Experiment 4: Kinetics of an Iodine Clock Reaction

I. Introduction

This experiment is designed to study the kinetics of a chemical reaction. The reaction involves the oxidation of iodide ions by bromate ions in the presence of acid. The objective of this lab is to determine the rate law and activation energy for the redox reaction between the iodide ion (I⁻), bromate ion (BrO₃⁻) ion, and hydrogen ion in water, for which the overall reaction is shown below (1).

Overall Reaction: \[ 6\text{I}^-(aq) + \text{BrO}_3^-(aq) + 6\text{H}^+(aq) \rightarrow 3\text{I}_2(aq) + \text{Br}^-(aq) + 3\text{H}_2\text{O}(l) \] (Reaction 1)

The reaction is somewhat slow at room temperature. The reaction rate depends on the concentration of the reactants and on the temperature. The rate law for the reaction is a mathematical expression that relates the reaction rate to the concentrations of reactants. If the rate of reaction is expressed as the rate of decrease in concentration of bromate ion, the general expression for the rate law, given this overall reaction, is:

\[
\text{Rate of reaction} = \frac{-\Delta[\text{BrO}_3^-]}{\Delta t} = k[\text{I}^-]^x[\text{BrO}_3^-]^y[\text{H}^+]^z
\] (Equation 1)

where the square brackets refer to the molar concentration of the indicated species. The rate is equal to the change in concentration of the bromate ion, \(-\Delta[\text{BrO}_3^-]\), divided by the change in time for the reaction to occur, \(\Delta t\). The term “\(k\)” is the rate constant for the equation, which changes as the temperature changes. The exponents \(x\), \(y\), and \(z\) are called the “orders” of the reaction with respect to the indicated substance, and show how the concentration of each substance affects the rate of reaction.

If reaction (1) is an elementary reaction, then the rate law would be:

\[
\text{Rate of reaction} = \frac{-\Delta[\text{BrO}_3^-]}{\Delta t} = k[\text{I}^-]^6[\text{BrO}_3^-]^1[\text{H}^+]^6
\] (Equation 2)

Because this is a thirteenth order rate law and would involve 13 molecules colliding simultaneously, it is unlikely that reaction (1) is an elementary reaction. One of the objectives of this experiment is, therefore, to determine the rate law for reaction (1).

A. The Method of Initial Rates

To find the rate of the reaction a method is needed to measure the rate at which one of the reactants is used up, or the rate at which one of the products is formed. In this experiment, the rate of reaction will be measured based on the rate at which iodine forms. Therefore we can equally say:

\[
\text{Rate of reaction} = \frac{-\Delta[\text{I}_2]}{3\Delta t} = k[\text{I}^-]^x[\text{BrO}_3^-]^y[\text{H}^+]^z
\] (Equation 3)

The way in which the rate law will be determined is using the “Method of Initial Rates”. In this method, a series of experiments is carried out in which the initial concentration of one of the reactants (let’s say \(\text{I}^-\)) is varied systematically while the concentrations of the other reactants (\(\text{H}^+\) and \(\text{BrO}_3^-\)) are left constant. This allows one to determine one of the coefficients (“\(x\)”). Then the process is repeated in a second series of experiments by systematically varying the other reactants’ concentrations.

In the method of initial rates, it is necessary to determine the initial rate of reaction for each experiment. The rate of reaction that is being measured is an average rate of reaction, and it’s a little hard to understand how it works. Let’s have a go:

The reaction will be carried out in the presence of thiosulfate ions, which will react with iodine as it forms:

\[ \text{I}_2(aq) + 2\text{S}_2\text{O}_3^{2-}(aq) \rightarrow 2\text{I}^-(aq) + \text{S}_4\text{O}_6^{2-}(aq) \] (Reaction 2)

There will also be excess iodide ions present. These will combine with iodine to form a triiodide ion (\(\text{I}_3^-\)). The presence of the product triiodide ion (from reaction 1 in the presence of excess iodide) in solution can be identified by an intensely colored starch-triiodide complex that forms when starch is present in the
solution. Because we add a few drops of starch to the reaction solution, if there is also any triiodide ion in solution, the solution will be very dark blue.

