Lab Report: Clock Reaction

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1 BACKGROUND

In this "iodine clock" chemical reaction, potassium iodate (KIO_3) and sodium hydrogen sulfite ($NaHSO_3$) undergo a reaction that results in the generation of elemental iodine (I_2). The subsequent equation illustrates the conversion of these two initial compounds into three resulting products, with iodide being the most significant among them.

$$IO_3{}^-(aq) + 3\,HSO_3{}^-(aq) \longrightarrow I^-(aq)_3{}^+ SO_4{}^{2-}(aq) + 3\,H^+(aq)$$
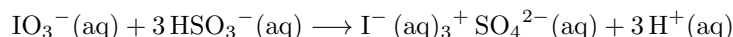
Once all the hydrogen sulfite has been completely consumed, it serves as the limiting reactant in this experiment. Subsequently, the surplus iodate combines with the iodide ions, leading to the production of iodine.

$$IO_3{}^-(aq) + 5\,I^-(aq) + 6\,H^+(aq) \longrightarrow 3\,I_2(s) + 3\,H_2O(l)$$
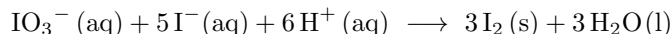
A solid iodine product doesn't appear since the iodine molecules quickly interact with starch to create a blue complex known as iodine-starch.

Because elemental iodine possesses a distinct color (unlike all the other reactants and products which are colorless), it is possible to monitor the reaction's progress by simply measuring the time it takes for the iodine color to become visible. To enhance the visibility of this color change, a small quantity of starch is introduced to intensify the iodine's color. Starch forms a vivid black/blue complex with iodine. Detecting the initial presence of iodine by itself would be challenging (as the solution would only have a very faint yellow color), but with starch present, the first few iodine molecules produced react with the starch, resulting in a much more noticeable and pronounced change in color.

The rate equation governing the reaction between hydrogen sulfite ion and iodate ion is anticipated to follow the general structure represented as

$$\text{Rate} = k \cdot [HSO_3{}^-]^x \cdot [IO_3{}^-]^y$$

, where x signifies the reaction's order concerning the concentration of hydrogen sulfite ion, and y signifies the reaction's order concerning the concentration of the iodate ion. It's important to emphasize that, despite knowing the stoichiometric coefficients in the balanced chemical reaction, these coefficients are not the same as the exponents in the rate equation. The specific orders must be established through experimental observation.

2   AIM

This practical is designed to conduct a quantitative study of the effect of concentration of reactants on clock reaction rate.

- Measure the time required to produce a specific amount of product in a chemical reaction.

- Explain how the concentration of reactants affects the rate of the reaction.

- Make a graph of the results of these rate studies.

- Determine the order of $[HSO_3^-]$ and $[IO_3^-]$ in the rate equation.

3   MATERIALS (PER PAIR)

0.02M potassium iodate (KIO_3)	0.001M sodium hydrogen sulfite ($NaHSO_3$)
2 100-mL beaker	1 plastic wash bottle
1 stirring rod	Soluble Starch
1 thermometer	Distilled water
1 stopwatch	5 80-mL beaker

4   PROCEDURE

1. Clean the beakers.

2. Fill 3 clean beakers with deionized water, KIO_3 and $NaHSO_3$, respectively.

3. Mix the solutions according to the table below:

Experiment label	IO_3^- (mL)	Water (mL)	HSO_3^- (mL)
1	10	0	10
2	8	2	10
3	6	4	10
4	4	6	10
5	2	8	10
6	10	0	10
7	10	2	8
8	10	4	6
9	10	6	4
10	10	8	2

Specifically, for experiments 1-5, mix IO_3^- and water, add 4 drops of starch to the mixture. Then measure 10 mL of HSO_3^- and pour it into the beaker, at the same time begin to record the time. (Similar for experiments 6-10.)

4. When the reaction starts, as a change in the color of the mixture has changed, the timing is stopped and the time is recorded in seconds. The same experiment is repeated for 3 times.

