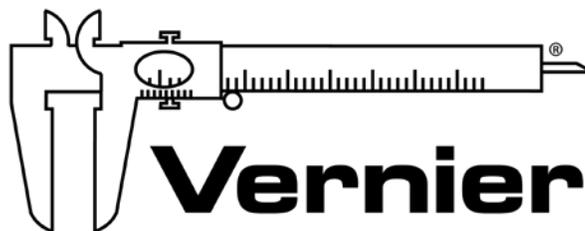


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HANDS-ON ACTIVITIES

Investigating Commercial Hand Warmers

- Go Direct Temperature Probe

Determination of an Equilibrium Constant

- Go Direct SpectroVis Plus Spectrophotometer

Investigating the Kinetics of a Crystal Violet Reaction

- Go Direct SpectroVis Plus Spectrophotometer

Acid/Base Titration

- Go Direct pH Sensor and Drop Counter

Investigating Commercial Hand Warmers

For centuries, humans have put various items into their mittens and boots to help keep their fingers and toes warm when the weather outside was very cold. There are several different types of products advertised as “hand warmers” or “foot warmers.” Some of these products are designed to retain warmth from an exothermic chemical reaction or a physical change.

One type of hand warmer is composed of a tightly sealed and insulated packet within which there are two smaller packets. One packet holds a volume of water and the other packet holds a powdered solid. You twist or squeeze the packet, and then shake it a little to mix the materials. As the solid dissolves, it releases heat.

In the Initial Investigation, you will use a Temperature Probe to determine the temperature change of two substances as they dissolve in water. After completing the Initial Investigation, you will test several substances to determine which substance is the best hand warmer based on two factors: enthalpy of solution and cost

PRE-LAB ACTIVITY

1. When baking soda (sodium hydrogen carbonate) dissolves in water, the temperature of the water decreases and the dissolution is spontaneous. Explain how and why this happens in terms of bond breaking and bond forming behavior.
2. When washing soda (sodium carbonate) dissolves in water, the temperature of the water increases and the dissolution is spontaneous. Explain how and why this happens in terms of bond breaking and bond forming behavior.
3. Measuring the enthalpy of solution of a substance is performed in a device called a calorimeter. A simple calorimeter can be made from a foam coffee cup because it insulates well, holds a reasonable volume of liquid, and it is inexpensive. Conduct research to find a method for calibrating a calorimeter. In other words, find a way to determine the amount of heat that is lost to the calorimeter itself.

INITIAL INVESTIGATION

In the Initial Investigation, you will measure the enthalpy of solution of magnesium sulfate, MgSO_4 , to gain practice in setting up a simple calorimeter, fashioned from foam coffee cups, and using the best lab techniques to achieve accurate and reliable temperature readings. It can be difficult to accurately measure the temperature change that occurs when a solid dissolves in water because of the many ways heat can be lost, including heat lost to the calorimeter itself.

Investigating Commercial Hand Warmers

You may consider running an extra set of tests to calibrate your calorimeter. If so, use your answer to Pre-Lab Question 3 to calibrate your calorimeter.

1. Obtain and wear goggles.
2. Connect the Temperature Probe to LabQuest and choose New from the File menu.
3. Set up the data-collection program.
 - a. Tap Rate.
 - b. Change the rate to 1 sample/s and the duration to 300 s.
4. Nest a foam coffee cup into a 400 mL beaker. Add 50.0 mL of distilled water to the cup.
5. Measure out about 5 g of solid magnesium sulfate, MgSO_4 . Record the mass of your sample. **WARNING:** *Solid magnesium sulfate, MgSO_4 : May cause damage to organs through prolonged or repeated exposure. Do not eat or drink when using this product.*
6. Immerse the tip of the Temperature Probe into the distilled water. Make sure the temperature of the water is stable and the probe is not touching the cup.
7. Start data collection. Collect water temperature readings for 10 seconds, to establish a baseline temperature, and then add all of the MgSO_4 at once. Use the Temperature Probe to gently stir the mixture.
8. When the temperature readings reach a maximum or minimum value and level off, stop data collection.
9. Analyze the graph to determine the temperature change.
10. Dispose of the mixture as directed by your instructor.
11. To save the file, choose Save from the File menu.
12. Repeat the necessary steps to measure the temperature change of a new 5 g sample of magnesium sulfate dissolving in 100.0 mL of distilled water.

PLANNING FOR THE FINAL INVESTIGATION

Discuss the results of the Initial Investigation with your group, to formulate a plan to determine the best substance to use in a commercial hand warmer. Your instructor will assign your group a set of three substances to evaluate for use in a hand warmer. The substances you will be testing will be drawn from Table 1.

