Determination of the Solubility Product of an Ionic Compound

PURPOSE

In this experiment you will determine the solubility and solubility product of a slightly soluble salt.

INTRODUCTION

When a salt of low solubility dissolves in water, equilibrium is established between the solid solute and the dissolved ions. There are two terms we use in discussing this condition. The first is solubility, which is the maximum amount of salt that will dissolve in a given amount of solvent (usually water) at a specified temperature. Solubility is usually expressed in units of molarity (moles/L), but sometimes g solute/g solvent is used.

The second term is solubility product, $K_{sp}$. This is a special equilibrium constant for the dissolving of a salt in water. By convention the equation for the reaction is always written with the solid as the reactant, and the ions as products. For a general salt, $A_mB_n$, the equation would be:

$$A_nB_n(s) \rightleftharpoons m A^+ (aq) + n B^- (aq)$$

The equilibrium constant for such a salt would be:

$$K_{sp} = [A]^m[B]^n$$

Note that the exponents in the solubility product expression match the coefficients in the equilibrium equation, and that the solid is not included in the $K_{sp}$ expression. Because of the way the expression is written the value of the constant is expressed without units. The magnitude of the constant expresses the degree to which the salt will dissolve (the degree to which the equilibrium lies to the right). Most $K_{sp}$ values are quite small, indicating that the salts are not very soluble. For example, the $K_{sp}$ for AgCl is $4 \times 10^{-11}$ at 10°C.

In this laboratory you will determine the $K_{sp}$ for the very slightly soluble salt calcium iodate, Ca(IO₃)₂. The equation for the dissolving of the salt is,

$$Ca(IO₃)_2 (s) \rightleftharpoons Ca^{+2} (aq) + 2 IO₃^- (aq) \quad (Eq. 1)$$

and the solubility product expression is,

$$K_{sp} = [Ca^{+2}][IO₃^-]^2$$

To determine the $K_{sp}$ you could measure the equilibrium concentration of $Ca^{+2}$ or $IO₃^-$. (Since the concentrations of the ions are stoichiometrically related, you only need to know one of the two.) In this experiment the concentration of the iodate ion will be determined through a titration with sodium thiosulfate, using starch as an indicator.
The ionic equations for the reactions are:

\[
\text{IO}_3^- + 8 \text{I}^- + 6 \text{H}^+ \rightleftharpoons 3 \text{I}_3^- + 3 \text{H}_2\text{O} \quad \text{(Eq. 2)}
\]

\[
3 \text{I}_3^- + 6 \text{S}_2\text{O}_3^{2-} \rightleftharpoons 9\text{I}^- + 3 \text{S}_4\text{O}_6^{2-} \quad \text{(Eq. 3)}
\]

**Combined**

\[
\text{IO}_3^- + 6 \text{S}_2\text{O}_3^{2-} + 6 \text{H}^+ \rightleftharpoons \text{I}^- + 3 \text{S}_4\text{O}_6^{2-} + 3 \text{H}_2\text{O}
\]

In the presence of starch the \(\text{I}_3^-\) forms a dark blue complex. To determine the concentration of iodate in the unknown solution you will add the starch indicator and then titrate with \(\text{Na}_2\text{S}_2\text{O}_3\) until the dark blue color disappears. For more information on the starch/iodine reaction, please go to: [http://www.elmhurst.edu/~chm/vchembook/548starchiodine.html](http://www.elmhurst.edu/~chm/vchembook/548starchiodine.html)

**PROCEDURE**

**Part I: Preparation of a Saturated Solution of Ca(IO\(_3\))\(_2\)**

Combine 100.0 mL of 0.20 M KIO\(_3\) with 10.0 mL of 1 M Ca(NO\(_3\))\(_2\) in a 250 mL beaker. Stir the mixture completely. A white precipitate of Ca(IO\(_3\))\(_2\) should form.

Let the mixture stand a few minutes, then filter it through a dry filter into a dry flask, using gravity filtration. Why is it important that the flask and the filter paper be dry? Do NOT add any water to rinse the precipitate!

**Part II - Determination of [IO\(_3^-\)] in the Saturated Solution of Ca(IO\(_3\))\(_2\)**

Set up a buret and fill it with the standard solution of \(\text{Na}_2\text{S}_2\text{O}_3\) that is available on the lab cart. Be sure to record its concentration in your laboratory notebook.

Pipet 10.0 mL of the filtrate from Part I, above, into a clean Erlenmeyer flask. Add about 20 mL of deionized water to the flask. (As you have in previous experiments, be sure to rinse the pipet prior to use to avoid any dilution of the saturated solution with water that might be in the pipet.)

Dissolve about 2 g of solid KI in the solution. Add 20 drops of 2 M HCl and mix thoroughly. The solution should take on a brown color.

Titrate the mixture with the \(\text{Na}_2\text{S}_2\text{O}_3\) until the brown color changes to yellow. Add several drops of starch to the solution, enough to turn the solution black. Continue titrating with the \(\text{Na}_2\text{S}_2\text{O}_3\) until the blue/black color just disappears. Repeat the titration twice for a total of three trials.
CALCULATIONS AND ANALYSIS

Using the known concentration of sodium thiosulfate, calculate the $[\text{IO}_3^-]$ in the filtrate. From the $[\text{IO}_3^-]$, calculate the solubility of $\text{Ca(IO}_3)_2$ as well as the $K_{sp}$ of $\text{Ca(IO}_3)_2$.

Look up the accepted value of the $K_{sp}$ and compare your results with the expected value. Do not expect exact agreement; ideally your results will be the same order of magnitude as the reported value. As usual, include a thoughtful discussion of your results in the Analysis section of your lab report. Report Sheet • Solubility Product ($K_{sp}$)

All of your calculations should be carried out in your lab notebook.

$[\text{Na}_2\text{S}_2\text{O}_3] = \underline{\underline{\text{M}}}$

Part II.

<table>
<thead>
<tr>
<th>Trial</th>
<th>$V_{\text{Ca(IO}_3)_2 \text{ soln}}$ (mL)</th>
<th>$V_{\text{H}_2\text{O added}}$ (mL)</th>
<th>mass$_{\text{KI added}}$ (g)</th>
<th># drops 2M HCl</th>
<th># drops starch</th>
<th>$V_{\text{Na}_2\text{S}_2\text{O}_3 \text{ added}}$ (mL)</th>
<th>mol$_{\text{Na}_2\text{S}_2\text{O}_3 \text{ m}}$</th>
<th>$[\text{IO}_3^-] \text{ (M)}$</th>
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Average

St. Dev.

$K_{sp}$ Calculation/Analysis.

Knowing:

$$\text{Ca(IO}_3)_2 \text{ (s)} \rightleftharpoons \text{Ca}^{2+} \text{ (aq)} + 2 \text{IO}_3^- \text{ (aq)}$$

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

1. Using the average $[\text{IO}_3^-]$, what is the $[\text{Ca}^{2+}]$?

2. Using the average $[\text{IO}_3^-]$ and the $[\text{Ca}^{2+}]$, what is the $K_{sp}$ for calcium iodate?

3. What is the literature value for the $K_{sp}$ of calcium iodate?

Source:

4. How does your calculated value compare to the literature value (use the formula for percent error)? Was your value too high or too low? What might be the source(s) of error?

$$\text{Percent Error} = \frac{|\text{Experimental} - \text{Literature}|}{\text{Literature}} \times 100$$