# REDOX TITRATION: Analysis of Commercial Bleach AP Chemistry

**Introduction:** Many commercial products are effective because they contain oxidizing agents. Some products that contain oxidizing agents are bleaches, hair coloring agents, scouring powders, and toilet bowl cleaners. The most common oxidizing agent in bleaches is sodium hypochlorite, NaClO. Commercial bleaches are created by bubbling chlorine gas into a sodium hydroxide solution. The amount of hypochlorite ion present in a solution of bleach can be determined by oxidation-reduction titration. One of the best methods is the iodine-thiosulfate titration procedure. Iodide ion, $I^-$, is easily oxidized by almost any oxidizing agent. In an acid solution, hypochlorite ions oxidize iodide ions to form iodine, $I_2$. The iodine that forms is then titrated with a standard solution of sodium thiosulfate.

Scientific Foundation:

The analysis takes place in a series of steps:

1. An acidified iodide ion is added to the hypochlorite ion solution, and the iodide is oxidized to iodine in the following *UNBALANCED* reaction.

$$ClO^-(aq) + I^-(aq) \text{--->} Cl^-(aq) + I_2(aq)$$

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

$$I_2(aq) + I^-(aq) \text{--->} I_3^-(aq)$$

3. The triiodide is titrated with a standard solution of thiosulfate ions, which reduces the iodine back to iodide ions:

$$I_3^-(aq) + 2S_2O_3^{2-}(aq) \text{--->} 3I^-(aq) + S_4O_6^{2-}(aq)$$

During this last reaction the red-brown color of the triiodide ion fades to yellow and then to the clear color of the iodide ion. It is possible to use the disappearance of the color of the $I_3^-$ ion as a method of determining the end point, but this is not a very sensitive procedure. Addition of starch to a solution that contains iodine or triiodide ion forms a reversible blue complex. The disappearance of this blue colored complex is a much more sensitive method of determining the end point. However, if the starch is added to the solution that contains a great deal of iodine, the complex that results may not be reversible. Therefore, the starch is not added until shortly before the end point is reached.

**Materials:**

| | |
|---|---|
| Transfer pipet, 5-mL, and bulb | Volumetric flask, 100 mL w/ stopper |
| Buret, Buret stand, and clamp | Erlenmeyer flask, 250 mL |
| Small beaker | 25 mL graduated cylinder |

**Purpose:** What you will do and how you will do it.

**Pre-lab Questions:**

1. In the introduction it was revealed how bleach was made. Balance this funky redox reaction in a basic medium using the ½ reaction method you learned in the aqueous chemistry unit. Please show all work! (8 points)

2. What is unique about the reaction you balanced in Q1? Make sure that you are being specific as to why and how you know this to be true. Provide evidence, definitions, etc. (4 points)

3. Why are strong oxidizing agents used as cleaners? What are oxidizing agents? What are some alternative oxidizing agents that could be used in place of bleach? What are the benefits and/or detractions of these alternatives? (8 points)

4. Find reaction #1 under the heading of *Scientific Foundation*. Balance the funky redox reaction in an acidic medium using the ½ reaction method you learned in the aqueous chemistry unit. Please show all work! (8 points)

5. What is the final, balanced, molar ratio between sodium thiosulfate and bleach? (2 points)

6. Why is it that iodide ions are so easily oxidized? (3 points)

7. The $I_3^-$ ion is sort of a strange one. What is the oxidation number for each I in this ion? Draw out the Lewis structure for this ion. Hint: The Is are in one chain, not a 3 membered ring. (4 points)

8. The reaction with thiosulfate ions produces the tetrathionate ion, $S_4O_6^{2-}$. Assuming O produces it's usual -2 oxidation state, calculate the oxidation state on the sulfur. The fact that the oxidation state comes out a fraction should tell you that the Ss in the ions do not all have the same oxidation state. In fact, the overall oxidation number is an average of the oxidation numbers on the sulfurs that make up the ion. With this in mind, what might be the logical break down of oxidation numbers on the four sulfurs? (8 points)

Create an Illustrated Procedure for the following steps:
*SAFETY ALERT – Concentrated bleach and hydrochloric acid are both damaging to skin, eyes, and clothing and they give off strong vapors. If you spill either solution on yourself, wash off with lots of water. Neutralize hydrochloric acid spills with baking soda.*

1. **Dilute the concentrated bleach**
   Use a pipet bulb and a 5-mL transfer pipet to measure out 5.00 mL of a commercial bleach solution into a 100 mL volumetric flask. Dilute to the mark with distilled water, stopper and mix well by inverting repeatedly.

2. **Measure the potassium iodide.**
   Weigh out approximately 2 g solid KI. This is a large excess over that which is needed.

3. **Oxidize the iodide ion with hypochlorite ion.**
   Carefully measure, using graduated cylinder, 25.00 mL of the dilute bleach into an Erlenmeyer flask. Add the solid KI and about 25 mL of distilled water. Swirl to dissolve the KI. Slowly, with swirling, add approximately 2 mL of 3 M HCL. The solution should turn dark yellow to red-brown from the presence of the $I_3^-$ complex.
   *SAFETY ALERT – Adding HCl to bleach may cause chlorine gas to be given off. Avoid smelling. You may be asked to do this step in the hood.*

3. **Titrate the iodine**
   Obtain about 60 mL of sodium thiosulfate solution (beaker). This should be enough for the entire experiment, including cleaning. Clean and prepare the buret appropriately (as you've done in earlier labs). Record the initial buret reading. Titrate with a standard 0.10 M sodium thiosulfate solution until the iodine color becomes light yellow. Add one dropper of starch solution (Don't add this at the beginning of the titration). The blue color of the starch-iodine complex should appear. Continue the titration until one drop of $Na_2S_2O_3$ solution causes the blue color to disappear. Record the final buret reading.
4. **Repeat**
   If time allots, repeat the titration beginning with step 2.
5. **Disposal**
   The solutions may be safely flushed down the drain with a large excess of water.

Data Table: Design your own! Only pertinent measurements for calculations are necessary.

Calculations:
1. Use the equations given to determine the mole ratio between moles of sodium thiosulfate and moles of sodium hypochlorite. SHOW WORK.
2. If you did more than one trial, calculate the average volume of $Na_2S_2O_3$ needed for the titration of 25.00 mL of diluted bleach.
3. Use the average volume and molarity of $Na_2S_2O_3$ to determine the molarity of the diluted bleach.
4. Calculate the molarity of the commercial (concentrated) bleach.
5. Assume that the density of the commercial bleach is 1.08 g/mL. Calculate the percent by mass of NaClO in the commercial bleach.
6. You will be given the actual percent by mass NaClO in commercial bleach according to the label. Calculate the percent error value (assuming the label is correct).

Post Lab Questions:
None