

Kinetics of a Reaction

Introduction

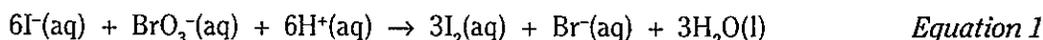
How fast will a chemical reaction occur? If a reaction is too slow, it may not be practical. If the reaction is too fast, it may explode. Measuring and controlling reaction rates makes it possible for chemists and engineers to make a variety of products, everything from antibiotics to fertilizers, in a safe and economical manner. The study of the speed of reactions and the processes which occur as the reaction takes place is known as kinetics. The purpose of this experiment is to investigate how the rate of a reaction can be measured and how reaction conditions affect reaction rates.

Concepts

- Rate law
- Activation energy
- Reaction rates
- Rate constant
- Catalyst

Background

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:



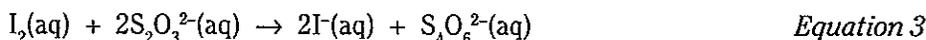
The reaction is somewhat slow at room temperature. The reaction rate depends on the concentration of the reactants and on the temperature. The rate law for the reaction is a mathematical expression that relates the reaction rate to the concentrations of reactants. If the rate of reaction is expressed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

$$\text{Rate} = \frac{-\Delta[\text{BrO}_3^-]}{\Delta t} = k[\text{I}^-]^x[\text{BrO}_3^-]^y[\text{H}^+]^z \quad \text{Equation 2}$$

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, Δt . 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.

The total rate law for the process is determined by measuring the rate, evaluating the rate constant, k , and determining the order of the reaction for each reactant (the values of x , y , and z).

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



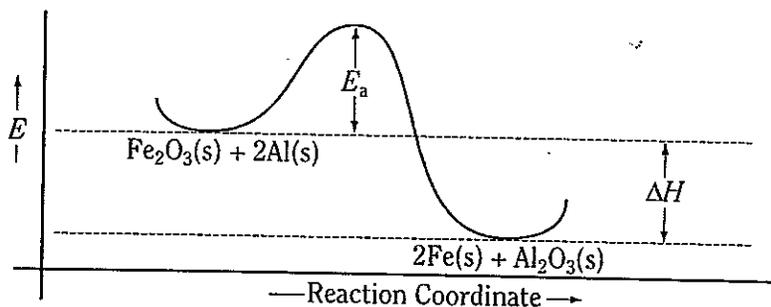
Reaction 1 is somewhat slow. Reaction 2 is extremely rapid, so that as quickly as iodine is produced in reaction 1, it is consumed in reaction 2. Reaction 2 continues until all of the added thiosulfate has been used up. After that, iodine begins to increase in concentration in solution. If some starch is present, iodine reacts with the starch to form a deep blue-colored complex that is readily apparent.

Carrying out reaction 1 in the presence of thiosulfate ion and starch produces a chemical “clock.” When the thiosulfate is consumed, the solution turns blue almost instantly.

Experiment 12

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. An examination of reactions 1 and 2 shows that six moles of $\text{S}_2\text{O}_3^{2-}$ are needed to react with the three moles of I_2 formed from the reaction of one mole of BrO_3^- . Knowing the amount of thiosulfate used, it is possible to calculate both the amount of I_2 that is formed and the amount of BrO_3^- that has reacted at the time of the color change. The reaction rate is expressed as the decrease in concentration of BrO_3^- ion divided by the time it takes for the blue color to appear.

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. This energy barrier is called the *activation energy*, E_a .



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:

$$k = Ae^{-E_a/RT}$$

Where A is the frequency constant and is related to the frequency of collisions; R is the universal gas constant; and T is the temperature in K.

Catalysts are substances that speed up a reaction, but are not consumed in the reaction. Catalysts work by lowering the overall activation energy of the reaction, thus increasing the rate of the reaction.

The experiment is designed so that the amounts of the reactants that are consumed are small in comparison with the total quantities present. This means that the concentration of reactants is almost unchanged during the reaction, and therefore the reaction rate is almost a constant during this time.

