|  |
| --- |
|  |
| Experiment 8: Kinetics of an Ionic Reaction |
| Lab Partners: Jeel Shah, Harshal Shah, Cosette Lim |

|  |
| --- |
| Max Shi  11-1-2021  I pledge my honor that I have abided by the Stevens Honor System. |

# Abstract

In this experiment, the kinetic salt effect was investigated in the context of the iodine clock reaction. The iodine clock reaction, due to being a reaction involving a color change when the first, rate-limiting step has reached completion, is a great candidate for the kinetic salt effect, as its reaction time is easily measured and it involves two negatively charged species, . The kinetic salt effect states that in a reaction with charged species of the same sign, introduction of a species of opposite sign will cause both reaction species to be attracted to the ion, reducing the overall distance between reaction species, and increasing reaction rate. To test this, the concentrations of the two reacting species were held constant and increasing concentrations of Mg2+were added to the solution. A reaction time reduction was observed, and when the square root of ionic strength against the log of reaction time was substituted into a best fit plot according to Debye-Hückle theory, a charge product of 1.786 was obtained, very close to the expected value of 2.

# Introduction

This experiment uses the “Iodine Clock” reaction, a reaction between iodine and sulfate anions in a hydrogen peroxide and sodium thiosulfate solution. Starch is added to the reaction as a color indicator for the I3- ion, and this drastic and predictable color change is used for lecture experiments. The reaction mechanism is as follows:

Equation : Iodine Clock Reaction Steps

In this reaction, the second reaction counts as the fast step, and the first reaction is the slow step. Thus, the rate of the entire reaction is determined by the concentration of the initial iodine ion and the concentration of sulfate ion.

Another interesting aspect of this reaction is the kinetic salt effect. As seen above, the reaction is between two anions in the solution, two species that would naturally repel each other due to having the same signed charge. The introduction of other ions, especially the Mg2+ cation in MgSO4, cause both anions in reaction to have an attraction to the positive charge. Their overall distances would be reduced, causing and overall reduction in repulsion as a result of introduction of the cation. According to Debye Hückel theory, the introduction of these ions changes the rate constant according to the following equation:

Equation : Debye-Hückel Theory of Ionic Solutions

where z represents the charge of the ions, zA and zB are the two reacting ions, and A is a constant equal to 0.509 at room temperature. Although we cannot measure k directly in this experiment, we can assume that the time is takes for the measurement to run is inversely proportional to the rate constant. Thus, with substitution, equation 3 is obtained, and applied to this experiment.

Equation : Debye-Hückel Theory with Time Substitution

By measuring the time it takes for the solution to change color, a rate can be determined to predict the overall charge of the ions in reaction.

# Experimental

To begin the experiment, two solutions were created. Solution A was a solution of 250 mL 1.00% by weight H2O2 diluted from 3% H2O2 and 0.00020 M Na2S2O3 diluted from 0.02 M. Solution B0 was a solution of 100 mL 0.040 M KI. Then, four solutions were made from solution B0, which was 100 mL each of 0.01, 0.04, 0.07 and 0.1 M MgSO4 along with the same concentration of KI as solution B0. For each solution, 50 mg of starch was added to a small beaker with 10 mL of solution A. Then, 10 mL of the first B solution was added to the beaker, and the stopwatch was started when the solutions were mixed. The reaction was mixed two more times, and the color change interval was recorded to obtain an average of the two closest values. This process was repeated for every solution created from B. Below is a table of the dilutions used to create each solution.

|  |  |  |
| --- | --- | --- |
| Solution A (250 mL) | 83.3 mL 3% H2O2 | 2.5 mL 0.02 M Na2S2O3 |
| Solution B0 (100 mL) | 0.004 mol KI \* 166 g/mol KI | 0.664 g KI |
| Solution B1 (100 mL) | 0.01 M \* 0.020 L \* 120.366 g/mol MgSO4 | 0.024 g MgSO4 + 0.664 g KI |
| Solution B2 (100 mL) | 0.04 M \* 0.020 L \* 120.366 g/mol MgSO4 | 0.096 g MgSO4 + 0.664 g KI |
| Solution B3 (100 mL) | 0.07 M \* 0.020 L \* 120.366 g/mol MgSO4 | 0.168 g MgSO4 + 0.664 g KI |
| Solution B4 (100 mL) | 0.1 M \* 0.020 L \* 120.366 g/mol MgSO4 | 0.240 g MgSO4 + 0.664 g KI |

