

## Experiment 5

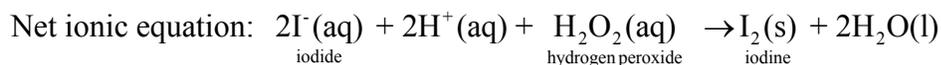
### Kinetics: The Oxidation of Iodide by Hydrogen Peroxide

#### Goals

To determine the differential rate law for the reaction between iodide and hydrogen peroxide in an acidic environment. To determine the activation energy and pre-exponential factor for the reaction. To determine the effect of a catalyst on the rate of reaction.

#### Discussion

When hydrogen peroxide is added to a solution of potassium iodide, the iodide ions are slowly oxidized according to the equation:



The rate law for this reaction should include the concentrations of iodide, hydrogen ion, and hydrogen peroxide. However, if the concentration of  $\text{H}^+$  is held constant throughout the experiment then its effect will not appear in the rate law. This results in a relatively simple rate law:

rate =  $k[\text{I}^-]^n[\text{H}_2\text{O}_2]^m[\text{H}^+]^p$  simplifies to:

rate =  $k'[\text{I}^-]^n[\text{H}_2\text{O}_2]^m$  where  $k' = k[\text{H}^+]^p$

#### Part 1: Determination of the Rate Law

The rate law for the reaction between iodide ions and hydrogen peroxide can be determined by carrying out experiments in which the concentrations of iodide and peroxide are varied. For example, consider the following reaction and data:



Experiment	[A] <sub>0</sub>	[B] <sub>0</sub>	Initial rate $\left(\frac{\text{M}}{\text{s}}\right)$
1	0.10	0.10	0.45
2	0.20	0.10	0.90
3	0.10	0.20	1.80

For this reaction, doubling the initial concentration of reactant A doubles the initial rate of reaction (experiments 1 and 2). This relationship corresponds to  $n=1$ . Likewise, doubling the initial concentration of B has the effect of quadrupling the initial rate (experiments 2 and 3). This corresponds to  $m=2$ .

We can now write the differential rate law for this reaction as: rate =  $k[\text{A}][\text{B}]^2$

The reaction is said to be first order in A and second order in B. Overall, the order of the reaction is 3.

The value of the rate constant  $k$  can be determined by using the known values of  $n$  and  $m$ :

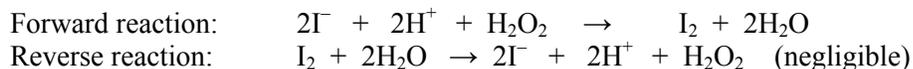
$$k = \frac{\text{Rate}}{[\text{A}][\text{B}]^2}$$

We can use the given initial concentrations and initial rate for each experiment and determine the value of  $k$  for each experiment. For the study we will perform, the reaction between iodide and hydrogen peroxide, there are several experimental requirements:

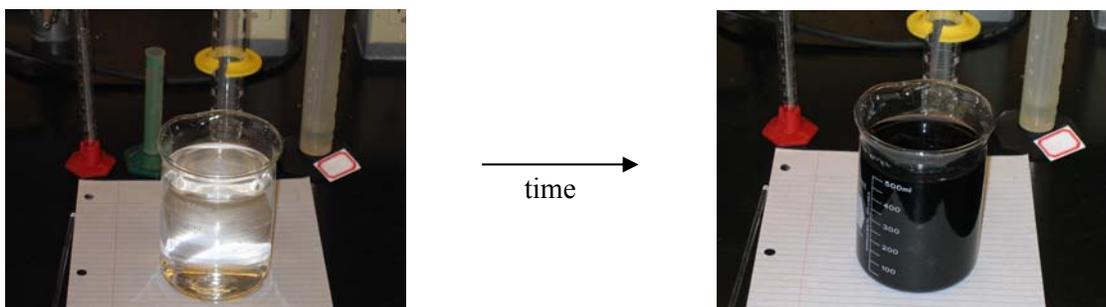
1. The acidity of the solution must be maintained at a constant level so that the concentration of  $\text{H}^+$  is constant.