Triiodide rapidly reacts with thiosulfate ion (reaction 3, below). The reaction of triiodide and thiosulfate will be used to monitor the overall reaction (reaction 1): it will tell us when all of the thiosulfate has been used up. Because reaction (3) is much faster than reaction (1), there is an insignificant concentration of triiodide ion (and insignificant starch-triiodide complex) in solution until the thiosulfate is completely consumed. When the thiosulfate is completely consumed, any further triiodide that forms reacts with the starch to turn the solution very dark blue.

Monitoring Reaction: \[ \text{Monitoring Reaction: } \quad \text{I}^-_{(aq)} + 2 \text{S}_2\text{O}_3^{2-}_{(aq)} \rightarrow 3 \text{I}^-_{(aq)} + \text{S}_4\text{O}_6^{2-}_{(aq)} \quad \text{(Reaction 3)} \]

Because reaction (3) is much faster than reaction (1) but depends upon the product of reaction (1) as a reactant of reaction (3), reaction (3) can only go as fast as reaction (1) can go. Therefore, the rate of reaction (3) is the same as the rate of reaction (1). (Read that a few times and see if it makes sense).

To summarize: The monitoring reaction uses thiosulfate to react away triiodide as it forms so that there is no color until there is no thiosulfate left in solution. Once the concentration of thiosulfate is zero, the solution turns very dark blue.

In this laboratory procedure, all of the reactions use the same quantity of thiosulfate ion. The blue color appears when all the thiosulfate is consumed. To determine the initial rate of the reaction for any given run, we will know the initial concentration of thiosulfate, \([\text{S}_2\text{O}_3^{2-}]_i\), based on the concentration put into the reaction solution. We will also know that the final concentration of thiosulfate, \([\text{S}_2\text{O}_3^{2-}]_f\), is zero when the solution turns blue. The initial time is zero and the final time is just how long it took the solution to turn blue. We can then calculate the initial rate of reaction (see Sample Calculation A).

Once “x”, “y”, and “z” have been determined by comparing the initial rates at different starting concentrations, the rate constant can be determined by plugging the rate and initial concentrations back into the rate law expression and solving for k.

It is important to note that the concentration of thiosulfate is much less than the concentrations of iodide ion, hydrogen ion, or bromate ion in the reaction mixture. Therefore, very little of reaction (1) has been completed when the change in color occurs. For this reason, we assume that the rate of reaction derived from this experiment is an initial rate of reaction.

B. Determining the Activation Energy

Another goal of this lab is to determine the activation energy for reaction (1). There is an energy barrier that all reactants must surmount for a reaction to take place. This energy can range from almost zero to many hundreds of kJ/mol. At the top of this energetic barrier, there is a transition state complex that forms instantaneously before the products are formed. This energy barrier is called the activation energy, \(E_a\).

Example of an activation energy reaction coordinate diagram.

Reactants need to possess this amount of energy both to overcome the repulsive electron cloud forces between approaching molecules and to break the existing bonds in the reacting molecules. In general, the
higher the activation energy, the slower the reaction. The activation energy is related to the rate constant by the Arrhenius equation:

\[ k = A e^{-\frac{E_a}{RT}} \]  

(Equation 4)

Where \( A \) is the frequency constant (a.k.a. pre-exponential factor); \( R \) is the universal gas constant; and \( T \) is the temperature in K.

The only way that we know of to determine the activation energy is to prepare an Arrhenius plot. An Arrhenius plot has an x-axis of \( \frac{1}{T} \), where T is temperature in Kelvin, and a y-axis of \( \ln k \), where \( k \) is the rate constant. Reaction rates will be determined in three different temperature ranges: room temperature, above room temperature and below room temperature.

The activation energy can be determined by using the Arrhenius Equation with the y intercept of the Arrhenius plot being equal to \( \ln A \) and the slope of the plot being equal to \( -\frac{E_a}{R} \).

\[ \text{Arrhenius Equation: } k = A e^{-\frac{E_a}{RT}} \quad \text{OR} \quad \ln k = \ln A + \left( -\frac{E_a}{R} \right) \left( \frac{1}{T} \right) \]

C. Adding a Catalyst

A third experimental goal of this lab is to see the effect that a catalyst can have on the rate of reaction for reaction (1). A catalyst is a substance that increases the rate of reaction (both in the forward and reverse directions) without being consumed in the reaction. Catalysts often allow the reactants to react using a reaction pathway (or mechanism) that is lower in activation energy than without the presence of the catalyst. The catalyst being used in this experiment is \( \text{Cu}^{2+} \) ion in copper sulfate. With the catalyst, the reaction can be written:

\[ \text{Overall Reaction: } 6\text{I}^-(aq) + 6\text{H}^+(aq) + \text{BrO}_3^-(aq) \rightarrow 3\text{I}_2(aq) + \text{Br}^-(aq) + 3\text{H}_2\text{O}(l) \]  