Figure : Solution Dilution Table

# Results

After the two closest trials for each solution were identified, their averages were taken, and the log of these values was calculated and recorded to form the left side of Equation 1. This data table is seen below.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Solution | Time 1 | Time 2 | Time 3 | Avg of 2 closest values | log(avg) |
| B0 | 11.78 | 12.84 | 12.4 | 12.62 | 1.101059355 |
| B1 | 6.2 | 5.79 | 5.27 | 5.995 | 0.777789187 |
| B2 | 4.08 | 4.01 | 4.35 | 4.045 | 0.606918526 |
| B3 | 3.5 | 3.45 | 3.88 | 3.475 | 0.540954809 |
| B4 | 3.25 | 2.9 | 2.61 | 2.755 | 0.440121603 |

Figure : Reaction Data and Calculation for Each Reaction Solution

To form the term for I, the concentrations of all ionic species were calculated, and plugged into equation 2. Due to the relatively dilute solutions, it was assumed that the density of the solution was 1 g/mL, which makes the conversion from molarity to molality 1:1. These calculations are seen below.

|  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- |
|  | Na+ | S2O3 2- | K+ | I- | Mg 2+ | SO4 2- | Ionic strength | I^1/2 |
| B0 | 0.0002 | 0.001 | 0.02 | 0.02 | 0 | 0 | 0.0221 | 0.14866069 |
| B1 | 0.0002 | 0.001 | 0.02 | 0.02 | 0.005 | 0.005 | 0.0421 | 0.20518285 |
| B2 | 0.0002 | 0.001 | 0.02 | 0.02 | 0.02 | 0.02 | 0.1021 | 0.31953091 |
| B3 | 0.0002 | 0.001 | 0.02 | 0.02 | 0.035 | 0.035 | 0.1621 | 0.40261644 |
| B4 | 0.0002 | 0.001 | 0.02 | 0.02 | 0.05 | 0.05 | 0.2221 | 0.47127487 |

Figure : Ionic Strength Calculations for Each Ion in Each Solution

The last column in figures 1 and 2 constitute the dependent and independent variable in equation 3, respectively, in the plot to determine the product of the charges of the two reacting ions. These data points were plotted on a graph, and a line of best fit was created.

Figure : Graph of Equation 3 with Reaction Data

The slope is analogous to the term in equation 3, and as such, we set them equal and solve for .

Equation : Calculation of Charge Product

# Discussion

Although this reaction was very prone to human error due to the measured reaction time being collected by a human with a stopwatch, as well as the reaction reaching completion being a subjective color-based determination by the person conducting the experiment, we were still able to recover a product of charges value of 1.786. This number represents the charges of the two reaction ions in the slow step, which were and , which would have a charge product of 2. Thus, despite the large human error possible in the experiment, the result of the experiment was remarkably close to the expected value, with a percent error of only 10.7%.

This experiment is also a great example of the kinetic salt effect in a reaction. As a greater concentration of MgSO4 was added to the solution, a decrease in reaction time was observed, especially noticeable in the halving of the reaction time between solution B0 and B1. This is strong evidence to support the idea of the kinetic salt effect in this reaction, where the two reacting ions are negatively charged, and introduction of a strong positive charge may reduce the average distance between two molecules due to mutual attraction to the opposite charge, improving the reaction rate. This increase in ionic strength of the solution is represented by an increase in the I1/2 term, and the reduction of reaction time gives a negative slope for the reaction. This leads to the charge product being positive, and the conclusion that the two reaction ions have the same charge.

This is an interesting observation, as this equation also suggests the opposite is possible when the two ions are opposite charges. If they are opposite charges, the charge product will be negative, which suggests an increase in ionic strength of the solution will exponentially increase the reaction time. This makes sense in the context of the kinetic salt effect, as the two reacting ions will not be as attracted to each other and attracted to other species in the reaction, reducing the overall collisions and the reaction rate. This would be something to further explore in a future experiment.

# Appendix

1. Excel spreadsheet for reaction data:

