

The catalyst potassium iodide biology essay



**ASSIGN
BUSTER**

This experiment revolves around the decomposition of hydrogen peroxide. There were several objectives for this experiment: Monitor the rate of a chemical reaction, determine the kinetic order, find the energy activation and understand how molecular components function in a rate limiting step of the reaction. Two reagents were examined were H_2O_2 and KI. 8 trials were carried out with 4 differing concentrations for each reagent and two different temperatures were used. The pressure and temperature was recorded using Logger Pro. The experiment found that the decomposition of hydrogen peroxide is increased from the addition of the catalyst KI. Specific calculations found the kinetic rate order, m , n , rate (m/s), rate constant and energy of activation.

INTRODUCTION:

Hydrogen peroxide is a vitally important oxidizing agent in the industry. It is used for various things and breaks down slowly over many years. The decomposition of hydrogen peroxide generates oxygen as a product. This oxygen can then be analyzed using a closed vessel by measuring the change in pressure.

Concepts

To examine the chemical kinetics behind the decomposition of H_2O_2 ; a major oxidizing agent in industry. By introducing a catalyst, such as KI, the decomposition will be sped up. This reaction releases O_2 . The change in pressure is then analyzed using a pressure monitor on Logger Pro. The kinetic order of both H_2O_2 and KI and activation energy to start the reaction can also be measured.

Objectives

The experiment was based on several objectives. First, monitor the rate of chemical reaction. Also, the experiment will focus on how to determine the kinetic order of a reaction from the dependence of the rate on reactions. Additionally, the experiment will help one understand how the molecular components of the rate limiting step of the reaction are determined by the kinetic order of the reaction. Lastly, to find the activation energy of the reaction from the temperature dependence of the reaction rate.

Purpose

The purpose of the experiment is to determine the kinetic order for both and KI. Additionally, the experiment will look to determine the energy of activation for the reaction.

METHODS:

Preparing Solutions

Before the experiment could begin solutions were prepared. 10 mL of 3% was placed into a test tube. A large beaker was then filled with 300-500 mL of room temperature water. The test tube was then placed in the beaker and a probe inserted to measure water temperature. 50 mL of . 5M KI was then measured out in a 100 mL graduated cylinder. The KI was transferred into a 125 mL Erlenmeyer flask. 10 mL of KI was measured and placed into a test tube which was placed into the same water bath as above. The remaining KI was covered and set aside.

Procedure

<https://assignbuster.com/the-catalyst-potassium-iodide-biology-essay/>

A 25 mL Erlenmeyer flask was obtained and secured in the water bath. 2 mL of was pipetted from the test tube into the flask. 1 mL of . 5M KI was quickly pipetted into the flask. The flask was swirled, connected to the pressure monitor and placed into the water bath. Using the logger pro device, data was collected for 500 seconds. Temperature and pressure were monitored during the experiment. The experiment was repeated 8 times with varying amounts of both reagents. A total of 4 different concentrations for each reagent. Water was added into the flask to achieve different concentrations. The last experiment changed the water bath temperature by 10 K. The data was then graphed on logger pro, plotting pressure by time.

Calculations

Before the experiment could begin, the amount of water added to the flask to dilute the reagent, needed to be calculated. This was done using the equation . Where M is the molarity and V is the volume. The gave the concentration needed for the reagent being examined. Multiplying showed the amount of water needed to be added for the right concentration to be achieved. Trial 1-4 changed the concentration of KI while trial 5-8 changed the concentration of . The data was then examined on logger pro and an average temperature was taken for each trial. A linear regression line calculated the average pressure on the graphs. With this pressure, the equation was used to calculate the M for . The reaction order for , expressed as m, and the reaction order for KI, expressed as n, were found graphically. The m value was found by plotting log rate (M/s) as a function of the log of concentration, while the n value was found by plotting log rate (M/s) as a function of the log of KI concentration. The rate law expresses m and n in the

<https://assignbuster.com/the-catalyst-potassium-iodide-biology-essay/>

equation . After m and n were determined graphically, the equation , was used to find the the rate constant(k). The standard deviation and mean of the rate constants were also calculated for trials run near room temperature. Finally, the activation energy was calculated for the two trials performed at different temperatures. The equation used was where $R= 8. 314 \text{ J/mol}\cdot\text{K}$.

RESULTS:

After all the data was collected. Results were analyzed. The following several pages outline the results of the breakdown of Hydrogen Peroxide by use of a catalyst, Potassium Iodide. (Graph 1) shows the pressure vs. time plot of for trial one of the experiment. The next figure (Table 1) outlines the rate of the reaction in M/s with partnered temperature and the concentrations of both and KI. The two graphs accompanied close to each other (Graph 2, Graph 3) detail the kinetic order of the reaction with respect to each reagent. These graphs calculated slope on a linear progression line which analyze the m and n value for and KI. The rate law is then shown where m and n were taken from (Graph 2, Graph 3). The next table (Table 2) gives the rate constant (k) for each trial. A standard deviation and average rate constant were calculated from trials 1-7. The last equation, energy of activation shows the energy of activation for the reaction run at two different temperatures

A plot of pressure vs. time for a single trial is shown below (Graph 1). This graph shows that as the time increased so did the pressure. This relationship seems to be linear with a positive correlation of . 9974. The slope of the graph is . 07959 kPa/s and a y-intercept of 99. 51 kPa.

Graph 1: Pressure vs. Time of . 88(mol/L) for trial number one of the data set. had a concentration of . 88 mol/L and KI was . 5 mol/L

The rates for each trial were calculated below (Table 1). Trial one is the foundation of all trials with the stock solutions used for the reactant and catalyst a final concentration of Hydrogen Peroxide and KI was . 59 mol/L and . 17 mol/L. For trials 2-4 the concentration of decreased while the concentration of KI stayed the same. Trials 5-7 saw a decreased concentration of KI while stayed constant . 59 mol/L. The rates for each trial are detailed on the far right column. The temperature is also shown in the second to last column in Kelvin. For trials 1-7 the temperature was kept around room temperature of 294 K. In trial 8 the temperature was increased by 10 K to 304 K. The concentration of both solutions is given in the middle two columns. The rate neared zero as concentrations of either or KI decreased except when concentration was reduced to . 15 mol/L. This shows that in order for the reaction to occur there needs to be a minimum concentration of either solution.

Trial

Final Concentration(mol/L)

KI Final Concentration(mol/L)

Temperature(K)

Rate(M/s)

1

. 59

. 17

294

3. 632E-5

2

. 29

. 17

294

2. 383E-5

3

. 19

. 17

294

1. 389E-5

4

. 15

. 17

294

9. 647E-6

5

. 59

. 08

294

3. 275E-5

6

. 59

. 06

294

2. 102E-5

7

. 59

. 04

295

1. 419E-5

8

. 59

. 17

303

5. 298E-5

Table 1: Final Concentrations of both and KI for each of the 8 trials.

Temperature is given in Kelvin with all but 1 trial around room temperature.

Rate of the reaction was calculated for each trial.

The $\log(\text{rate})$ vs. $\log(\text{concentration})$ of for trials 1-4 was used to find the m value for for (Graph 2). The m value is the kinetic order of the reaction for Hydrogen peroxide. The slope of the graph is . 947. With respect to the kinetic order; this was rounded to 1. This means that the m value for is 1. The correlation for this graph was very high meaning that the relationship between the two is interdependent on one another.

Graph 2: Log (Rate) vs. Log (Concentration) of pertaining to trials 1-4. The slope of the graph gives the m value for the kinetic order of the equation

The $\log(\text{Rate})$ vs. $\log(\text{Concentration})$ of KI for trials 1, 5-7 is shown below in (Graph 3). This graph was used to find the n value for KI with respect to the kinetic order of the rate equation. The graph has a positive slope of . 6406.