(Reaction 4)

and the general form of the rate law becomes:

\[ \text{rate of reaction} = k [\text{I}^-]^x [\text{H}^+]^y [\text{BrO}_3^-]^z [\text{Cu}^{2+}]^p \]

Though it is not necessarily true, we will assume that the exponents “x”, “y”, and “z” don’t change with the addition of the catalyst.

Copper ion is not even that good of a catalyst. It lowers the activation energy, but not nearly as much as most biological enzymes lower the activation energy. Because the decrease in activation energy with the catalyst is small, it is hard to measure accurately even though the difference in rate with the catalyst is easy to see.

D. Recording Reaction Times and Accuracy

Most students will record the reaction times using stopwatches that report time to the nearest one hundredth of a second (0.01 sec). Considering that there is at least half a second when the chemicals are being poured and mixing and that there is at least half a second error in the determination of when the solution turns blue, we will report our times to the nearest second. To prevent errors in using rounded values in further calculations, we will record the average time to the nearest tenth of a second.

E. A Friendly Reminder about Dilution

Because all of the experiments in today’s lab involve the mixing of three different beakers of solutions, each chemical will be diluted by the others. It is important to consider this fact in your calculations.

II. Sample Calculations

A. Determination of the Reaction Rate

For the sample calculation, assume that at the start of the reaction, the concentration of \( \text{S}_2\text{O}_3^{2-} \) is \( 1.00 \times 10^{-3} \text{ M} \) (not the correct concentration as your calculations will show). As reaction (1) proceeds, the
I₂ is consumed in reaction (2) until there is no S₂O₃²⁻ remaining. At this point, the reaction mixture becomes blue and the time of reaction is noted. If the time is noted to be 220 seconds, then:

\[
\text{(initial) reaction rate} = -\frac{1}{2} \frac{\Delta [S_2O_3^{2-}]}{\Delta t} = -\frac{1}{2} \frac{[S_2O_3^{2-}]_i - [S_2O_3^{2-}]_f}{t_i - t_f} = -\frac{1}{2} \frac{(-1.00 \times 10^{-3})}{220 \text{ sec}} = 2.27 \times 10^{-6} \text{ M/sec}
\]

It is very important to point out that it is the concentration of S₂O₃²⁻ that is used for this calculation must be the concentration of the ion in the reaction mixture, not the original solution.

B. Determination of the Rate Law

The general form of the rate law for reaction (1) is:

\[
\text{rate} = k[I^-]^x[BrO}_3^\text{2}[^y][H^+]^z
\]

where k is the rate constant and “x”, “y” and “z” are the exponents associated with the reactant concentrations. Assuming that temperature is constant (and therefore k is constant), by holding [I⁻] and [H⁺] constant, the effect of [BrO₃⁻] on reaction rate can be determined from two experiments with different values of [BrO₃⁻]. In the following expression, everything cancels out except the different rates and the different [S₂O₅²⁻].

\[
\frac{\text{rate}_{\text{exp} 1}}{\text{rate}_{\text{exp} 2}} = \frac{k[I^-]^x[BrO}_3^\text{2}[^y][H^+]^z}{k[I^-]^x[BrO}_3^\text{2}[^y][H^+]^z} \rightarrow \text{canceling out } k[I^-]^m \rightarrow \frac{\text{rate}_{\text{exp} 1}}{\text{rate}_{\text{exp} 2}} = \frac{[BrO}_3^\text{2}[^y]}{[BrO}_3^\text{2}[^y]}
\]

\[
\frac{2.28 \times 10^{-6}}{1.38 \times 10^{-7}} = \frac{(8.00 \times 10^{-2})^y}{(4.00 \times 10^{-2})^y}
\]

\[
16.5 = \frac{(8.00 \times 10^{-2})^y}{(4.00 \times 10^{-2})^y} \rightarrow 16.5 = (2)^y \quad \ln(16.5) = n \cdot \ln(2.00) \rightarrow n = 4.04
\]

For the calculation of all exponents in this lab, you must perform this calculation multiple times. Retain the value of the exponent to 3 significant figures until you have averaged all three values and then round to the nearest whole number.