2. The temperature of the reactants must be the same for all runs. The rate constant is dependent on the temperature of the solution. Since the heat of reaction is relatively small for this reaction the temperature should remain relatively constant throughout the process. If this were not the case then we would need to place the reaction in a constant temperature bath.

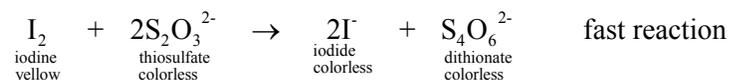
3. The reverse reaction must be suppressed.



If the concentration of iodine ( $\text{I}_2$ ) is allowed to build up as the reaction proceeds then the rate of the reverse reaction will increase and an equilibrium mixture will result. Any rate measured for the reaction under such conditions will not represent the rate of the forward reaction but the net difference between the forward and reverse rates. In this reaction, however, we can ignore the reverse reaction because the equilibrium position is so far on the  $\text{I}_2$  side that the reaction above does not take place to any appreciable extent.



4. We need a method to accurately measure the rate of reaction. The addition of thiosulfate ions ( $\text{S}_2\text{O}_3^{2-}$ ) allows an accurate measurement of the rate at which the peroxide-iodide reaction is taking place. Suppose that you add a small and known amount of thiosulfate ion to the original mixture of peroxide and iodide. Iodine is produced slowly by the reaction between peroxide and iodide ions and the thiosulfate ions immediately react with the iodine as it is produced:



As long as excess thiosulfate ions are present in the solution, no free iodine can accumulate because it is immediately turned into iodide ions which are colorless.

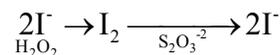
The thiosulfate ions are the limiting reagent. So once all the thiosulfate ions are consumed, iodine starts to form in the solution. Iodine is a pale yellow. If starch is added to the solution then a more dramatic blue solution is formed by the complex of starch-iodine. The color change is sharp, and the time elapsed to this point is determined simply by use of a timer. The time from the addition of the peroxide solution to the appearance of the blue color is  $\Delta t$  for the reaction. Since the stoichiometry of the thiosulfate-iodine and the peroxide-iodide reactions is known, it is not difficult to calculate how many moles of peroxide were reduced in the known interval of time. Consequently, the average rate (moles of hydrogen peroxide consumed per liter per second) of the reaction during this period can be calculated.

$$\text{Rate} = -\frac{\Delta[\text{H}_2\text{O}_2]}{\Delta t}$$

← change in peroxide concentration, calculated from known amounts of reactants  
← change in time, measured with a stopwatch

For the peroxide-iodide reaction the average rate of the reaction over the period taken for the measurement is a good approximation of the initial reaction rate.

An additional and helpful consequence of the iodine–thiosulfate reaction is that the concentration of iodide ion in the solution remains constant at its original value. Even though the peroxide–iodide reaction tends to consume iodide ions, the thiosulfate ion immediately returns the iodine to the iodide ion form:



As long as there are excess thiosulfate ions present, there is no change in the molarity of the iodide ion. This means that the rate of the reaction will change only because the concentration of the  $\text{H}_2\text{O}_2$  decreases.

If the amount of  $\text{H}_2\text{O}_2$  which reacts during the period taken for the measurement is sufficiently small, no significant change in concentration of the  $\text{H}_2\text{O}_2$  will occur. Under these conditions, the average rate of the reaction determined for this period will be a close approximation of the reaction rate corresponding to the initial concentration of reactants which is the initial rate of the reaction:

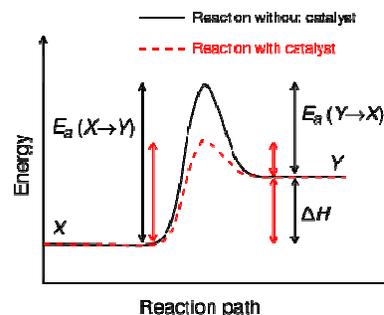
$$\text{Initial rate} = k[\text{I}^-]_0^n[\text{H}_2\text{O}_2]_0^m \quad \text{where } [\text{I}^-]_0 = \text{initial concentration of iodide} \\ [\text{H}_2\text{O}_2]_0 = \text{initial concentration of hydrogen peroxide}$$

