Iodination of acetone lab report pdf

Introduction: Can we help with your assignment? Let us do your homework! Professional writers in all thematic areas are available and will meet their assignment deadline. Free correction and editing of copies included. Kinetics in chemistry deal with the rate at which a chemical reaction occurs. This rate, known as the reaction rate, is defined as the change in the concentration of a reagent or product over time, and is measured in M/s. The speed of a reaction is proportional to the concentration of reagents. An equation called the rate law expresses the relationship of the reaction rate to the constant rate, k, and concentrations of reagents elevated to some experimentally found powers, x and y. The fee law is expressed as, rate at k [A]x[B]y. The constant k is equal to the rate divided by the concentration of a given substance. The purpose in this lab was to experimentally determine the constant rate k, as well as the exponential values of x and y in the rate law. Experimental: The procedure for this lab was obtained from the student laboratory course website. Results: Table 1: Reagent Volumes Trial I2 Acetone H+ dH2O Total Volume 1 0.5 mL 0.8 mL 0.8 mL 1.9 mL 4.0 mL 2 0.5 mL 1.6 mL 0.8 mL 1.1 mL 4.0 mL 3 0.5 mL 0.8 mL 1.6 mL 1.1 mL 4.0 mL Table 2: Initial Concentrations, Times, and Rate for Each Trial Trial [I2] [Acetone] [H+] Time (sec) Rate [I2] k 1 6.25e-2 M 0.68 M 0.2 M 184 sec 3.40e-6 M/sec 2.50e-5 M-1s-1 2 6.25e-2 M 1.36 M 0.2 M 125 sec 5.00e-6 M/sec 1.84e-5 M-1s-1 3 6.25e-2 M 0.68 M 0.4 M 88 sec 7.10e-6 M/sec 2.62e-5 M-1s-1 Sample Calculations: Discussion: To conduct this experiment, the groups placed 1.9 mL of distilled water , 0.8 ml H+ (HCl) and 0.8 ml of an acetone solution in a 4.0 ml. 0.5 ml bucket of I2 was added to the bucket, and the group immediately began to time the reaction. They mixed the contents of the solution by reversing the bucket several times before placing it in a calibrated spectrometer. The absorbance rate was monitored at 400 nm until it reached a nominal zero value. Time stopped and groups were able to determine the rate. Two more tests were carried out, first doubling the original volume of acetone and keeping the H+ in a constant, then in test 3, doubling the volume of the H+ and using the original volume of acetone – 0.8 mL. Constantly keeping the volume of I2 at 0.5 ml, and duplicating one solution while maintaining the other constant made it possible to calculate the constant rate value, k. The chemical reaction being studied was chemical kinetics, the rate at which I2 disappeared. To determine the disappearance rate of I2 in the reaction, equation M1V1-M2V2 was used to find the concentration of I2. So, that value for the time elapsed for Rate. When I2 was first added to the bucket, it was dark red. As the reaction progressed, the solution lost its color and became apparent, consuming the I2 completely. At this point, the spectrophotometer showed that at 400 nm, zero light was being absorbed into the solution. The initial concentrations were varied according to the design of the experiment in order to calculate the exponents of the rate law. The fee law does not include [I2] because I2 does not affect the chemical reaction rate under the selected conditions. It is not included because the disappearance rate of I2 was what it was being resolved for. The rate law determined from this laboratory is the k[Acetone][H+] rate. By using the values of [Acetone], [H+], and the rate, and taking the average of the three trials, it was found that the value of k was 2.32e-5 M-1s-1. READ: Daniel J. Levinson's Seasons Of A Man's Life ExplainedThe goal of determining the values of k, the exponents x and y, and the disappearance rate of I2 were successfully met. The expected results match the results obtained: the concentration of acetone and H+ are directly related to the reaction rate. Both [Acetone] and [H+] are first-order reactions, resulting in a second-order general reaction. The main source of error in the laboratory came from not measuring the substances correctly, resulting in a very unlinked time and reaction rate. Having incorrect amounts of each solution in the bucket directly affected the rate at which I2 disappeared, which in turn made the results not as clear or concecisive. Conclusion: The objective of this laboratory was to experimentally determine how acetone and H+ concentrations affect the rate at which I2 disappears in a reaction by calculating the values of k, the exponents x and y, and putting those values in the equation of the law of fees. These numbers were found by altering the amounts of acetone or H+ used in each trial, which made it possible to use an equation provided above to solve the value of k, and exponents, x and y, in the equation of the rate law. It was found that both acetone and H+ have a direct effect on the reaction rate of I2. The type law for acetone iodine is rate-k[Acetone][H+]. The mean value of k calculated from the three trials was found to be approximately 2.32e-5 M-1s-1. Written by Elizabeth Introduction: The