

17A. A Solubility Product Constant

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 Ca^{2+} ions or 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 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 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 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 I_2 reversibly to form a dark blue color. 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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Lei Huang

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Table of Contents Solubility Product Constant Lab 17a Answers

1. Understanding the eBook Solubility Product Constant Lab 17a Answers
 - The Rise of Digital Reading Solubility Product Constant Lab 17a Answers
 - Advantages of eBooks Over Traditional Books
2. Identifying Solubility Product Constant Lab 17a Answers
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Solubility Product Constant Lab 17a Answers
 - User-Friendly Interface
4. Exploring eBook Recommendations from Solubility Product Constant Lab 17a Answers
 - Personalized Recommendations
 - Solubility Product Constant Lab 17a Answers User Reviews and Ratings
 - Solubility Product Constant Lab 17a Answers and Bestseller Lists
5. Accessing Solubility Product Constant Lab 17a Answers Free and Paid eBooks

- Solubility Product Constant Lab 17a Answers Public Domain eBooks
 - Solubility Product Constant Lab 17a Answers eBook Subscription Services
 - Solubility Product Constant Lab 17a Answers Budget-Friendly Options
6. Navigating Solubility Product Constant Lab 17a Answers eBook Formats
- ePub, PDF, MOBI, and More
 - Solubility Product Constant Lab 17a Answers Compatibility with Devices
 - Solubility Product Constant Lab 17a Answers Enhanced eBook Features
7. Enhancing Your Reading Experience
- Adjustable Fonts and Text Sizes of Solubility Product Constant Lab 17a Answers
 - Highlighting and Note-Taking Solubility Product Constant Lab 17a Answers
 - Interactive Elements Solubility Product Constant Lab 17a Answers
8. Staying Engaged with Solubility Product Constant Lab 17a Answers
- Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Solubility Product Constant Lab 17a Answers
9. Balancing eBooks and Physical Books Solubility Product Constant Lab 17a Answers
- Benefits of a Digital Library
 - Creating a Diverse Reading Collection Solubility Product Constant Lab 17a Answers
10. Overcoming Reading Challenges
- Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Solubility Product Constant Lab 17a Answers
- Setting Reading Goals Solubility Product Constant Lab 17a Answers
 - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Solubility Product Constant Lab 17a Answers
- Fact-Checking eBook Content of Solubility Product Constant Lab 17a Answers
 - Distinguishing Credible Sources
13. Promoting Lifelong Learning
- Utilizing eBooks for Skill Development

- Exploring Educational eBooks

14. Embracing eBook Trends

- Integration of Multimedia Elements
- Interactive and Gamified eBooks

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