III. Experimental Procedure

A. Equipment needed

**In lab:** 100-1000μL digital pipets, 10mL beakers
Class sets of hot plates and thermometers are stored in the lab. Ask your instructor for the location.

**Chemicals in lab:**
- DI water
- 0.0100 M KI
- 0.00100 M Na₂S₂O₃
- 0.0400 M KBrO₃ or NaBrO₃
- 1% starch solution
- 0.100M CuSO₄

**From stock room:** cotton swabs, bucket of ice, 1.5 dram (~5ml) glass vials with caps.
B. Waste Disposal

All solutions except rinses must be disposed of in the proper waste container in the hood.

C. Before Starting Experimental Work (Before Class)

1. In your notebook, enter the experiment title, date, your name and name of partner.
2. Write the purposes of the lab. (There are several!) Include all of the different kinetic values that will be determined in this lab (for example, “x”, “y”, etc. …)
3. Write out the overall and monitoring reactions.
4. Describe in your own words how the reaction rate is determined using the monitoring reaction
5. Write an executive summary of the procedures described in this lab. Include all major procedural steps that you will need to follow.
6. Record the information from Tables 1, 2 & 3 in your notebook or tape them in your notebook.

D. Procedure 1. Finding of the Reaction Rate Law in the absence of a catalyst. (Exp# A1 to A7)

Table 1: Experimental parameters for determining the reaction rate law.

<table>
<thead>
<tr>
<th>Exp.</th>
<th>Initial Reagents</th>
<th>Final Reagent</th>
<th>Temp.</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>KI 0.0100 M</td>
<td>DI H₂O 0.100 M</td>
<td>Starch 1%</td>
</tr>
<tr>
<td>A1</td>
<td>0.300 mL 0.600 mL</td>
<td>0.300 mL 0.150 mL</td>
<td>0.150 mL</td>
</tr>
<tr>
<td>A2</td>
<td>0.600 mL 0.300 mL</td>
<td>0.300 mL 0.150 mL</td>
<td>0.150 mL</td>
</tr>
<tr>
<td>A3</td>
<td>0.900 mL ---</td>
<td>0.300 mL 0.150 mL</td>
<td>0.150 mL</td>
</tr>
<tr>
<td>A4</td>
<td>0.300 mL 0.300 mL</td>
<td>0.300 mL 0.150 mL</td>
<td>0.150 mL</td>
</tr>
<tr>
<td>A5</td>
<td>0.300 mL ---</td>
<td>0.300 mL 0.150 mL</td>
<td>0.150 mL</td>
</tr>
<tr>
<td>A6</td>
<td>0.300 mL 0.300 mL</td>
<td>0.600 mL 0.150 mL</td>
<td>0.150 mL</td>
</tr>
<tr>
<td>A7</td>
<td>0.300 mL ---</td>
<td>0.900 mL 0.150 mL</td>
<td>0.150 mL</td>
</tr>
</tbody>
</table>

*Note: NaBrO₃ may be used in place of KBrO₃.

1. In 6 separate clean 10mL beakers, obtain enough of the solutions listed in Table 1 to fill each beaker about 2/3 full.
2. You will need 7 pipet tips that fit the adjustable 100-1000μL digital pipets. Set them up as in the picture below so that you can dedicate one tip to each reagent that you will use during the experiment.
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3. Measure the temperature of the starch solution. Assume that all reagents are at the same room temperature and record this temperature in Table A for runs A1 thru A7.

4. Obtain a clean 5mL glass vial and label it A1.

5. Fill the vial with the “Initial Reagents” for Experiment A1 using an adjustable digital pipet.

6. One partner should get ready to start a timer. Another partner should obtain the listed amount of KBrO₃ in a digital pipet.

7. At the same time, one partner should start the timer while the other one adds the KBrO₃ to vial A1. Cap the vial, shake to mix the reagents, and place the vial on a white surface to make it easy to see color changes.

8. When the solution first becomes colored (usually blue, sometimes brown), stop the timer.

9. In Table A, record the volumes used and the time it took (in seconds, to the nearest second) for the solution to turn colored.

10. Label a 250-500 mL beaker “Waste”. Empty the contents of the vial into the waste beaker. Rinse it 3-4 times with DI water. (The rinses can go down the sink.) Drain the vial upside down on a paper towel for 30 seconds and then dry the interior with a cotton swab.