The experiment utilizes a microscale procedure. Only 12 drops of reactants delivered from capillary droppers are used for each measurement. The steps involved are as follows:

- Part 1. Measure the volume of a drop of solution.** This must be done to determine the number of moles of thiosulfate ion in one drop. This will allow the moles of bromate ions that react to be calculated.
- Part 2. Determine the reaction rate and calculate the rate law.** This is done by carrying out an experiment at specific concentrations of each of the reactants and measuring the reaction rate. The concentration of one reactant is then changed and the reaction rate change is observed. This is repeated for each reactant. This data allows the calculation of the order of each reactant. Once the orders are known, the value of the rate constant can be determined.
- Part 3. Determine the activation energy.** Reaction rates generally increase as the temperature goes up. By measuring how the rate changes as the temperature is varied, the activation energy, E_a , for the reaction can be calculated. The natural log of the Arrhenius equation is:

$$\ln k = \frac{-E_a}{RT} + \ln A$$

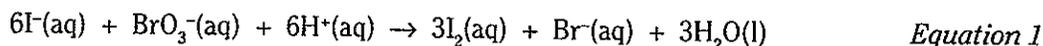
where $\ln k$ is the natural logarithm of the rate constant, E_a is the activation energy, R is the gas constant, $8.314 \text{ J/mol}\cdot\text{K}$, and T is the temperature on the kelvin scale. A is the frequency factor.

This equation follows the straight line relationship: $y = mx + b$. A plot of the natural logarithm of k versus $1/T$ will give a straight line graph. The slope of the graph is $-E_a/R$. By determining the slope, the activation energy can be calculated.

- Part 4. Observe the effect of a catalyst on the rate of the reaction.** The catalyst used is copper(II) nitrate solution.

Experiment Overview

The purpose of this experiment is to utilize a microscale technique to determine the total rate law for the oxidation of iodide ions by bromate ions in the presence of acid:

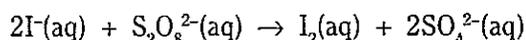


There are several steps in the experiment. First, the order for each of the reactants is found by varying the concentration of each reactant individually. Once the orders are known, the rate constant is calculated. Second, the activation energy is found by repeating the experiment at several different temperatures, measuring the rate, and calculating the rate constants at the different temperatures. A graph of the reciprocal of absolute temperature versus the natural logarithm of the rate constant allows the calculation of the activation energy. Last, a catalyst is added and the change in reaction rate is observed.

Experiment Number	KI, 0.010 M	Distilled H ₂ O	HCl, 0.10 M	Starch, 2%	Na ₂ S ₂ O ₃ , 0.0010 M	KBrO ₃ , 0.040 M
1	2 drops	4 drops	2 drops	1 drop	1 drop	2 drops
2	4 drops	2 drops	2 drops	1 drop	1 drop	2 drops
3	6 drops	0 drops	2 drops	1 drop	1 drop	2 drops
4	2 drops	2 drops	2 drops	1 drop	1 drop	4 drops
5	2 drops	0 drops	2 drops	1 drop	1 drop	6 drops
6	2 drops	2 drops	4 drops	1 drop	1 drop	2 drops
7	2 drops	0 drops	6 drops	1 drop	1 drop	2 drops

Pre-Lab Questions (Use a separate sheet of paper to answer the following questions.)

1. Another version of the iodine clock reaction involves reaction of iodide ions with persulfate ions.



The following rate data was collected by measuring the time required for the appearance of the blue color due to the iodine–starch complex.

Trial	[I ⁻]	[S ₂ O ₈ ²⁻]	Reaction Time
1	0.040 M	0.040 M	270 sec
2	0.080 M	0.040 M	138 sec
3	0.040 M	0.080 M	142 sec

- In each trial, the blue color appeared after 0.0020 M iodine (I₂) had been produced. Calculate the reaction rate for each trial by dividing the concentration of iodine formed by the reaction time.
 - Compare trials 1 and 2 to determine the order of reaction with respect to iodide ions. How did the concentration of iodide ions change in these two trials, and how did the rate change accordingly? What is the reaction order for iodide?
 - Which two trials should be compared to determine the order of reaction with respect to persulfate ions? What is the reaction order for persulfate?
 - Write the rate law for this version of the iodine clock reaction. Could the rate law have been predicted using the coefficients in the balanced chemical equation? Explain.
2. An experiment is designed to study the rate of the reaction: $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$

Exp. No.	[NO] _{initial}	[O ₂] _{initial}	-Δ[NO]/Δt _{initial}
1	0.0125 M	0.0250 M	0.0282 M/s
2	0.0250 M	0.0250 M	0.112 M/s
3	0.0125 M	0.0500 M	0.0560 M/s

Write the rate law including the value for the specific rate constant, *k*.