Table 1		
Substance	Formula mass (g/mol)	Cost per 500 g (\$)
Sodium carbonate, Na ₂ CO ₃	105.99	6.25
Calcium chloride, CaCl ₂	110.98	6.65
Sodium chloride, NaCl	58.44	4.10
Lithium chloride, LiCl	42.39	90.00
Sodium acetate, NaCH ₃ COO	82.03	14.50
Ammonium chloride, NH ₄ Cl	53.49	11.10

You will be given 15 g of each substance, thus plan carefully.

You will measure the enthalpy of dissolution for each substance. As you plan your Final Investigation, consider the following issues:

- What is the best way to evaluate cost effectiveness?
- A hand warmer should be warm, but not hot. How warm is too warm?
- A hand warmer has an optimum size (fits comfortably in a glove or mitten) which limits the amount of “stuff” you can use to generate and hold heat.
- Any product involving chemicals needs to be safe to use and not pose a disposal problem.

HAZARD ALERTS

WARNING: *Solid anhydrous sodium carbonate, anhydrous, Na₂CO₃: May be harmful if swallowed. Causes eye irritation and mild skin irritation.*

WARNING: *Solid calcium chloride, CaCl₂: Do not eat or drink when using this product—harmful if swallowed. Causes serious eye irritation.*

WARNING: *Solid sodium chloride, NaCl: May be harmful if swallowed. Treat as a non-food-grade chemical.*

WARNING: *Lithium chloride causes skin irritation, eye irritation, and respiratory tract irritation.*

WARNING: *Solid anhydrous sodium acetate, CO₂CH₃Na: May be harmful if swallowed. Causes eye irritation and mild skin irritation.*

WARNING: *Solid ammonium chloride, NH₄Cl: Do not eat or drink when using this product—harmful if swallowed. Causes skin irritation.*

FINAL INVESTIGATION

As you carry out your approved plan, consider the following questions:

- What are the optimum number of data-collection runs needed to achieve the best data?
- What is the insulating efficiency of the coffee cup calorimeter for best temperature readings?
- How should the samples be handled and measured to achieve the best data?

ANALYZING RESULTS

When preparing your report, include

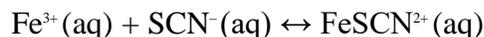
- A statement of the results: What is the best and most cost-effective substance to use as a hand warmer, of the substances assigned to your group?
- A description of the procedure that you used in the investigation, including any changes that were made to the method that was used during the Initial Investigation
- An analysis of the graphs and supporting calculations

Additional items to consider including in your report

- A comparison of your results with the results from other groups testing different substances
- Recommended modifications to the procedure that would increase accuracy, save time, or ensure that liquids are handled more efficiently and safely
- Recommended methods and necessary materials to construct a prototype hand warmer

The Determination of an Equilibrium Constant

The equilibrium state of a chemical reaction can be characterized by quantitatively defining its equilibrium constant, K_{eq} . In this experiment, you will determine the value of K_{eq} for the reaction between iron (III) ions and thiocyanate ions, SCN^- .



When you mix amounts of Fe^{3+} and SCN^- , a reaction occurs to produce FeSCN^{2+} , but not all of the reactants react. Thus, your beaker (or flask or cauldron) will contain some of each of these three species, which is your equilibrium system. To learn more about the system, we need to figure out a way to count the number of different ions in the reaction mixture. That is the major objective of this experiment, and to achieve this objective you will take advantage of something about FeSCN^{2+} — in aqueous solution it has a reddish color. The two reactants, Fe^{3+} and SCN^- , are essentially colorless in solution, thus the red color you will see when you conduct the reaction is produced by the FeSCN^{2+} ions.

One of the more important numbers that help us understand an equilibrium system is called the equilibrium constant, K_{eq} . For the reaction between Fe^{3+} and SCN^- , the K_{eq} is defined by the equation

$$K_{eq} = \frac{[\text{FeSCN}^{2+}]}{[\text{Fe}^{3+}][\text{SCN}^-]}$$

To find the value of K_{eq} at a given temperature, it is necessary to determine the molar concentration of each of the three species in solution at equilibrium. You will determine the concentrations by using a Spectrometer to measure the amount of light of a specific wavelength that passes through a sample of the equilibrium mixtures. The amount of light absorbed by a colored solution is proportional to its concentration. The red FeSCN^{2+} solution absorbs blue light, so you will determine an appropriate wavelength based on the absorbance spectrum of the solution. The wavelength will be close to, but not exactly, 470 nm.

