Experiment #6
The Rate and Order of a Chemical Reaction

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CHE-1210L, Section 89226, Spring Semester 2024
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Goal of the Experiment

This experiment uses various concentrations of potassium iodide (KI) and iron (III) chloride ($FeCl_3$) to determine the order of the reaction. When these two compounds are mixed and the reaction proceeds, there is a color change which means the samples can be measured by a spectrometer. Using the graph generated by the spectrometer software, the initial rate can be calculated and used with the known concentrations of the compounds to find the order of the reaction.

Chemistry Principle

By using a stock solution of potassium iodide and iron (III) chloride, samples of various concentrations can be made. Using different concentrations of the samples will show the effect of each compound on the rate of reaction. This information is used to create the rate law for the reaction.

The known rate law is $rate = k[FeCl_3]^1[KI]^0$, meaning the order of the reaction with respect to iron (III) chloride is 1, and the order with respect to potassium iodide is 0. The overall order of the reaction is 1. This is consistent with other common reaction orders, and the expectation is that the experimental results will be close to the known rate law.

Rate law for a chemical reaction can only be found experimentally. The best way to determine the rate of the reaction is by using the initial rate. The initial rate is nearly linear and easily measurable across multiple samples because the measurement will be taken from the start of the reaction to the point where the reaction no longer resembles a linear function. That same interval can be copied and used across multiple trials.

Results and Discussion

The experiment consisted of five trials with different concentrations of $FeCl_3$ and KI. Based on the known order of the reaction, changing the concentration of $FeCl_3$ would change the rate of the reaction. Changing the concentration of KI should have no effect on the rate of the reaction.

In practice, the rate of the reaction did decrease when $[FeCl_3]$ was reduced. However, the decrease was not consistent with a rate order of 1. The rate also changed when [KI] changed. This happened both when the concentration of $FeCl_3$ was held constant at .01M and .005M.

The rate law determined by the experimental results was *rate* = $k[FeCl_3]^{.585}[KI]^{.645}$.

Conclusion

After five trials, the rate law that was experimentally determined did not match the known rate law. Sources of error may have included not getting a reading on the sample fast enough to accurately capture the initial rate, an inaccurate starting concentration of either compound, or an inaccurate stock solution. No matter the source of error, more trials should be performed for a more accurate result.