Archer G11

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Analysis of a Commercial Bleach

Purpose: The purpose of this lab is to determine the amount of sodium hypochlorite (NaClO) in commercial bleach. This can be done by forming triiodide ions. To make the measurement more accurate, starch was added to help determine the endpoint of the solution. The significance of this lab is that industry can use these techniques to determine the amount of NaClO in the bleach of the rival industry and improve it.

Hypothesis: The hypothesis is that, an accurate determination of NaClO in commercial bleach can be done. By mixing the acidified iodide ion to the hypochlorite solution, the iodide is oxidized to iodine which forms complex triiodide ions that give the red-brown color to the solution. Because the endpoint of the titration of triiodide is hard to determine, starch is added to give the solution a dark blue color. If starch was not added, the color would be turning from yellow to clear. It is quite hard to distinguish between the two colors and therefore, starch was added so the color would turn from yellow to dark blue. This makes it a lot easier to determine the endpoint.

Materials:

Materials	Quantities
Distilled water	100 mL
5% bleach (NaClO)	5 mL
3 M Hydrochloric acid (HCl)	6 mL
0.100 M Sodium thiosulfate ($Na_2S_2O_3$)	100 mL
2% Starch solution	3 dropperful
Potassium iodide (KI)	6 g
Burette	1 burette
Burette clamp	1 burette clamp
Ring stand	1 ring stand
Small funnel	1 funnel
Pipette	6 pipettes
0.0001-g precision balance	1 balance
Wash glass	3 wash glasses
Spatula	1 spatula
Rubber spatula	1 rubber spatula
125-mL Erlenmeyer flask	6 flasks
Size 2 rubber stopper	1 rubber stopper
Stirring rod	1 stirring rod
10-mL graduated cylinder	1 cylinder
25-mL beaker	1 beaker

Procedures:

1.) Measure out 5 mL of NaClO into a 125-mL Erlenmeyer flask

2.) Add 95 mL of distilled water to the flask
3.) Mass out 2 g of solid KI
4.) Transfer 25 mL of the diluted NaClO in the flask to another 125-mL Erlenmeyer flask
5.) Repeat step 3 and 4 for two more times
6.) Label the flask: Trial 1, Trial 2, Trial 3
7.) Add 2 g of solid KI to the Trial 1 flask
8.) Swirl the flask to dissolve the KI
9.) Working in the fume hood, stir the solution
10.) Add 2 mL of 3 M HCl to the flask while stirring
11.) The solution should become red-brown color
12.) Rinse a burette with distilled water
13.) Re-rinse the burette with 0.100 M $Na_2S_2O_3$
14.) Attach the burette to the ring stand
15.) Fill the burette with 0.100 M $Na_2S_2O_3$ until the volume goes above the 0-mL line
16.) Put a 25-mL beaker under the burette
17.) Turn the burette cork so that the solution flows into the 25-mL
18.) Turn the cork off when the volume of the $Na_2S_2O_3$ reaches the 0-mL line
19.) Titrate the solution in the Trial 1 flask until the solution turns from red-brown to light yellow
20.) Add a dropperful of starch solution to the Trial 1 flask
21.) The solution should become blue
22.) Swirl the flask
23.) Titrate the solution in the Trial 1 flask until the blue color disappear
24.) Record the final burette reading
25.) Repeat step 7 to 24 for Trial 2 and Trial 3 flasks

Results: During the first titration, the solution turns from red-brown to orange-brown before becoming yellow. The starch solution cannot be stored for a long period, and it feels somewhat sticky. The blue color after adding the starch solution is a dark shade of blue. The solution looks similar to ink. During the second titration, the blue color turns to orange then to yellow if more $Na_2S_2O_3$ is added.