### The Temperature-Dependence of the Rate Constant

In general, the rate of a reaction increases with increasing temperature. This is consistent with our everyday experiences that cooking food at higher temperatures requires less time than at lower temperatures or that garbage decays faster in the summer than it does in the winter. These observations indicate that the rate constant ( $k$ ) itself depends on temperature. The relationship between the rate constant and temperature is reflected in the Arrhenius equation:

$$k = A e^{-\frac{E_a}{RT}} \quad \text{where } k \text{ is the rate constant, } A \text{ is called the Arrhenius}$$

pre-exponential constant,  $E_a$  is the activation energy,  $R$  is the gas constant, and  $T$  is the temperature (in units of kelvin). An important point regarding the Arrhenius equation is that the rate constant ( $k$ ) depends on temperature and on the value of the activation energy ( $E_a$ ). The activation energy can be thought of as a barrier to the formation of products and its value is useful in understanding the energetics of the reaction. Experimentally, we can determine the value of the activation energy by measuring the value of the rate constant at several measured temperatures. A discussion of this strategy is given in the procedure and calculations sections.



### The Effect of Catalysts

Aside from increasing the concentrations of reactants or raising the temperature of a reaction mixture, there is another way to increase the rate of a reaction. Some chemical species can increase the rate of a reaction. Some chemical species can speed-up a reaction without being consumed themselves. We call these species catalysts. Examples of catalysts include enzymes in biology and catalytic converters in automobiles. Catalysts operate by decreasing the value of the activation energy for the reaction. The decrease in activation energy resulting from the addition of a catalyst is shown in the figure above. An effective catalyst exists for the reaction of iodide with hydrogen peroxide. The catalyst is iron (II) ions dissolved in the reaction mixture. By adding these ions to the reaction mixture we can increase the rate of the reaction.

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### Procedure

Eye protection must be worn throughout this experiment.

**Equipment and Materials.** Beakers–3000 mL and/or 4000 mL, 600 mL, 250 mL, 100 mL. Graduated cylinders–100 mL, 10 mL. Electric timer, hot plate, 1 M HCl, 1 M KI, 0.04 M  $\text{Na}_2\text{S}_2\text{O}_3$ , 0.2 M  $\text{H}_2\text{O}_2$ , starch solution, ice, and deionized water. The precise molarity of each solution will appear on reagent bottles. Work in groups of 2-3.

### Part 1: Determining the Rate Law for the Reaction of Hydrogen Peroxide with Iodide

1. Rinse glassware with tap water and then three times with a small quantity of distilled or deionized water.
2. A plastic carboy contains deionized water equilibrated to room temperature that should be used in the reaction mixture. Don't use this water for rinsing since its supply is limited.
3. There are five mixtures (experiments) that we will study in this part. Charts 1 and 2 show the volumes of each solution that should be added to the reaction mixture. There are two options (Chart 1 or Chart 2) for volumes of solutions that we'll use. The instructor will suggest which set of mixtures we'll prepare and use. These mixtures possess varying amounts of hydrogen peroxide and potassium iodide. Notice that the amounts of hydrochloric acid and sodium thiosulfate are constant.
4. Perform experiment #1 first. Then do experiment #2, then experiment #3, then experiment #4, and finally, experiment #5.
5. Measure each solution (deionized water, HCl, KI, starch,  $\text{Na}_2\text{S}_2\text{O}_3$ ,  $\text{H}_2\text{O}_2$ ) using graduated cylinders, not beakers.

