Determination of Sodium Hypochlorite Levels in Bleach

Household bleach is a solution of sodium hypochlorite (NaOCl) and water. It is widely used as a disinfectant and in the bleaching of textiles and paper (if you have ever spilled a little bleach on your clothes, you probably noticed this immediately). Its ability to whiten textiles was discovered in 1787 by the French Chemist Berthollet. Louis Pasteur discovered the disinfectant properties of NaOCl in the late ninteenth century. Sodium hypochlorite effectively kills bacteria, fungi, and viruses.

The mechanism of the disinfectant activity of chlorine bleach is not completely known. It seems that NaOCl can pass through bacterial cell walls. Once inside the cell, the oxidizing action of sodium hypochlorite and byproducts could result in the destruction of many things inside the cell that would lead to cell death (enzymes, nucleic acids, etc).

In the past, most consumer chlorine bleach was sold in a 5.25% solution ([kg NaOCl/kg solution] x 100%). Today, some more concentrated solutions are being sold and touted as an improvement over the less concentrated bleaches. The increased concentration itself is not an improvement (the less concentrated bleaches are usually diluted themselves). The improvement is that less is needed and that will save space and money.

In this experiment, the amount of sodium hypochlorite in two bottles of bleach and the relative cost of the two bleaches will be determined.

Overview of the Experiement

In order to determine the amount of bleach in each bottle, a red-ox titration will be performed. This red-ox titration is similar to an acid-base titration, a reagent of known concentration is added to the unknown until a color change occurs. In this redox equation, the color change is from brown-yellow (I_2) to colorless (I^-).

The buret in the titration will be filled with $Na_2S_2O_3$ and will act as a reducing agent. When the $S_2O_3^{2-}$ ion reacts with the I_2 in the flask, I^- is formed (equation). When all of the I_2 reacts, the solution becomes clear.

$$I_{2(brown)} + 2S_2O_3^{2-} \text{------>} S_4O_6^{2-} + 2I^-_{(clear)} \qquad (1)$$

Excercise: Break this equation into it's two half reactions. Write them in the space below. If you add the two half reactions together, do the electrons cancel ?

So where did the I_2 come from?

Reacting the bleach directly with the $S_2O_3^{2-}$ solution would result in reduction of the OCl^- to Cl^- according to equation 2. As you can see, both

$$OCl^-_{\text{(faint yellow)}} + 2S_2O_3^{2-}{}_{\text{colorless)}} \text{---->} Cl^-_{\text{(colorless)}} + S_4O_6^{2-}{}_{\text{(colorless)}} \quad (2)$$

solutions are mostly colorless. It would be difficult to see the transition from faint yellow to colorless. For this reason, the bleach solution is added to an I^- solution (equation 3)

$$OCl^- + 2I^-_{\text{(clear)}} \text{---->} I_{2\text{(brown)}} + Cl^- \quad (3)$$

So 1.00 mole of bleach results in the formation of 1.00 moles of I_2. According to equation 1, 0.500 moles of I_2 then reacts with 2.00 mole of $S_2O_3^{2-}$. The result is that for 1.00 mole of OCl^- in the flask, 2.00 moles of $S_2O_3^{2-}$ are required to reach the endpoint.

Sample Calculation.

Assume you added 1.00 mL of bleach to the I^- solution in an Erlenmeyer flask. The initial reading on the buret of $S_2O_3^{2-}$ is 0.15 mL and the final reading is 35.45 mL. The volume of $S_2O_3^{2-}$ solution is :

$$35.45 \text{ mL} - 0.15 \text{ mL} = 35.30 \text{ mL (or 0.03530 L)} \quad (4)$$

If the concentration of the $S_2O_3^{2-}$ solution is 0.250 M, the amount of $S_2O_3^{2-}$ used is

$$\frac{0.03530L}{1} \times \frac{0.250moles}{L} = 8.83 \times 10^{-3} moles \quad (5)$$