In order to successfully evaluate this equilibrium system, it is necessary to conduct two separate tests. In Part I of the experiment, you will prepare a series of standard solutions of FeSCN^{2+} from solutions of varying concentrations of SCN^- and constant concentrations of H^+ and Fe^{3+} that are in stoichiometric excess. The excess of H^+ ions will ensure that Fe^{3+} engages in no side reactions (to form FeOH^{2+} , for example) which could interfere with your measurements. In an excess of Fe^{3+} ions, the SCN^- ions will be the limiting reagent, thus all of the SCN^- will form FeSCN^{2+} ions. The FeSCN^{2+} complex forms slowly, taking at least one minute for the color to develop. It is best to take absorbance readings after a specific length of time has passed, between two and four minutes after preparing the equilibrium mixture. Do not wait much longer than 5 minutes to take readings, however, because the mixture is light sensitive and the FeSCN^{2+} ions will slowly decompose.

The Determination of an Equilibrium Constant

In Part II of the experiment, you will prepare a new series of solutions that have varied concentrations of the SCN^- ions and constant concentrations of H^+ ions and Fe^{3+} ions. You will use the results of this test to accurately evaluate the equilibrium concentrations of each species and calculate the K_{eq} of the reaction.

OBJECTIVES

- Prepare and test standard solutions of FeSCN^{2+} in equilibrium.
- Determine the molar concentrations of the ions present in an equilibrium system.
- Determine the value of the equilibrium constant, K_{eq} , for the reaction.

MATERIALS

Chromebook, computer, **or** mobile device
Vernier Spectral Analysis app
Go Direct SpectroVis Plus
plastic cuvette
four 10 mL graduated cylinders
one 50 mL graduated cylinder
seven small beakers (100–250 mL)
several plastic Beral pipets
0.200 M iron (III), Fe^{3+} , solution in 1.0 M HNO_3
0.0020 M iron (III), Fe^{3+} , solution in 1.0 M HNO_3
0.0020 M thiocyanate, SCN^- , solution
distilled water
lint free tissues or Kim® Wipes

PRE-LAB EXERCISE

For the solutions that you will prepare in Step 2 of Part I below, calculate the $[\text{FeSCN}^{2+}]$. Presume that all of the SCN^- ions react. In Part I of the experiment, mol of $\text{SCN}^- = \text{mol}$ of FeSCN^{2+} . Record these values in the following table:

Beaker number	$[\text{FeSCN}^{2+}]$
1	
2	
3	
4	

PROCEDURE

Part I Prepare and Test Standard Solutions

- Obtain and wear goggles.
- Label four small beakers 1–4. Obtain small volumes of 0.200 M $\text{Fe}(\text{NO}_3)_3$, 0.0020 M SCN^- , and distilled water. Prepare four solutions according to the chart below. Use graduated cylinders to measure the solutions. Mix each solution thoroughly. Record the temperature of one of the solutions as the temperature for the equilibrium constant, K_{eq} . **DANGER:** Iron (III) nitrate solution, $\text{Fe}(\text{NO}_3)_3 \cdot 9\text{H}_2\text{O}$: Causes skin and eye irritation. Do not breathe mist, vapors, or spray. **WARNING:** Potassium thiocyanate solution, KSCN : Causes eye irritation and mild skin irritation.

Important: The mixtures you will prepare are light sensitive. You need to measure the absorbance of these four mixtures within 2–5 minutes of preparing them.

Beaker	0.200 M $\text{Fe}(\text{NO}_3)_3$ (mL)	0.0020 M SCN^- (mL)	H_2O (mL)
1	5.0	4.0	41.0
2	5.0	3.0	42.0
3	5.0	2.0	43.0
4	5.0	1.0	44.0

- Prepare a *blank* by filling a cuvette 3/4 full with 0.200 M $\text{Fe}(\text{NO}_3)_3$. To correctly use cuvettes, remember:
 - Wipe the outside of each cuvette with a lint-free tissue.
 - Handle cuvettes only by the top edge of the ribbed sides.
 - Dislodge any bubbles by gently tapping the cuvette on a hard surface.
 - Always position the cuvette so the light passes through the clear sides.
- Launch Spectral Analysis. Connect the Go Direct SpectroVis Plus to your Chromebook, computer, or mobile device. Click or tap Absorbance vs. Concentration.
- To calibrate the Spectrometer, place the blank cuvette in the Spectrometer and select Finish Calibration. **Note:** If necessary, wait for the Spectrometer to warm up before selecting Finish Calibration.
- Determine the optimum wavelength for the equilibrium mixture.
 - Empty the water from the blank cuvette. Using the solution in Beaker 1, rinse the cuvette twice with ~1 mL amounts and then fill it 3/4 full. Wipe the outside with a tissue and place the cuvette in the Spectrometer.
 - The live graph will update with the spectrum of the sample. Click or tap the desired wavelength or enter the Wavelength. Click or tap Done. **Note:** The λ max should be 400–480 nm.

The Determination of an Equilibrium Constant

7. Collect absorbance-concentration data for the four standard equilibrium mixtures.
 - a. Leave the cuvette, containing the Beaker 1 mixture, in the device.
 - b. Click or tap Collect to start data collection. After the absorbance reading stabilizes, click or tap Keep and enter the value for the concentration of FeSCN^{2+} from your Pre-Lab calculations. Click or tap Keep Point. The absorbance and concentration values have now been saved for the first solution.
 - c. Discard the cuvette contents as directed. Using the solution in Beaker 2, rinse the cuvette twice with ~1 mL amounts, and then fill it 3/4 full. Place the cuvette in the device, wait for the value displayed on the screen to stabilize, and click or tap Keep. Enter the value for the concentration of FeSCN^{2+} in Beaker 2, then click or tap Keep Point.
 - d. Repeat Part c of this step to measure the absorbance of the solutions in Beakers 3 and 4. **Note:** Wait until Step 10 to test the unknown.
 - e. When you have finished testing the standard solutions, click or tap Stop to stop data collection.
8. Display a graph of absorbance vs. concentration with a linear regression curve.
 - a. Click or tap Graph Tools, , and choose Edit Graph Options.
 - b. Enter 0 for both the x-axis Left value and the y-axis the Bottom value.
 - c. Dismiss the Graph Options box. Your graph will now include the origin (0,0).
 - d. Click or tap Graph Tools, , and choose Apply Curve Fit.
 - e. Select Linear as the curve fit. Click or tap Apply. The graph should indicate a direct relationship between absorbance and concentration, a relationship known as Beer's law. Record the linear fit equation in your data table. **Important:** Do not remove the curve fit. You will use the best-fit line equation in Part II.

Part II Prepare and Test Equilibrium Systems

9. Label three new small beakers A–C. Prepare the solutions according to the chart below. Use 10.0 mL graduated cylinders to measure the solutions. Mix each solution thoroughly. **Note:** You are using 0.0020 M $\text{Fe}(\text{NO}_3)_3$ in this test. **WARNING:** Iron (III) nitrate solution, $\text{Fe}(\text{NO}_3)_3 \cdot 9\text{H}_2\text{O}$: Causes skin and eye irritation. Do not breathe mist, vapors, or spray.

Beaker	0.0020 M $\text{Fe}(\text{NO}_3)_3$ (mL)	0.0020 M SCN^- (mL)	H_2O (mL)
A	3.00	3.00	4.00
B	3.00	4.00	3.00
C	3.00	5.00	2.00

10. Collect absorbance-concentration data for the three beakers of equilibrium mixtures.
 - a. Using the solution in Beaker A, rinse the cuvette twice with ~1 mL amounts and then fill it 3/4 full. Wipe the outside with a tissue and place the cuvette in the device. Write down, in your data table, the absorbance of the sample in Beaker A. Click or tap Graph Tools, , and enable Interpolate to determine the concentration of the sample. Dismiss the Graph Tools box and click or tap the graph to interpolate.
 - b. Click or tap the linear regression line to find the FeSCN^{2+} concentration for the sample in Beaker A. Write down the concentration in your data table.
 - c. Discard the cuvette contents as directed. Rinse and fill the cuvette with the solution in Beaker B and place it in the device. After the reading stabilizes, write down the absorbance in your data table. Click or tap the linear regression line to find the FeSCN^{2+} concentration for the sample in Beaker A. Write down the concentration in your data table.
 - d. Repeat Step c for the mixtures in Beaker C.

DATA TABLE

Part I

Temperature: _____ °C

Beaker	$[\text{FeSCN}^{2+}]$	Absorbance
1		
2		
3		
4		

Linear regression equation	
----------------------------	--

Part II

Beaker	Absorbance	$[\text{FeSCN}^{2+}]$ at equilibrium
A		
B		
C		

Calculating Equilibrium Concentrations

A common method that is used to organize and calculate the concentrations of the species in an equilibrium system is colloquially known as an I.C.E. chart. “I.C.E.” stands for *Initial*

The Determination of an Equilibrium Constant

concentration, Change in concentration, and the Equilibrium concentration. The initial concentrations of the Fe^{3+} and the SCN^- ions can be calculated from the mixing chart in Part II, Step 10. You have already determined the equilibrium concentration of the FeSCN^{2+} ions by completing the analysis in Part II. The rest is a little bit of math.

