Experiment 23: The Kinetics of the Crystal Violet and Sodium Hydroxide Reaction

Purpose To study the kinetics of the reaction between crystal violet (CV) and sodium hydroxide (NaOH). This will involve the determination of the rate law equation; finding the values of the exponents and the value of the rate constant.

Background

The SDS for crystal violet has several warnings; health hazard, exclamation mark, and environmental hazard are some of them. However, crystal violet has been used for several commercial purposes; bacterial and fungal infection-treatment creams, as well as dye for wood, silk, and paper *(ref. PubChem, Gentian Violet, accessed 4/16/2021)*.

This crystal violet compound has an intense, violet color, which makes it ideal for study with a spectrophotometer. The CV compound consists of the CV cation and chloride anion. When a solution of crystal violet is mixed with a solution of hydroxide ion, a reaction occurs in which the product formed is colorless:

$$CV^{1+}_{(aq)} + OH^{1-}_{(aq)} \rightarrow CVOH_{(aq)}$$

violet *Colorless* *Colorless*

The CV solutions you will use for this kinetics experiment will start as violet colored solutions, and when mixed with the sodium hydroxide solution, will fade to colorless. You will use the spectrophotometer to record the absorbance of 590 nm light by the CV^{1+} ion for several minutes, for several trials. The absorbance values will decrease, since the concentration of CV^{1+} decreases over time (reactants get used).

The rate law equation has a general format of reaction rate = $k\,[A]^m\,[B]^n$
For this experiment, the more specific format is reaction rate = $k\,[CV^{1+}]^m\,[OH^{1-}]^n$

To determine the exponents in the rate law equation, several trials of the reaction will be done. The experiment is set up so we can study how the change in concentration of only one reactant affects the rate of the reaction. Therefore, one reactant's initial concentration is different for two of the trials, while the other reactant's initial concentration is kept the same.

When the initial concentration of the hydroxide ion is the same for two trials, and the CV^{1+} initial concentration is different for those two trials, this allows for the determination of the CV^{1+} exponent.
When the initial concentration of CV^{1+} is the same for two trials, and the OH^{1-} initial concentration is different for those two trials, this allows for the determination of the OH^{1-} exponent.

The general format for the mathematical equation for the calculation of the exponent is:

(ratio of initial concentrations)^m = (ratio of reaction rates)

Be sure to ratio the appropriate two trials for each reactant. Solve for the exponent.

Chemicals
0.10 M NaOH solution ***(You will only need approximately 20 mL of NaOH solution.)***
2.0 x 10^{-5} M CV solution ***(You will only need approximately 20 mL of CV solution.)***

Equipment
Small test tubes that fit in the sample compartment of the spectrophotometer
5 mL pipets (three), pipettor
Parafilm squares, timer, three beakers (for the reactants and DI water)

Procedure
Part A: Preparation of the Standard Solutions for the Calibration Curve
Use the CV stock solution, 2.0 x 10^{-5} M, and pipet into the test tubes. Pipet in the DI water, cover with Parafilm, and invert three times to mix.

Standard 1: Pipet 1 mL CV stock and 4 mL DI water

Standard 2: Pipet 2 mL CV stock and 3 mL DI water

Standard 3: Pipet 3 mL CV stock and 2 mL DI water

Standard 4: Pipet 4 mL CV stock and 1 mL DI water

Set the wavelength on the spectrophotometer to 590 nm, use deionized water as the blank, and measure and record the absorbance of each of the standard solutions. This data will be used to prepare a calibration curve.

Part B: Determining the Exponents in the Rate Law Equation
Use the crystal violet stock solution, 2.0 x 10^{-5} M, and the NaOH stock solution, 0.10 M

Table 1: Volumes of Reactants for the Rate Experiment

Trial Number	Volume CV soln, mL	Volume DI H_2O, mL	Volume NaOH soln, mL
1	1.0	1.0	3.0
2	2.0	0.0	3.0
3	2.0	2.0	1.0

Do one trial at a time; start to finish. Put the reactants into the test tube before placing the test tube into the spectrophotometer sample compartment.

a) Add deionized water to a test tube and use it as a blank to zero the spectrophotometer.
b) Pipet the crystal violet solution into a test tube, then pipet the deionized water into the test tube.
c) Wipe the outside of the test tube with a Kimwipe.
d) Pipet the NaOH into the test tube, and start the timer when you first start to pipet the NaOH into the test tube.
e) Place a piece of Parafilm over the test tube and invert three times to mix.

f) Remove the Parafilm, and immediately place the test tube into the sample compartment of the spectrophotometer, and then close the lid.
g) Record the absorbance every 10 seconds for 240 seconds. The first time point that you will probably be able to record data for is the 20 or 30 second time.
h) Perform steps a-g for all three trials, using the volumes indicated in Table 1.

Repeat Trials 1 – 3 for a second time, which will let you average the absorbances. There will be trial 1a, 1b, 2a, 2b, 3a, 3b, and then the calculated averages: 1_{ave}, 2_{ave}, 3_{ave}.

Calculations: *(Each calculation is relatively easy, but there are a lot of them.)*

1) Calculate the concentration of CV^{1+} in each standard solution. (Mc*Vc = Md*Vd)
2) Calculate the initial concentration of CV^{1+} in each rate experiment trial. (Mc*Vc = Md*Vd) The total volume in each test tube was 5.0 mL.
3) Calculate the initial concentration of NaOH in each rate experiment trial. (Mc*Vc = Md*Vd) The total volume in each test tube was 5.0 mL.
4) Calculate the average absorbance for each time point for each trial. (Use Excel or Google Sheets for these calculations, watch the video in Blackboard for directions.)
5) Use Excel or Google Sheets to make a calibration curve with the concentrations of the standard solutions and their absorbances. Get the equation of the best-fit line, which will be in the y = mx + b format. The R^2 value will show you how well you made your standard solutions.
6) Calculate the average concentration for each time point, for each trial, using the equation of the best-fit line from the calibration curve and the ave. absorbance values. Use Excel or Google Sheets to do these calculations (watch the video in Blackboard).
7) Calculate the initial reaction rate for each trial; use $\Delta[CV^{1+}] / \Delta t$ with the first two time points you were able to record absorbance data for.
8) Determine the exponents in the rate law equation, by determining the ratio of concentrations and ratio of initial reaction rates, using the appropriate two trials from the rate experiments.
9) Confirm the order of CV^{1+} you calculated by graphing (ln $[CV^{1+}]$ vs. time (s)); t is the x-axis. A straight line confirms a first order for CV. Make a graph of ($(1/[CV^{1+}])$ vs. time (s)). A straight line confirms a second order for CV.
10) Calculate the value of the rate constant, k, using the rate law equation you have determined so far and each trial in the rate experiments. Calculate the average k value.
11) Report out the rate law equation that you determined; use the average k value and the exponents you calculated.