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Iodine clock reaction pre lab answers

Objectives To determine the effects of temperature and a catalyst on the rate of a chemical reaction. We have studied the effects of concentration on the rate of the reaction in last week's experiment. This week we will look at other factors that may affect the rate; temperature and catalysts. The reaction we will be studying is: $\text{Ce}(\text{H}_2\text{O}_2) + \text{Ce}(\text{I}^-) + \text{Ce}(\text{H}_2\text{O}_2) \rightarrow \text{Ce}(\text{I}_2) + \text{Ce}(\text{H}_2\text{O})$ There is another reaction involved that changes color to let us know when Reaction 1 is complete, but it does not play a role in the kinetics. Previous determinations have found that the rate law for this reaction is: $\text{Rate} = k[\text{H}_2\text{O}_2][\text{I}^-]$ That is, the reaction is first order with respect to both hydrogen peroxide and iodide. In this experiment, we will examine the effects of both temperature (cold and hot) and the effect of a catalyst on the rate of the reaction. Note that since the clock reaction is relatively fast, the $\text{Ce}(\text{I}_2\text{aq})$ will be consumed by the clock reaction as quickly as it can be produced by the reaction of interest thus holding the $\text{Ce}(\text{I}_2)$ concentration at a very low value close to zero. Only after the $\text{Ce}(\text{S}_2\text{O}_3^{2-})$ has been entirely consumed will the concentration of $\text{Ce}(\text{I}_2)$ start to increase. After the $\text{Ce}(\text{S}_2\text{O}_3^{2-})$ has been consumed the concentration of $\text{Ce}(\text{I}_2)$ will increase rapidly allowing the $\text{Ce}(\text{I}_2)$ to react with a starch indicator resulting in the solution turning to a deep blue color. Catalysts increase the rate of a chemical reaction by providing an alternate pathway or mechanism through which a reaction can proceed. Since the alternate mechanism has a lower activation energy than that of the uncatalyzed reaction the effect of the catalyst is to increase the rate of the reaction. To demonstrate the effect of a catalyst on the reaction rate you will repeat one of the trials from Part A in the presence of an ammonium molybdate, $\text{Ce}(\text{NH}_4)_2\text{MoO}_4$, a catalyst. Safety and Waste Disposal Solutions from the catalyst portion must be collected and placed in the hazardous waste container. In a 100 mL beaker, add (using your graduated cylinder) 8.0 mL of water, 3.0 mL of acetate buffer, 3.0 mL of potassium iodide solution, 1.0 mL of starch solution and 1.0 mL of sodium thiosulfate solution. Mix well. Measure the temperature of the solution. Obtain 1.0 mL of hydrogen peroxide in a different graduated cylinder than you used to measure the reagents in step 1. Simultaneously add the hydrogen peroxide to the beaker and start the stop watch. Observe the solution. When you see a change, stop timing and record the time. Repeat steps 1-4 until you have two times that are within 10 seconds of each other. Obtain and mix the reagents in the same volumes as in step 1 for the room temperature solution. Place the beaker in a water bath with ice. Measure the temperature on the solution. Leave the solution in the bath until the temperature is at least (10°C) colder than the room temperature solution. Add the hydrogen peroxide as in step 2 of the room temperature solution and record the time at which the solution changes. Repeat until your times are no more than 30 seconds apart. Repeat the Cold Temperature procedure using a hot water bath rather than an ice bath. The temperature should be at least (10°C) warmer than the room temperature solution. Times should be no more than 8 seconds apart. Warning Do this part last as the catalyst may contaminate your other runs. Mix the same reagents as in Step 1 of the room temperature trial. Add 1.0 mL ammonium molybdate solution to the beaker. Stir well. Obtain 1.0 mL hydrogen peroxide. Add the peroxide and time as before. Dispose in a waste container and wash your beaker out well. Repeat until you have times that are no more than 5 seconds apart. Since we are not changing the concentrations, we can approximate the rate to equal 1/time. Calculate the two closest rates for each set of reactions (room temperature, hot, cold and catalyst) and average the results. Calculate the ratio of the room temperature reaction to the rate of the reaction in the cold water bath. $\frac{\text{Rate (RT)}}{\text{Rate (Cold)}}$ Calculate the ratio of the reaction in the hot water bath to the rate of the room temperature reaction. $\frac{\text{Rate (Hot)}}{\text{Rate (RT)}}$ Calculate the ratio of the catalyst reaction to the rate of the room temperature reaction. $\frac{\text{Rate (Catalyst)}}{\text{Rate (RT)}}$ In order to continue enjoying our site, we ask that you confirm your identity as a human. Thank you very much for your cooperation. Transcribed image text: Kinetics of an Iodine Clock Reaction Pre-Lab Questions and Calculations You will complete this quiz in Canvas 1 hour before your lab period. This page whe Canvas quiz. will not be turned in or graded. You may use this page to set up your calculations before you take the Canvas quiz. 1. If 5 drops of 0.15 M KI is added to 40 drops of Na2S2O8, what is the final concentration of KI? 2. Which substance(s) is/are used in the experiment to keep ion concentrations and volumes constant but do not appear in the equations? What substance must be used up (limiting reactant) in the clock reaction to cause the solution to turn blue? 3. are used the solution to turn blue. 4. What is the rate of reaction? 5. What is the rate of reaction using the values from question #4? The rate law for the reduction reaction of peroxodisulfate (PODS) by iodide: $\text{S}_2\text{O}_8^{2-}(\text{aq}) + 2 \text{I}^- (\text{aq}) \rightarrow \text{I}_2(\text{aq}) + 2 \text{SO}_4^{2-}(\text{aq})$ will be determined. The orders of reaction with respect to PODS and to iodide will be found by measuring rates for various concentrations of the reactants. Rates will also be measured at three additional temperatures, such that the Arrhenius equation can be used to calculate the activation energy and the pre-exponential factor for the reaction. • Understand and explain common factors that influence the rates of chemical reactions. • Apply linear fitting methods to find relationships between dependent and independent variables. • Quantitatively and qualitatively evaluate experimental data to determine the reaction orders, rate constant, Arrhenius factor, and activation energy of an iodine clock reaction. • Identify and discuss factors or effects that may contribute to deviations between theoretical and experimental results and formulate optimization strategies. • relevant reference and textbook information on kinetics and dilutions Thermodynamics describes the equilibrium phases and concentrations of reactants and products under a given set of conditions. It gives no information on how quickly the equilibrium concentrations are reached. Chemical kinetics is the study of chemical reaction rates. Understanding how fast a reaction occurs, along with the factors that affect the rate, often helps to determine exactly how a reaction is occurring (the mechanism of the reaction). The speed with which a reaction occurs is described by the reaction rate. For a simple reaction like the reaction rate can be expressed as concentration change with respect to change in time. Since the stoichiometry of the reaction above involves coefficients of 1 only, there is a simple relationship between the rates of formation of each product and the rates of consumption of each reactant. The negative signs on the reactant terms ensure that the calculated rate is positive for the forward reaction (the reactant concentration change is negative when the reaction proceeds from left to right). (1) Rate = change in concentration / time interval = - - - Reaction rates often depend on reactant concentrations. In a number of simple cases, the rate is proportional to a reactant's concentration, raised to some power. The overall dependence is the product of these terms. (2) Rate $\propto$ [a][m][b]n or Rate = k[a][m][b]n Here, the reaction is mth order in a and nth order in b; its overall order is m + n. The proportionality constant, k, which is called the reaction rate constant, depends on temperature and is given by the Arrhenius equation. (3) k = Ae-Ea/RT so Rate = Ae-Ea/RT[a][m][b]n Where A is a constant called the pre-exponential factor; Ea the activation energy for the reaction; R is the universal gas constant in units of J/K · mol; and, T is the absolute temperature in Kelvin. You will determine the rate law for the PODS/I- reaction (m, n, A, and Ea). This enables the rate for any concentrations of PODS and I- at any temperature to be calculated. In this reaction, two iodide ions provide one electron each to reduce the peroxodisulfate ion, $\text{S}_2\text{O}_8^{2-}$, to form two stable sulfate ions and molecular iodine. The overall redox reaction is shown below. (4) $\text{S}_2\text{O}_8^{2-}(\text{aq}) + 2 \text{I}^- (\text{aq}) \rightarrow \text{I}_2(\text{aq}) + 2 \text{SO}_4^{2-}(\text{aq})$ The rate law can be written as follows. (5) Rate = - - - = - k[S2O82-][m][I-]n = Ae-Ea/RT[S2O82-][m][I-]n The stoichiometric coefficients for iodide and for sulfate appear in the denominators for their rate expressions. I- is consumed twice as fast as PODS; SO_4^{2-} is formed twice as fast as I_2 . The reaction is mth order in $\text{S}_2\text{O}_8^{2-}$, nth order in I-, and (m + n)th order overall. The specific rate constant, k, for this reaction, as it is balanced, contains the rate's dependence on temperature. This reaction is quite slow, making its rate ideal to measure in a lab setting. It is also convenient to study because molecular iodine, I_2 , is easy to detect. I_2 reacts with starch to form a deeply blue-colored complex. By adding starch to the reaction mixtures, a deep blue color is observed as soon as iodine is produced by the reaction. However, the blue color appears almost immediately and it gets extremely intense very quickly. In order to control how quickly the blue color appears, a second reaction is used. A small amount of thiosulfate ions, $\text{S}_2\text{O}_3^{2-}$, is added to the initial mixture. The $\text{S}_2\text{O}_3^{2-}$ ions react immediately with any I_2 present to regenerate two I- ions. This prevents the solution from turning blue until all of the $\text{S}_2\text{O}_3^{2-}$ is consumed. (6) $\text{I}_2(\text{aq}) + 2 \text{S}_2\text{O}_3^{2-}(\text{aq}) \rightarrow \text{S}_4\text{O}_6^{2-}(\text{aq}) + 2 \text{I}^- (\text{aq})$ Once goes to zero, the I_2 produced will no longer be removed and the solution turns blue. By using a known quantity of $\text{S}_2\text{O}_3^{2-}$, the rate can be calculated in terms of (7) Rate = - - - = k[S2O82-][m][I-]n Because the coefficient of $\text{S}_2\text{O}_3^{2-}$ in the I_2 scavenging reaction is 2 (while the coefficient for I_2 is 1), there is a 2 in the denominator of this term. The change in $\text{S}_2\text{O}_3^{2-}$ concentration is the negative of the [S2O32-]initial because [S2O32-]final = 0. The change in time is simply the measured time, t, from the beginning of the experiment because tinitial = 0. The initial $\text{S}_2\text{O}_8^{2-}$ and I- concentrations change very little during the reaction because the amount of $\text{S}_2\text{O}_3^{2-}$ added is relatively small, so their concentrations can be calculated using the dilution equation. Use Eq. 7Rate = - - - = k[S2O82-][m][I-]n for your experiments. Each run uses the same initial concentration of $\text{S}_2\text{O}_3^{2-}$. • To determine m, use different concentrations of PODS. • To determine n, use different concentrations of I-. • Once you have found m and n, determine the average value of the rate constant k at room temperature. To determine A and Ea, data at different temperatures are required. By doing trials with constant [S2O82-] and [I-] at several temperatures, you can determine k values under these conditions. The activation energy Ea and pre-exponential factor A will be determined graphically. Once m, n, Ea, and A are known, the reaction rate for any concentrations of PODS and I- at any temperature can be calculated. Notes: Bring stop watches or digital watches for recording reaction times in seconds. You will do this experiment with a partner - one person will mix while the other times. Half of the pairs will do Part 1a and half will do Part 1b (unless otherwise instructed). Every pair will do Part 2 (unless otherwise instructed). Every pair will put their times on the board so that the entire class will have good data. Before starting Part 1, set up the solutions for Part 2 (but do not mix them). The solutions need time to reach the temperature of their water bath. Note: Wash your glassware before you use it, and return it clean. Dedicate one pipet to each solution (don't mix and match). There are five runs in this series. The following table gives the volumes of reagents that must be accurately pipetted into a clean, dry 150-mL beaker for each trial. Once you have mixed the first four reagents, initiate the experiment. 2 Obtain 5 mL of 0.20 M KI solution in a volumetric pipet. 3 Start to record the time halfway through the addition of the KI. 4 You may want to record the draining time to get an approximate error in reaction times. 5 Gently swirl the beaker to keep the solution mixed. 6 Record the time of the first appearance of the dark blue color in seconds. There are five runs in this series. The following table gives the volumes of reagents that must be accurately pipetted into a clean, dry 150-mL beaker for each trial. Once you have mixed the first four reagents, initiate the experiment. 2 Obtain 5 mL of 0.10 M $\text{K}_2\text{S}_2\text{O}_8$ solution in a volumetric pipet. 