5   Data Collection

$V_{IO_3^-}$ (mL, ± 0.05 mL)	V_{H_2O} (mL, , ± 0.05 mL)	Time 1 (s)	Time 2 (s)	Time 3 (s)
10	0	10.37	10.43	9.89
8	2	13.55	14.64	14.65
6	4	18.56	18.03	18.19
4	6	27.67	27.54	29.89
2	8	54.9	50.37	52.63

Table 1: Raw data collection for experiment 1-5.

6   Example of Data Processing

Here we take the 2nd experiment with 8 mL of IO_3^-, 2 mL of H_2O and 10 mL of HSO_3^- as an example.

Consider the second experiment with 8 mL of IO_3^-, 2 mL of H_2O and 10 mL of HSO_3^-, the concentration of IO_3^- is

$$[IO_3^-] = \frac{8}{8+2+10} \times 0.02 = 0.008 \text{mol L}^{-1}$$

The concentration of HSO_3^- is constant,

$$[HSO_3^-] = \frac{10}{20} \times 0.001 = 0.0005 \text{mol L}^{-1}$$

The limiting reagent is HSO_3^- and therefore, when calculating the rate, the rate of reaction is the concentration of hydrogen sulfite changes with respect to time

$$\text{Rate} = \frac{\Delta[HSO_3^-]}{3t} = \frac{[HSO_3^-]}{3t}$$

Here Δt should be considered as the average time

$$t = \text{average time} = \frac{\text{time 1} + \text{time 2} + \text{time 3}}{3} = \frac{13.55 + 14.64 + 14.65}{3} = 14.28$$

Therefore, the rate of reaction

$$\text{Rate} = \frac{[HSO_3^-]}{3t} = \frac{0.0005}{3 \times 14.28} = 1.16713 \times 10^{-4} \text{mol L}^{-1} \text{s}^{-1}$$

Then, to calculate the error, first we consider the error of Δt which is equal to the standard deviation of time 1, time 2 and time 3.

$V_{HSO_3^-}$ (mL, ± 0.05 mL)	Water (mL, ± 0.05 mL)	Time 1 (s)	Time 2 (s)	Time 3 (s)
10	0	10.37	10.43	9.89
8	2	13.3	16.7	16.6
6	4	22.8	20.68	21.08
4	6	47.34	39.23	42.69
2	8	109.32	102.7	103.49

Table 2: Raw data collection for experiment 6-10.

$$\Delta t = \text{Standard Deviation} = \sqrt{\frac{\sum_{i=1}^{3}(t_i - \Delta t)^2}{3}} = 0.516204094$$

The measure of the volume of different solutions has a fixed error of ± 0.05 mL.

Concentration of HSO_3^- is $\dfrac{V_{HSO_3^-}}{V_{total}}$. Therefore, the percentage error of the concentration of $HSO_3^- = $ percentage error of HSO_3^-'s volume + the percentage error of total volume.

$$\begin{aligned}\frac{\Delta[HSO_3^-]}{[HSO_3^-]} &= \frac{\Delta V_{HSO_3^-}}{V_{HSO_3^-}} + \frac{\Delta V_{total}}{V_{total}} \\ &= \frac{0.05}{10} + \frac{0.05 \times 3}{20} \\ &= 0.0125\end{aligned}$$

Similarly, the percentage error of $[IO_3^-]$

$$\begin{aligned}\frac{\Delta[IO_3^-]}{[IO_3^-]} &= \frac{\Delta V_{IO_3^-}}{V_{IO_3^-}} + \frac{\Delta V_{total}}{V_{total}} \\ &= \frac{0.05}{8} + \frac{0.05 \times 3}{20} \\ &= 0.01375\end{aligned}$$

The error of $[IO_3^-]$

$$\begin{aligned}\Delta[IO_3^-] &= [IO_3-] \times \frac{\Delta[IO_3^-]}{[IO_3^-]} \\ &= 0.008 \times 0.01375 \\ &= 1.1 \times 10^{-4}\end{aligned}$$