11. Repeat steps 4-10 for vials A2 thru A7.

E. Procedure 2. The Effect of Copper Ion as a Catalyst. (Experiments B1-B-3)

Experiment B1 thru B3: These use the same reagents as experiment A1 with the addition of a Cu²⁺ catalyst and less water.

1. Repeat steps D4-D10 for vials B1 thru B3

Table 2: Experimental Parameters for Evaluation of a Catalyst.

<table>
<thead>
<tr>
<th>Exp.</th>
<th>Initial Reagents</th>
<th>Final Reagent</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>KI 0.0100 M</td>
<td>DI H₂O 0.100 M</td>
</tr>
<tr>
<td>B1</td>
<td>0.300 mL</td>
<td>0.450 mL</td>
</tr>
<tr>
<td>B2</td>
<td>0.300 mL</td>
<td>0.300 mL</td>
</tr>
<tr>
<td>B3</td>
<td>0.300 mL</td>
<td>0.150 mL</td>
</tr>
</tbody>
</table>

F. Procedure 3. Determination of the Activation Energy via the Effect of Temperature on Reaction Rate. (Experiments C1 thru C2)

Table 3: Experimental Parameters for Determining the Activation Energy.

<table>
<thead>
<tr>
<th>Exp.</th>
<th>Initial Reagents</th>
<th>Final Reagent</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>KI 0.0100 M</td>
<td>DI H₂O 0.100 M</td>
</tr>
<tr>
<td>C1</td>
<td>0.300 mL</td>
<td>0.600 mL</td>
</tr>
<tr>
<td>C2</td>
<td>0.300 mL</td>
<td>0.600 mL</td>
</tr>
</tbody>
</table>

*Note: NaBrO₃ may be used in place of KBrO₃.

Experiment C1: Use the same reagents as experiment A1, but at a lower temperature, between 0–2°C.

1. Obtain a clean 5mL glass vial and label it C1.
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2. Fill the vial with the “Initial Reagents” for Experiment C1 using an adjustable digital pipet.

3. Place Vial C1 and the small beaker of 0.0400M KBrO₃ in an ice-water bath in a medium beaker. Keep the level low to prevent the beaker from tipping over. Allow them to come to temperature for at least 5 minutes.

4. One partner should get ready to start a timer. The second partner should obtain the listed amount of KBrO₃ in a digital pipet from the now cold reagent.

5. At the same time, one partner should start the timer while the other one adds the KBrO₃ to vial C1. Cap the vial, shake to mix the reagents, and place the vial back in the ice bath near the edge where the color can be easily observed.

6. When the solution becomes colored (blue to brown), stop the timer.

7. In Table A, record the volumes used and the time it took for the solution to turn dark.

8. Measure the temperature of the solution and record it in Table A.

9. Empty the contents of the vial into the waste container. Rinse it 3-4 times with DI water. (The rinses can go down the sink.) Drain the vial upside down on a paper towel for 30 seconds and then dry the interior with a cotton swab.

**Experiment C2:** Use the same reagents as experiment A1, but at a higher temperature, between 39-41°C.

   1. Obtain a clean 5mL glass vial and label it C2.

   2. Fill the vial with the “Initial Reagents” for Experiment C2 using an adjustable digital pipet.

   3. Place Vial C2 and the small beaker of 0.0400M KBrO₃ in warm water bath in a medium beaker on a hot plate. (Start with hot water out of the sink tap. It will usually already be close to the correct temperature.) Keep the level low to prevent the beaker from tipping over. Allow them to come to temperature for at least 5 minutes.

   4. One partner should get ready to start a timer. The second partner should obtain the listed amount of KBrO₃ in a digital pipet from the now warm reagent.

   5. At the same time, one partner should start the timer while the other one adds the KBrO₃ to vial C2. Cap the vial, shake to mix the reagents, and place the vial on a white surface to make it easy to see color changes.

   6. When the solution becomes colored (blue to brown), stop the timer.

   7. In Table A, record the volumes used and the time it took for the solution to turn dark.

   8. Measure the temperature of the solution and record it in Table A.

   9. Empty the contents of the vial into the waste container. Rinse it 3-4 times with DI water. (The rinses can go down the sink.) Drain the vial upside down on a paper towel for 30 seconds and then dry the interior with a cotton swab.

**G. Cleanup**

   1. Rinse all pipet tips with water and share with your partner to use during the rest of the semester. Do not throw them away.

