

GENERAL CHEMISTRY II LAB MANUAL



Vincent Hradil
Triton College

General Chemistry II CHM 141 Lab Manual

Vincent P. Hradil, PhD

Triton College

This text is disseminated via the Open Education Resource (OER) LibreTexts Project (<https://LibreTexts.org>) and like the thousands of other texts available within this powerful platform, it is freely available for reading, printing, and "consuming."

The LibreTexts mission is to bring together students, faculty, and scholars in a collaborative effort to provide an accessible, and comprehensive platform that empowers our community to develop, curate, adapt, and adopt openly licensed resources and technologies; through these efforts we can reduce the financial burden borne from traditional educational resource costs, ensuring education is more accessible for students and communities worldwide.

Most, but not all, pages in the library have licenses that may allow individuals to make changes, save, and print this book. Carefully consult the applicable license(s) before pursuing such effects. Instructors can adopt existing LibreTexts texts or Remix them to quickly build course-specific resources to meet the needs of their students. Unlike traditional textbooks, LibreTexts' web based origins allow powerful integration of advanced features and new technologies to support learning.



LibreTexts is the adaptable, user-friendly non-profit open education resource platform that educators trust for creating, customizing, and sharing accessible, interactive textbooks, adaptive homework, and ancillary materials. We collaborate with individuals and organizations to champion open education initiatives, support institutional publishing programs, drive curriculum development projects, and more.

The LibreTexts libraries are Powered by [NICE CXone Expert](#) and were supported by the Department of Education Open Textbook Pilot Project, the California Education Learning Lab, the UC Davis Office of the Provost, the UC Davis Library, the California State University Affordable Learning Solutions Program, and Merlot. This material is based upon work supported by the National Science Foundation under Grant No. 1246120, 1525057, and 1413739.

Any opinions, findings, and conclusions or recommendations expressed in this material are those of the author(s) and do not necessarily reflect the views of the National Science Foundation nor the US Department of Education.

Have questions or comments? For information about adoptions or adaptations contact info@LibreTexts.org or visit our main website at <https://LibreTexts.org>.

The LibreTexts name, logo, and book covers are trademarked by LibreTexts, Inc. (a 501(c)(3) not-for-profit organization) and protected against unauthorized use.

This text was compiled on 10/15/2026

TABLE OF CONTENTS

Licensing

0: Lab Safety Orientation

- 0.3: Lab Safety Quiz
- 0.4: Additional Resources
- 0.1: Lab Safety Videos
- 0.2: Lab Safety Contract

1: Beer's Law

- 1.1: Beer's Law - Experiment
- 1.2: Beer's Law - Pre-lab
- 1.3: Beer's Law - Data and Report

2: Determination of an Equilibrium Constant

- 2.1: Determination of an Equilibrium Constant - Experiment
- 2.2: Determination of an Equilibrium Constant - Pre-lab
- 2.3: Determination of an Equilibrium Constant - Data and Report

3: Acid Dissociation Constant

- 3.1: Acid Dissociation Constant - Experiment
- 3.2: Acid Dissociation Constant - Pre-lab
- 3.3: Acid Dissociation Constant - Data and Report

4: Titration of a Diprotic Acid

- 4.1: Titration of a Diprotic Acid - Experiment
- 4.2: Titration of a Diprotic Acid - Pre-lab
- 4.3: Titration of a Diprotic Acid - Data and Report

5: Investigating Buffers

- 5.1: Investigating Buffers - Experiment
- 5.2: Investigating Buffers - Pre-lab
- 5.3: Investigating Buffers - Data and Report

6: Solubility Product

- 6.1: Solubility Product - Experiment
- 6.2: Solubility Product - Pre-lab
- 6.3: Solubility Product - Data and Report

7: Thermodynamics of Borax

- 7.1: Thermodynamics of Borax - Experiment
- 7.2: Thermodynamics of Borax - Pre-lab
- 7.3: Thermodynamics of Borax - Data and Report