But what is needed is the concentration of OCl^-. For every one mole of OCl^-, two moles of $S_2O_3^{2-}$ is required, so

$$(8.83 \text{ x } 10^{-3} \text{ moles } S_2O_3^{2-})\text{x}\frac{1moleOCl^-}{2molesS_2O_3^{2-}} = 1.77 \text{ x } 10^{-2} \text{ moles } OCl^-$$

Because 1.0 mL (0.0010 L) of OCl^- solution was initially added (what came from the bottle):

$$\text{concentration of } OCl^- \text{ solution} = \frac{1.77 \times 10^{-2} molesOCl^-}{0.0010L} = 17.7 \text{ M } OCl^-$$

Experimental

The solutions used include:

•Two types of bleach (one normal and one ‘concentrated”)

•An acidic I^- solution prepared from 6.3 g of NaI (or 7.0 g of KI) and 20 mL glacial acetic acid diluted to 1L with water

•A 0.100 M solution of $Na_2S_2O_3$

Part 1.

To begin the titration, wash out the buret first with water. After you have drained the water out, add ~ 5 mL of the $Na_2S_2O_3$ solution to the buret and twist the buret while pouring slowly to allow the $Na_2S_2O_3$ to contact the entire inner surface of the buret.

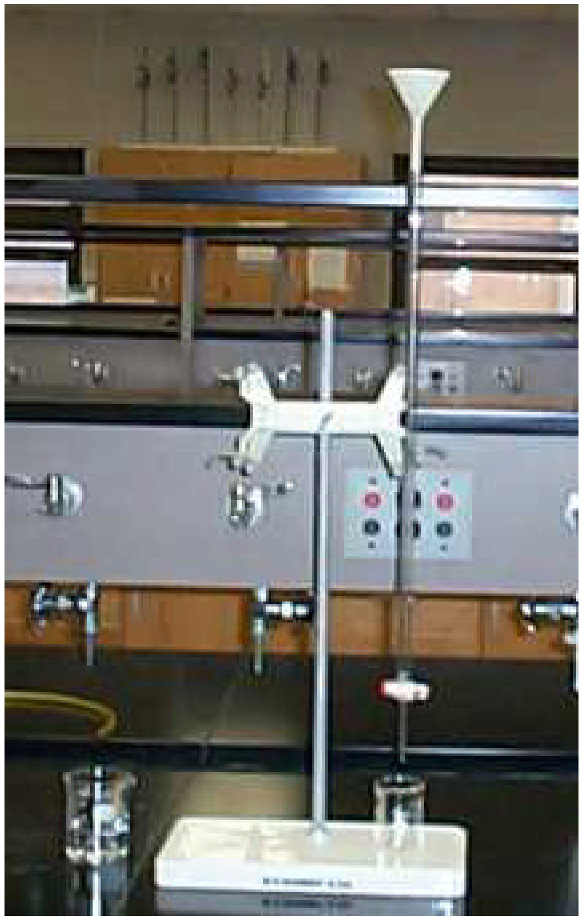


Figure 1. The buret with funnel for filling.

Make sure the stopcock is closed (perpendicular to the buret) and fill the buret close to the 0 mL mark (but not above) with the $Na_2S_2O_3$ solution. With a beaker under the stopcock, drain some of the solution out the stopcock until the air bubbles in the tip are gone (you may need to add a bit more $Na_2S_2O_3$ to the buret afterwards),

