BLEACH ANALYSIS

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INTRODUCTION

Many commercially available liquid household bleaches are dilute solutions of the oxidizing agent, sodium hypochlorite. Bleach is used as a disinfectant because of its ability to oxidize the cell membranes of bacteria. It is also used to remove stains in clothing. The colors of many dyes and stains are due to the presence of multiple (double and triple) bonds in organic molecules. The color of the dye or stain is due to the ability of electrons in these multiple bonds to absorb and emit electromagnetic radiation in the visible region of the electromagnetic spectrum. When sodium hypochlorite in bleach comes in contact with these organic molecules, it oxidizes them, forming new products. The new products no longer have multiple bonds, so the color of the dye or stain looks whiter.

The percent "available chlorine" is used to denote the strength of a bleaching agent. It is actually the ratio of the mass of Cl_2 to the mass of the bleach mixture expressed as a percentage, even if Cl_2 is not actually present in the bleach. If sodium hypochlorite is the bleaching agent in a bleach solution, the percent "available chlorine" is determined using the following equation that shows how sodium hypochlorite can be produced from chlorine and sodium hydroxide.

$$Cl_2(aq) + 2\ NaOH(aq) \rightarrow NaClO(aq) + NaCl(aq) + H_2O(l)$$

Suppose a bleach solution is 8.75%, by mass, sodium hypochlorite. The percent "available chlorine" could be calculated as follows:

$$?\,g\,Cl_2 = 100\,g\,soln\ \frac{8.75\,g\,NaOCl}{100\,g\,soln}\ \frac{1\,mol\,NaOCl}{74.44\,g\,NaOCl}\ \frac{1\,mol\,Cl_2}{1\,mol\,NaOCl}\ \frac{70.90\,g\,Cl_2}{1\,mol\,Cl_2} = 8.33\,g\,Cl_2$$

and the percent "available chlorine" would be 8.33% Cl_2.

EXPERIMENTAL OVERVIEW

In this lab, you will use a standard potassium iodate, KIO_3, solution to standardize a sodium thiosulfate, $Na_2S_2O_3$, solution.

Since potassium iodate is available in its pure crystalline form, we can provide you with a standard potassium iodate solution. We don't provide you with a standard sodium thiosulfate solution because sodium thiosulfate pentahydrate, which is used to prepare the solution, absorbs water from moist air, releases water in warm, dry air, and cannot be dried (mp 48^{o}C) without some decomposition. A solution having an approximate concentration can be prepared by weighing out and dissolving the appropriate amount of solute in water, but the exact concentration of the solution must be determined by titration.

$$IO_3^-(aq) + 5\ I^-(aq) + 6\ H^+(aq) \rightarrow 3\ I_2(aq) + 3\ H_2O(l)$$

$$I_2(aq) + 2\ S_2O_3^{2-}(aq) \rightarrow 2\ I^-(aq) + S_4O_6^{2-}(aq)$$

Once the sodium thiosulfate solution has been standardized, it can be used to determine the oxidizing power ("available chlorine") of a bleach sample.

$$ClO^-(aq) + 2\ I^-(aq) + H_2O(l) \rightarrow I_2(aq) + Cl^-(aq) + 2\ OH^-(aq)$$

$$I_2(aq) + 2\ S_2O_3^{2-}(aq) \rightarrow 2\ I^-(aq) + S_4O_6^{2-}(aq)$$

If you examine both of the above pairs of reactions, you should notice that iodine, I_2, is generated in the first reaction of each pair and used as an **indicator** in the second reaction of each pair.

An iodine solution has to be stabilized if it is to be useful as an indicator. Iodine has a high vapor pressure (sublimes easily) and low solubility in aqueous solution. Thus, iodine would be lost as it is generated (by oxidation of iodide) yielding poor results. To get around this problem, excess iodide, I^-, is added to the solution in the form of KI. The addition of excess iodide (from KI) favors the formation of the stable, soluble triiodide ion, I_3^-.

$$I_2(aq) + I^-(aq) \leftrightarrow I_3^-(aq)$$

During the titration with the sodium thiosulfate solution, equilibrium favors the release of more iodine from triodide until nearly all I_2 is titrated. I_2 **at very low concentrations** (near the endpoint of the titration) forms a blue complex with starch:

$$\underset{\text{colorless}}{I_2(aq)} + starch(aq) \leftrightarrow \underset{\text{blue}}{I_2{\cdot}starch\ (aq)}$$

After addition of the starch, sodium thiosulfate can continue to be used to react with the I_2 to provide an accurate, visible endpoint (disappearance of blue color) for the reaction. Effectively, this reaction at the end of the titration is a shift in the equilibrium of the above equation to the left; note that as the I_2 is consumed, the equilibrium shifts dropping the $I_2{\cdot}$starch concentration, causing the solution to turn from blue to colorless at the endpoint.