**Chart 1: Volumes of solutions for each experiment (Total volume = 500 mL). Use a 600 mL beaker for the reaction.**

Expt #	Deionized Water (mL)	HCl (mL)	KI (mL)	Starch (mL)	$\text{Na}_2\text{S}_2\text{O}_3$ (mL)	$\text{H}_2\text{O}_2$ (mL)
1	425	10	10.0	5	25.0	25.0
2	415	10	20.0	5	25.0	25.0
3	405	10	30.0	5	25.0	25.0
4	380	10	30.0	5	25.0	50.0
5	330	10	30.0	5	25.0	100.0

**Chart 2: Volumes of solutions for each experiment (Total volume = 200 mL). Use a 250 mL beaker for the reaction.**

Expt #	Deionized Water (mL)	HCl (mL)	KI (mL)	Starch (mL)	$\text{Na}_2\text{S}_2\text{O}_3$ (mL)	$\text{H}_2\text{O}_2$ (mL)
1	170	4	4.0	2	10.0	10.0
2	166	4	8.0	2	10.0	10.0
3	162	4	12.0	2	10.0	10.0
4	152	4	12.0	2	10.0	20.0
5	132	4	12.0	2	10.0	40.0

6. Add the solutions, except for the peroxide, in the order indicated in the chart, to the beaker (250 mL or 600 mL). Measure and record the temperature of mixture 1. Place the beaker on a piece of white paper so that the color change may be observed as easily as possible.
7. One student operates the electric timer so that timing can begin at the exact moment when the other student begins to add the peroxide. A good technique is to begin stirring before the addition of the peroxide, rapidly pour in the peroxide while stirring, and continue stirring for about 10 seconds. Once the solution is well-mixed, further stirring is unnecessary.
8. Watch the mixture attentively for the first appearance of a blue coloration and stop the timer. This should happen rather suddenly, and the time required should be measured to the nearest second. Times to color change vary depending on the solution—the range can be from as short as 30 s to as long as 600 s. Record the result immediately in the Data Table. Pour the solution in the beaker into the sink.

9. Rinse the beaker with tap water and deionized water. Prepare and measure the rate of reaction for the next mixture in a similar manner. Since you will be using the same graduated cylinders for the same reagents each time, there is no reason to rinse them between experiments.

### Part 2: The Effect of Temperature on the Reaction Rate

Data will be shared between groups for this part of the experiment. The mixture described in this section should be prepared and placed into the appropriate bath at the beginning of the laboratory.

To study the effect of temperature on reaction rate we need to perform the experiment at temperatures above room temperature and at temperatures below room temperature. Since we are measuring the effect of temperature and not concentration we need to keep the concentrations of all reactants constant for this part. To do this we will choose the concentrations used in experiment #3 from part 1.

Expt #	Deionized Water (mL)	HCl (mL)	KI (mL)	Starch (mL)	Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> (mL)	H <sub>2</sub> O <sub>2</sub> (mL)
3	405	10	30.0	5	25.0	25.0

or

Expt #	Deionized Water (mL)	HCl (mL)	KI (mL)	Starch (mL)	Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> (mL)	H <sub>2</sub> O <sub>2</sub> (mL)
3	162	4	12.0	2	10.0	10.0

1. Each group should prepare one #3 solution, without peroxide, in a beaker.
2. Place 25.0 mL (or 10.0 mL) of peroxide should be in a 100 mL graduated cylinder.
3. Place the mixture into the cold bath or hot bath, as directed by your instructor.
4. Once the temperature of the mixture has equilibrated with the bath, measure the rate of reaction as you did in part 1. The temperature and time data should be recorded in the data sheet.

### Part 3: The Effect of a Catalyst on the Reaction Rate

1. Prepare a #3 solution, without peroxide, in a 600 mL beaker.
2. Place 25.0 mL of peroxide should be in a 100 mL graduated cylinder.
3. Add the peroxide to the mixture at the same time as you add 2 mL of iron (II) ions to the mixture. Measure and record the time of reaction.