3. The following data are for the decomposition of N₂O₅(g). Graph the data as described in the experimental directions. Determine the slope. Calculate the energy of activation.

Temperature	Specific rate constant	Temperature (K)	1/T (K ⁻¹)	ln <i>k</i>
25 °C	3.7 × 10 ⁻⁵ s ⁻¹	_____	_____	_____
35 °C	1.3 × 10 ⁻⁴ s ⁻¹	_____	_____	_____
45 °C	5.0 × 10 ⁻⁴ s ⁻¹	_____	_____	_____
55 °C	1.4 × 10 ⁻³ s ⁻¹	_____	_____	_____

Show calculations for:

Slope _____

Activation energy, *E*_a _____

Materials

Chemicals

Copper(II) nitrate solution, $\text{Cu}(\text{NO}_3)_2$, 0.1 M, 5 mL

Hydrochloric acid solution, HCl , 0.10 M, 5 mL

Potassium iodide solution, KI , 0.010 M, 5 mL

Potassium bromate solution, KBrO_3 , 0.040 M, 5 mL

Sodium thiosulfate solution, $\text{Na}_2\text{S}_2\text{O}_3$, 0.0010 M, 5 mL

Starch solution, 2%, 5 mL

Water, distilled or deionized

Equipment

Balance, 0.001-g or 0.0001-g precision

Beaker, 10-mL or 50-mL

Beral-type pipet with microtip, 7

Cassette tape case

Cotton swabs for cleaning reaction strip wells

Label tape, for pipets

Marking pen

Reaction strip, 12-well

Thermometer, 0 °C–100 °C

Timer, seconds

Toothpicks for stirring

Trough for hot and cold water baths, shared

Safety Precautions

Dilute hydrochloric acid solution is severely irritating to skin and eyes and is slightly toxic by ingestion and inhalation. Dilute copper(II) nitrate solution is irritating to skin, eyes, and mucous membranes and slightly toxic by ingestion. Dilute potassium bromate solution is irritating to body tissue and slightly toxic by ingestion. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory.

Procedure

Part 1. Find the Volume of One Drop of Solution

1. Obtain a microtip Beral-type pipet. Fill the pipet with approximately 3 mL of distilled or deionized water.
2. Mass a small beaker using an analytical balance. Record the mass in Data Table 1.
3. Holding the pipet vertically, deliver five drops of water into the beaker, and find the total mass. Record your data in Data Table 1.
4. Add an additional five drops of water into the beaker, and again determine the mass. Record this value in Data Table 1.
5. Deliver five more drops and again find the mass. Record the data in Data Table 1.
6. Repeat steps 1 to 5 for two more determinations.

Part 2. Determine the Reaction Rate and Calculate the Rate Law

It is necessary to use consistently good technique to obtain reproducible data. Hold pipets vertically and be sure no air bubbles are introduced. Since such small quantities of reagents are used, it is very easy to repeat measurements. Calculation of the orders of reactants are all based on the values obtained for the first experiment, so be sure to get reproducible data from the beginning. All other experiments should be carried out at least twice.

The table that follows shows the reagent quantities to be used in carrying out the reactions needed. It is important to use care in measuring out the solutions. Since the total solution volume is quite small, even one extra drop can cause a substantial change in concentrations.

Experiment 12

Table 1. Reagent Quantities for Experiments

Experiment Number	KI, 0.010 M	Distilled H ₂ O	HCl, 0.10 M	Starch, 2%	Na ₂ S ₂ O ₃ , 0.0010 M	KBrO ₃ , 0.040 M
1	2 drops	4 drops	2 drops	1 drop	1 drop	2 drops
2	4 drops	2 drops	2 drops	1 drop	1 drop	2 drops
3	6 drops	0 drops	2 drops	1 drop	1 drop	2 drops
4	2 drops	2 drops	2 drops	1 drop	1 drop	4 drops
5	2 drops	0 drops	2 drops	1 drop	1 drop	6 drops
6	2 drops	2 drops	4 drops	1 drop	1 drop	2 drops
7	2 drops	0 drops	6 drops	1 drop	1 drop	2 drops

A study of Table 1 shows that all experiments contain the same total number of drops of solution. Only one drop of sodium thiosulfate, Na₂S₂O₃, and one drop of starch solution are added to each well. In Experiments 1, 2, and 3, the concentration of potassium iodide, KI, is gradually increased while all other solutions' volumes remain constant. Experiments 1, 4, and 5 have an increasing concentration of potassium bromate, KBrO₃. Experiments 1, 6, and 7 show an increase in the concentration of hydrochloric acid, HCl.

