

ANALYSIS OF COMMERCIAL BLEACH

To determine the amount of **sodium hypochlorite** (NaClO) in commercial bleach by means of an oxidation-reduction "back-titration".

Some commercial products that contain oxidizing agents are bleaches, hair coloring agents, scouring powders, and toilet bowl cleaners. The most common oxidizing agent in bleaches is **NaClO** . Commercial bleach is made by bubbling **chlorine gas** into a **sodium hydroxide** solution. Some of the chlorine oxidizes to the **hypochlorite ion** ClO^- , and some reduces to the **chloride ion** Cl^- . The solution remains strongly basic. The **net ionic equation** for this process is:



The hypochlorite ion ClO^- contains the **chlorine atom** in the **+1** oxidation state. Since the **chloride ion** Cl^- is much more stable (less potential energy) than the Cl^{+1} atom in the hypochlorite ion, the Cl^{+1} readily reduces to Cl^- and thus serves to "oxidize" other substances during the bleaching process.

The amount of **hypochlorite ion** present in a bleach solution is determined by an oxidation-reduction titration called a "back-titration". While the ClO^- ion cannot be titrated directly, it easily oxidizes the iodide ion I^- in **KI** to form iodine I_2 which is then titrated with a standard solution of **sodium thiosulfate** $\text{Na}_2\text{S}_2\text{O}_3$.

The analysis takes place in a series of steps:

- 1) **Acidified iodide ion** is added to **hypochlorite ion** solution, and iodide oxidizes to iodine.



- 2) **Iodine**, only slightly soluble in water, dissolves very well in an aqueous solution of iodide ions, in which it forms a **colored complex ion** called the triiodide ion I_3^- . Triiodide ion is a combination of a neutral I_2 molecule with an I^- ion. The $\text{I}_3^-_{(aq)}$ ion is yellow in dilute solution, and dark red-brown when concentrated.



- 3) The **triiodide ion** is titrated with a standard solution of thiosulfate ions, which reduces the iodine back to the colorless iodide ions.



During the last reaction, the red-brown color of the triiodide ion fades to yellow and then to the colorless iodide ion. While it is possible to use the disappearance of the color of the $\text{I}_3^-_{(aq)}$ ion as the method of determining endpoint, this is not a very sensitive procedure. **Addition of starch to a solution containing iodine (or triiodide ion) forms a reversible blue complex**. The **disappearance of this blue complex** is a much more sensitive method of determining endpoint. However, if the starch is added to a great deal of iodine, the complex which forms may not be reversible. Therefore, the starch is not added until shortly before the endpoint is reached. The quantity of thiosulfate used in step (3) is directly related to the amount of hypochlorite initially present. The $\text{S}_4\text{O}_6^{2-}$ ion is called the tetrathionate ion.

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Table of Contents Analysis Of Commercial Bleach Lab Answers

1. Understanding the eBook Analysis Of Commercial Bleach Lab Answers
 - The Rise of Digital Reading Analysis Of Commercial Bleach Lab Answers
 - Advantages of eBooks Over Traditional Books
2. Identifying Analysis Of Commercial Bleach Lab 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 Analysis Of Commercial Bleach Lab Answers
 - User-Friendly Interface
4. Exploring eBook Recommendations from Analysis Of Commercial Bleach Lab Answers
 - Personalized Recommendations

- Analysis Of Commercial Bleach Lab Answers User Reviews and Ratings
- Analysis Of Commercial Bleach Lab Answers and Bestseller Lists
- 5. Accessing Analysis Of Commercial Bleach Lab Answers Free and Paid eBooks
 - Analysis Of Commercial Bleach Lab Answers Public Domain eBooks
 - Analysis Of Commercial Bleach Lab Answers eBook Subscription Services
 - Analysis Of Commercial Bleach Lab Answers Budget-Friendly Options
- 6. Navigating Analysis Of Commercial Bleach Lab Answers eBook Formats
 - ePub, PDF, MOBI, and More
 - Analysis Of Commercial Bleach Lab Answers Compatibility with Devices
 - Analysis Of Commercial Bleach Lab Answers Enhanced eBook Features
- 7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Analysis Of Commercial Bleach Lab Answers
 - Highlighting and Note-Taking Analysis Of Commercial Bleach Lab Answers
 - Interactive Elements Analysis Of Commercial Bleach Lab Answers
- 8. Staying Engaged with Analysis Of Commercial Bleach Lab Answers
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Analysis Of Commercial Bleach Lab Answers
- 9. Balancing eBooks and Physical Books Analysis Of Commercial Bleach Lab Answers
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Analysis Of Commercial Bleach Lab Answers
- 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
- 11. Cultivating a Reading Routine Analysis Of Commercial Bleach Lab Answers
 - Setting Reading Goals Analysis Of Commercial Bleach Lab Answers
 - Carving Out Dedicated Reading Time
- 12. Sourcing Reliable Information of Analysis Of Commercial Bleach Lab Answers
 - Fact-Checking eBook Content of Analysis Of Commercial Bleach Lab Answers

- Distinguishing Credible Sources
- 13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
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 - Integration of Multimedia Elements
 - Interactive and Gamified eBooks

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