Therefore, the percentage error of rate is the same as the percentage error of concentration of HSO_3^- + percentage error of t

$$\begin{aligned}\frac{\Delta \text{Rate}}{\text{Rate}} &= \frac{\Delta[HSO_3^-]}{[HSO_3^-]} + \frac{\Delta t}{t} \\ &= 0.0125 + \frac{0.516204094}{14.28} \\ &\approx 4.86 \times 10^{-2}\end{aligned}$$

and the error of rate

$$\Delta \text{Rate} = 4.86 \times 10^{-2} \times 1.16713 \times 10^{-4} = \pm 5.67796 \times 10^{-7} \text{mol L}^{-1}\text{s}^{-1}$$

Therefore, for the second experiment, the independent variable $[IO_3^-]$ is $(8 \pm 0.11) \times 10^{-3}\text{mol L}^{-1}$, and the dependent variable rate of reaction is $1.17 \pm 0.06 \times 10^{-5}\text{mol L}^{-1}\text{s}^{-1}$.

7   DATA PROCESSING

By the same action we took for experiment 2, we can calculate the rate of reaction and the concentration for all experiments and the results are shown in the table below.

7.1 Data without Error

$[\text{IO}_3{}^-]$ (mol L^{-1})	$[\text{HSO}_3{}^-]$ (mol L^{-1})	Average Time t (s)	Rate $(\text{mol L}^{-1}\,\text{s}^{-1})$
0.01	0.0005	10.23	1.63×10^{-5}
0.008	0.0005	14.28	1.17×10^{-5}
0.006	0.0005	18.26	9.13×10^{-6}
0.004	0.0005	28.37	5.88×10^{-6}
0.002	0.0005	52.63	3.17×10^{-6}

Table 3: Processed data for experiments 6-10 with concentration of $\text{HSO}_3{}^-$ as its independent variable.

$[\text{IO}_3{}^-]$ (mol L^{-1})	$[\text{HSO}_3{}^-]$ (mol L^{-1})	Average Time t (s)	Rate $(\text{mol L}^{-1}\,\text{s}^{-1})$
0.01	0.0005	10.23	1.63×10^{-5}
0.01	0.0004	15.53	8.58×10^{-6}
0.01	0.0003	21.52	4.65×10^{-6}
0.01	0.0002	43.09	1.55×10^{-6}
0.01	0.0001	105.17	3.17×10^{-7}

Table 4: Processed data for experiments 6-10 with concentration of $\text{HSO}_3{}^-$ as its independent variable.

7.2 Error Data

Doing error calculation by similar action in determining the second experiment, the following errors are calculated in order to plot the error bar.

8 Conclusion and Evaluation

Therefore, by the above datas, we can draw the diagram of rate of reaction against the concentration of $\text{HSO}_3{}^-$ and $\text{IO}_3{}^-$.

From the 2 diagrams in Figure 1, we can see that after using the power function to fit with the scatter points, the power for rate of reaction against $\text{HSO}_3{}^-$ is 2.4406, close to 2. And the power for rate of reaction against $\text{IO}_3{}^-$ is 0.9986, close to 1.

Therefore, the rate of reaction$= k \cdot [\text{HSO}_3{}^-]^2 \cdot [\text{IO}_3{}^-]$, corresponding with the theoretical value.

In this experiment, we delved into the intriguing clock reaction between sulfite ions ($\text{HSO}_3{}^-$) and iodate ions ($\text{IO}_3{}^-$). Characterized by its striking color change from colorless to a vivid blue, this reaction provided an avenue for investigating reaction kinetics and chemical dynamics. The primary objectives were to understand the factors influencing the reaction rate and to determine the rate law governing this distinctive clock reaction.

Our experimental findings revealed several notable insights:

Rate Expression: Through our data analysis, we observed that the rate of the reaction closely aligns with the rate expression: Rate $= [\text{HSO}_3{}^-]^2 \times [\text{IO}_3{}^-]$. This power relationship, particularly the square dependence on $\text{HSO}_3{}^-$, suggests an intriguing aspect of the reaction kinetics that warrants further investigation.

