

## **EXPERIMENT C3: SOLUBILITY PRODUCT & COMMON ION EFFECT**

### **Learning Outcomes**

Upon completion of this lab, the student will be able to:

- 1) Measure the solubility product constant for a sparingly soluble salt.
- 2) Evaluate the effect of the presence of common ions on the solubility of sparingly soluble salts.

### **Introduction**

The primary objective of this experiment is to study the solubility of sparingly soluble salts and determining the molar solubility and solubility product constant of a sparingly soluble salt. The effect of the presence of common ions on the solubility of the salts will also be examined.

### **Solubility Product and Solubility**

According to the general solubility rules, several ionic salts are considered to be insoluble or more accurately sparingly soluble. In these instances, only a small amount of the salt is able to dissolve in water. This is due to the fact that, in order for an ionic compound to be able to dissolve in water, it must dissociate (or form a complex, as will be discussed in a future experiment). The energy necessary to dissociate is often compensated for by the energy released when the resulting ions are surrounded by water. However, if the overall process is energetically unfavorable, then the salt will not be soluble. An example of one such ionic salt is calcium iodate ( $\text{Ca}(\text{IO}_3)_2$ ).

When calcium iodate is dissolved in water, the following equilibrium results:



Equation 1 is referred to as the solubility equilibrium for calcium iodate. The equilibrium constant for the above reaction is also known as the solubility product constant and is given by the following expression.

$$K = K_{\text{SP}} = [\text{Ca}^{2+}] \times [\text{IO}_3^{-}]^2 \quad \text{Equation 2}$$

Recall that the solubility,  $s$ , of any sparingly soluble salt refers to the maximum amount (generally in grams per mL or molar concentration) of the salt that dissociates into ions in the aqueous medium. The solubility can therefore be used to calculate the  $K_{\text{SP}}$  for a salt (and vice versa).

Assume for instance that, “s” represents the molar solubility of calcium iodate. Then according to Equation 1:

$$[\text{Ca}^{2+}] = s \text{ and } [\text{IO}_3^-] = 2s$$

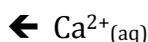
$$\text{Therefore } K_{\text{SP}} = s \times (2s)^2 = 4s^3$$

**Equation 3**

### **Common Ion Effect**

According to Le Châtelier’s principle, the equilibrium of an ionic substance can be influenced by the presence of a *common ion* – an ion that is present in the ionic compound itself. The presence of a common ion in the medium of an aqueous solution of an ionic substance shifts the equilibrium to the left, since that common ion would be considered a product, thereby lowering the solubility of the ionic substance. This effect is referred to as the common ion effect.

Consider the calcium iodate equilibrium shown in Equation 1 above. According to the common ion effect, if the same solution contains  $\text{Ca}^{2+}_{(\text{aq})}$  from a different source, say from calcium nitrate or calcium chloride, the additional  $\text{Ca}^{2+}_{(\text{aq})}$  would force the equilibrium shown in Equation 1 to shift to the left.



This leads to the formation of more solid calcium iodate, implying a decreased solubility of the salt.

A similar effect would result if the same solution were to contain additional  $\text{IO}_3^{-}_{(\text{aq})}$  from a different source, say from potassium iodate or ammonium iodate.

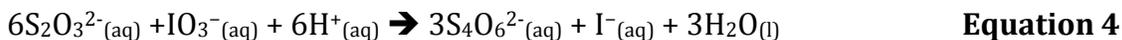
In summary, the presence of a common ion decreases the solubility of a sparingly soluble salt.

NOTE: A decrease in solubility, “s”, as a result of the presence of common ion does not indicate a change in the solubility product constant,  $K_{\text{SP}}$ .  $K_{\text{SP}}$ , as the name indicates, is a constant and, as is the case with all equilibrium constants, only changes with a change in temperature. A decrease in solubility, “s”, only implies that a smaller amount of the ionic solid, in this example the calcium iodate, dissolves and forms ions in the aqueous medium.

### **Determination of Solubility**

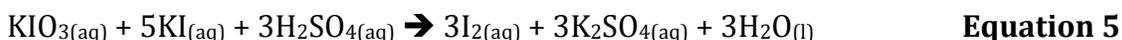
In the examination of the solubility of calcium iodate, the measurement of either the concentration of calcium ions or the concentration of iodate ions in a saturated solution of calcium iodate is necessary.

The concentration of iodate, as was discussed in experiment A9, can be determined by a redox titration using a standardized thiosulfate solution. The chemical equation for the reaction between iodate and thiosulfate is shown below:



According to Equation 4, six moles of thiosulfate are needed to completely react with one mole of iodate.

Equation 4 does not provide the complete picture of the chemistry of this reaction. The iodate is first combined with iodide ( $\text{I}^{-}$ ) to facilitate the electron transfer process. The chemical reaction between iodate and iodide is as follows:



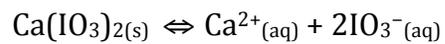
The  $\text{I}_2$  formed in Equation 5 reacts with the thiosulfate according to the following reaction:



When Equation 5 and Equation 6 are combined, the overall reaction obtained is Equation 4. Since three moles of  $\text{I}_2$  are formed in Equation 5 and each mole of  $\text{I}_2$  is titrated with two moles of  $\text{S}_2\text{O}_3^{2-}$  according to Equation 6, it then follows that six moles of  $\text{S}_2\text{O}_3^{2-}$  are needed to completely react with one mole of  $\text{IO}_3^{-}$ .

Since this is a redox reaction, there is no indicator present in the traditional (acid/base) sense. As mentioned above, the endpoint of the titration is the disappearance of the color due to the presence of  $\text{I}_2$ . However, starch is used to enhance the visualization of the endpoint. The color of the mixture in the Erlenmeyer flask after the initial reaction of  $\text{IO}_3^{-}$  and  $\text{I}^{-}$  (Equation 5) will be reddish brown. Once the titration begins the  $\text{I}_2$  is converted to  $\text{I}^{-}$  (according to Equation 6), the color will become lighter and lighter yellow. Since it can be difficult to determine whether a solution is colorless or very faintly yellow, starch is added at this point to reversibly complex with the remaining  $\text{I}_2$  in the flask and gives the solution a deep blue to violet color. Once all the  $\text{I}_2$  has been converted to  $\text{I}^{-}$  by the  $\text{S}_2\text{O}_3^{2-}$  the solution will become colorless, as the starch is no longer complexed with  $\text{I}_2$ .

Assume that the concentration of the  $\text{IO}_3^-$  thus determined is "x". According to Equation 1:



**Equation 1**

$$[\text{IO}_3^-] = x$$

$$\text{Therefore } [\text{Ca}^{2+}] = \frac{x}{2}$$

$$K_{sp} = [\text{Ca}^{2+}][\text{IO}_3^-]^2$$

$$K_{sp} = \left(\frac{x}{2}\right)(x)^2 = \frac{x^3}{2}$$

$$\text{And solubility} = [\text{Ca}^{2+}] = \frac{x}{2}$$

## Experimental Design

A precipitate of calcium iodate will be prepared first by combining calcium nitrate with potassium iodate. This precipitate will be divided into two parts. One part of the precipitate will be dissolved in deionized water. The solubility and  $K_{SP}$  of the calcium iodate will be determined by titrating the saturated solution of the calcium iodate with a standardized thiosulfate solution. The second part of the precipitate is dissolved in a solution containing calcium nitrate. The solubility of the calcium iodate in the presence of the common ion will be determined by titrating this solution with standardized thiosulfate.

## Reagents and Supplies

0.00500 M  $KIO_3$  standard, 0.02 M  $Na_2S_2O_3$  (to be standardized by students), 0.2 M  $KIO_3$ , 0.1 M  $Ca(NO_3)_2$ , starch solution, 2 M sulfuric acid, 10% potassium iodide solution

Filter flask, Hirsch funnel, filter paper

(See posted Material Safety Data Sheets)

## Procedure

### PART 1: PREPARATION OF CALCIUM IODATE

1. Obtain 9.0 mL of 0.2 M  $\text{KIO}_3$  (from the brown glass bottle).
2. Obtain 10 mL of 0.1 M  $\text{Ca}(\text{NO}_3)_2$
3. Combine the above two solutions in a 25-mL Erlenmeyer flask and let the mixture stand for at least 10 minutes.
4. Set up the vacuum filtration apparatus using the filter flask, Hirsch funnel, and filter paper.
5. Filter the mixture from the Erlenmeyer flask (in Step 3). Rinse any precipitate with cold deionized water and transfer completely to the filtration apparatus.
6. Allow the solid precipitate to dry under vacuum for five minutes.
7. Remove the filter paper containing the calcium iodate precipitate from the filtration apparatus carefully.
8. Divide the precipitate into two parts (approximately). Transfer each part to two small beakers labeled A and B.
9. To beaker A: add 10 mL of deionized water.
10. To beaker B: add 9 mL of deionized water and 1 mL of 0.1 M  $\text{Ca}(\text{NO}_3)_2$ .
11. Thoroughly stir both solutions for one minute each then allow both beakers to rest.
12. Allow the mixtures in beakers A and B to rest until clear supernatant results. Use the supernatant solutions from beakers A and B, respectively, for Parts 3 and 4 of the experiment.
13. Discard all other solutions into appropriate waste containers as directed by the instructor.

PART 2: STANDARDIZATION OF THE THIOSULFATE

1. Obtain about 10 mL of each of the following solutions: 1% starch, 2 M sulfuric acid, 10% aqueous potassium iodide, sodium thiosulfate, standard potassium iodate (from the wash bottle, record the exact concentration indicated on the bottle). Clearly label each container.
2. Clamp two microburettes to a ring stand. Label one microburette as the thiosulfate burette and the other as iodate burette.
3. Rinse, condition, and fill each burette with the appropriate reagent.
4. Record the initial burette readings of both the burettes.
5. Add about 1 mL of iodate from the burette into a 25-mL Erlenmeyer flask. Record the final burette reading.
6. Add the following reagents into the above Erlenmeyer flask:
  - a. 5 mL of deionized water
  - b. Six drops of 2 M  $\text{H}_2\text{SO}_4$
  - c. 2 mL of 10% KI
7. The solution in the Erlenmeyer flask should be reddish brown in color. Begin titrating this solution with the thiosulfate.
8. If any reagent is stuck to the sides of the flask, rinse the flask with deionized water to ensure mixing of all the reagents.
9. When the solution in the Erlenmeyer flask is pale yellow, add 2 mL of deionized water and three drops of 1% starch solution. This solution should now be deep blue to violet in color.
10. Continue titrating with the thiosulfate until the solution turns colorless. Record the final burette reading of the thiosulfate burette.
11. Repeat steps 4 to 10 two to three more times.
12. Discard all other solutions into appropriate waste containers as directed by the instructor.

PART 3: SOLUBILITY OF CALCIUM IODATE

1. Obtain about 10 mL of each of the following solutions: 1% starch, 2 M sulfuric acid, 10% aqueous potassium iodide, sodium thiosulfate, iodate solution from beaker A (Part 1). Clearly label each container.
2. Clamp two microburettes to a ring stand. Label one microburette as the thiosulfate burette and the other as iodate burette.
3. Rinse, condition, and fill each burette with the appropriate reagent.
4. Record the initial burette readings of both the burettes.
5. Add about 1 mL of iodate from the burette into a 25-mL Erlenmeyer flask. Record the final burette reading.
6. Add the following reagents into the above Erlenmeyer flask:
  - a. 5 mL of deionized water
  - b. Six drops of 2 M  $\text{H}_2\text{SO}_4$
  - c. 2 mL of 10% KI
7. The solution in the Erlenmeyer flask should be reddish brown in color. Begin titrating this solution with the thiosulfate.
8. If any reagent is stuck to the sides of the flask, rinse the flask with deionized water to ensure mixing of all the reagents.
9. When the solution in the Erlenmeyer flask is pale yellow, add 2 mL of deionized water and three drops of 1% starch solution. This solution should now be deep blue to violet in color.
10. Continue titrating with the thiosulfate until the solution turns colorless. Record the final burette reading of the thiosulfate burette.
11. Repeat steps 4 to 10 two to three more times.
12. Discard all other solutions into appropriate waste containers as directed by the instructor.

