

# 17A. A Solubility Product Constant

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## Introduction

The interaction of a slightly soluble ionic compound with its dissolved ions leads to another type of equilibrium (Ebbing/Gammon, Chapter 17). The equilibrium constant for this kind of equilibrium has a special name. It is called the *solubility product constant*.

## Purpose

In this experiment, you will determine the solubility of  $\text{Ca}(\text{IO}_3)_2$  (calcium iodate) and calculate its solubility product constant. You will also demonstrate that the solubility of  $\text{Ca}(\text{IO}_3)_2$  is independent of the amount of water that is used to obtain a saturated solution of this substance.

## Concept of the Experiment

If the solubility of a slightly soluble ionic compound is measured, the solubility product constant for the substance can be calculated (Ebbing/Gammon, Example 17.3). This is the method that you will adopt for  $\text{Ca}(\text{IO}_3)_2$  in this experiment. The equilibrium between solid  $\text{Ca}(\text{IO}_3)_2$  and its ions in a saturated solution is



If some analytical technique is used to determine the concentration of either  $\text{Ca}^{2+}$  ions or  $\text{IO}_3^-$  ions in the saturated solution, the solubility of  $\text{Ca}(\text{IO}_3)_2$  will be known, and the solubility product constant ( $K_{sp}$ ) can be calculated.

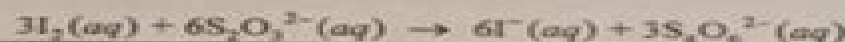
In this experiment, you will determine the concentration of  $\text{IO}_3^-$  ions through a titration with a standardized solution of  $\text{Na}_2\text{S}_2\text{O}_3$  (sodium thiosulfate) in the presence of KI (potassium iodide), using starch as an indicator. A definition of standardization can be found in the experiment "An Acid-Base Titration Curve." Potassium iodide will react with the iodate ions to give  $\text{I}_2$  as the sole product containing iodine. The reaction is



The molecular iodine reacts with  $\text{S}_2\text{O}_3^{2-}$  ions during the titration according to



where  $\text{S}_4\text{O}_6^{2-}$  is the tetrathionate ion. When these two equations are combined, multiplying the second reaction by 3 so the intermediate  $\text{I}_2$  is eliminated, the net reaction occurring in the titration is obtained.



Starch is used as an indicator in this titration because it reacts with  $\text{I}_2$  reversibly to form a dark blue color.  $\text{I}_2$  is consumed in the titration, so the color fades as the titration progresses. You will know when a stoichiometric volume of the  $\text{Na}_2\text{S}_2\text{O}_3$  solution has been added because at that point, one drop of that solution will cause the disappearance of the last trace of the blue color. A trial titration will enable you to find the approximate volume that is required before you do the first of two exact titrations.

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**David W. Oxtoby, H. P. Gillis, Norman H.  
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