8: Redox Reactions

8.1: Redox Reactions - Experiment

8.2: Redox Reactions - Pre-lab

8.3: Redox Reactions - Data and Report

9: Voltaic Cells

9.1: Voltaic Cells - Experiment

9.2: Voltaic Cells - Pre-lab

9.3: Voltaic Cells - Data and Report

10: Decomposition of Hydrogen Peroxide

10.1: Decomposition of Hydrogen Peroxide - Experiment

10.2: Decomposition of Hydrogen Peroxide - Pre-lab

10.3: Decomposition of Hydrogen Peroxide - Data and Report

11: Determination of an Activation Energy

11.1: Determination of an Activation Energy - Experiment

11.2: Determination of an Activation Energy - Pre-lab

11.3: Determination of an Activation Energy - Data and Report

12: Radioactivity Simulation

12.1: Radioactivity Simulation - Experiment

12.2: Radioactivity Simulation - Pre-lab

12.3: Radioactivity Simulation - Data and Report

13: Appendices

13.1: Vernier LabQuest 3 Instructions

13.2: Tables

13.2.1: The Periodic Table

13.2.2: Fundamental Physical Constants

13.2.3: Water Properties

13.2.4: Composition of Commercial Acids and Bases

13.2.5: Standard Thermodynamic Properties for Selected Substances

13.2.6: Ionization Constants of Weak Acids

13.2.7: Ionization Constants of Weak Bases

13.2.8: Solubility Products

13.2.9: Formation Constants for Complex Ions

13.2.10: Standard Electrode (Half-Cell) Potentials

13.2.11: Half-Lives for Several Radioactive Isotopes

Index

Glossary

Detailed Licensing

Detailed Licensing

Licensing

A detailed breakdown of this resource's licensing can be found in **Back Matter/Detailed Licensing**.

This page titled [Licensing](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

0: LAB SAFETY ORIENTATION

0.3: Lab Safety Quiz

0.4: Additional Resources

0.1: Lab Safety Videos

0.2: Lab Safety Contract



Figure 0.1: The four principles of safety (Copyright; [ACS](#))

0.3: Lab Safety Quiz

Multiple-Choice Questions

1. **What should you do if a chemical splashes into your eyes?**
 - A. Rub your eyes vigorously
 - B. Rinse your eyes with water for at least 15 minutes
 - C. Close your eyes and wait for the pain to subside
 - D. Apply a cold compress
2. **Which of the following is NOT appropriate lab attire?**
 - A. Closed-toe shoes
 - B. Long pants
 - C. Loose, flowing clothing
 - D. Safety goggles
3. **What is the first thing you should do if you spill a chemical on your skin?**
 - A. Wipe it off with a cloth
 - B. Rinse the affected area with water immediately
 - C. Apply a bandage
 - D. Ignore it if it doesn't hurt
4. **Where should you dispose of broken glassware?**
 - A. Regular trash bin
 - B. Recycling bin
 - C. Designated glass disposal container
 - D. Sink
5. **Which of the following statements about lab safety is TRUE?**
 - A. Eating and drinking are allowed in the lab
 - B. You should always work alone in the lab
 - C. Always read labels and safety data sheets before using chemicals
 - D. It's okay to leave a Bunsen burner unattended

True/False Questions

6. **You should always wear safety goggles when handling chemicals.**
 - True
 - False
6. **It's safe to taste chemicals in the lab to identify them.**
 - True
 - False
6. **You should know the location of safety equipment such as fire extinguishers, eyewash stations, and safety showers.**
 - True
 - False
6. **Long hair should be tied back when working in the lab.**
 - True
 - False
6. **It's acceptable to use your phone in the lab as long as you're careful.**
 - True

- False

0.4: Additional Resources

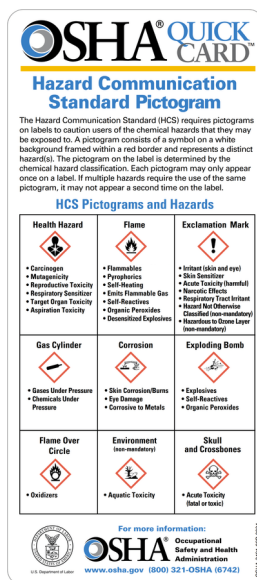


Figure 0.4.1: OSHA Hazard Communication Standard Pictogram QuickCard

0.1: Lab Safety Videos

Safety Orientation Videos

Students will find below the eight Lab safety orientation videos. The first six are part of the College Lab Safety Series by the American Chemical Society. Students should watch the first seven videos before signing the safety contract and attempting the safety quiz. The eighth video is an older version of the College Lab Safety video by the American Chemical Society, which should be watched, time permitting, as it provides detailed dos and don'ts in a college lab setting.

