

Investigation 7: Reaction Kinetics: Effect of Temperature

Focus Questions: How does temperature effect the rate of a reaction? How can we determine the rate law for a reaction as a function of temperature? How can we apply this to a real life situation?

Pre-lab required reading

Chemistry; an Atoms-Focused Approach: [Sections 13.4](#)

Primers:

- [Keeping a Laboratory Notebook](#)
- [Using a micropipette](#)
- [Volumetric glassware – general](#)
- [Volumetric glassware – volumetric flask](#)
- [Pseudo-first order reactions: the method of isolation](#)
- [Spectrophotometry](#)
- [SpectroVis usage](#)

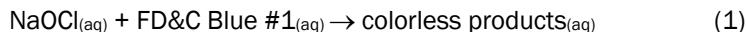
Safety and Waste Disposal

- Eye protection should be worn at all times. Bleach is a corrosive oxidizing agent that will react with eyes and skin and can remove color from clothing.
- Gloves should be worn when handling the bleach.
- Dispose of all solutions down the drain with plenty of water.

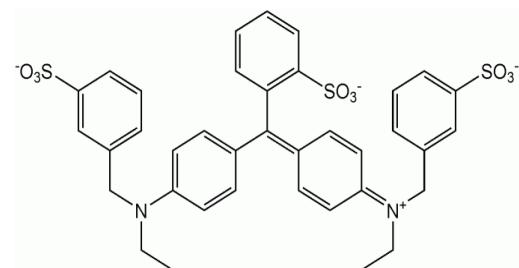
Background

The Reaction

In this experiment we will be studying the reaction between blue food dye #1, erioglaucine, and bleach, $\text{NaOCl}_{(\text{aq})}$, in solution. The overall reaction equation is:



The structure of aqueous FD&C Blue #1 is shown below:



Observing the reaction

The FD&C Blue#1_(aq) ion is colored. Thus a solution of the reactants will be colored. However, the products are not colored and a solution of the products will be colorless. The progress of the oxidation-reduction reaction represented by equation 1 above can be observed by watching the color of the solution disappear. Quantitative information will be obtained by using the absorbance of the solution at the wavelength of maximum absorption for the FD&C Blue#1 ion (630 nm). We will use the SpectraVis spectrometers to record the absorbance as a function of time.

The general rate law for this reaction (1) is

$$\text{Rate} = k[\text{NaOCl}]^x[\text{Blue}\#1]^y \quad (2)$$

In Investigation 4 (CHEM120), the purpose of the experiment was to fully define the rate law shown in equation (2) by determining experimental values for the variables x, y and k. The *method of isolation* was used with an excess of NaOCl. If we assume that the rate law developed during that investigation holds as a variety of temperatures, we can now investigate the effect of temperature on this reaction.

Procedure

Part one:

Part A: Preparation of materials:

Obtain 10-20 mL of Blue#1 Stock solution. Volumetrically prepare 50mL of diluted Blue#1 (concentration of $\sim 1.7 \times 10^{-5} M$ is desired). Be sure to record the exact value of the concentration of solution prepared. Place some of this diluted solution, ~20mL distilled water, and ~20mL bleach solution in separate clean, dry, labeled test tubes. You may need more than one test tube for each. Use the label information (i.e. 6% NaOCl) to calculate the concentration of NaOCl in your bleach solution. Obtain a clean dry cuvette and three clean micropipettes (with tips) – these will be used to measure all the solutions and also the distilled water. Be careful to keep the micropipette clean and dry by not setting it sidewise on the counter, use the clips provided. Be sure to use the same tip with the appropriate solution at all times so that your solutions do not become contaminated (if not sure – use a new tip!).

Part B: Determining the effect of temperature

Using a hot plate, make a hot water bath in a two thirds-filled 400-mL beaker. Heat the bath and the separate solutions of diluted blue #1, bleach and water until they are approximately the 60 °C. Try to maintain this temperature to within 1-2 °C by taking the beaker on or off the hot plate as necessary. Be careful when handling the hot beaker!

Optimizing Experimental Conditions:

Determine which of the trials used last semester has the optimal concentrations at this higher temperature. You may want to begin by using the volumes and concentrations of a trial that was relatively slow at room temperature. Think about what you would want to change if this reaction is still too fast at the higher temperature. What concentration should be decreased?

Be sure to blank the spectrometer with the cuvette containing warmed water before adding the Blue#1 and Bleach. The Blue#1 and bleach should be added simultaneously after the water aliquot is in the cuvette. This method should sufficiently mix the solution quickly. The spectrometer should be set up and ready to record the instant that the solutions are added. Record the temperature in the cuvette, as it will cool slightly upon removal from the warm water bath. Be careful to not immerse the temperature probe in the path of the light. The spectrometer should be set to collect absorbance as a function of time. Record the absorbance as a function of time until the absorbance completely levels off. This may take a few trials and may vary with each spectrometer. Once you have optimized the conditions, you may begin your temperature trials.

Collecting Variable Temperature Trials:

Be sure to use the SAME concentrations (volumes of each solution) for each trial, so that you are ONLY changing temperature. This is why you are optimizing the conditions. First conduct a trial at the warmest temperature reached (around 60°C), then allow the solutions to slowly cool by turning off the hot plate or by carefully removing the bath from the hotplate. Consistently add the Blue#1 and bleach simultaneously to mix the solutions and record the absorbance as a function of time until the absorbance completely levels off for each trial. Repeat this procedure such that you collect 8-10 trials at different temperatures. You may need to cool the solutions below room temperature using an ice water bath.

Transforming the raw data:

In order to compare the different temperature trials to each other, one should normalize the absorbance values for each trial. To do this, the “zero-point” absorbance (the absorbance at the end of the trial when it has leveled off completely) should be subtracted from the absorbance at every other point. Then, transform the plots of normalized absorbance as a function of time obtained for each temperature into the linear form for a pseudo-first order reaction. Use the rate constants obtained from those plots to make an Arrhenius Plot of the data.

References

Atkins, P.; Jones, L. "Chemical Principles: The Quest for Insight", 5th ed.; Freeman: New York. **2010**.

Henary, M.M.; Russell, A.A. *J. Chem. Ed.* **2007**, 84, 3.