3 Start to record the time half way into the addition of the 5 mL of $\text{K}_2\text{S}_2\text{O}_8$. 4 You may want to record the draining time to get an approximate error in recorded reaction times. 5 Gently swirl the beaker to keep the solution mixed. 6 Record the time of the first appearance of the dark blue color in seconds. Note: Put your results on the board and compare your results with those of others. 7 Calculate to two significant figures using to find M2 for each trial. Note that V2 is the final volume (13 mL). In Part 1a (trials 1a1 - 1a5), varied; in Part 1b (trials 1b1 - 1b5), varied. 8 For each run, determine the rate to two significant figures using Rate = [S2O32-] / t stays the same (use two significant figures); calculate it using Time, t, is the only value that changes trial to trial. 9 To find m, use part 1a data. Take the logarithm of the rate equation. (8a) $\log(\text{Rate}) = m \log [\text{S}_2\text{O}_8^{2-}] + \log(k[\text{I}^-]^n)$ y = mx + b 10 Plot log(rate) vs. log Your log value should have two decimal places. Graph your data, draw the best-fit straight line, and determine the slope (which equals m). Round m to the nearest integer. 11 To find n, use part 1b data. Take the logarithm of the rate equation. (8b) $\log(\text{Rate}) = n \log [\text{I}^-] + \log(k[\text{S}_2\text{O}_8^{2-}]^m)$ y = mx + b 12 Plot log(rate) vs. log Your log values should have two decimal places. Graph your data, draw the best-fit straight line, and determine the slope (which equals n). Round n to the nearest integer. 13 To find the rate constant at room temperature for each trial, use a rearrangement of the rate equation. (9) k = [S2O32-] / (2t[S2O82-][m][I-]n Substitute in your values of m and n, along with the correct time and concentrations for each trial. Calculate values of k for each of the ten room-temperature runs to two significant figures; then find the average value of k and its standard deviation. Do the values of k differ very much? Should they? You already have the data for run #2 of Part 1a (room temperature). Use the concentrations listed in Part 1a, trial #2 to find the rate at 0°C, 35°C, and 45°C. 1 Into one test tube (labeled with contents and your name) pipet the reagents given in the table for run #2 of Part 1a. 2 Into another test tube (similarly labeled) pipet 5 mL of 0.20 M KI solution. 4 Place the pair in an ice water bath made in a 250 or 400 mL beaker. 5 Repeat the process with two test tubes and place them in the 35°C temperature bath. 6 Repeat the process once more and place the tubes in the 45°C temperature bath. 7 Remove a pair of tubes from each bath. 8 Begin to record the time as you add the 5 mL of KI from one test tube into the other. 9 Stopper the tube and gently invert it to mix. Put it back into the water bath. 10 Record the time when you see the onset of the blue color. The low-temperature run may take 20 minutes or more - keep your eye on it, but do other things while you wait. 11 Repeat for both the other baths. 12 Record room temperature. Note: Put your results on the board and compare your results with those of others. 2 You now have four data points consisting of k and the corresponding T (in Kelvin). To perform the graphical analysis, take the ln of both sides of the Arrhenius equation and rearrange the following. (10) $k = Ae-Ea/RT$ $\ln k = -Ea/RT$ $\ln k = mx + b$ 3 Plot $\ln k$ vs $1/T$; $\ln k$ should contain two decimal places, $1/T$ should contain three. You should have four points. The slope of the best-fit line should be straight and equal to -Ea/R; report your slope to three significant figures. The y-intercept should be $\ln A$. 4 Determine the activation energy, Ea, to three significant figures in kJ/mol; R = 8.314 J/K · mol. Calculate the pre-exponential factor, A to two significant figures. 5 Write the complete rate law, substituting your experimental values for m, n, Ea, and A. Waste disposal: follow your TA's instructions. Nothing goes down the sink. Complete your lab summary or write a report (as instructed). Tables including concentrations, times, temperatures and rates Report the Arrhenius Equation for the reaction (including reaction orders, pre-exponential factor, and activation energy) part 1a: rate, log(rate), slope determination part 1b: rate, log(rate), slope determination part 1: individual room temperature rate constants with average and error part 2: k, ln k, 1/T, slope of graph, A determination What did you find out and how? What could be done to improve the accuracy of the experiment?

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