

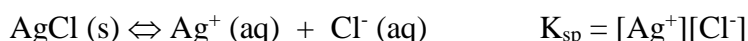
You are encouraged to carefully read the following sections in Tro (3rd ed.) to prepare for this experiment: Sec 16.5, pp 783-788 (Solubility Equilibria and the Solubility Product Constant).

Objectives: You will observe the common ion effect on the K_{sp} and molar solubility of a slightly soluble salt, as determined from the hydrogen ion concentration.

Background:

K_{sp} : K_{sp} is the **solubility product constant** for an ionic compound. Recall in our textbook that a series of rules are given in Table 4.1, p. 161, to determine whether an ionic compound was soluble or insoluble in water. The solubility product constant is the numerical value that explains the entries in the table. The larger the K_{sp} value the greater the degree of dissociation of the ionic compound in water. The K_{sp} of an ionic compound can be calculated using the equilibrium expression for the dissociation of that compound and the molar solubility of the compound. The **molar solubility** is the solubility of the compound in moles per liter.

Example: Dissociation of Silver Chloride



	[Ag ⁺]	[Cl ⁻]
Initial	0	0
Change	+x	+x
Equilibrium	x	x

In the table, $x = [\text{Ag}^+] = [\text{Cl}^-]$ = molar solubility of AgCl, the concentration of dissolved AgCl in a saturated solution of this solid. Since the K_{sp} for AgCl is 1.77×10^{-10} , then:

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.77 \times 10^{-10} = (x)(x)$$

$$1.77 \times 10^{-10} = x^2$$

$$1.33 \times 10^{-5} \text{ M} = x = \text{molar solubility of AgCl in pure water}$$

Common Ion Effect: The Common Ion Effect is observed when an ionic compound is dissolved in a solution that already contains one of the ions found in the salt. Since K_{sp} is a constant that depends on the temperature of the solution but not on the concentrations of the components in solution and following Le Châtelier's Principle, then the presence of a common ion causes the solubility of the ionic compound to be lowered compared to that found in pure water.

Example: Dissociation of Silver Chloride in 0.1 M NaCl

NaCl is a very soluble salt and will dissociate completely into $\text{Na}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$. The concentration of each ion will be equal to 0.1 M. The presence of the $\text{Cl}^-(\text{aq})$ ion means, that when AgCl is added to a solution containing the NaCl, the initial concentration of the $\text{Cl}^-(\text{aq})$ ion is 0.1 M.



	[Ag ⁺]	[Cl ⁻]
Initial	0	0.1
Change	+y	+y
Equilibrium	y	0.1 + y

In the table, y represents the molar solubility (not in water) but now in 0.1 M NaCl(aq). Additionally, K_{sp} for AgCl is still 1.77×10^{-10} since it is a constant. So:

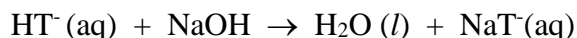
$$K_{sp} = [Ag^+][Cl^-] = 1.77 \times 10^{-10} = (y)(0.1 + y)$$

$$1.77 \times 10^{-10} = 0.1(y) + y^2$$

$$\text{Rearrangement: } 0 = y^2 + 0.1y - 1.77 \times 10^{-10}$$

Using the Quadratic Equation: $1.76 \times 10^{-9} = y =$ molar solubility of AgCl in 0.1 M NaCl. Note that y , solubility in NaCl(aq), is *much* lower than the solubility in pure water (x from above) as predicted by LeChatelier's principle.

In this laboratory, you will observe the effect of the presence of a common ion on the molar solubility and K_{sp} of potassium hydrogen tartrate, or KHT. KHT(s) dissociates into the potassium ion, K^+ (aq), and hydrogen tartrate ion, HT^- (aq) in solution. The HT^- concentration can then be determined by titrating it against the hydroxide ion, OH^- .



Then by using an ICE chart along with the solubility of KHT (which equals the concentration of HT^- since the mole ratio of KHT: HT^- is 1:1), K_{sp} can be calculated for KHT at the temperature for the experiment.



Titration of a Saturated Potassium Hydrogen Tartrate (KHT) Solution Against Standardized NaOH

Preparation of the Saturated KHT Solution for Titration

1. Gather:
 - a. 3 125-mL Erlenmeyer flasks (washed and rinsed with and distilled water)
 - b. 1 100-mL graduated cylinder (washed and rinsed with distilled water)
 - c. 1 250-mL beaker (washed and rinsed with distilled water)
 - d. 1 funnel (washed and dried)
 - e. 1 25-mL volumetric pipet (washed)
 - f. Pipet bulb
 - g. Stirring rod (washed and dried)
 - h. 1 piece of filter paper
 - i. 1 250-mL Erlenmeyer flask (washed and dried)
2. Weigh 1.4 g of KHT.
3. Transfer the KHT into a clean 250-mL beaker.
4. Using graduated cylinders, add the volumes of NaCl and KCl indicated by your instructor to the KHT.
5. Stir the mixture with a stir rod for approximately 10 minutes.