PROCEDURE

Each **pair** of students will titrate one brand of commercial liquid bleach to find the amount of "available chlorine" and the percent, by mass, of sodium hypochlorite in a commercial liquid bleach.

Part A: Standardization of $Na_2S_2O_3$

1. Fill a buret with the provided $Na_2S_2O_3$ solution.

2. Pipet 25 mL of the standard KIO_3 solution into a 125 ml Erlenmeyer flask. Add about 2 g of KI and 10 mL of 0.5 M H_2SO_4.

$$IO_3^-(aq) + 5\,I^-(aq) + 6\,H^+(aq) \rightarrow 3\,I_2(aq) + 3\,H_2O(l)$$

3. Titrate the I_2 solution with the $Na_2S_2O_3$ solution.

$$I_2(aq) + 2\,S_2O_3^{2-}(aq) \rightarrow 2\,I^-(aq) + S_4O_6^{2-}(aq)$$

   In order to detect the endpoint of this titration you need to do a few things and make some careful observations:

   a. When performing the titration, add the $Na_2S_2O_3$ until the color of the solution is **PALE YELLOW.** At this point, you are very close to the endpoint.

   b. Once you have reached the pale yellow solution, add 2 mL of the starch solution. This will allow you to better observe the endpoint. After you have added the starch, the solution will turn a deep blue indicating that there is still $I_2(aq)$ present in the solution. Continue the titration until the blue color completely disappears. Perform this part of the titration **slowly** and carefully since you are already very close to the endpoint.

4. Repeat the titration at least two more times. Do not add the KI and H_2SO_4 until you are ready to titrate that individual sample.

Part B: Analysis of a Liquid Bleach Sample

1. Pipet 10 mL of your assigned bleach sample into a 100 mL volumetric flask and dilute to the mark with water, being sure to thoroughly mix the sample. Pipet 25 mL of this diluted solution into a 250 mL Erlenmeyer flask. Add about 2 g of KI and 10 mL of 0.5 M H_2SO_4.

2. Titrate the diluted bleach sample with the sodium thiosulfate solution. Follow the same procedure you used when standardizing the sodium thiosulfate solution. Here the $I_2(aq)$ is produced from the reaction of ClO^- with I^-.

$$ClO^-(aq) + 2\,I^-(aq) + H_2O(l) \rightarrow I_2(aq) + Cl^-(aq) + 2\,OH^-(aq)$$

$$I_2(aq) + 2\,S_2O_3^{2-}(aq) \rightarrow 2\,I^-(aq) + S_4O_6^{2-}(aq)$$

3. Repeat the titration at least two more times. Do not add the KI and H_2SO_4 until you are ready to titrate that individual sample.

DATA AND ANALYSIS SHEET: BLEACH ANALYSIS

Name: ____

Date ____________ Lab Partner ___________________________________

Bleach sample identification: ______________

Concentration of the prepared KIO_3 solution: ____________________

Part A: Standardization of $Na_2S_2O_3$

	Trial 1	Trial 2	Trial 3
mL of KIO_3 solution			
mol KIO_3 titrated			
final buret reading			
initial buret reading			
mL $Na_2S_2O_3$ added			
mol $Na_2S_2O_3$ added			
molarity of $Na_2S_2O_3$ solution			
average molarity of $Na_2S_2O_3$ solution			

Sample calculations (Trial 1):

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Name: ____

Part B: Analysis of a Liquid Bleach Sample

	Trial 1	Trial 2	Trial 3
volume of original bleach sample (mL) titrated	2.5		
mass, in g, of original bleach sample titrated*			
final buret reading			
initial buret reading			
mL $Na_2S_2O_3$ added			
mol $Na_2S_2O_3$ added			
mol ClO^- reacted			
mol "available chlorine" in titrated sample			
grams "available chlorine" in titrated sample			
percent "available chlorine" in original sample			
average percent "available chlorine" in original sample			
average percent, by mass, of NaOCl in original sample			

*assume the density of the bleach sample is 1.084 g/mL

Sample calculations (Trial 1).

Calculation of percent, by mass, of NaOCl in your bleach sample.