

Chunyang Ding

Mr. Rierson

AP/IB Chemistry SL

28 January 2013

### Clocking the Effect of Molarity on Speed of Reaction

In basic levels of chemistry, most of the experimenter's attention is on the reaction and the products that stem from it, while much less attention is placed on the actual speed of the reaction. While most people do assume that the temperature of the solution is often the most deciding factor in the rate of production, it is possible that there are many other factors that would have a more powerful, and less intuitive, effect. In order to study the effect of a molarity of solution on the time of reaction, the reactants of the "Clock reaction", or the reaction between  $KIO_3$  and  $NaHSO_3$ , will be modified.

Prior to the experiment, there was some basic knowledge of how an optimal balance of chemicals would result in all chemicals being consumed, and therefore produce the largest amount of result. However, what was not considered was the speed at which this should happen. Before this experiment, it could have been assumed that the reaction would take place almost immediately, as most reactions occur as soon as one compound meets with another compound. The largest discrepancies would be if aqueous solutions did not mix easily, therefore leading to a large portion of unreacted solution on either sides, but that does not seem to be the case in this experiment. Instead, it seems that there might be multiple reactions that would ultimately control the production of Iodide!

Although there is very little information about these reactions, it is possible to use our experiment to take an educated guess at what is happening within the reaction. In order to do this, we have designed the following lab.

**How does the molarity of  $KIO_3$  have an impact on the time for the color changing reaction between Iodine and starch?**

If we increase the molarity of  $KIO_3$ , there will be an optimal spot for the minimum speed, and then gradually get slower, as modeled by an upwards facing parabolic curve of some sort. This makes sense because there would be an optimal balance between  $KIO_3$  and  $NaHSO_3$ .

In order to execute the experiment, dilutions were made of  $KIO_3$  with regular water, in order to change its molarity. This would provide the independent variable that was changed between experiments. The dependent variable was the time it took for the color to change, and therefore for the Iodine to form. This variable was timed as when the color was thoroughly saturated throughout the solution, not at the first hint of a dark color. Controlled variables included the temperature of the solutions, as it had been previously predicted that a difference in temperature might independent result in a difference in the rate of reaction. The volume of total solution was also kept constant, so that there was not particularly more water in any trial, and so there could be no argument that the differing volume created a larger space for the reaction to take place, and thus result in a slower result. Finally, another controlled variable was concentration of the  $NaHSO_3$ . While this seems like a rather obvious control, it is worth mentioning solely because if this was allowed to change, there would be a remarkably ineffective experiment with multiple independent variables, and no clear trend may be found between any concentrations with the time of reaction.

If our hypothesis is correct, when we graph the data we will see a clear point where there is the “optimal” mixture of concentration of one reactant to another. Before and after that point, you would be able to see an increase in the amount of time needed for the reaction, as this would only exist at one point.

Materials:

The materials we used included:

**Chemicals:**

- 125 mL of  $NaHSO_3$ , 0.0500M
- 125 mL of  $KIO_3$ , 0.0500M

**Equipment:**

- 5 plastic cups
- 3 250 mL Beakers
- 2 10mL Pipets
- 1 Pipet seal
- Stopwatch
- Flask Stopper

**Diagram**

**Data Tables:**

Concentration of $KIO_3$ (M)	Time of Reaction (s)				
	Trial 1	Trial 2	Trial 3	Trial 4	Trial 5
0.01					
0.02					
0.03					
0.04					
0.05					

**Procedure:**

- 1) Obtain all chemicals in Erlenmeyer flasks. Keep the stopper on the beaker with the chemical  $NaHSO_3$  in it.
- 2) Fill one 250 mL beaker with distilled water, and label as “water”.
- 3) Label a pipet as  $KIO_3$ , and another as  $NaHSO_3$ .
- 4) Label a beaker as “dilute solution”.
- 5) Using the  $KIO_3$  pipet, draw out 5 mL of  $KIO_3$ , and place into the beaker labeled “dilute solution”.
- 6) Using the same pipet, draw out 20 mL of water and place into the “dilute solution” beaker.
- 7) Mix the “dilute solution” beaker slightly.
- 8) Using the same “ $KIO_3$ ” pipet, dispose 5 mL of the solution into a paper cup.
- 9) Repeat step 8 for each of the paper cups.
- 10) Place 20 mL of regular water into every paper cup.
- 11) Take off the stopper on the  $NaHSO_3$  flask.
- 12) Using the  $NaHSO_3$  pipet, draw out 5mL of  $NaHSO_3$  and place into the first paper cup, while starting the stopwatch.