Read the entire procedure before beginning the experiment.

Rate Order of KI

- Obtain six microtip pipets and fold an adhesive label around the stem of each pipet (see Figure 1). Label the pipets KI, H₂O, HCl, Starch, Na₂S₂O₃, and KBrO₃.
- Fill each pipet with about 2 mL of the appropriate liquid.
- Place the pipets in an opened cassette case for storage (Figure 1).
- Obtain a clean, 12-well reaction strip and arrange it so that the numbers can be read from left to right.
- The first determination will vary the concentration of only the KI. Using Table 1 as a guide, fill number one well in the first reaction strip with the appropriate number of drops of the first five reagents listed in Experiment #1. Mix the solution in well with a new toothpick. Since each experiment in Table 1 will be run in triplicate, repeat this step for the next two wells in order.
- In the next three wells of the reaction strip, repeat step 5 using the first five reagent amounts listed for Experiment #2.
- Fill wells 7, 8, and 9 with the first five reagent amounts listed in Experiment #3.
- Fill wells 10, 11, and 12 with the first five reagents amounts listed in experiment #4.
- To well 1, begin the timer as the two drops of KBrO₃ are added. Stir the solution with a toothpick.
- Record in Data Table 2 the time required for the first tint of blue color to appear.
- Repeat steps 8 and 9 for well 2, the well 3.

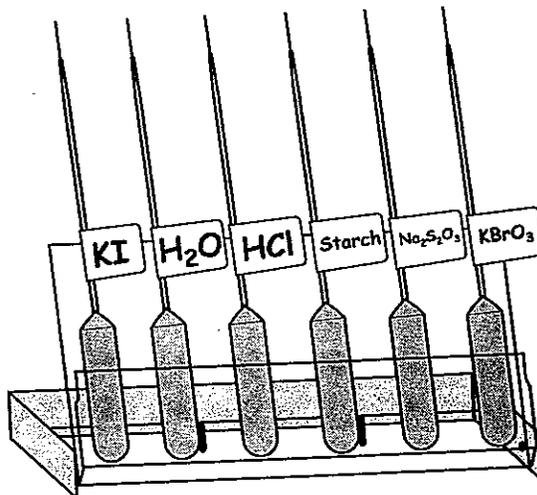


Figure 1.

12. Take the temperature of one of the reaction solutions. Record this temperature for all the reactions in Data Table 2.
13. Repeat steps 8 and 9 for wells 4, 5, and then 6. Record these times in Data Table 2 for experiment #2.
14. Repeat steps 8 and 9 for wells 7, 8, and then 9. Record these times in Data Table 2 for experiment #3.
15. Repeat steps 8 and 9 for wells 10, 11, and then 12. Record these times in Data Table 2 for experiment #4.
Note: Be sure to use four drops of KBrO_3 solution.
16. Rinse the contents of the well strips with warm water. Use a cotton swab to dry the inside of each well.
17. Repeat the entire process (steps 4–13) for the reagent combinations listed in Table 1 that cover experiments #5, #6 and #7, respectively.

Part 3. Determine the Activation Energy

In this part of the experiment, the reaction will be carried out at several different temperatures using the concentrations given in Part 2 for Experiment 1. The temperatures will be about $40\text{ }^\circ\text{C}$, $20\text{ }^\circ\text{C}$, and $0\text{ }^\circ\text{C}$. Use data from Experiment 1 at room temperature for the second measurement.

1. Prepare a shallow warm water bath of about $40\text{ }^\circ\text{C}$.
2. Using the amounts listed for Experiment #1 in Table 1, fill each of the first six wells in the reaction strip with the appropriate number of drops of the first five reagents listed. Mix the solutions well with a new toothpick.
3. Place the reaction strip in the warm temperature bath.
4. Fill the Beral pipet labeled KBrO_3 half-full with 0.040 M KBrO_3 solution.
5. Place this pipet in the warm temperature water bath for at least five minutes.
6. Measure the temperature of the water bath with a thermometer and record the value in Data Table 3.
7. Take the pipet out of the water bath and dry the outside of the pipet.
8. With the reaction strip still in the warm temperature bath, add two drops of KBrO_3 solution to the first well, stir, and immediately start the timer. Place the pipet back in the warm temperature bath.
9. Record the time, in seconds, when the first blue color appears.
10. Repeat steps 7–9 for the reaction solutions in wells 2 and 3.
11. Remove the reaction strip and the pipet from the warm temperature bath.
12. Add ice cubes and water to create a cold temperature water bath.
13. Place both the reaction strip and Beral-type pipet containing the KBrO_3 solution into the cold temperature water bath.
14. Measure the temperature of the water bath with a thermometer and record the value in Data Table 3.
15. Repeat steps 7–9 for wells 4, 5, and 6. Record the time, in seconds, for each reaction in Data Table 3.

