Analyzing the Relationship Between the Concentration of KIO_3 and the Speed of the $NaHSO_3$-based Iodine Clock Reaction

Tarik Onalan

31 March 2015

1   Introduction

1.1   Purpose

1.2   Hypothesis

Reaction rate will decrease (increasing the time taken to react) as KIO_3 concentration decreases, as there will be fewer KIO_3 molecules available to react at any one time.

1.3   Variables

Independent Variable

- KIO_3 concentration [M]

Dependent Variable

- Reaction speed [s]

Controlled Variables

- Molarity of $NaHSO_3$ solution [0.020 M]

- Volume of KIO_3 solution [10.000 mL]

- Volume of $NaHSO_3$ solution [10.000 mL]

- Temperature of solutions [25.200 °C]

- Distilled Water

2   Materials

- 75 mL 0.020 M KIO_3

- 100 mL 0.020 M $NaHSO_3$

- 9 · test tubes

- 2 · 100 mL beaker
- 2 · 10 mL graduated cylinder
- 1 · 2 mL pipette
- 1 · timer
- 1 · thermometer

3 Procedure

1. Pour KIO_3 and $NaHSO_3$ solutions into beakers
2. Measure 10 mL KIO_3 solution in graduated cylinder
3. Measure 10 mL $NaHSO_3$ solution in graduated cylinder
4. If KIO_3 solution is not 10 mL, fill up to 10 mL with distilled water
5. Pour both solutions into test tube
6. Start timer
7. Wait until solution turns blue/black
8. Stop timer
9. Record time taken to react
10. Repeat steps $2-9$ 2 times with 8 mL and 6 mL KIO_3 solution, respectively
11. Repeat steps $1-10$ as necessary for data collection

4 Data

Table 1: Data

Molarity	Trial 1 s	Trial 2 s	Trial 3 s	avg
0.020±0.50	8.030±0.50	9.180±0.50	9.510±0.50	8.91±0.50
0.016±0.50	11.000±0.50	11.120±0.50	11.590±0.50	11.24±0.50
0.012±0.50	18.460±0.50	15.820±0.50	15.450±0.50	16.58±0.50

Reaction Speed Relative to the Molarity of KIO_3 Solution
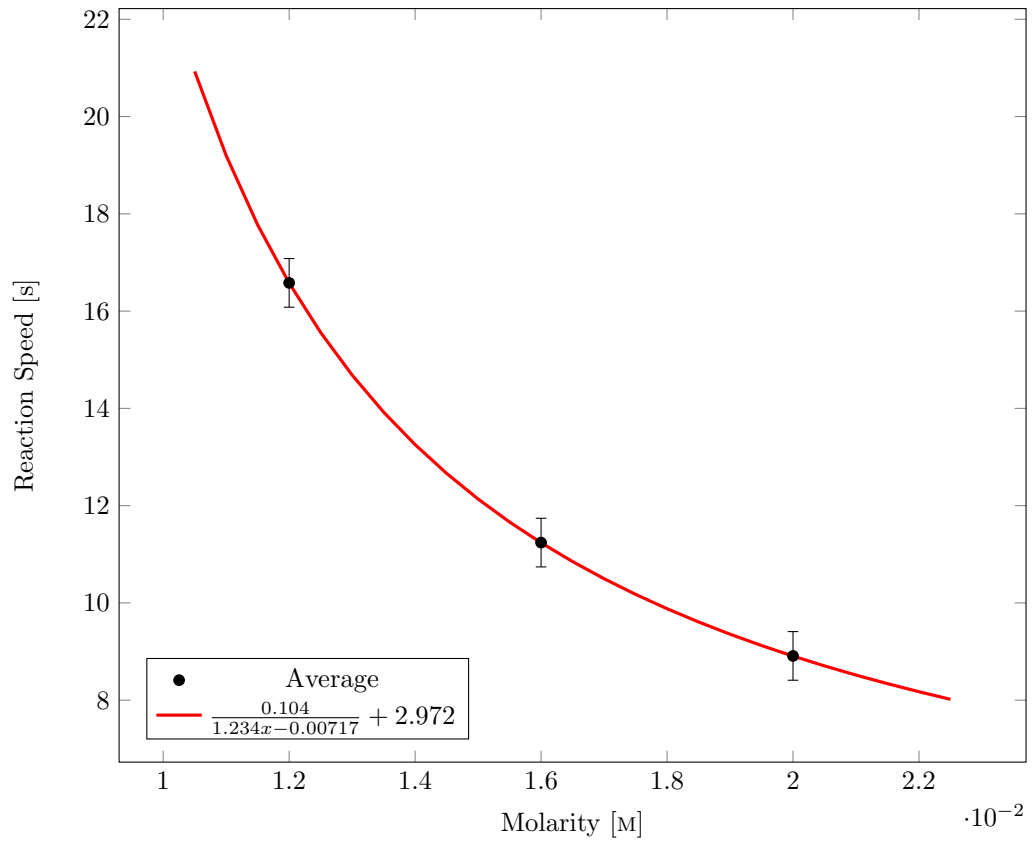
5 Calculations