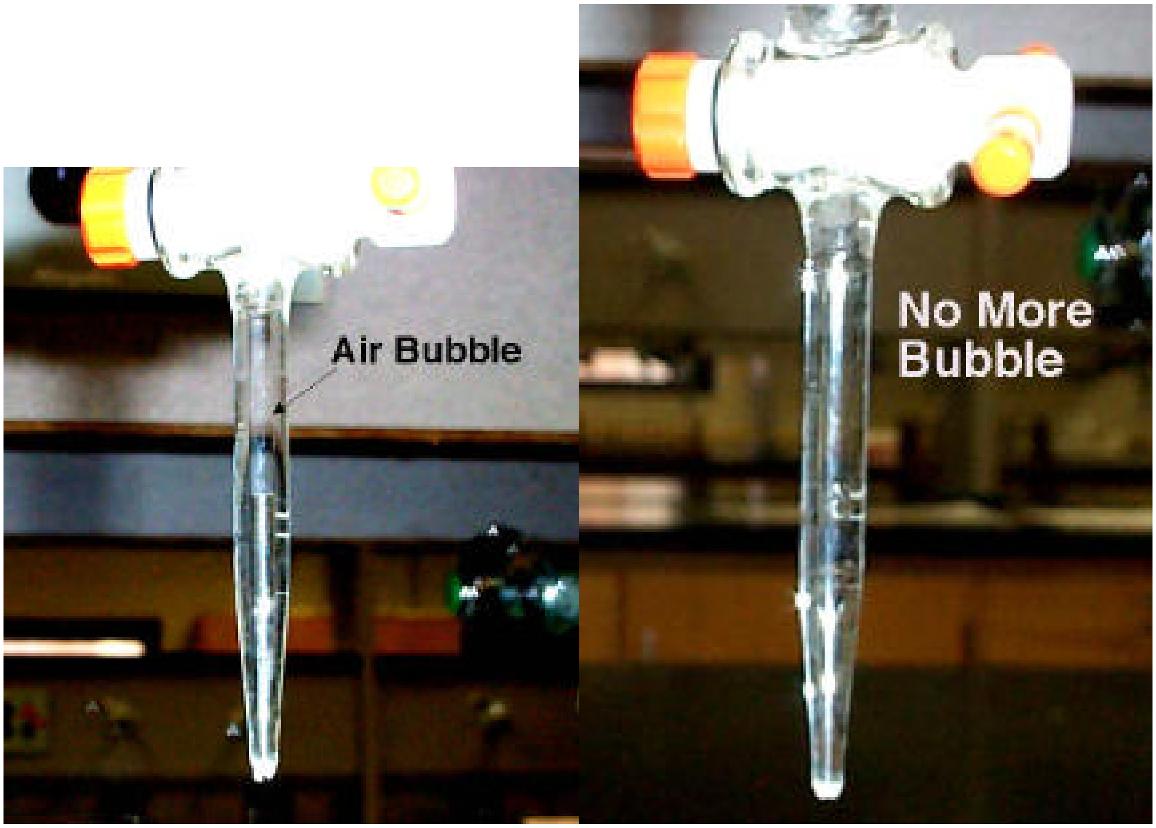


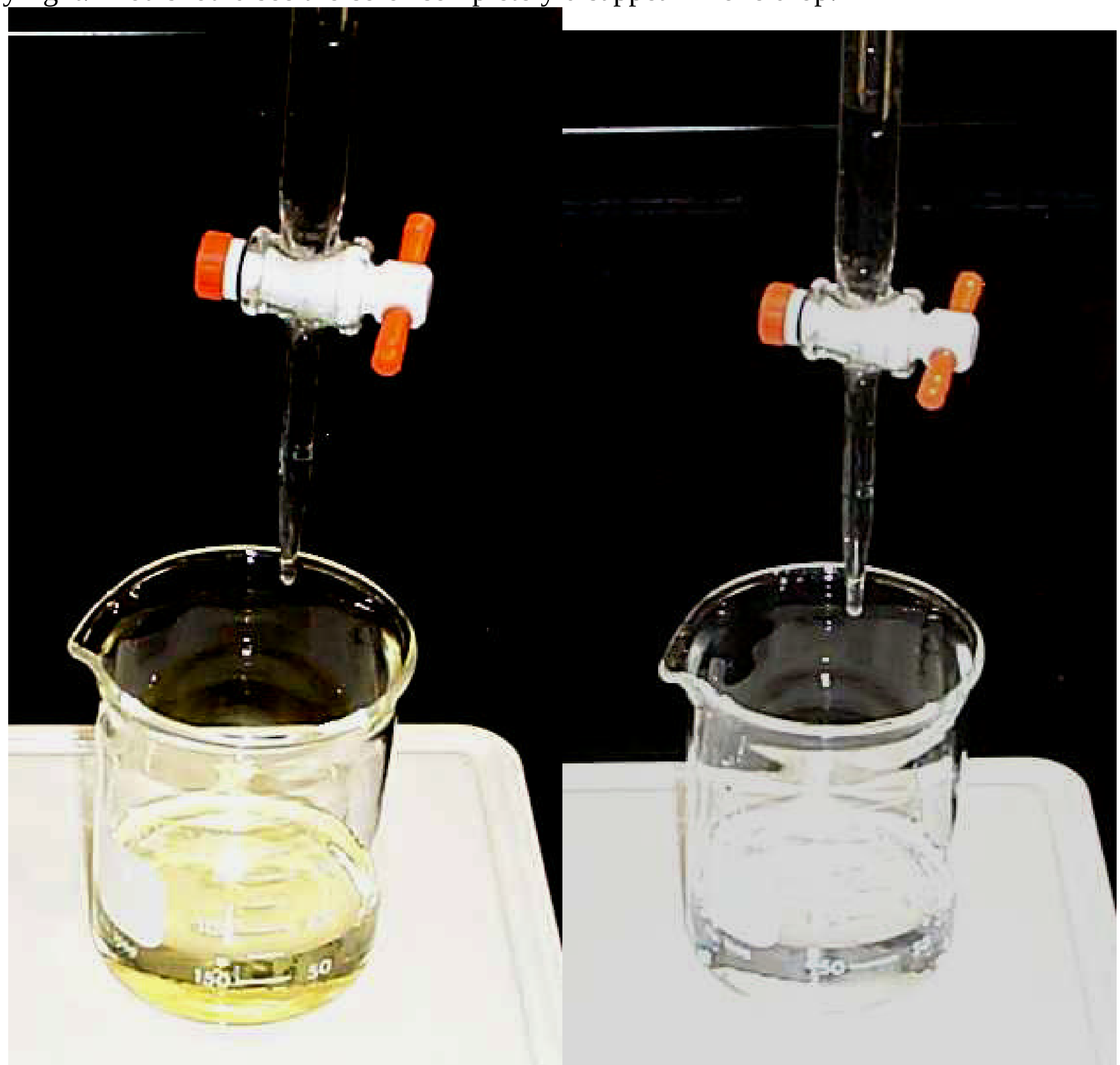
Figure 2. Buret tip with bubble, and without bubble

Add 50 mL of the acid/I^- solution to a 125 mL Erlenmeyer flask or 250 mL beaker. Add 1.0 mL of your bleach to flask as well. When the bleach is added, the solution should turn brown.

(before) (after)

Figure 3. Iodide solution before addition of bleach and after addition of bleach

Record the initial volume of the $S_2O_3^{2-}$ in the buret. Slowly add the $S_2O_3^{2-}$ solution to the bleach in the flask or beaker. After you each mL or so, stop and swirl the solution. Do not add the solution too quickly (but you don't need to add it drop by drop yet). The color of the iodine will get lighter and lighter. You will need to add the solution dropwise as the solution becomes very light. You should see the color completely disappear in one drop.



(a) (b)

Figure 4. Titration soution two drops before the endpoint (a) and at the endpoint (b)

Calculations

(1) Record the volume of the buret at the endpoint.

(2) Subtract the initial volume from the final volume to determine the volume of $Na_2S_2O_3$ used

(3) Using the concentration of the $Na_2S_2O_3$ (given by your instructor), determine the number of moles of $S_2O_3^{2-}$ used.

(4) Determine the number of moles of NaOCl that were in the flask or beaker (see sample calculations above)

(5) Using the volume of bleach you added to the flask or beaker, determine the concentration of the NaOCl in the bleach.

Part 2. Repeat the steps in part 1 with the other bleach.

Part 3. Comparison

Using the molarity of the two bleaches (from parts 1 and 2) determine the number of moles of NaOCl in each of the bleach bottles (when full). Using the formula weight of NaOCl, determine the mass of NaOCl in each bottle. If you have the price of the two bottles, divide the price for each bottle by the number of moles of NaOCl in each bottle (you units should be $/mole). Which bottle is cheaper? Is the concentrated bleach a good buy? Are there any hazards to using the more concentrated bleach?