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

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

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

$$2I^+(aq) + ClO^-(aq) + 2S_2O_3^{2-}(aq) \rightarrow Cl^-(aq) + H_2O(l) + S_4O_6^{2-}(aq)$$

The ratio of sodium hypochlorite ($NaClO^-$) to sodium thiosulfate ($Na_2S_2O_3$) is 1:2

Titration of Iodine Solution			
Molarity of $Na_2S_2O_3$ (M)	0.100		
Molar mass of NaClO (g)	74.44		
	Trial 1	**Trial 2**	**Trial 3**
Volume of original bleach (mL)	5		
Volume of diluted bleach (mL)	25	25	25
Mass of KI (g)	2.0021	2.0007	2.0000
Initial burette reading (mL)	0	14.5	29
Final burette reading (mL)	14.5	29	41.5
Volume of $Na_2S_2O_3$ added (mL)	14.5	14.5	12.5
Average volume added (mL)	13.8		
Moles of $Na_2S_2O_3$ added (mols)	1.45×10^{-3}	1.45×10^{-3}	1.25×10^{-3}
Average moles added	1.38×10^{-3}		
Moles of ClO^- in diluted bleach	6.92×10^{-4}		
Average molarity of ClO^- in diluted bleach (M)	0.0277		
Molarity of the original bleach (M)	0.553		
Mass of NaClO	0.2		
Mass of commercial bleach	5.4		
Percent of NaClO in the commercial bleach (%)	3.81		
Standard deviation of each trial (mL)	0.89		
Percent error (%)	23.7		

Volume if $Na_2S_2O_3$ added = (Final burette reading) – (Initial burette reading)

Trial 1: 14.5 – 0 = 14.5 mL

Trial 2: 29 – 14.5 = 14.5 mL

Trial 3: 41.5 – 29 = 12.5 mL

Average volume added = Σ(Volume of $Na_2S_2O_3$ added) ÷ 3

(14.5 + 14.5 + 12.5) ÷ 3 = 13.83 mL

Moles of $Na_2S_2O_3$ added = (Volume of $Na_2S_2O_3$ added) × (Molarity of $Na_2S_2O_3$)

Trial 1: $0.0145 \times 0.100 = 1.45 \times 10^{-3}$ moles $Na_2S_2O_3$

Trial 2: $0.0145 \times 0.100 = 1.45 \times 10^{-3}$ moles $Na_2S_2O_3$

Trial 3: $0.0125 \times 0.100 = 1.25 \times 10^{-3}$ moles $Na_2S_2O_3$

Average moles added = (Average volume added) × (Molarity of $Na_2S_2O_3$)

$0.01383 \times 0.100 = 1.383 \times 10^{-3.}$ moles $Na_2S_2O_3$

Moles of ClO^- in diluted bleach = (Average moles added) ÷ 2

$(1.383 \times 10^{-3}) \div 2 = 6.915 \times 10^{-4}$ moles ClO^-

Average molarity of ClO^- in diluted bleach = (Moles of ClO^- in diluted bleach) ÷ (Volume of diluted bleach)

$(6.915 \times 10^{-4}) \div 0.025 = 0.02766$ M ClO^-

Molarity of the original bleach = [(Average molarity of ClO^- in diluted bleach) × (0.100 L)] ÷ (Volume of original bleach)

$(0.02766 \times 0.100) \div 0.005 = 0.5532$ M ClO^-

Mass of NaClO = [(Molarity of the original bleach) × (Volume of original bleach)] × (Molar mass of NaClO)

$(0.5532 \times 0.005) \times 74.44 = 0.2059$ g NaClO

Mass of commercial bleach = (Density of commercial bleach) × (Volume of original bleach)

$1.08 \times 5 = 5.4$ g

Percent of NaClO in the commercial bleach = [(Mass of NaClO) ÷ (Mass of commercial bleach)] × 100%

$(0.2059 \div 5.4) \times 100\% = 3.813\%$

Standard deviation of each trial = Σ|(Volume of $Na_2S_2O_3$ added) – (Average volume added)| ÷ 3

$(|14.5 - 13.83| + |14.5 - 13.83| + |12.5 - 13.83|) \div 3 = 0.89$ mL

Percent error = {[(Actual percent of NaClO in the commercial bleach) – (Percent of NaClO in the commercial bleach)] ÷ (Actual percent of NaClO in the commercial bleach)]} × 100%