- 13) After the color change seems to have penetrated through the solution, stop the stopwatch.
- 14) Record the time in the data table.
- 15) Repeat steps 12 through 14 five times, for each trial.
- 16) Place the stopper back onto the  $NaHSO_3$  flask.
- 17) Clean out the plastic cups by dumping the solution, and rinsing out with water.\*
- 18) Repeat steps 5 through 17 with differing concentrations of  $KIO_3$ , as determined by a varying ration of  $KIO_3$  and water in the “dilute solution” beaker. (10:15 for 0.02M; 15:10 for 0.03M; 20:5 for 0.04M; 25:0 for 0.05M)

\*Note: Be careful when rinsing out the cups. In our experiment, during the cleaning process for the third set of experimental data, one of the plastic cups was accidentally cracked and therefore unusable. A 250 mL beaker was used instead, leading to a possible error that is covered in the error analysis.

Through our experiment, we found it easiest to have one person be in charge of dealing with the dilutions and the administration of the chemicals, while another was in charge of clean up, timing, and other tasks. This allowed our group to be reasonably efficient and not lag behind.

### **Data Collection:**

Trials (Time of rxn, seconds)					
Molarity of $KIO_3$ (M)	1	2	3	4	5
0.01	142.33	181.42	179.59	182.89	178.60
0.02	56.68	55.03	58.91	56.29	61.62
0.03	31.38	31.86	32.29	30.91	31.33
0.04	22.76	19.58	20.85	21.56	27.37
0.05	18.09	18.25	16.70	18.29	18.13

Tracking of uncertainties:

### **Molarity Calculation:**

Molarity calculated by  $\frac{\text{Molarity of } KIO_3 * \text{Number of Liters of } KIO_3}{\text{Number of liters of dilute solution}}$  so that the uncertainty can be calculated by adding together all the percent uncertainties, and then multiplying by the final result. This number differs depending on the concentration, so steps will not be shown for all different dilutions, but for only one test case.

$$\frac{\text{Uncert } M}{M} = \frac{0.0001M}{0.0500M}$$

$$\frac{(\text{Uncert Volume})}{\text{Total Volume}} = \frac{0.0001 L}{0.0050L}$$

$$\frac{(\text{Uncert Volume Dilute})}{\text{Total Volume}} = \frac{0.0001L}{0.0250L}$$

$$\% \text{Uncert } M + \% \text{Uncert } V_1 + \% \text{Uncert } V_2 = 0.026\%$$

$$\text{Actual Uncertainty} = \text{Percent Uncertainty} * \text{Actual Result}$$

