EXPERIMENT 3   *The Solubility Product of Calcium Iodate*

Introduction

The solubility product constant is simply the equilibrium constant for the solution of a slightly soluble salt in water. When the generic salt M_xA_y dissolves in water there are produced x moles of M^{+y} and y moles of A^{-x} for every mole of M_xA_y that dissolves according to

$$M_xA_{y(s)} \rightleftharpoons xM^{+Y}_{(aq)} + yA^{-x}_{(aq)} \qquad \textbf{(EQ 3.1)}$$

and the equilibrium constant for the reaction (omitting M_xA_y because it is a solid) would be,

$$K_{sp} = [M^{+y}]^x[A^{-x}]^y \qquad \textbf{(EQ 3.2)}$$

For example, the solubility of MgF_2 is 2.6×10^{-4} *M* at room temperature.

$$MgF_{2(s)} \rightleftharpoons Mg^{+2}_{(aq)} + 2F^{-}_{(aq)} \qquad \textbf{(EQ 3.3)}$$

A saturated solution of MgF_2 is therefore 2.6×10^{-4} molar in Mg^{2+} and 5.2×10^{-4} *M* in F^- and

$$K_{sp} = [Mg^{2+}][F^-]^2 = (2.6 \times 10^{-4})(5.2 \times 10^{-4})^2 = 7.0 \times 10^{-11} M^3 \qquad \textbf{(EQ 3.4)}$$

According to Le Châtelier's principle, addition of either Mg^{2+} or F^- should depress the solubility of MgF_2. Let us work the following example

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EXAMPLE 3.1

What would be the solubility of MgF_2 in a solution of 0.10 M NaF?

According to the K_{sp} expression,

$$K_{sp} = [Mg^{2+}][F^-]^2 \qquad \textbf{(EQ 3.5)}$$

If the [F⁻] is initially 0.10 M, then

$$K_{sp} = 7.0 \times 10^{-11} = [Mg^{2+}](0.10 + [F^-])^2 \qquad \textbf{(EQ 3.6)}$$

Since [F⁻] is expected to be small compared to 0.1, we can say that,

$$K_{sp} = [Mg^{2+}](0.10)^2 \qquad [Mg^{2+}] = \frac{7.0 \times 10^{-11}}{(0.10)^2} = 7.0 \times 10^{-9} \quad \textbf{(EQ 3.7)}$$

The [F⁻] from solution of MgF_2 is therefore 1.4×10^{-8} and is much less than 0.1.

K_{sp} of $Ca(IO_3)_2$

In this experiment you will determine the K_{sp} of $Ca(IO_3)_2$ and observe the common ion effect of added $Ca(NO_3)_2$ on the solubility of $Ca(IO_3)_2$. The concentration of iodate ion in solution will be determined by reducing the iodate to iodine with excess iodide solution, and subsequently titrating the iodine with thiosulfate. Iodine binds to starch to form a very dark purple complex, so starch is used as an indicator in the titration

In the presence of an excess of iodide and acid the iodate is converted quantitatively to iodine.

$$IO_3^- + 5I^- + 6H^+ \rightarrow 3I_2 + 3H_2O \qquad \textbf{(EQ 3.8)}$$

Three moles of iodine are produced per mole of iodate.

The iodine is then reduced to iodide by thiosulfate ion, according to reaction 3.9

$$I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-} \qquad \textbf{(EQ 3.9)}$$

Two moles of thiosulfate reduce one mole of iodine.

It is seen that the moles of iodate in the initial solution will be one sixth the number of moles of thiosulfate used to titrate the iodine.

Procedure

Saturated solutions of calcium iodate in water and in 0.10 *M* calcium nitrate have been prepared, as well as standardized 0.05 *M* sodium thiosulfate solution. *Make sure that you record the actual molarity of these solutions.* This is found on the container label.

Solubility Product of $Ca(IO_3)_2$

1. To a 250 mL Erlenmeyer flask, add approximately 2.0 g of KI and 50 mL of water.
2. Then pipet 10.00 mL of the saturated calcium iodate solution into the flask and add 100 mL of 1.0 *M* HCl.
3. Titrate the brown solution with standardized 0.05 *M* sodium thiosulfate to a light amber color.
4. To the light amber solution add 5 mL of prepared starch solution. Titrate the purple solution until it turns colorless.
5. Record the temperature of the solution.
6. Repeat the procedure two more times for a total of three trials.

Common Ion Effect

Duplicate the procedure in "Solubility Product of $Ca(IO_3)_2$" (above) with 10.00 mL samples of the saturated solution of calcium iodate prepared in 0.100 *M* calcium nitrate solution.

Calculations

1. Calculate the solubility of $Ca(IO_3)_2$ in water and in 0.10 *M* $Ca(NO_3)_2$ solution
2. Calculate the value of K_{sp} for $Ca(IO_3)_2$ in each solution.
3. Calculate the average value of K_{sp} and the % error from the value in the CRC.

Questions

1. K_{sp} for Ag_2SO_4 is 1.2×10^{-5} at 25 °C. Calculate the expected molar solubility of Ag_2SO_4 at 25°C
   a. in water
   b. in 0.20 *M* $AgNO_3$
   c. in 0.20 *M* Na_2SO_4
   d. in 0.20 *M* $NaNO_3$

Pre-Laboratory Questions

1. A saturated solution of the slightly soluble salt MA_3 is $1.4 \times10^{-3}M$ in MA_3. What is its K_{sp} value?
2. A 20.0 mL sample of $NaIO_3$ was treated with an excess of KI and HCl and the iodine produced titrated with 0.05125 *M* $Na_2S_2O_3$. 19.35 mL, of $Na_2S_2O_3$ solution were required to reach a starch-iodine endpoint. What was the concentration of the $NaIO_3$?