**Experiment 5. Kinetics: The Oxidation of Iodide by Hydrogen Peroxide**

Name \_\_\_\_\_

Name(s) of Partner(s) \_\_\_\_\_

\_\_\_\_\_

**Stock solution Concentrations\***

concentration of KI \_\_\_\_\_

concentration of  $\text{Na}_2\text{S}_2\text{O}_3$  \_\_\_\_\_concentration of  $\text{H}_2\text{O}_2$  \_\_\_\_\_

\*From the reagent bottles.

**Part 1: Effect of Concentration**

Expt #	Time (s)
1	
2	
3	
4	
5	

**Part 2: Effect of Temperature\***

	Temperature (°C)	Time (s)
cold bath		
cold bath		
hot bath		
hot bath		

\*concentrations are the same as experiment 3 in Part 1.

**Part 3: Effect of a Catalyst**

Time (s)\* \_\_\_\_\_

\* concentrations are the same as experiment 3 in Part 1.

**Experiment 5. Kinetics: The Oxidation of Iodide by Hydrogen Peroxide**

Name \_\_\_\_\_

**Part 1**

Complete the calculations in Tables 1–3.

**Table 1—Initial Moles\***

The number of moles of each reactant is found by multiplying the volume of reactant solution added by the molarity of the reactant. For example:

$$(0.01\text{L KI})\left(1.0\frac{\text{mol}}{\text{L}}\text{KI}\right)=0.01\text{mol KI}$$

Expt #	[KI] mol/L	volume KI (L)	moles KI	[Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> ] mol/L	volume Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> (L)	moles Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub>	[H <sub>2</sub> O <sub>2</sub> ] mol/L	volume H <sub>2</sub> O <sub>2</sub> (L)	moles H <sub>2</sub> O <sub>2</sub>
1	1.0	0.01	0.01						
2									
3									
4									
5									

\*The molarities used in this table are those of the stock solutions.

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**Experiment 5. Kinetics: The Oxidation of Iodide by Hydrogen Peroxide**

Name \_\_\_\_\_

**Table 2—Initial Molarity\***

To determine the rate law we need to know the initial concentrations of the reactants. This is found by dividing the initial number of moles for each reactant by the total volume of the solution. The initial number of moles was calculated in Table 1 for each reactant.

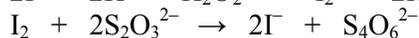
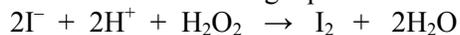
Total volume of solution (mL) \_\_\_\_\_ Total volume of solution (L) \_\_\_\_\_

Expt #	initial moles KI	[KI] <sub>0</sub> mol/L	initial moles Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub>	[Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> ] <sub>0</sub> mol/L	initial moles H <sub>2</sub> O <sub>2</sub>	[H <sub>2</sub> O <sub>2</sub> ] <sub>0</sub> mol/L
1						
2						
3						
4						
5						

**Table 3—Initial Rates**

To calculate the rate of reaction we need to determine what the change in concentration of one of the reactants or products was in each experiment. We'll choose hydrogen peroxide. To determine the change in concentration we need to know how many moles were consumed (because peroxide is a reactant) and how long it took those moles of peroxide to be consumed). Only some of the hydrogen peroxide was consumed, so we can't directly determine how many moles were consumed. But we do know that all of the thiosulfate was consumed. Since all of the thiosulfate was consumed we just need to calculate how many moles of thiosulfate we initially put into the solution. The number of moles of peroxide consumed is related to this number of moles of thiosulfate by the molar ratios for the equations in the reactions:

Consider the following equations:



we see that the molar ratios are:

$$\frac{1 \text{ mol H}_2\text{O}_2}{1 \text{ mol I}_2} \quad \text{and} \quad \frac{2 \text{ mol S}_2\text{O}_3^{2-}}{1 \text{ mol I}_2}$$

$$\text{so} \quad \frac{1 \text{ mol H}_2\text{O}_2}{2 \text{ mol S}_2\text{O}_3^{2-}}$$