This was rounded to 1. The 1 value represents the n value in the rate equation. As the $\log(\text{concentration})$ increased so did the $\log(\text{rate})$ showing a

positive correlation between the two. The correlation was very strong at. 9062 portraying a strong relationship between the two

Graph 3: Log (Rate) vs. Log (Concentration) of KI for trials 1, 5-7. The slope gives the n value for KI with respect to kinetic order of the reaction

The rate law for the decomposition of catalyzed by KI was found graphically from (Graph 2, Graph 3). The equation is . This means that it is a 1: 1 reaction order for both reagents.

The rate constants, k, for each trial are given in the table (Table 2). The average rate constant and standard deviation of the rate constant are given on the bottom of the table. They pertain only to the trials carried out near room temperature of 294 K. Rate constant is highest in trial 5. The starting concentrations for and KI were . 88 mol/L and . 25 mol/L respectively. At this point, was kept to its stock concentration while KI was halved. The standard deviation of the rate constant is 1. 26E-04 is very small, meaning the calculations and experiment was carried out with very little error.

Trial

Rate Constant(k)

1

3. 61E-04

2

4. 83E-04

3

4. 30E-04

4

3. 78E-04

5

6. 94E-04

6

5. 94E-04

7

6. 01E-04

8

5. 28E-04

Temperature(K)

Average of Constant

Standard Deviation of Constant

294

4. 90E-04

1. 26E-04

Table 2: The calculated rate constant (k) for each trial. The average rate constant and standard deviation of rate constant for the room temperature values for trials 1-7. Trial 8 had a temperature of 304 K

The energy of activation was 28, 252 J/mol and 28. 252 kJ/mol. The equation was rearranged to find . . This is the minimum amount of energy required to bring about the reaction

DISCUSSION:

KI increases the rate of decomposition of hydrogen peroxide as seen in this experiment. It does not undergo a chemical change as first thought. Since KI is basically recycled it never has to undergo a net change in its chemical composition. It is consumed in the reaction but reproduced each time. The ion is the catalyzer in the reaction of hydrogen peroxide. The ion is consumed in one reaction and reproduced in a series of steps in the next reaction. The rate law helps to explain the phenomenon of rate limiting step of the reaction.

Since the rate order of the reaction is first order, it can be explained that the first step is the rate limiting step of the reaction. Before hydrogen peroxide can decompose it must be catalyzed by the KI. By definition, the activated complex is a species at energy peak that must fall apart to form products or reform reactants. In order for the rate limiting step to occur, the hydrogen peroxide and potassium iodide must fall apart to reform reactants. This step is the slowest because in order for this to happen a buildup of energy must occur. Once the energy reaches its peak, the reaction picks up speed and the

reactants are reformed over and over until the reaction exhausts itself. At this point the products are formed, water and oxygen.

In terms of possible chemical steps for this catalysis that are consistent with the observed kinetics; the elementary steps of the reaction coincide with the slow decomposition of hydrogen peroxide, reformation of KI and chemical exhaustion of both reagents to form water and oxygen.

CONCLUSION:

This experiment was interesting due to the fact that it involved various steps and equations. I was able to master several equations and learn numerous concepts. I learned how to calculate the energy of activation, which is needed in order for the reaction to occur. Additionally, the change in concentration of either the reactant or catalyst can greatly affect the rate at which a reaction occurs. Also, I found out how the kinetic order ties into the basic foundation for the reaction and that this can be found graphically by plotting the log (rate) vs. the log (concentration).

I think that this experiment was very helpful not only in master certain skills within the laboratory but also succeeding in understand the basis for all chemical reactions. I did not know chemical kinetics was such an important part of reactions and it will definitely help me to further better my understanding of chemistry in general at a molecular level.