PART 4: EFFECT OF  $Ca^{2+}$  ON THE SOLUBILITY OF CALCIUM IODATE

1. Obtain about 10 mL of each of the following solutions: 1% starch, 2 M sulfuric acid, 10% aqueous potassium iodide, sodium thiosulfate, iodate solution from beaker B (Part 1). Clearly label each container.
2. Clamp two microburettes to a ring stand. Label one microburette as the thiosulfate burette and the other as iodate burette.
3. Rinse, condition, and fill each burette with the appropriate reagent.
4. Record the initial burette readings of both the burettes.
5. Add about 1 mL of iodate from the burette into a 25-mL Erlenmeyer flask. Record the final burette reading.
6. Add the following reagents into the above Erlenmeyer flask:
  - a. 5 mL of deionized water
  - b. Six drops of 2 M  $H_2SO_4$
  - c. 2 mL of 10% KI
7. The solution in the Erlenmeyer flask should be reddish brown in color. Begin titrating this solution with the thiosulfate.
8. If any reagent is stuck to the sides of the flask, rinse the flask with deionized water to ensure mixing of all the reagents.
9. When the solution in the Erlenmeyer flask is pale yellow, add 2 mL of deionized water and three drops of 1% starch solution. This solution should now be deep blue to violet in color.
10. Continue titrating with the thiosulfate until the solution turns colorless. Record the final burette reading of the thiosulfate burette.
11. Repeat steps 5 to 10 two to three more times.
12. Discard all other solutions into appropriate waste containers as directed by the instructor.

## Data Table

### PART 2: STANDARDIZATION OF THE THIOSULFATE

#### Iodate solution

	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>	<b>Trial 4</b>
Initial burette reading (mL)				
Final burette reading (mL)				

#### Thiosulfate solution

	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>	<b>Trial 4</b>
Initial burette reading (mL)				
Final burette reading (mL)				

### PART 3: SOLUBILITY OF CALCIUM IODATE

#### Iodate solution

	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>	<b>Trial 4</b>
Initial burette reading (mL)				
Final burette reading (mL)				

#### Thiosulfate solution

	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>	<b>Trial 4</b>
Initial burette reading (mL)				
Final burette reading (mL)				

PART 4: EFFECT OF  $\text{Ca}^{2+}$  ON THE SOLUBILITY OF CALCIUM IODATE**Iodate solution**

	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>	<b>Trial 4</b>
Initial burette reading (mL)				
Final burette reading (mL)				

**Thiosulfate solution**

	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>	<b>Trial 4</b>
Initial burette reading (mL)				
Final burette reading (mL)				

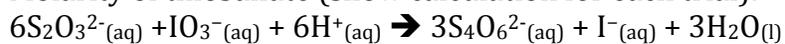
## Data Analysis

### PART 2: STANDARDIZATION OF THE THIOSULFATE

Molarity of  $\text{KIO}_3$  (provided on the reagent bottle)=

	<b>Trial 1</b>	<b>Trial 2</b>	<b>Trial 3</b>	<b>Trial 4</b>
Volume of $\text{IO}_3^-$ (liters)				
Volume of $\text{S}_2\text{O}_3^{2-}$ (liters)				

Molarity of thiosulfate (show calculation for each trial):