$[(5 - 3.81) \div 5] \times 100 = 23.74\%$

Analysis: The hypothesis could not be verified true. This is because the results acquired from the experiment have the percent error of 23%, which shows that it is not accurate. Oxidation refers to the losing of electron, resulting in an element having a more positive oxidation number. Reduction, on the other hand, refers to the gaining of electron, resulting in an element having a more negative oxidation number.

$$\mathrm{Cl_2}(g) + 2\mathrm{OH^-}(aq) \rightarrow \mathrm{ClO^-}(aq) + \mathrm{Cl^-}(aq) + \mathrm{H_2O}(l)$$

$$\text{Oxidation half} - \text{reaction: } \mathrm{Cl_2}(g) + 4\mathrm{OH^-}(aq) \rightarrow 2\mathrm{ClO^-}(aq) + 2\mathrm{H_2O}(l) + 2\mathrm{e^-}$$

Cl is being oxidized from the oxidation number of 0 to 1+

$$\text{Reduction half} - \text{reaction: } \mathrm{Cl_2}(g) + 2\mathrm{e^-} \rightarrow 2\mathrm{Cl^-}(aq)$$

Cl is being reduced from 0 to 1-

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

$$\text{Oxidation half} - \text{reaction: } 3I^-(aq) \rightarrow I_3^-(aq) + 2e^-$$

I^- is being oxidized from 1- pt 1/3-

$$\text{Reduction half} - \text{reaction: } 3I_2(aq) + 2e^- \rightarrow 2I_3^-(aq)$$

I_2 is being reduced from 0 to 1/3-

The advantage of using an aliquot is that it helps the molarity of the bleach to be closer to the molarity of the $Na_2S_2O_3$. The use of the bleach with similar molarity to the $Na_2S_2O_3$ use for titrating allows for an accurate measurement. Roughly, four aliquot can be measured from a 100-mL volumetric flask. However, the last one would probably be slightly less than 25 mL due to the solution sticking to the inside of the flask. S in $S_4O_6^{2-}$ probably all has the same oxidation number of 2.5. If the pipet was rinsed with distilled water immediately before being use, the percent of NaClO in the commercial bleach would be too low. This is because the distilled water left in the pipet from the rinsing would dilute the commercial bleach even further. Thus, lower amount of $Na_2S_2O_3$ would need to be use to titrate the solution. If three grams of KI were used instead of two grams, it would not affect the percent of NaClO in the commercial bleach. This is because the I^- ions would simply react with the ClO^- until it is the reaction goes to completion and the rest of the I^- would simply be an excess. If some I^- ions were to sublime from the solution, it would not affect the solution as long as enough I^- ions remain for the reactions. A major source of experimental error is the difficulty in determining the endpoint during the titration due to the gradual change in color of the solution.

Conclusion: The hypothesis could, to some extent, be confirmed. The results did not match with the hypothesis completely. There are possible reasons that could cause this. An error could be that the titration was not done correctly. Because of the gradual change in color during the titration, the endpoint was hard to determine. During the first part of titration, the solution gradually changes from red-brown to dark orange, light orange, then to yellow. Thus, it is possible that the titration was stopped before the endpoint was reached. This would have caused the amount of $Na_2S_2O_3$ added to be lower than normal. This would results in a decrease in the amount of NaClO determined, leading to a decrease in percent of NaClO in the commercial bleach. Another error could be that the dilution of the commercial bleach was not done precisely. The dilution of the commercial bleach was done using a 250-mL Erlenmeyer flask which might not have the accurate label of volume. This could have caused too much distilled water to be added and diluted the bleach too much. Therefore, this may have caused the commercial bleach to be too diluted and, therefore, decrease the amount of $Na_2S_2O_3$ needed to titrate the solution. This would lead to a lower percentage of NaClO in the commercial bleach. The first error could be solved by using the pH meter. The endpoint is mark by the rapid change in pH. Therefore, using the pH meter could help in accurate determination of the endpoint of the titration. The second error could be solved by using a graduated cylinder rather than the volumetric flask to dilute the solution. This

would help because the graduated cylinder has more accurate volume-measurement-lines than the volumetric flask.