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Data Tables and Calculations

Part 1 Data Table. 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
Trial 2	Mass of beaker plus 10 drops of water (c)	g
	Mass of second 5 drops of water (c) – (b)	g
	Average mass of 1 drop of water	g
Trial 3	Mass of beaker plus 15 drops of water (d)	g
	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

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

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

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Part 3 Data Table. Determine the Activation Energy.

				Time of Reaction, seconds		
Approximate Temperature, °C	Measured Temperature, °C	Measured Temperature, K	Measured Temperature^{-1}, K^{-1}	Trial 1	Trial 2	Average Time
0						
20						
40						

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

	Reaction Time, seconds
Uncatalyzed Reaction	
Catalyzed Reaction	

Calculations

Part 1.

Calculate the volume of one drop of solution. Assume the density of water to be 1.00 g/mL.

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

Calculations

Part 2A. Calculate the Rate.

The rate will be expressed as $-\Delta[\mathrm{BrO_3^-}]/\Delta\tau$. In each reaction there is one drop of 0.0010 *M* $\mathrm{Na_2S_2O_3}$ solution. Calculate the number of moles of $\mathrm{S_2O_3^{2-}}$ present in one drop:

$$\text{Volume of 1 drop (in L)} \times \frac{0.0010\ \text{mol Na}_2\text{S}_2\text{O}_3}{\text{L}} \times \frac{1\ \text{mol S}_2\text{O}_3^{2-}}{1\ \text{mol Na}_2\text{S}_2\text{O}_3} = \text{moles S}_2\text{O}_3^{2-}\ \text{ions}$$

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 $\mathrm{BrO_3^-}$ which react as all of the $\mathrm{S_2O_3^{2-}}$ ion is used up:

$$\text{mol S}_2\text{O}_3^{2-} \times \frac{1\ \text{mol I}_2}{2\ \text{mol S}_2\text{O}_3^{2-}} \times \frac{1\ \text{mol BrO}_3^-}{3\ \text{mol I}_2} = \text{mol BrO}_3^-\ \text{reacted}$$

The value of $-\Delta[\mathrm{BrO_3^-}]$ in all reactions, since all experiments have a total volume of 12 drops, is:

$$-\Delta[\mathrm{BrO_3^-}] = \frac{\text{mol BrO}_3^-\ \text{reacted}}{\text{volume of 12 drops}}$$

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The rate of each reaction can be found by dividing $-\Delta[BrO_3^-]$ by the number of seconds for the reaction to take place.

$$\text{Rate} = \frac{-\Delta[BrO_3^-]}{\Delta \text{ time}}$$

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

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:

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

where $V_{concentrated}$ and $M_{concentrated}$ 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.

For example, in experiment 1 the initial $[I^-]$ is found as follows:

$$[I^-] = \frac{2 \text{ drops} \times 0.010 \text{ M KI}}{12 \text{ drops solution}} = 0.0017 \text{ M}$$

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Find the initial concentration of each reactant and record in the data table below.

	Initial Concentrations, Moles/Liter		
	$[I^-]$	$[BrO_3^-]$	$[H^+]$
Exp. 1			
Exp. 2			
Exp. 3			
Exp. 4			
Exp. 5			
Exp. 6			
Exp. 7			

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:

$$\text{Rate} = k[I^-]^x[BrO_3^-]^y[H^+]^z$$

Exp. 1: Rate_1 ______ $= k[\quad]^x\,[\quad]^y\,[\quad]^z$

Exp. 2: Rate_2 ______ $= k[\quad]^x\,[\quad]^y\,[\quad]^z$

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

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

Divide and 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.

For example: $8 = 2^n \qquad \log 8 = n \log 2 \qquad n = \dfrac{\log 8}{\log 2} = 3$

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.

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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 J/mol•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.

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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?