purpose of this experiment is to determine the rate law for the reaction of iodine with acetone. To determine the rate law, we will use the initial rates. Since this is an iodine of the acetone experiment, the initial rate would be the time it takes for the brown color of the iodine to become clear. In the end we should expect to see clear solution and a concentration of zero for iodine. The things that could affect the rate of iodine would be reagent concentrations. The formula used to calculate the rate of this reaction reaction provided by the laboratory manual and is equation 1 below: Reaction rate K [acetone]a[iodine]i[HCl]h When 'a' is the order with respect to acetone, 'I' is the order with respect to iodine and 'h' the order with respect to HCl. After the initial rates are determined, we can find the value of the rate constant, K. We now conduct an experiment to determine the initial rates of six tests, each time varying the concentration of one reagent at a time because it is known that rates could depend on reagent concentrations. Experimental: We did our experiment three times, but there seemed to be something wrong with spectrometers at the time, so the data used in this lab report was given. Here's what the experiment was supposed to be: A- Spectrometer Configuration. 1- A spectrometer was used to observe iodine. The spectrometer was set to a wavelength of 410 nm and then calibrated using a bucket filled with water. 2- For water, we put it in 100% transmission and 0% absorption because it is colorless, and for the dark iodine color solution, it was 100% absorbency, 0% transmission. B- Preparation of solutions. 1- Six reactions were carried out, each using a different amount of reagents. The amount of each reagent to be used was given to us in a table for each of the six tests. 2- The table given to us for the amount of each reagent is below. This is Table 1. Reaction number 4.0 M Acetone(mL) 1.0M HCl (mL) Water(mL) 0.0050M Iodine (ml) 1 3 3 8 8 4 2 6 3 5 4 3 9 3 2 4 4 3 6 5 4 5 3 2 2 4 6 3 4 4 8 Table 1. Number of reagents to be used in each test 3- After the reagents are mixed, they should be placed in the bucket and spectrometer as quickly as possible. A graduated pipette should be used for iodine solution and a different pipette for the rest of the reagents. After pipetting acetone, the pipette should be rinsed with HCl before pipetting the HCl, and the same goes for water. C- Procedure. 1- For reaction number one: Pipette in a beaker 3.00 mL acetone, 3.00 mL of HCl and 8.00 mL of water and in another beaker, pipette 4.00 mL of iodine. 2- Pour the iodine solution into the beaker containing acetone, HCl and water. Mix quickly. 3- Fill the bucket with the solution and place it on the spectrometer. Notice the %T that increases as iodine reacts. When this goes The reaction is over. Record the time at which this is constant. 4- Repeat for reactions 2-6. 5- A seventh reaction is made to confirm the order with respect to iodine and discussion results: Table 2 below describes the time it took for %T to go constant for each of the reactions. All this was done at room temperature (72oF which is 22.2oC). Reaction number 4.0 M Acetone(mL) 1.0M HCl (mL) Water(mL) 0.0050M Iodine (ml) Time(s) 1 3 3 8 4 266 2 6 6 5 5 4 138 3 9 3 2 4 100 4 3 6 5 5 4 133 5 3 2 4 4 85 3 3 4 8 425 Table 2. The time it takes each of the reactions to achieve a constant %T Table 3 below shows the corrected concentrations of each of the reagents, the rate at [I2]/s and the constant K speed for each of the reactions. I'll present the data first and then explain and write down the calculations made to get it. Reaction Number Acetone (M) HCl (M) Water (mL) Iodine (M) Rate ([I2]/s x 105) Rate Constant (k x 105) 1 6.66×10-1 1.66×10-1 8.00 1.11×10-3 4.17×10-1 3.76×105 2 1.33 1.66×10-1 5.00 1.11×10-3 8.04×10-1 3.64×105 3 2.0 1.66×10-1 2.00 1.11×10-3 1.11 3.34×105 4 6.66×10-1 3.33×10-1 5.00 1.11×10-3 8.34×10-1 3.76×105 5 6.66×10-1 5.0×10-1 2.00 1.11×10-3 1.30 3.9×105 6 6.66×10-1 1.66×10-1 4.00 2.22×10-3 5.22×10-1 4.72×105 7 6.66×10-1 1.66×10-1 0 3.33×10-3 4.44×10-1 4.01×105 Table 3: Corrected concentrations of reactants, the rate and rate constant. The calculations are as follows: A- Calculation of corrected concentrations. Using the following formula allows us to calculate the corrected concentrations of reagents to be used. Equation 2: M1V1-M2V2 When M1 is the molarity of the given substance, V1 is the volume used of that substance, M2 is the total molarity used and V2 is the total volume used. They are all calculated in the same way, but I will demonstrate what we did with acetone: Taking reaction number 1, we use 3 ml of an acetone solution 4.0 M. Those are V1 and M1 respectively. M2 is what we need to find and V2 is the total volume of reagents used. Here we use 4.00 ml of iodine, 8.00 ml of water, 3.00 ml of HCl and 3.00 ml of acetone. Adding together gives us 18 ml which is V2. Now plug them into equation 2 gives us: M1V1 m2V2 (4.0)(3.00)-M2(18) M2o (4.0 x 3.00)/18 M2-6.66 X10-1 This is the corrected concentration of acetone in reaction number 1. Doing exactly the same for all of them gives us the corrected concentrations for each. B- Calculation of the rate. The rate