   2. Rinse vials and caps with DI water and return to equipment area.

   3. Transfer the contents of your waste beaker to the central waste container.
4. Clean all remaining items used and return them to your equipment locker.

5. Wipe down your work area with a damp paper towel.

IV. Further Instructions

A. Required Calculations and Graphs

1. **Data Sharing:** Obtain data for Experiments A1 thru C2 from 5 other student groups.

2. **Determination of average reaction times:** If the times recorded for a given experiment vary significantly between the groups, perform a Q-test on any suspect values. **Average** any values retained by the Q-test. **Record** values in Table 1 to one decimal place for use in further calculations in order to prevent rounding errors. **Show** a sample calculation for Q and the average for the first experiment that required a Q-test.

3. **Determination of initial rates of reactions:** Calculate the initial rate of reaction for each of the experiments that was performed. This will require that you first calculate the concentration of each reactant in the reaction mixture. Remember that all reagents dilute each other when you mix them. Create Table B in your notebook. **Record** your calculated values in Table B. **Show** a sample calculation for experiment A1.

4. **Determination of rate law exponents:** Calculate the value for each exponent in the rate law (x, y, z, and p for the catalyst) using the method of initial rates. For each determination of exponents x, y, & z, there are three possible combinations of experiments that can be used to estimate the exponent. For the exponent of I (“x”), use experiments A1/A2, A1/A3, and A2/A3 to estimate the exponent three times and average the values you find. Then, choose the closest whole number. **Show the full calculation for all of the values of “x” to 3 sig figs as shown on page 4, average the three values of the exponent and report the average rounded to the nearest whole number.**

   Also, determine three estimates of the exponent of BrO$_3^-$ (“y”) and H$^+$, (“z”) to 3 sig figs and one value for the copper catalyst exponent, (“p”) to 3 sig figs. You do not have to show all explicit calculations, but include the ratio of the rates (rate2/rate1), the ratio of the concentrations (conc.2/conc.1), the natural logs of these two values, and the value of the exponent that you get by dividing the two natural log values. All values should be summarized in Table C. Determine the average value for x, y, z and p and then round the final values to the nearest whole number. (Note: -0.5 to +0.5 rounds to 0; 0.5 to 1.5 rounds to 1; 1.5 to 2.5 rounds to 2.)

5. **Determination of the rate constants:**

   Once the rate law is determined, any experiment can be used to determine the rate constant k by simply plugging in the appropriate values for an experiment.

   a) **Calculate** the value of the rate constant **without** a catalyst for the first 7 experiments (A1-A7). **Show** the first calculation. **Calculate and report** the average of the calculated values for the rate constant k at room temperature.

   b) **Calculate** the value of the rate constant without catalyst at low temperature. (Exp C1)

   c) **Calculate** the value of the rate constant without catalyst at elevated temperature. (Exp C2)

   d) **Calculate** the value of the rate constant **with** a catalyst for the 3 experiments that used the catalyst (B1 – B3). **Show** all calculations. **Calculate and report** the average of the calculated values for the rate constant k.

6. **Determination of the Rate Laws:**

   a) **Write the three rate laws** for the overall reaction without the catalyst, using the values of x, y, z, and k determined above. (i.e., at low temp., room temp., and high temp.)

   b) **Write the rate law** for the overall reaction with the catalyst at room temp, using the values of x, y, z, p, and k determined above.
7. **Determination of the Activation Energy**

Using the three k values obtained in calculations 5a (average), 5b and 5c, **Fill in Table D.**

Graph: **Prepare** an Arrhenius plot of ln k versus 1/T. This should lead to a linear plot, with the slope being $-E_a/R$. You must prepare your plot in such a way (using Excel) that the computer fits a best-fit line through the data and reports the slope for you to determine $E_a$. This can be done by selecting Chart/Add Trendline from the menu. On the “Type” tab of the dialog box, select “linear”. On the “options” tab, be sure to check the box to include the equation of the line on the graph. Do NOT set the intercept to zero. **Determine** the activation energy (in kJ/mol) from the linear least squares fit slope of the graph. **Show** your calculation. **Do not use the two-point slope formula!**

**B. In-Class Work**

1. Complete Table A (Rows A1-A3) showing all calculations for row A1.
2. Complete the first row in Table B (A1), showing all calculations.
3. Complete the 2nd & 3rd row in Table B (A2 & A3). (You do not need to show the calculations for these rows or any of the other rows.)
4. Calculate “x” (The exponent on $I$) using experiments A2 & A1 to 3 sig figs (with all work shown and all intermediate values shown in the first row of table C). Have your instructor verify your work.
5. Calculate “x” twice more to 3 sig figs using the pairs of experiments shown in the second and third rows of Table C. Only the intermediate values required by the table need to be shown, not the entire calculation.
6. Average the three values of “x”. Record these in the last column of table C and then round to the nearest integer.

**C. At-Home Work**

1. Complete all tables and calculations not completed during class. The Since all groups were required to perform the experiments within a 2°C range, use the temperatures from the class average posted on the board to do your calculation. If you do not have this data, use your group’s temperatures to perform all calculations.
2. Prepare the Excel Graph and do the calculations associated with the determination of the activation energy.

**D. Experimental Discussion and Summary:**

After all calculations are complete:

1. Explain why the reaction rate changes as concentration changes.
2. Explain why the reaction rate changes as temperature changes.
3. Explain why the reaction rate changes when a catalyst is added.
4. In the method of initial rates that was used in this experiment, why do we pick the “initial” portion of the reaction to study?
5. Write an executive summary that addresses the purpose of this lab, describing what was done in general terms and what was learned in this experiment, including any issues that were encountered and dealt with.
6. Present a table that summarizes what was determined from this experiment. (e.g., rate laws, rate constants & activation energy under varying conditions.)
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**TableA: Data Collection for Iodine Clock Reaction Kinetics Experiments**

<table>
<thead>
<tr>
<th>Chemical Formula</th>
<th>KI</th>
<th>DI H₂O</th>
<th>HCl</th>
<th>Starch</th>
<th>Na₂S₂O₃</th>
<th>CuSO₄</th>
<th>KBrO₃</th>
<th>Total Volume</th>
<th>Temp.</th>
<th>Reaction Times</th>
<th>Avg. Time*</th>
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</thead>
<tbody>
<tr>
<td>Conc. M</td>
<td>N/A</td>
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<td>total (mL)</td>
<td>(ºC)</td>
<td>Group 1 (yours)</td>
<td>Group 2</td>
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<td>Exp A1</td>
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*Note: Only include values accepted by the Q-test in your reported average and record one extra (non-significant) digit in your average (to avoid rounding errors when you use these averages in future calculations).
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### Table B: Data table for determining the rate law.

<table>
<thead>
<tr>
<th>Expt.</th>
<th>([S_2O_3^{2-}], M)</th>
<th>([I^-], M)</th>
<th>([BrO_3^-], M)</th>
<th>([H^+], M)</th>
<th>([Cu^{2+}], M)</th>
<th>Avg Rxn Time from Table A (s)</th>
<th>Temp. °C</th>
<th>Initial Rate of reaction, M/s</th>
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</thead>
<tbody>
<tr>
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</table>

These are concentrations in the mixture.

### Table C: Rate Law Calculations Summary

<table>
<thead>
<tr>
<th>Exp Pair</th>
<th>Rate Ratio (value of rate2/rate1)</th>
<th>Conc. Ratio (value of conc.2/conc.1)</th>
<th>Natural log of (Rate Ratio)</th>
<th>Natural log of (Conc. Ratio)</th>
<th>Value of x, y, z, p</th>
<th>Avg of x, y, z, p</th>
<th>Exponent on I⁻</th>
<th>Exponent on BrO₃⁻</th>
<th>Exponent on H⁺</th>
<th>Exponent on Cu²⁺</th>
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</thead>
<tbody>
<tr>
<td>A2/A1</td>
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<td>x ⇒___</td>
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<td>y = unrounded</td>
<td>y ⇒___</td>
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<td>p = unrounded</td>
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<td>B3/B1</td>
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<td>B3/B2</td>
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</table>

### Table D: Data table for determining the activation energy.

<table>
<thead>
<tr>
<th>Experiment(s)</th>
<th>Temp (°C)</th>
<th>Temp (K)</th>
<th>1/T (K⁻¹)</th>
<th>Rate const. k</th>
<th>ln k</th>
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</thead>
<tbody>
<tr>
<td>A1-A7 avg (Room temp.)</td>
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<tr>
<td>C1 (low temp.)</td>
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<td>C2 (high temp.)</td>
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</table>