$$= 0.026 * 0.01M$$

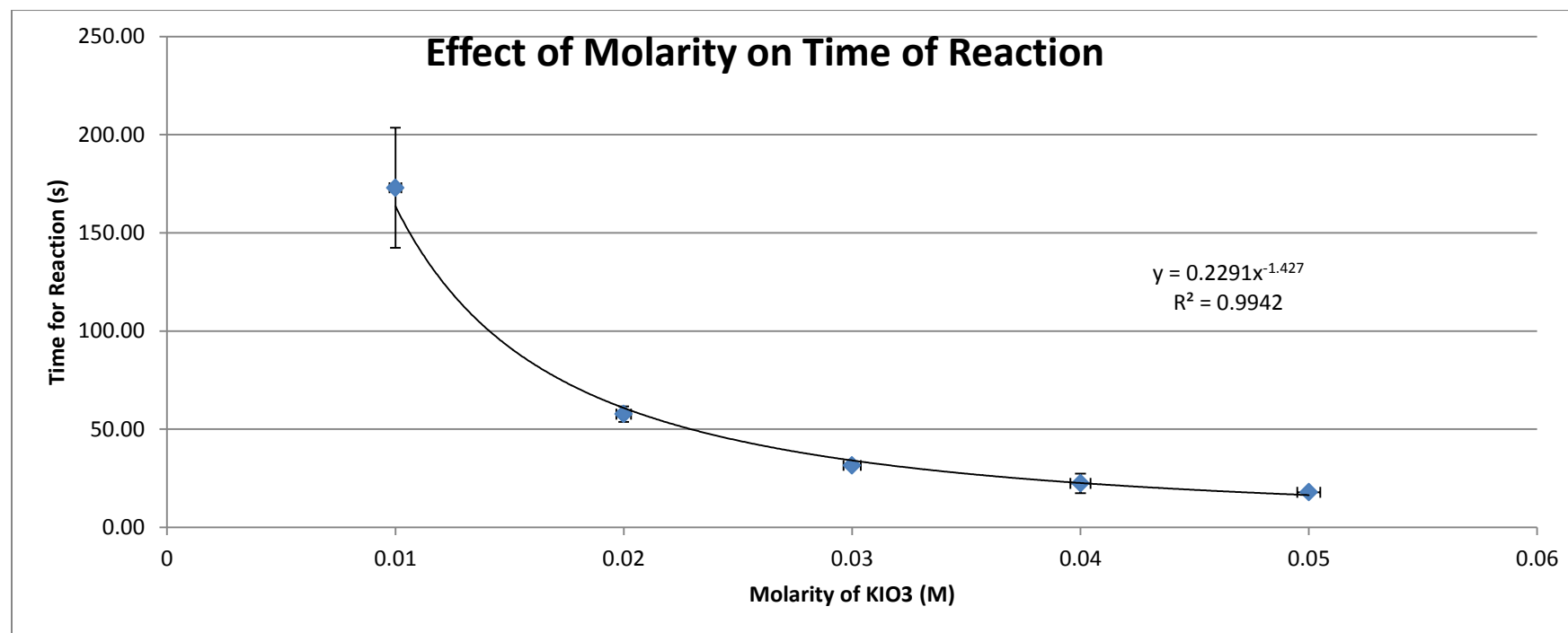
$$= 0.00026$$

### **Calculations**

In order to arrive at the average, we naturally added up each trial and divided by five, as seen in:  $Avg = \frac{T_1 + T_2 + T_3 + T_4 + T_5}{5}$ , and to get the uncertainty associated with this average, the most deviant number was found and the absolute difference between the most deviating and the average was found, as shown in:  $Uncertainty = |T_{deviating} - Avg|$ .