BEAKER A

	Fe^{3+}	SCN^-	FeSCN^{2+}
Initial			0.00
Change			
Equilibrium			

BEAKER B

	Fe^{3+}	SCN^-	FeSCN^{2+}
Initial			0.00
Change			
Equilibrium			

BEAKER C

	Fe^{3+}	SCN^-	FeSCN^{2+}
Initial			0.00
Change			
Equilibrium			

DATA ANALYSIS

- (Part II) Use your data to determine the $[\text{Fe}^{3+}]$, $[\text{SCN}^-]$, and $[\text{FeSCN}^{2+}]$ at equilibrium for each of the mixtures that you prepared in Part II. Complete the table below and give an example of your calculations.

	A	B	C
$[\text{FeSCN}^{2+}]$			
$[\text{Fe}^{3+}]$			
$[\text{SCN}^-]$			

- Calculate the value of K_{eq} for the reaction. Explain how you used the data to calculate K_{eq} .

Investigating the Kinetics of a Crystal Violet Reaction

Spectroscopy can be used to determine the concentration of a substance in an aqueous solution. Spectroscopy can also be used to measure the progress of a chemical reaction. In this investigation, the objective is to use spectroscopy to learn more about a chemical reaction. Further, the results of your investigation should provide sufficient evidence to describe the kinetics of a reaction and write the *rate law*.

The rate law of a chemical reaction describes, in equation form, the effect each reactant has on the progress of the reaction. For a generic reaction equation, $A + B \rightarrow C + D$, the rate law is written as $\text{rate} = k[A]^x[B]^y$, where the exponents x and y refer to the order of the reaction with respect to the given reactant. The order of a reactant is established by determining the effect of concentration of a reactant on the rate at which the reaction proceeds. Commonly, order is determined to be zero, first, or second, and expressed in the rate law as an exponent of 0, 1, or 2.

When a solution of the organic indicator crystal violet reacts with a solution of sodium hydroxide, one of the results of the reaction is a color change. The purple color of the crystal violet solution disappears as the reaction proceeds to conclusion. How and why a chemical reaction proceeds in a certain manner is part of the study of chemical reactions called *kinetics*.

In the Initial Investigation, you will use a Vernier SpectroVis Plus Spectrophotometer to measure the absorbance of a series of standard solutions of crystal violet. With that information, you will devise a plan for conducting the reaction between solutions of crystal violet and sodium hydroxide, from which the rate law, with respect to crystal violet, can be written.

PRE-LAB ACTIVITY

Consider the graph of the absorbance spectrum of a 2.5×10^{-5} M crystal violet solution shown in Figure 1.

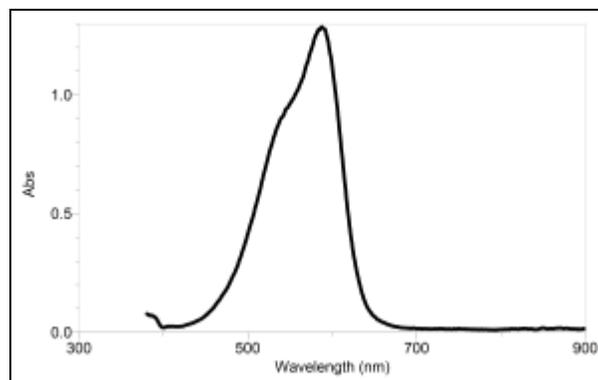


Figure 1

Investigating the Kinetics of a Crystal Violet Reaction Experiment

1. What is the best (optimal) wavelength, or wavelength range, at which to measure the absorbance of a set of five crystal violet solutions, if the maximum concentration is $2.5 \times 10^{-5} \text{ M}$?
2. In the Initial Investigation, you will prepare five solutions of crystal violet and measure the absorbance of the solutions at the wavelength you selected in Question 1. You will need 10 mL of each solution. Complete the Table 1 to help you plan for the Initial Investigation.

Concentration (mol/L)	Vol. $2.5 \times 10^{-5} \text{ M}$ crystal violet (mL)	Vol. distilled water (mL)
2.5×10^{-5}	10.0	0.0

3. Conduct research to find and write the balanced equation for the reaction between aqueous solutions of crystal violet and sodium hydroxide.
4. Conduct research to find and write the rate law for the reaction between aqueous solutions of crystal violet and sodium hydroxide. Explain all the components of the rate law.

INITIAL INVESTIGATION

In the Initial Investigation, you will measure the absorbance of a set of crystal violet solutions at a given wavelength.