1. ACS (American Chemical Society) College Lab Safety Series

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

[View this YouTube Video online](#)



YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

[View this YouTube Video online](#)



YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

[View this YouTube Video online](#)



YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

[View this YouTube Video online](#)



YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

[View this YouTube Video online](#)



YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

[View this YouTube Video online](#)



2. Lab Safety and Techniques- Crash course safety

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

[View this YouTube Video online](#)



3. Safety Video by American Chemical Society (1991)

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

[View this YouTube Video online](#)



License

Creative Commons Attribution license (reuse allowed)

This page titled [0.1: Lab Safety Videos](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

0.2: Lab Safety Contract

Triton College Chemistry Division

Student Safety Contract

At Triton College, safety is our number one priority. A list of rules has been developed and provided to you in this student safety contract to ensure a safe science laboratory experience. These rules must always be followed. Please be advised that you will be asked to leave the laboratory if you violate these rules. **No exceptions!!**

Download, sign, and date both contract pages before you participate in the laboratory sessions. Your signature indicates that you have read this contract and agree to abide by the rules. You will keep a copy for your records and upload or submit one to your instructor via your Blackboard or Learning Management System (LMS) as directed.

I, _____, dated _____ agree that I will comply with the following safety rules. I will:

1. Always wear safety goggles in the laboratory.
2. Always wear appropriate clothing in the laboratory. This includes long pants, shirts with sleeves, and closed shoes (i.e., no sandals). Also, I understand that long hair must be tied back, and loose or baggy clothing or jewelry must be secured.

**Note: Students violating rules one and/or two will be asked to leave the lab immediately. No second chances!! NO EXCEPTIONS!!*

3. Read laboratory instructions beforehand and ask the instructor about any part of the lab I am unsure about.
4. Use proper equipment for laboratory procedures.
5. Never perform unauthorized experiments or work alone in the laboratory.
6. Never remove chemicals or other materials/equipment from the laboratory area.
7. Read the labels on reagent bottles carefully, remove only small amounts of reagent with the proper tools, and never return unused chemicals to the bottle.
8. Dispose of broken glass and waste chemicals in the appropriate waste container.
9. Immediately inform the instructor of a chemical spill or accident (even minor ones) in the laboratory.
10. Never taste or smell any chemicals.
11. Know the location of eyewash fountains, fire extinguishers, fire blankets, safety showers, first aid kits, and exits.
12. Do not bring food or drink into the lab; refrain from eating or drinking in the laboratory. This includes chewing gum and eating candy.
13. Keep aisles clear. Books, purses, backpacks, and other personal belongings will be kept outside or under the lab benches.
14. I will responsibly conduct myself at all times in the laboratory. Horseplay, practical jokes, and pranks are dangerous and prohibited.
15. Clean my work area before leaving the lab.
16. Wash my hands before I leave the lab.

Agreement

I, _____ have read and agree to follow all the safety rules outlined in this contract. I realize I must obey these rules to ensure my safety and that of my fellow students. I will fully cooperate with my instructor and fellow students to maintain a safe lab environment. I will also closely follow all oral and written instructions provided by the instructor, the science department, and the college. I am aware that any violation of this safety contract, resulting in unsafe conditions in the laboratory or misbehavior on my part, may lead to my removal from the laboratory, a failing grade, and/or dismissal from the course.

Student Signature: _____

Date: _____

Course section and number: _____

Questions

1. Are you color blind? Yes / No
2. Do you have any allergies? Yes / No
If yes, please list specific allergies: _____
3. Do you wear contact lenses? Yes / No
4. If you are pregnant, breastfeeding, or trying to conceive, avoid working with heavy metal compounds and organic solvents. Notify your instructor and request a list of chemicals used in this course. Use this list to consult with your physician before proceeding with the course. Under your physician's supervision, you may need to purchase a respirator, which a doctor must fit and adjust to your needs.
5. Please inform your instructor of any other medical conditions or situations you feel we should be aware of to ensure your safety and the safety of others while working in the laboratory.

Please be aware that anything written in this contract will remain confidential.

1: BEER'S LAW

PURPOSE

- To determine the concentration of an unknown copper(II) sulfate solution.
- To understand and apply Beer's Law, which relates the absorbance of a solution to its concentration.
- To prepare and measure the absorbance of standard copper(II) sulfate solutions to construct a standard calibration curve.
- To use the standard curve—or its linear regression equation—to calculate the molar concentration of the unknown solution.

INTRODUCTION

The goal of this experiment is to determine the concentration of an unknown copper(II) sulfate solution by measuring how strongly it absorbs light. In colored solutions, absorbance increases as concentration increases; a more concentrated solution absorbs more light and transmits less than a more dilute one. This predictable behavior provides the basis for quantitative analysis using visible spectroscopy.