5.1 Average

$$\langle s(M) \rangle \tag{1}$$

$$\frac{\sum_i s(M)}{i} \tag{2}$$

$$\frac{s(M_1) + s(M_2) + ... + s(M_{i-1}) + s(M_i)}{i} \tag{3}$$

3

5.2 Error and Minimization

Error was calculated using sum squared error, which is defined as follows:

$$E = \sum_{i=1}^{n} (y_i - f(x_i))^2 \tag{4}$$

A program carried out the minimization of the above function, which entails tracing the derivative of the function with respect to a given parameter —defined here as k—to the local minima of the function

$$E' = \frac{\partial}{\partial k} \left(\sum_{i=1}^{n} (y_i - f(x_i))^2 \right) \tag{5}$$

5.3 Line of Best Fit

When choosing my line of best fit, I first tried to piece together all of the information that I could gather without data. First I knew that my function $?(x)$ had to satisfy the following conditions for the following scenarios:

$$?(x) = \begin{cases} \text{Zero Order:} & [A] = -kt + [A]_0 \\ \text{First Order:} & \ln [A] = -kt + \ln [A]_0 \\ \text{Second Order:} & \frac{1}{[A]} = kt + \frac{1}{[A]_0} \\ \text{nth Order:} & \frac{1}{[A]^{n-1}} = (n-1)kt + \frac{1}{[A]_0^{n-1}} \end{cases} \tag{6}$$

which I simplified to

$$?(x) = \begin{cases} f(x) = a(bx + c) + d \\ g(x) = a \ln(bx + c) + d \\ h(x) = \frac{a}{bx+c} + d \end{cases} \tag{7}$$

I ended up having to assume that the rate of reaction was less than or equal to 2 primarily due to the fact that ternary (third-order or more) reactions are extremely rare, and testing for one would subsequently be impractical. Judging by the appearance of the graphed data—steep slope for smaller x, small slope for larger x—I had an intuition that the best model would be the second order reaction:

$$h(x) = \frac{a}{bx + c} + d \tag{8}$$

Testing my prediction proved me correct, as the error was much lower for the function $h(x)$—representing a second order reaction—than for $g(x)$ or $f(x)$—representing first and zeroeth order reactions, respectively. These results are shown in the table below.

Error		
$f(x)$	$g(x)$	$h(x)$
$1.510\,02$	$13.954\,68$	$6.513\,84 \cdot 10^{-8}$

6 Conclusion

My hypothesis, that the iodine clock reaction would take longer the lower the concentration of KIO_3, was correct. The data I collected shows—throughout

all 3 trials—that as KIO_3 concentration decreases, the reaction takes longer to "complete", per se. The average, as shown by the graph, has a definite downward trend, starting at $(0.012, 16.58)$ and ending at $(0.02, 8.91)$. Another proof could be the fact that the data almost perfectly fits the line of best fit, with a sum squared error of $6.513\,84 \cdot 10^{-8}$; the best-fit model is a power function with an exponent of -1, which means that it is monotonically decreasing on the interval $(0, +\infty)$, so it would be prudent to assume that any further data points—further dilutions pushing the concentration to the "left" of the graph— would follow the model, increasing the time taken to react.

The only major difficulty in this lab was making sure that the KIO_3 and $NaHSO_3$ solutions did not contaminate each other. I did experience minor issues in trial 1, molarity 0.02 and trial 1, molarity 0.012. In the former, I dropped a trace amount of KIO_3 into my $NaHSO_3$, causing the reaction to proceed faster than normal. In the latter reaction, I did the opposite, dropping a trace amount of $NaHSO_3$ into my KIO_3, reacting and further lowering the already low concentration of KIO_3. Even so, my conclusion is still valid, as the other trials $(2-3)$ display the same inverse correlation between molarity and time taken for the solutions to react.

In the future, I should make sure that all of my solutions are spaced sufficiently apart, so there is little likelihood that I drop solution into a graduated cylinder or test tube before I mean to mix the two solutions. Given that this is still a "likelihood", however, I could also run more trials and increase the number of test cases—both of which were 3 for this experiment—both to compensate for any possible execution errors, and to have a more robust dataset by which to make my conclusions, as I found that 3 points over 3 trials was much too sparse to make a certain conclusion with any level of comfort.