1. Obtain and wear goggles.
2. To prepare the samples, obtain small volumes of $2.5 \times 10^{-5} \text{ M}$ crystal violet solution $2.5 \times 10^{-5} \text{ M}$ and distilled water. Prepare a set of serial dilutions of the stock crystal violet solution according to Table 1 that you prepared in Pre-Lab Question 2.
WARNING: *Aqueous crystal violet: May be harmful if swallowed. May cause skin irritation and eye damage.*
3. Prepare a *blank* by filling a cuvette 3/4 full with distilled water. To correctly use cuvettes, remember
 - Wipe the outside of each cuvette with a lint-free tissue.
 - Handle cuvettes only by the top edge of the ribbed sides.
 - Dislodge any bubbles by gently tapping the cuvette on a hard surface.
 - Always position the cuvette so the light passes through the clear sides.

4. Set up the spectrophotometer and data-collection software.
 - a. Launch Spectral Analysis.
 - b. Connect the Go Direct SpectroVis Plus Spectrophotometer to your Chromebook, computer, or mobile device.
 - c. Click or tap Absorbance vs. Concentration.
5. Calibrate the spectrophotometer.
 - a. Place the blank cuvette in the spectrophotometer.
 - b. Select Finish Calibration. **Note:** If necessary, wait for the spectrophotometer to warm up before selecting Finish Calibration.
6. Determine the optimal wavelength for crystal violet.
 - a. Remove the blank cuvette from the spectrophotometer. Refill the cuvette with the stock solution of crystal violet (solution of greatest molar concentration).
 - b. The live graph will update with the spectrum of the sample. Click or tap the desired wavelength or enter the Wavelength. Click or tap Done.
7. Collect absorbance-concentration data for the crystal violet standard solutions.
 - a. With the cuvette of stock crystal violet solution still in the spectrophotometer, start data collection.
 - b. After the value displayed on the screen has stabilized, select Keep and enter the concentration in mol/L. Select Keep Point. The absorbance and concentration values have now been stored for the first crystal violet solution.
 - c. Remove the cuvette, pour out the stock solution, and rinse the cuvette. Refill the cuvette with a new crystal violet standard solution. Place the cuvette in the spectrophotometer.
 - d. After the value displayed on the screen has stabilized, select Keep and enter the concentration in mol/L. Select Keep Point.
 - e. Repeat the necessary steps to store concentration values for the remaining standard solutions.
 - f. After the final standard solution has been measured, stop data collection.

PLANNING FOR THE FINAL INVESTIGATION

Based on the results of the Initial Investigation, develop a method for measuring the change in the concentration of a crystal violet solution as it reacts with sodium hydroxide. The objective is to collect data that is sufficiently valid and reliable to help describe and quantify the kinetics of the reaction. Keep in mind that the color of the solution will slowly fade as the reaction proceeds. Consider the following issues as you develop your plan:

- What is the linear fit (best fit line) equation for the set of crystal violet standards measured in the Initial Investigation? What does the slope of this line tell you about the standard solutions?
- What is the relationship between the molar concentration of a crystal violet solution and its absorbance of a specific wavelength of visible light?

Investigating the Kinetics of a Crystal Violet Reaction Experiment

- How does the testing of a series of standard solutions help describe what will happen during a reaction where the color of a solution is disappearing?
- How quickly, or slowly, should a reaction proceed in order to achieve the best readings of its change in color?
- In the Initial Investigation, the molar concentration of the NaOH solution was orders of magnitude more concentrated than the crystal violet solution. How does this affect the tests you need to run to determine the order of the reaction with respect to NaOH?

Your plan should provide sufficient data to write the rate law for the reaction between aqueous solutions of crystal violet and sodium hydroxide. **WARNING:** *Sodium hydroxide solution, NaOH: Causes skin and eye irritation.*

FINAL INVESTIGATION

As you carry out your approved plan, consider the following points

- You should collect data in absorbance vs. time mode. To change modes, click or tap File, , and choose New Experiment. Click or tap Absorbance vs. Time.
- The default parameters for time-based data collection are a duration of 200 seconds and a rate of 2 seconds per sample. It is wise to lengthen the duration to ensure that all the necessary absorbance readings are collected.
- Confirm that you are using the correct wavelength for your Final Investigation. If you need to change the wavelength, click or tap the Absorbance meter. The live graph will update with the spectrum of the sample. Click or tap the desired wavelength or enter a value for the Wavelength. Click or tap Done.

ANALYZING RESULTS

When preparing your report, include

- A statement of the results: What is the rate law for the reaction between solutions of crystal violet and sodium hydroxide?
- A description of the testing method used to determine the order of the reaction with respect to crystal violet
- A description of the method used to determine the order of the reaction with respect to sodium hydroxide
- An analysis of the data and calculations leading to the determination of the value of the rate law constant, k

Additional items to consider including in your report

- A comparison of your results with those of other groups
- Recommended modifications to the procedure that would increase accuracy, save time, or ensure that liquids are handled more efficiently and safely

Investigating Acid-Base Titrations

(Drop Counter version)

A titration is a laboratory process used to determine the volume of a solution needed to react with a given amount of another solution. One of the most common titrations performed in a Chemistry lab is an acid-base titration. In the Initial Investigation, you will be assigned an acid solution to titrate with a solution of the strong base sodium hydroxide, NaOH. The concentration of the NaOH solution is given and you will determine the concentration of the acid solution.

Your assigned acid may be strong, such as hydrochloric acid, HCl, or weak, such as acetic acid, CH₃COOH. When titrating a strong acid, the main objective is to determine the *equivalence point* of the titration. Near the equivalence point, the pH increases very rapidly, as shown in Figure 1. The graph of a strong acid-strong base titration is analyzed as accurately as possible to determine the exact amount of NaOH needed to neutralize all the acid,

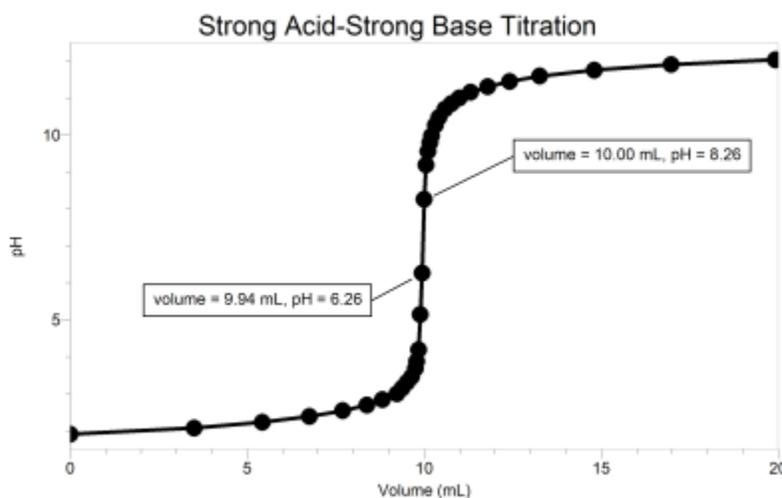


Figure 1

In the case of titrating a weak acid, a second objective is to determine an *equilibrium constant*. In reacting with a base, a weak acid will establish an equilibrium that can be evaluated mathematically, and is abbreviated as K_a . A K_a value is a unique quantity, which helps identify the acid. A weak acid may also have more than one K_a , depending on how many dissociation steps it undergoes in releasing ionizable hydrogen to form H_3O^+ ions.

After completing the Initial Investigation, you will be given a weak acid solution of unknown identity and concentration to test by titration. It will be helpful to use reference sources to find out more about acids and bases, acid-base titrations, and weak acid equilibria before planning and conducting your investigation.

PRE-LAB ACTIVITY

1. If you are given equal volumes of 0.25 M hydrochloric acid solution and 0.25 M acetic acid solution, will one acid require more of a sodium hydroxide solution, at a specific molar concentration, to be neutralized? Explain your answer.
2. In the scenario posed in #1 above, will the pH values at the equivalence points of the titrations be the same? Explain your answer.
3. If you are given a structural drawing, or 3-D model, of a compound and told that it is an acid, what features of the structure could you use to determine if the acid is strong or weak?

INITIAL INVESTIGATION

In the Initial Investigation, you will conduct a titration between a solution of sodium hydroxide, NaOH, and an acid solution to which your group has been assigned. Your instructor will provide you with the molar concentration of the NaOH solution; the molar concentration of the acid will be unknown.

1. Obtain and wear goggles.
2. Obtain about 25 mL of the acid solution assigned to your group. Measure out 10 mL of the acid solution and 50 mL of distilled water to a 250 mL beaker. **DANGER:** *Unknown acid solution: Causes severe skin and eye burns and damage. Do not breathe mist, vapors, or spray. Harmful inhaled. Do not eat or drink when using this product.*
3. Launch Graphical Analysis. Connect the pH Sensor and Drop Counter to your Chromebook, computer, or mobile device.
4. Lower the Drop Counter onto a Stir Station.
5. Obtain the plastic 60 mL reagent reservoir. Close both valves by turning the handles to a horizontal position. Follow the steps below to set up the reagent reservoir for the titration.
 - a. Rinse the reagent reservoir with a few mL of the 0.10 M NaOH solution and pour the rinse NaOH into an empty 250 mL beaker.
 - b. Use a utility clamp to attach the reservoir to the Stir Station.
 - c. Fill the reagent reservoir with slightly more than 60 mL of the 0.10 M NaOH solution.
 - d. Place the 250 mL beaker, which contains the rinse NaOH, beneath the tip of the reservoir.
 - e. Drain a small amount of the NaOH solution into the 250 mL beaker so that it fills the reservoir's tip. To do this, turn both valve handles to the vertical position for a moment, then turn them both back to horizontal.
 - f. Discard the drained NaOH solution in the 250 mL beaker as directed.



