buret Reading</td>
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<td>mL</td>
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<tr>
<td>Final Buret Reading</td>
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<td>mL</td>
</tr>
<tr>
<td>2-3 Volume of NaOH Used</td>
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<tr>
<td>4-5 Molarity of HT</td>
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<td>M</td>
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<tr>
<td>6 Molar Solubility of KHT</td>
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<td>M</td>
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<tr>
<td>7 Average Molar Solubility of KHT</td>
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<td>M</td>
</tr>
<tr>
<td>8 $K_{sp}$ of KHT</td>
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<td>PART E</td>
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<tr>
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<td>Volume of KHT/KCl Solution</td>
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<tr>
<td>Molarity of NaOH Solution</td>
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<tr>
<td>Initial Buret Reading</td>
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<tr>
<td>Final Buret Reading</td>
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<tr>
<td>9-10 Volume of NaOH Used</td>
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<tr>
<td>11-12 Molarity of HT'</td>
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<td>13 Molar Solubility of KHT in KCl</td>
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<tr>
<td>14 Average Molar Solubility of KHT in KCl</td>
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<tr>
<td>15 $K_{sp}$ of KHT</td>
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</tbody>
</table>

**CALCULATIONS**

1.  

2.  

3.  


QUESTIONS

1. How did the molar solubility of KHT in pure water compare to its molar solubility in the KCl solution? Are these results what you would expect? Why?

_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________

2. How did the solubility product constant of KHT calculated from the molar solubility in pure water compare to its solubility product constant calculated from the molar solubility in the KCl solution? Are these results what you would expect? Why?

_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________

3. If all of the undissolved KHT was not filtered out from the solution prior to titrating, would the calculated molar solubility of KHT be greater or less than the actual molar solubility? Explain based upon your calculation in either Box 4 or Box 11.

_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________

4. When cleaning the buret, if the buret is rinsed with water but then is not rinsed several times with the NaOH solution before actually titrating, would the calculated molar solubility of KHT be greater or less than the actual molar solubility? Explain based upon your calculation in either Box 4 or Box 11.

_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________
_________________________________________________________________________________