## Experiment 5. Kinetics: The Oxidation of Iodide by Hydrogen Peroxide

Name \_\_\_\_\_

If we calculate the moles of  $S_2O_3^{2-}$  that were in the mixture and multiply this number by 1/2, we then have the moles of peroxide consumed. Remember that when all of the  $S_2O_3^{2-}$  is used the color changes to blue, the timer is stopped, and the reading on the timer is  $\Delta t$ . The calculation is:

$$\text{initial moles } S_2O_3^{2-} \left( \frac{1 \text{ mol } H_2O_2}{2 \text{ mol } S_2O_3^{2-}} \right) = \text{moles } H_2O_2 \text{ used}$$

Expt #	moles $H_2O_2$ used	volume (L)*	$[H_2O_2]$ used mol/L	time (s)	rate (M/s)**
1					
2					
3					
4					
5					

\*This is the total volume of the mixture.

\*\*The rate is  $\text{Rate} = -\frac{\Delta[H_2O_2]}{\Delta t}$  where the change in the concentration is  $[H_2O_2]$  used and  $\Delta t$  is the time that it took for the solution to change color.

### Table 4– Logarithms

Directions

Complete Table 4. From the calculations in the table generate two graphs. The first graph is  $\log(\text{Rate})$  versus  $\log[H_2O_2]$ . The second graph is  $\log(\text{Rate})$  versus  $\log[I^-]$ . The purpose behind these graphs is given below.

Expt #	$\log(\text{rate})$	$[H_2O_2]_0^*$	$\log [H_2O_2]_0$	$[I^-]_0$	$\log [I^-]_0^*$
1					
2					
3					
4					
5					

\*From Table 2

We wish to find the orders of the reactants, m and n:

$$\text{Initial rate} = k'[\text{I}^-]_0^n [\text{H}_2\text{O}_2]_0^{nm}$$

By taking the log of both sides of the rate law we can do this.

$$\log(\text{Initial rate}) = \log(k') + m\log[\text{H}_2\text{O}_2]_0 + n\log[\text{I}^-]_0$$

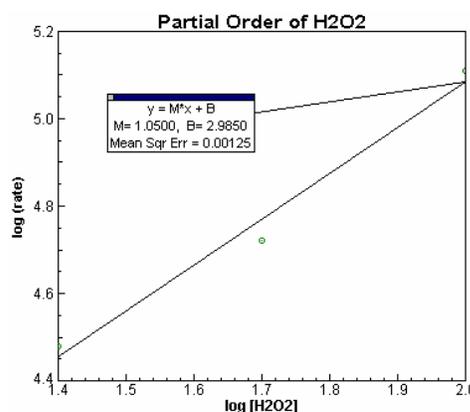
Now, suppose that we make multiple measurements of the rate in which we vary the initial concentration of  $\text{H}_2\text{O}_2$  but keep the initial concentration of  $\text{I}^-$  constant. The rate will change because  $[\text{H}_2\text{O}_2]_0$  changes. But any change in the rate depends only on the change in initial  $\text{H}_2\text{O}_2$  concentration. The rate depends on  $\text{H}_2\text{O}_2$  only. We can generate a graph that shows this to be the case. If we make a graph of  $\log(\text{Rate})$  versus  $\log[\text{H}_2\text{O}_2]_0$  then we get the following result:

$$\log(\text{Initial rate}) = \log(k') + m\log[\text{H}_2\text{O}_2]_0 + n\log[\text{I}^-] \quad \text{where the constant is equal to } \log(k') + n\log[\text{I}^-]$$

$$\log(\text{Initial rate}) = m\log[\text{H}_2\text{O}_2]_0 + \text{Constant}$$



This result is the familiar linear relationship. In this case, the order of peroxide (m) is the slope of the graph. The y-intercept is unimportant to us since we merely want the order n. Note that this graph only works for the case where the concentration of iodide is constant. If the iodide and peroxide concentrations change simultaneously then we can't use this analysis.