was calculated using the given formula just below it in Table 3.I will also indicate it here as equation 3. Rate [I2]/s x105 Once again, I will demonstrate what we did using reaction 1. The concentration of iodine in one became 1.11×10-3 using equation 2. The time it took %T to be consistent in reaction one was 266 seconds. Plug those numbers into equation 3 da: Rate [I2]/s x105 Rate((1.11×10-3)/266) x 105 Rate 4.17×10-1 Doing this for each reaction gave us the corresponding rates. C- Calculation of the speed constant, K. To calculate the velocity constant, we need to go back to equation 1. Equation one was as follows: Reaction rate-K [acetone]a[iodine]i[HCl]h We calculated the concentrations of acetone, iodine and HCl in Part A, and calculated the reaction rate for each of the reactions in Part B. All values are in Table 3. To find k, all we need here are orders a, i and h. Orders can be found using the following equation: Equation 4. (Rate 2)/(Rate 1)-(M2/M1)x When Rate 2 is the highest of the two rates chosen, Rate 1 is the smallest, M2 is the Corrected concentration corresponding to Rate 2 and M1 is the corrected concentration corresponding to Rate 1. Superscript x is the order with respect to the specific reagent used. Now the next step is to find out which equations to use. To find the order of HCl, I took solutions 1 and 4 because HCl is different while the others are the same. For the order of acetone, I took solutions 1 and 2 because the concentrations of HCl and I2 are the same. And for the order of iodine, I took solutions 1 and 6 because the concentrations of HCl and acetone are the same. Now, starting with HCl, the use of equation 4 and the connection of the corresponding values gives: ([8.34×10-1] / [4.17×10-1]) á ([3.3×10-1] /[a 11.66×10-1])h Log(8.34/4.17) á h Log (3.3/1.66) h to 1.01 Order h with respect to HCl is 1.01. En segundo lugar, la acetona: El uso de la ecuación 4 y la conexión de los valores correspondientes da: Log([8.04×10-1]/[4.17×10-1]) - un registro ([1.133]/[6.66×10-1]) a(0.947) a a-0.9 La orden a con respecto a la acetona es 0,9 Tercio, Yodo: El uso de las ecuaciones 4 y la conexión en los valores correspondientes da: Log([5.22×10-1]/ [4.17×10-1]) - i Log ([2.22×10-3]/ [1.11×10-3]) 0.98 301) i a 0,3 El orden a con respecto a la acetona es 0,3 En este laboratorio, decidimos utilizar enteros y por lo tanto redondeamos estas órdenes hacia arriba o hacia abajo para hacer un número entero de la siguiente manera: a 0,9 se redondeó hasta 1 h 1,01 se redondeó a 1 i 0,3 redondeado a 0 Ahora que tenemos todo lo que necesitamos para encontrar la tasa constante se redondeó a 1 i 0,3 se redondeó a 0 Ahora que tenemos todo lo que necesitamos para encontrar the constant rate was rounded to 1 i 0.3 was rounded to 0 Now that we have everything we need to find the constant rate was rounded to 1 i x 0.3 was rounded to 0 Now that we have everything we need to find the constant rate is to 1 i 0.3 was rounded to 0 Now that we have everything we need to find the constant rate was rounded to 1 i 0.3 was rounded to 0 Now that we have everything we need to find the constant rate, K, now we can connect everything in equation one which again is as follows: Reaction rate K [acetone]a[iodine]i[HCl]h The velocity constants for all reactions are in Table 3. I'll write the math for one of them because they're all the same. That's how I got the speed constant for reaction one. Reaction Rate-K [acetone]a[iodine]i[HCl]h 4.17×10-1o K [6.66×10-1]1 [1.66×10-10-1 1]1 [1.11×10-3]0 4.17×10-1o K(0.111) K á (4.17×10-1)/ (4.17×10-1)/ K- 3.756 And since in the table it is given that the velocity constant is Kx105 then each K value obtained is multiplied by 105 So the velocity constant for reaction 1 is: 3.76×105 For the seventh reaction, the order is going to be the same because all the other values are the same so we're going to have to take the same solutions 1 and 6 because those are the same and the iodine is different. This confirms the order with respect to iodine and fulfills the purpose of the 7th test. Conclusion: Now that we find the rates and the constant speed, we can say which reaction was faster and which concentrations were best to use according to reaction rates. The results show that reaction 4 had the highest rate and therefore the optimal concentrations that should be used to react at that rate would be 6.66×10-1 M of acetone, 3.33×10-1 M of HCl, 5.00 ml of water and 1.11×10-3M of iodine. All of this reacts to a speed of 8.34×10-1. The velocity constant is higher in equation six, which means that the reagents react at a constant rate of 4.72×105 however, the rate of that reaction is lower than the reaction rate 4. The temperature here is not a variable because all of these were carried out at room temperature, therefore this shows what was indicated above in the introduction, that rates vary depending on the concentration of the reagents. Reagents.

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