Trial 1

Trial 2

Trial 3

Trial 4

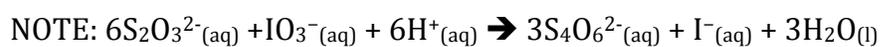
	<b>Molarity of <math>\text{S}_2\text{O}_3^{2-}</math></b>
Trial 1	
Trial 2	
Trial 3	
Trial 4	
<b>Average</b>	

PART 3: SOLUBILITY OF CALCIUM IODATEMolarity of  $S_2O_3^{2-}$  (average value from Part 2)=

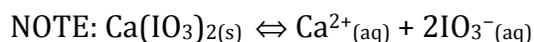
	Trial 1	Trial 2	Trial 3	Trial 4
Volume of $IO_3^-$ (liters)				
Volume of $S_2O_3^{2-}$ (liters)				

For each trial calculate the following:

- a. Molarity of iodate (show calculation for each trial)



- b. Molarity of  $Ca^{2+}$   
 c. Solubility of  $Ca(IO_3)_2$   
 d.  $K_{SP}$  of  $Ca(IO_3)_2$



$$[IO_3^-] = x$$

$$\text{Therefore } [Ca^{2+}] = \frac{x}{2}$$

$$K_{SP} = [Ca^{2+}][IO_3^-]^2$$

$$K_{SP} = \left(\frac{x}{2}\right)(x)^2 = \frac{x^3}{2}$$

$$\text{And solubility} = [Ca^{2+}] = \frac{x}{2}$$

Trial 1

Trial 2

Trial 3

Trial 4

	<b>Molarity of <math>\text{IO}_3^-</math></b>	<b>Molarity of <math>\text{Ca}^{2+}</math></b>	<b>Solubility of <math>\text{Ca}(\text{IO}_3)_2</math>, M</b>	<b><math>K_{\text{SP}}</math></b>
Trial 1				
Trial 2				
Trial 3				
Trial 4				
<b>Average</b>				

**Solubility of  $\text{Ca}(\text{IO}_3)_2$  in water =**

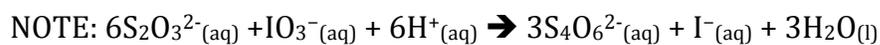
**$K_{\text{SP}}$  of  $\text{Ca}(\text{IO}_3)_2$  at room temperature =**

PART 4: EFFECT OF  $\text{Ca}^{2+}$  ON THE SOLUBILITY OF CALCIUM IODATEMolarity of  $\text{S}_2\text{O}_3^{2-}$  (average value from Part 2)=

	Trial 1	Trial 2	Trial 3	Trial 4
Volume of $\text{IO}_3^-$ (liters)				
Volume of $\text{S}_2\text{O}_3^{2-}$ (liters)				

For each trial calculate the following:

- a. Molarity of iodate (show calculation for each trial)



- b. Molarity of  $\text{Ca}^{2+}$   
 c. Solubility of  $\text{Ca}(\text{IO}_3)_2$



$$[\text{IO}_3^-] = x$$

$$\text{Therefore } [\text{Ca}^{2+}] = \frac{x}{2}$$

$$K_{sp} = [\text{Ca}^{2+}][\text{IO}_3^-]^2$$

$$K_{sp} = \left(\frac{x}{2}\right)(x)^2 = \frac{x^3}{2}$$

$$\text{And solubility} = [\text{Ca}^{2+}] = \frac{x}{2}$$

Trial 1

Trial 2

Trial 3

Trial 4

	<b>Molarity of <math>\text{IO}_3^-</math></b>	<b>Molarity of <math>\text{Ca}^{2+}</math></b>	<b>Solubility of <math>\text{Ca}(\text{IO}_3)_2</math>, M</b>
Trial 1			
Trial 2			
Trial 3			
Trial 4			
<b>Average</b>			

**Solubility of  $\text{Ca}(\text{IO}_3)_2$  in 0.01 M  $\text{Ca}(\text{NO}_3)_2$  =**