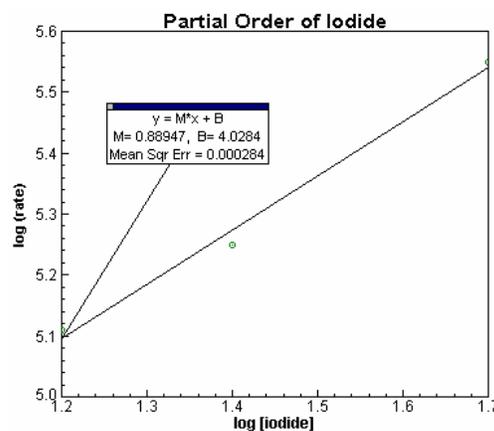
is

If we want to find the order of iodide (n) we can do the same analysis. But we need to pick data in which the initial iodide concentration changes and the initial peroxide concentration remains constant. If we review the initial concentration data from Table 2 we can find data in which this is the case. We can then generate a second graph

$$\log(\text{Initial rate}) = \log(k') + m\log[\text{H}_2\text{O}_2]_0 + n\log[\text{I}^-]$$

$$\log(\text{Initial rate}) = n\log[\text{I}^-] + \text{constant} \quad \text{where the constant is } \log(k') + m\log[\text{H}_2\text{O}_2]_0$$

Again, the slope of the graph gives us the order of the reactant, in this case, iodide.



### Table 5– Rate Constants

Using the values of m and n to 2 significant figures (from the graphs), the initial rate, and the initial concentrations of the peroxide and iodide, calculate the rate constant k for each of the five runs. The values should be the same because k is only a function of temperature for a given reaction. However, due to experimental error the five values of k that you calculate will probably be somewhat different. The equation to calculate k is:

$$k = \frac{\text{initial rate}}{[\text{H}_2\text{O}_2]^m [\text{I}^-]^n}$$

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Name \_\_\_\_\_

The values of m and n should be entered to two significant figures. The average value of k is found by adding the five values and dividing by five.

Expt	[H <sub>2</sub> O <sub>2</sub> ] <sub>0</sub> <sup>*</sup>	[H <sub>2</sub> O <sub>2</sub> ] <sub>0</sub> <sup>m</sup>	[I <sup>-</sup> ] <sub>0</sub> <sup>*</sup>	[I <sup>-</sup> ] <sub>0</sub> <sup>n</sup>	initial rate (M/s)	k
1						
2						
3						
4						
5						
Value of m (2 sig fig) =			Value of n (2 sig fig) =			
average value of k =						
standard deviation of k =						

\*From Table 2.

Calculate the standard deviation of your values of k. The standard deviation s, is a measure of how close the measured values are to one another and is defined as:

$$s = \sqrt{\frac{(k_1 - k)^2 + (k_2 - k)^2 + \dots + (k_n - k)^2}{n-1}}$$

where k<sub>1</sub>, k<sub>2</sub>, k<sub>3</sub>, k<sub>4</sub>, k<sub>5</sub> are the five rate constants and k is the average value of the rate constant. n = the number of measurements, in this case n = 5. The standard deviation is a measure of the variability of your measurements. If the only errors made in the experiment are random errors, then the smaller the standard deviation, the less the random error in your measurements.

**Part II: The Temperature Dependence of the Rate Constant and Determination of the Activation Energy and Pre-exponential Factor**

In the first part of this experiment, we took data which allows us to calculate k, the rate constant for the reaction, and m and n, the partial orders of the reactants. But k, while not sensitive to concentration, is temperature dependant, i.e., it will change if you change the temperature of the reaction. We can take advantage of this temperature dependence of k. Consider the Arrhenius equation:

$$k = Ae^{\frac{-E_a}{RT}}$$

where A is known as the frequency factor or pre-exponential factor, e is the base of the natural logarithm system, E<sub>a</sub> is the activation energy in joules/mole, R = 8.31  $\frac{\text{J}}{\text{Kmol}}$  is the ideal gas law constant, and T is the temperature in kelvin.