% Error [IO3-]	% Error [HSO3-]	Δt (s)	% Error Rate	y Error Bar $(\text{mol L}^{-1}\,\text{s}^{-1})$	x Error Bar (mol L^{-1})
0.0125	0.0125	0.242	0.0361	5.89×10^{-7}	0.000125
0.0138	0.0125	0.516	0.0486	5.68×10^{-7}	0.00011
0.0158	0.0125	0.222	0.0247	2.25×10^{-7}	0.000095
0.02	0.0125	1.08	0.0505	2.97×10^{-7}	0.00008
0.0325	0.0125	1.85	0.0476	1.51×10^{-7}	0.000065

Table 5: Errors to calculate the vertical and horizontal error bars for experiment 1-5.

% Error [HSO3-]	Δt (s)	% Error Rate	y Error Bar $(\mathrm{mol\,L^{-1}\,s^{-1}})$	x Error Bar $(\mathrm{mol\,L^{-1}})$
0.0125	0.242	0.0361	5.89×10^{-7}	0.00000625
0.01375	1.58	0.115	9.91×10^{-7}	0.0000055
0.0158	0.920	0.0586	2.72×10^{-7}	0.00000475
0.02	3.32	0.0971	1.50×10^{-7}	0.000004
0.0325	2.95	0.0606	1.92×10^{-8}	0.00000325

Table 6: Errors to calculate the vertical and horizontal error bars for experiment 6-10.

(a) Rate of reaction against $\mathrm{HSO_3}^-$
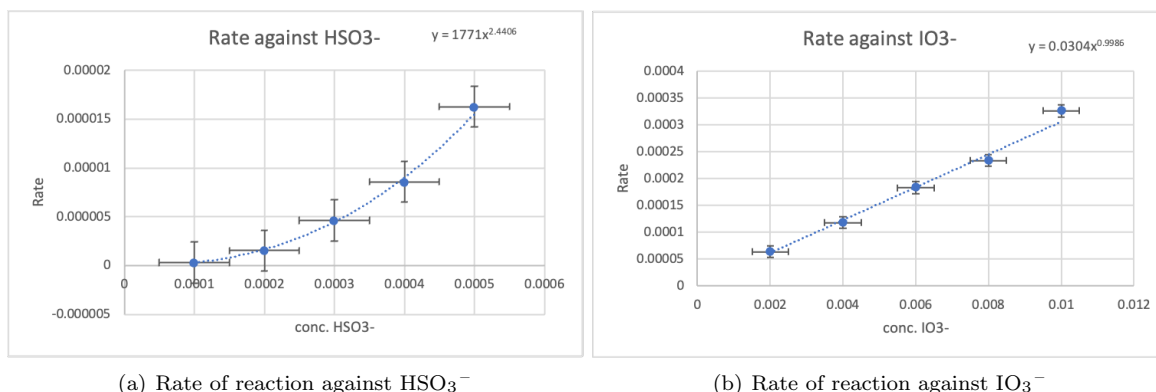

(b) Rate of reaction against $\mathrm{IO_3}^-$

Figure 1: Diagrams of the rate of reaction against reactants' concentration.

One notable challenge in our experiment was accurately determining the reaction time. The transformation of color from colorless to blue is swift and may not be easy to discern with precision using a stopwatch, particularly given the abrupt nature of the reaction. This highlights the importance of obtaining more advanced time measurement equipment or exploring alternative methods to enhance the accuracy of the rate determination. The inaccuracy led to a less precise value of the power of $[\mathrm{HSO_3}^-]$ in our calculation compared to the theoretical value.

In conclusion, our exploration of the clock reaction between $\mathrm{HSO_3}^-$ and $\mathrm{IO_3}^-$ has provided valuable insights into the kinetics and rate expression of this unique reaction. We acknowledge the challenge of accurately measuring the reaction time using conventional stopwatches, which may have introduced some uncertainty into our results. The rate expression, Rate $= [\mathrm{HSO_3}^-]^2 \times [\mathrm{IO_3}^-]$.

This conclusion addresses the calculation of the rate expression, mentions the timing challenge, and emphasizes the potential for future research to refine the understanding of the reaction kinetics.