

## ANALYSIS OF COMMERCIAL BLEACH

To determine the amount of **sodium hypochlorite** ( $\text{NaClO}$ ) 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  **$\text{NaClO}$** . Commercial bleach is made by bubbling **chlorine gas** into a **sodium hydroxide** solution. Some of the chlorine oxidizes to the **hypochlorite ion**  $\text{ClO}^-$ , and some reduces to the **chloride ion**  $\text{Cl}^-$ . The solution remains strongly basic. The **net ionic equation** for this process is:



The hypochlorite ion  $\text{ClO}^-$  contains the **chlorine atom** in the **+1** oxidation state. Since the **chloride ion**  $\text{Cl}^-$  is much more stable (less potential energy) than the  $\text{Cl}^{+1}$  atom in the hypochlorite ion, the  $\text{Cl}^{+1}$  readily reduces to  $\text{Cl}^-$  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  $\text{ClO}^-$  ion cannot be titrated directly, it easily oxidizes the iodide ion  $\text{I}^-$  in **KI** to form iodine  $\text{I}_2$  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  $\text{I}_3^-$ . Triiodide ion is a combination of a neutral  $\text{I}_2$  molecule with an  $\text{I}^-$  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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