Exponential equations like the Arrhenius equation produce curved lines when plotted. It is desirable to convert the equation into that of a straight line, because then the activation energy and the pre-exponential factor can be found. The way to convert an exponential equation into a straight line equation is to take the natural logarithm (ln) of both sides of the equation:

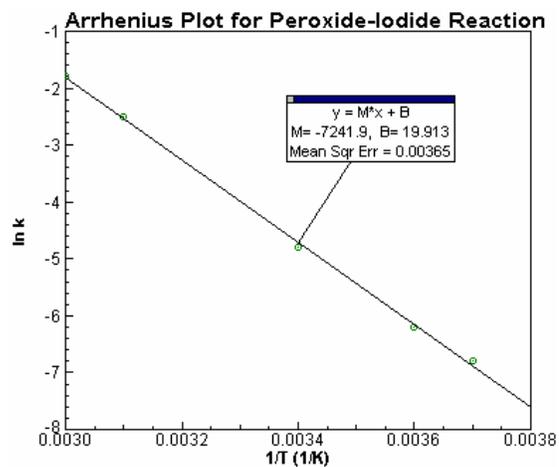
$$\ln(k) = -\frac{E_a}{R} \left( \frac{1}{T} \right) + \ln(A)$$

$\uparrow$                      $\uparrow$     $\uparrow$     $\uparrow$   
 y                    =   m   x   +   b

Notice that taking the natural log of both sides results in a linear equation (just as in Part 1). If the natural logarithm of k (ln(k)) is plotted versus the reciprocal of the temperature (1/T) a straight line is produced. The slope, m of the line, is then the negative of the activation energy divided by the ideal gas constant, and the y-intercept, b, is the natural logarithm of the pre-exponential factor:

$$m = -E_a/R \qquad b = \ln(A)$$

$$E_a = -mR \qquad A = e^b$$



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**Experiment 5. Kinetics: The Oxidation of Iodide by Hydrogen Peroxide**

Name \_\_\_\_\_

**Table 6– k at Different Temperatures**

Reaction conditions	T (°C)*	T (K)**	time (s)*	rate (M/s)	k	ln k	1/T (K <sup>-1</sup> )
hot bath							
hot bath							
room temp							
cold bath							
cold bath							
E <sub>a</sub> =				A =			

\*From the Data Table.

\*\*T<sub>K</sub> = T<sub>°C</sub> + 273Determine the activation energy E<sub>a</sub> and the frequency factor A.**Part 3: The Effect of a Catalyst on the Rate of Reaction**

Using the remaining room temperature solution #3 that was prepared at the beginning of the lab, your instructor will perform an experiment at room temperature after one milliliter of a catalyst, a solution of Fe(III) ions, has been added. Record your results in the table below.

**Table 7–The Effect of A Catalyst**

Reaction of Expt #3 solution	Solution temperature (°C)	Time with catalyst (s)*	Time without catalyst (s)**
Reaction with catalyst of 1 mL iron (II) solution			

\*From Data Table for Part 3

\*\*From Experiment #3 of Part 1.



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**Experiment 5. Kinetics: The Oxidation of Iodide by Hydrogen Peroxide****Results**

Name \_\_\_\_\_

**Part 1**

Write the differential rate law for the reaction between iodide and hydrogen peroxide.

\_\_\_\_\_

For most purposes, if a small percentage of the reactants are used (less than 10%), the concentration change may be ignored when comparing the initial rate to the average rate. Show, by calculation, that for the present experiments, the number of moles of peroxide used in each run and then calculate what percent of the original hydrogen peroxide was used.

**Part 2**

What are the values of the activation energy and the Arrhenius pre-exponential factor for the reaction between iodide and hydrogen peroxide? Be sure to include the correct units and number of significant figures.

 $E_a$  \_\_\_\_\_

A \_\_\_\_\_

**Part 3**

Compare the rate of the catalyzed reaction to the uncatalyzed reaction. Would you say the iron(II) ions are a good catalyst for this reaction? Why?

\_\_\_\_\_  
\_\_\_\_\_  
\_\_\_\_\_