Solution #	Volume of 0.10 M NaCl in mL	Volume of 0.10 M KCl in mL
1	100.0	0.0
2	90.0	10.0
3	80.0	20.0
4	70.0	30.0
5	60.0	40.0
6	50.0	50.0
7	40.0	60.0
8	30.0	70.0
9	20.0	80.0
10	10.0	90.0
11	0.0	100.0

6. Using a 250-mL Erlenmeyer flask, funnel, and filter paper, prepare a gravity filtration apparatus. Don't wet the filter paper and make sure the funnel is dry to avoid adding water to the filtrate.
7. Filter your solution (note: the filtrate must be clear).
8. Using a 25-mL volumetric pipet, withdraw 25 mL of saturated KHT solution and discard in order to rinse the pipet.
9. Using the 25-mL volumetric pipet, transfer 25-mL of the saturated KHT solution into a 125-mL Erlenmeyer flask and record on the data sheet as volume of KHT saturated solution titrated, to the nearest 0.01 mL.
10. Repeat 1 more time into an additional 125-mL flask.
11. Measure the temperature of the two solutions and record the temperature on your data sheet.
12. Add two drops of phenolphthalein indicator to each of the saturated KHT solutions in the 125-mL Erlenmeyer flasks.

Preparation of the Standard NaOH Solution for Titration

Remember that standardized NaOH is valuable and time consuming to make. Take only what you need from the bottle, and be certain the stock bottle is tightly sealed after you use it.

1. Rinse the buret with a small portion (~5 mL) of the NaOH solution, discard this rinse.
2. Fill the buret with NaOH and open the stopcock to filling the tip of the buret with the NaOH solution. Discard this small amount of solution.

Titration

1. Record the indicated molarity of the standardized NaOH solution from the bottle (should be ≈ 0.05 M) onto your data sheet.
2. Record the initial buret reading on your data sheet to the nearest 0.01 mL.
3. Titrate the solution with the standardized NaOH solution until the light pink end point persists for at least 30 s.
4. Record the final buret reading on your data sheet to the nearest 0.01 mL.
5. Refill your buret with additional NaOH.
6. Titrate a second 25-mL portion of your filtrate.
7. Continue titrating samples until your instructor has approved two titrations.

Data

	Run 1	Run 2
Volume of 0.10 M KCl		
Volume of 0.10 M NaCl		
Volume of KHT saturated solution titrated		
Temperature of KHT solution		
Molarity of standardized NaOH from bottle		
Initial Buret Reading of NaOH (mL)		
Final Buret Reading of NaOH (mL)		
Instructor's initials		

Results



	Run 1	Run 2
Volume of NaOH used, mL		
Moles of NaOH required to reach endpoint		
Moles of HT^-		
Concentration of HT^- in saturated solution, M		
Molar Solubility of KHT, M		
Average Molar Solubility of KHT, M		
Concentration of K^+ before KHT dissolves (don't forget dilution calculation), M		
Concentration of K^+ from dissolved KHT, M		
Concentration of K^+ in saturated solution, M		
K_{sp} of KHT		
Average K_{sp} of KHT		

Class Results

Solution #	[K ⁺] from KCl	Avg. [K ⁺] in saturated solution	Avg. [HT ⁻] (molar solubility)	Avg. K _{sp}
1				
2				
3				
4				
5				
6				
7				
8				
9				
10				
11				

Post Laboratory Problems

1. A student determined the molar solubility of $\text{Mg}(\text{OH})_2$ under a variety of conditions. In each of the following scenarios would the student expect to observe a decreased molar solubility due to the common ion effect? Explain your reasoning.
 - a. $\text{Mg}(\text{OH})_2$ was dissolved in a solution of 0.1 M MgCl_2 .
 - b. $\text{Mg}(\text{OH})_2$ was dissolved in a solution of KOH .
 - c. $\text{Mg}(\text{OH})_2$ was dissolved in a solution of NaCl .
2. From the Class Results, how was the molar solubility affected by the change in $[\text{K}^+]$ concentration from KCl ? Explain your answer.
3. From the Class Results, did the K_{sp} change significantly depending on the $[\text{K}^+]$ concentration from KCl ? Explain your answer.

4. An experiment (similar to the one performed in this lab) involved dissolving lead (II) chloride, PbCl_2 , in a 0.10 M solution of lead (II) nitrate, $\text{Pb}(\text{NO}_3)_2$. The chloride ion was then detected using Fajans Method. Fajans Method involves titrating Cl^- against silver nitrate to make AgCl . The endpoint of the titration is observed when a dichlorofluorescein indicator changes from yellow to pink.

a. Write the reaction for the dissociation of $\text{PbCl}_2(\text{s})$.

b. Write the K_{sp} expression for the dissociation of PbCl_2 .

c. Develop an ICE Chart for the dissociation of PbCl_2 in 0.10 M $\text{Pb}(\text{NO}_3)_2$.

d. According to Fajans method the $[\text{Cl}^-] = 0.00527 \text{ M}$ in the saturated solution. Determine the K_{sp} of PbCl_2 .

Pre-Laboratory Problems

1. Write the reaction for the dissociation of KHT(s) in water.
2. Write the equilibrium expression for the dissociation of KHT in water.
3. Develop an ICE Chart for the dissociation of KHT in water.
4. During the Weak Acid Laboratory, you determined the HT^- concentration for the dissociation of KHT in water to be approximately 0.0387 M. Determine the K_{sp} for the dissociation of KHT in pure water.
5. Why is it important to add enough KHT to the KCl-NaCl solution in this lab to ensure solid remains?