Figure 2

6. Calibrate the Drop Counter so that a precise volume of titrant is recorded in units of milliliters.
 - a. Click or tap the Volume meter and choose Calibrate.
 - If you have previously calibrated the drop size of your reagent reservoir and want to continue with the same drop size, select Manual. Enter the drops per mL from your previous calibration. Select Apply and proceed directly to Step 7.
 - If you want to perform a new calibration, continue with the this step.
 - b. Place a 10 mL graduated cylinder directly below the slot on the Drop Counter, lining it up with the tip of the reagent reservoir.
 - c. Open the bottom valve on the reagent reservoir (vertical). Keep the top valve closed (horizontal).
 - d. Slowly open the top valve of the reagent reservoir so that drops are released at a slow rate (~1 drop every two seconds). You should see the drops being counted on the screen.
 - e. When the volume of NaOH solution in the graduated cylinder is between 9 and 10 mL, close the bottom valve of the reagent reservoir.
 - f. Enter the precise volume of NaOH in the graduated cylinder.
 - g. Record the number of drops per mL for future titrations using this reservoir.
 - h. Click or tap Apply.
 - i. Discard the NaOH solution in the graduated cylinder as indicated by your instructor and set the graduated cylinder aside.
7. Assemble the apparatus.
 - a. Insert the pH Sensor through the large hole in the Drop Counter.
 - b. Adjust the positions of the Drop Counter and reagent reservoir so they are both lined up with the center of the Stir Station.
 - c. Lift up the pH Sensor, and slide the beaker containing the acetic acid solution onto the Stir Station. Lower the pH Sensor into the beaker.
 - d. Place a magnetic stirring bar in the beaker, and adjust the position of the pH Sensor so the probe will not be struck when the stirring bar is stirring.
 - e. Adjust the reagent reservoir so its tip is just above the Drop Counter slot.
 - f. Turn the Stir Station on so that the stirring bar is stirring at a fast rate.
8. You are now ready to perform the titration.
 - a. Start data collection. No data will be collected until the first drop goes through the Drop Counter slot.
 - b. Fully open the bottom valve. The top valve should still be adjusted so drops are released at a rate of about 1 drop every 2 seconds. When the first drop passes through the Drop Counter slot, check the graph to see that the first data pair was recorded.
 - c. Watch your graph to see when a large increase in pH takes place—this is the equivalence point of the reaction. When this jump in pH occurs, allow about 3 more milliliters of NaOH solution to be added.
 - d. Close the bottom valve of the reagent reservoir.
 - e. Stop data collection to view a graph of pH vs. volume.
9. Dispose of the reaction mixture as directed by your instructor.

PLANNING FOR THE FINAL INVESTIGATION

As you analyze and discuss the data from your Initial Investigation, consider the following points:

- At the *equivalence point*, equimolar amounts of the acid and NaOH have been mixed. What is the most accurate way to determine the equivalence point?
- Calculate the number of moles of NaOH used to neutralize the acid.
- Calculate the molar concentration of the acid solutions.

In the Final Investigation, you will be given a new, unknown, weak acid solution. In planning to test the unknown acid solution, consider the following:

- What is the best method for calculating the K_a of a weak acid, using titration data?
- If an unknown acid has more than one equivalence point, is there reason to expect a specific K_a (K_{a1} or K_{a2} , or possibly K_{a3}) to be a better value to identify the acid?
- How important is the concentration of the NaOH solution to achieving good results? Is there reason to use a solution of NaOH that is similar to that of your unknown weak acid?
- Can suitably accurate results be achieved by titrating the unknown weak acid solution with a weak base solution, such as ammonium hydroxide?

FINAL INVESTIGATION

As you carry out your approved plan, consider the following questions:

- What is the optimum number of data-collection runs needed to achieve the best data?
- What is the best volume of the unknown to titrate to achieve the best data?
- What sampling and measuring techniques should be used to achieve the best data?

ANALYZING RESULTS

When preparing your report, include

- A statement of the results: What was the identity and the molar concentration of the unknown weak acid assigned to your group?
- A description of the procedure that you used in the investigation, including any changes that were made to the method that was used during the Initial Investigation
- An analysis of the graphs and supporting calculations, including any K_a values

Additional items to consider including in your report

- A comparison of your results with the results from other groups testing different weak acids
- Recommended modifications to the procedure that would increase accuracy, save time, or ensure that liquids are handled more efficiently and safely
- A discussion of any unanticipated difficulties your group faced in testing the unknown and/or determining its identity and molar concentration