Part 4. Observe the Effect of a Catalyst on the Rate

Repeat the procedure given in Part 2 for Experiment 1 only, but this time add 1 drop of 0.1 M copper(II) nitrate solution, $\text{Cu}(\text{NO}_3)_2$, and only 3 drops of distilled or deionized water to the mixture. Fill only the first reaction wells. The total volume will still be 12 drops. Record the reaction times in Data Table 4.

Disposal

Your teacher will provide disposal and cleanup instructions.

Experiment 12

Data Tables

○ Data Table 1. Find the Volume of One Drop of Solution.

	Mass of empty beaker (a)		g
Trial 1	Mass of beaker plus 5 drops of water (b)		g
	Mass of first 5 drops of water (b) – (a)		g
	Average mass of 1 drop of water		g
	Mass of beaker plus 10 drops of water (c)		g
Trial 2	Mass of second 5 drops of water (c) – (b)		g
	Average mass of 1 drop of water		g
	Mass of beaker plus 15 drops of water (d)		g
Trial 3	Mass of third 5 drops of water (d) – (c)		g
	Average mass of 1 drop of water		g
	Average mass of 1 drop of water (Trials 1–3)		g

Data Table 2. Determine the Reaction Rate and Calculate the Rate Law.

Experiment No.	Time, seconds				Temp. °C
	Trial 1	Trial 2	Trial 3	Average	
1					
2					
3					
4					
5					
6					
7					

Data Table 3. Determine the Activation Energy.

Approximate Temperature, °C	Measured Temperature, °C	Measured Temperature, K	Measured Temperature ⁻¹ , K ⁻¹	Time of Reaction, seconds		
				Trial 1	Trial 2	Average Time
0						
20						
40						

Data Table 4. Observe the Effect of a Catalyst on the Rate.

	Reaction Time, seconds
Uncatalyzed Reaction	
Catalyzed Reaction	

Post-Lab Calculations and Analysis

(Use a separate sheet of paper to answer the following questions.)

Part 1. Calculate the Volume of One Drop of Solution

Assume the density of water to be 1.00 g/mL.

$$\text{Volume of one drop} = \frac{\text{mass 1 drop (g)}}{1.00 \text{ g/mL}} \times \frac{1 \text{ L}}{1000 \text{ mL}}$$

Part 2A. Calculate the Rate

The rate will be expressed as $-\Delta[\text{BrO}_3^-]/\Delta t$. In each reaction there is one drop of 0.0010 M $\text{Na}_2\text{S}_2\text{O}_3$ solution. Calculate the number of moles of $\text{S}_2\text{O}_3^{2-}$ present in one drop:

$$\text{moles } \text{S}_2\text{O}_3^{2-} \text{ ions} =$$

The blue color begins to appear when all the thiosulfate ion is consumed. Examination of reactions 1 and 2 allows us to calculate the moles of BrO_3^- which react as all of the $\text{S}_2\text{O}_3^{2-}$ ion is used up:

$$\text{mol } \text{BrO}_3^- \text{ reacted} =$$

The value of $-\Delta[\text{BrO}_3^-]$ in all reactions, since all experiments have a total volume of 12 drops, is:

$$-\Delta[\text{BrO}_3^-] = \frac{\text{mol } \text{BrO}_3^- \text{ reacted}}{\text{volume of 12 drops}}$$

The rate of each reaction can be found by dividing $-\Delta[\text{BrO}_3^-]$ by the number of seconds for the reaction to take place.

$$\text{Rate} = \frac{-\Delta[\text{BrO}_3^-]}{\Delta \text{ time}}$$

Calculate the rate of reaction in each experiment and enter the results into the following table. Use the average time for each experiment.