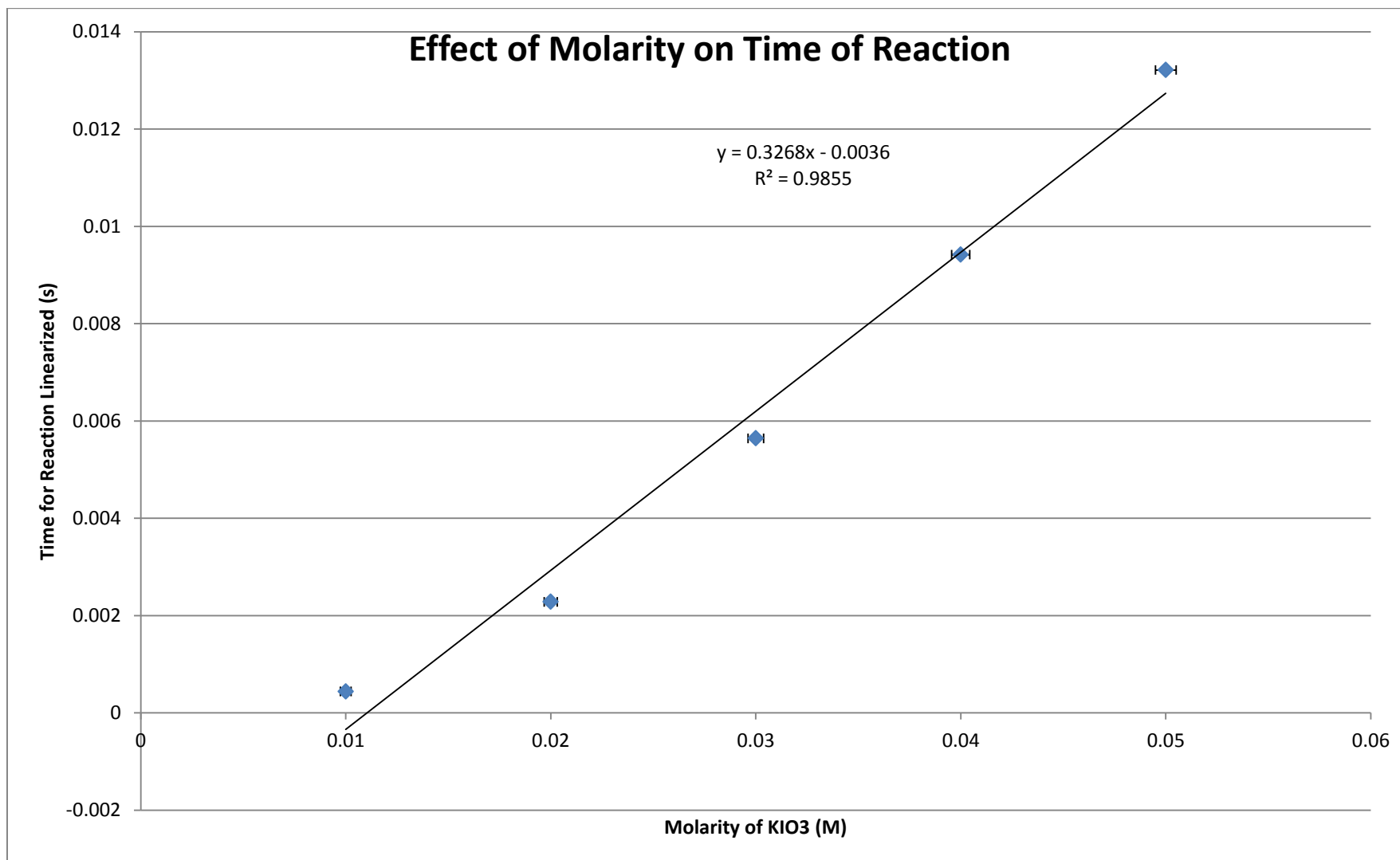
**Processed Data:**

Molarity of KIO <sub>3</sub> (M)	Trials (Time of rxn, seconds)					Uncertainty in Time (s)	Average Time (s)	Uncertainty in Molarity	Linearized moles (x <sup>-1.5</sup> )
	1	2	3	4	5				
0.01	142.33	181.42	179.59	182.89	178.60	30.64	172.97	0.00026	0.000439601
0.02	56.68	55.03	58.91	56.29	61.62	3.91	57.71	0.00032	0.002281227
0.03	31.38	31.86	32.29	30.91	31.33	0.74	31.55	0.00038001	0.005641809
0.04	22.76	19.58	20.85	21.56	27.37	4.95	22.42	0.00044	0.009417386
0.05	18.09	18.25	16.70	18.29	18.13	1.19	17.89	0.0005	0.013213311

**Graphs**

**Note:** Error bars do exist in both the x and the y axis, but due to their remarkably small size, they are not very visible on the graph.

Linearized Graph:





Note: Uncertainty in the y direction is so small ( $<1E-6$ ) that it was not calculated nor included.

These graphs provide quite a vexing result, as it implies that the speed of the reaction is inversely related to the cube of the square of the molarity of  $KIO_3$ . This result was quite unexpected, but one thing that is clear is that there is no clear “optimal” point. Instead, from the data that we collected, it seems as if the speed of the reaction will always increase with higher concentrations of  $KIO_3$ , but will increase slower and slower as time goes on. The y-intercept in the linearized graph likely points to a point of molarity where no reaction would be able to seen, for some reason.

### **Conclusion**

Through performing this experiment, the data has shown empirical proof that there exists an inverse relationship between the molarity of  $KIO_3$  and the time for the reaction to finish. Therefore, our original hypothesis is unsupported. There is no “optimal” point, but instead, a continuous decrease in reaction time. This result is quite odd and does not exactly fit in with the chemistry that has been taught up to this point in time, but the possibility that this reaction to produce iodine may in fact be multiple conflicting oscillating reactions, suggest that there may be more to learn about the kinetics of the reaction.

Through this experiment, most of our data was quite carefully controlled, with a few exceptions. For one, as noted above, was the breaking of one of the plastic cups halfway through the experiment. Replacing it with a different sized beaker may have possible slightly altered the results, but as seen in our uncertainty for the time, this is still relatively small. Another error occurs in the first trial of our 0.01M condition. This was one of the earlier trials that we did, and the experimenters were not entirely sure of what to expect. Therefore, the timing of that trial may

have resulted in a clear outlier. Finally, the time at which the stopwatch was halted may have been slightly ambiguous between trials. As we did not stop at the first sign of color change, due to the presence of some slight Iodine-starch complexes on the bottom of the cup, and waited for some time before the clock was stopped, there may have been discrepancies as to when we stopped. However, these are quite minor and are favorably reflected in our minimal uncertainty reports.

The remarkable level of precision and accurate data, along with the high curve fit, implies that the experiment was carried out mostly successfully with very few sources of error. Therefore, there is little that can be fixed, other than a larger attention to detail and noting the time at which the clock is stopped.

This experiment has introduced us to multiple complex reactions, and even though the nature of these reactions is still unclear, it is clear that there is a negative trend in the time taken to finish.