	Reaction Rate, M/s
Exp. 1	
Exp. 2	
Exp. 3	
Exp. 4	
Exp. 5	
Exp. 6	
Exp. 7	

Experiment 12

Part 2B. Calculate Initial Concentrations

Calculate the initial concentration of each reactant for each experiment. These are the concentrations of each reactant after all the reactants have been mixed, but before any reaction has taken place. This will not be the same as the concentration of the starting solution because combining the reactants dilutes all of the solutions. On dilution, the number of moles of reactant stays the same. Therefore:

$$\text{no. moles} = V_{\text{concentrated}} \times M_{\text{concentrated}} = V_{\text{dilute}} \times M_{\text{dilute}}$$

where $V_{\text{concentrated}}$ and $M_{\text{concentrated}}$ are the volume and molarity of the starting, concentrated solutions, and V_{dilute} and M_{dilute} are the volume and molarity of the diluted reaction mixtures. Since volumes will be proportional to the number of drops of solution used, the number of drops substitute for volumes.

Find the initial concentration of each reactant. Show the calculations.

Part 2C. Calculate the Order of Each Reactant

Next, the values for the exponents x , y , and z need to be determined. The experiment is designed so that the concentration of one ion changes while the others remain constant. Comparing values in Experiments 1, 2, and 3, we see that Experiment 2 has double the I^- concentration as Experiment 1, and Experiment 3 has triple the I^- concentration as Experiment 1.

Substitute the concentration values for Experiments 1 and 2 into the equation:

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

$$\text{Exp. 1: Rate}_1 \text{ ______} = k[\text{ ______}]^x[\text{ ______}]^y[\text{ ______}]^z$$

$$\text{Exp. 2: Rate}_2 \text{ ______} = k[\text{ ______}]^x[\text{ ______}]^y[\text{ ______}]^z$$

Divide the first equation by the second. Notice that most of the terms will cancel out and the ratio reduces to:

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \frac{[\text{ ______}]^x}{[\text{ ______}]^x}$$

Solve for x . Report the value of x to the nearest integer. Repeat the calculations using Experiments 1 and 3 to confirm the value for x . *Note:* To solve for an exponential value, take the logarithm of both sides of the equation.

$$\text{For example: } 8 = 2^n \quad \log 8 = n \log 2 \quad n = \frac{\log 8}{\log 2} = 3$$

Next use the same procedure with Experiments 1, 4, and 5 to find the value of y . Lastly, use Experiments 1, 6, and 7 to find the value of z . Show how the calculations are carried out.

Write the experimentally determined rate law.

Part 2D. Find the Rate Constant

Substitute data from each experiment into the rate law equation to find the value of k . Report the average value of k . Don't forget to include proper units for k .

Experiment	1	2	3	4	5	6	7
Value of k							

Average value of k =

Part 3. Calculate the Activation Energy, E_a

Using the data from Part 3, calculate the values listed in the table below for each measured temperature.

Measured Temperature, K	Measured Temperature ⁻¹ , K ⁻¹	Average Time, s	Rate of Reaction, M/s	Rate Constant, k, (with units)	ln k

Graph the data with natural logarithm of the rate constant, $\ln k$, on the vertical axis versus $1/T$ (temperature in the Kelvin scale) on the horizontal axis. Draw the straight line that is closest to the most points, and determine the slope of the line. The slope = $-E_a/R$, where E_a is the activation energy and $R = 8.314 \text{ J/mol}\cdot\text{K}$. Calculate the activation energy for the reaction. Give a title to the graph, and label the axes appropriately. Indicate the points used to determine the slope of the line.

Post-Lab Questions (Use a separate sheet of paper to answer the following questions.)

1. Why does the reaction rate change as concentrations of the reactants change?
2. Explain the general procedure used to find the rate law.
3. Why does reaction rate change as temperature changes?
4. Explain the general procedure used to determine the activation energy.
5. Differentiate between reaction rate and specific rate constant.
6. Comment on the effect of the catalyst. Predict how the activation energy changes when a catalyst is added to the reaction.
7. Make a general statement about the consistency of the data as shown by calculating the orders of reactants, and by the graphical analysis which leads to activation energy. Were the calculated orders close to integers? Did the check of the order give the same value for the order? Were the points on the graph close to a straight line?
8. Write out the "two-point" form of the Arrhenius equation which relates rate constants, temperatures and activation energy.
9. How could you improve the data?