To obtain accurate measurements, it is first necessary to identify the wavelength at which the solution absorbs most strongly. This wavelength, known as λ_{max} , provides the greatest sensitivity because small changes in concentration produce measurable changes in absorbance.

Beer's Law describes the relationship between absorbance and concentration. Once a set of standard solutions with known concentrations is prepared, their absorbance values are measured using a spectrophotometer. Each absorbance reflects the amount of light that passes through the solution and reaches a photocell inside the instrument.

Plotting absorbance versus concentration for these standard solutions produces a linear calibration curve, commonly called a Beer's Law plot or standard curve. The slope of this line captures the proportionality between absorbance and concentration and can be used to determine the concentration of an unknown solution. A strong linear fit—indicated by a coefficient of determination (R^2) close to 1.000 and a y-intercept (b) near zero—demonstrates reliable adherence to Beer's Law.

In this experiment, you will:

- Prepare a series of standard copper(II) sulfate solutions.
- Measure their absorbance values at λ_{max} .
- Construct a standard calibration curve using your spectrophotometer data.
- Measure the absorbance of an unknown copper(II) sulfate solution and determine its concentration using the standard curve or regression equation.

DATA PREP: THE "BEST FIT" LINE

In this lab, you will calculate the concentration of an unknown. To do this accurately, your calibration curve must be linear.

- **The Diagnostic:** We use the **Coefficient of Determination (R^2)** to grade the straightness of the line.
- **The Goal:** An R^2 of 1.000 is perfect. In this lab, if your $R^2 < 0.98$, it implies your pipetting technique was inconsistent, and your calculated unknown concentration will likely be wrong.

KEY EQUATIONS

Beer's Law:

$$A = \varepsilon bc \quad \text{or} \quad A = kc \quad (\text{for constant } \varepsilon \text{ and path length } b) \quad (1.1)$$

Dilution Equation:

$$M_1 V_1 = M_2 V_2 \quad (1.2)$$

[1.1: Beer's Law - Experiment](#)

[1.2: Beer's Law - Pre-lab](#)

[1.3: Beer's Law - Data and Report](#)

1.1: Beer's Law - Experiment

SAFETY PRECAUTIONS

- Chemical splash goggles must be worn at all times in the laboratory.
- Copper(II) sulfate solution is a localized irritant. Avoid contact with skin and eyes. Clean up spills immediately.

EQUIPMENT AND CHEMICALS NEEDED

Table 1.1.1: Materials Needed

Equipment	Chemicals
UV-Vis spectrophotometer (SpectroVis Plus) & LabQuest interface	0.10 M CuSO ₄ stock solution
Two optical cuvettes & lint-free wipes (Kimwipes)	Unknown CuSO ₄ solution
Five 16 × 150 mm test tubes & test tube rack	Deionized water
10-mL graduated volumetric pipette & pipette pump/bulb	Aqueous Heavy Metal Waste container
10-mL volumetric flask & two 50-mL beakers	

EXPERIMENTAL PROCEDURE

Part A: Prepare Standard Solutions

1. Obtain approximately 30 mL of 0.10 M CuSO₄ stock solution in a clean 50-mL beaker.
2. Label five clean test tubes 1–5 and place them in a test tube rack.
3. Prepare five standard solutions using a 10-mL graduated pipette and a 10-mL volumetric flask according to Table 1.1.2 :
 - a. Pipette the specified volume of 0.10 M CuSO₄ stock solution directly into the 10-mL volumetric flask.
 - b. Dilute to the mark with deionized water. Stopper and invert the volumetric flask several times to mix thoroughly.
 - c. Transfer the prepared solution into the corresponding labeled test tube.
 - d. Rinse the volumetric flask with small portions of deionized water before preparing the next standard.

Table 1.1.2: Standard Solution Dilution Series

Test Tube	0.10 M CuSO ₄ (mL)	Deionized Water (mL)	Final CuSO ₄ Concentration (M)
1	2.0	8.0	0.020
2	4.0	6.0	0.040
3	6.0	4.0	0.060
4	8.0	2.0	0.080
5	10.0	0.0	0.100

Part B: Calibrate Spectrophotometer

4. Prepare a blank cuvette by filling it $\frac{3}{4}$ full with deionized water. Handle cuvettes strictly by the ribbed sides, wipe the optical clear sides clean with a lint-free tissue, and ensure no air bubbles adhere to the walls.
5. Insert the blank cuvette into the spectrometer slot and perform the baseline calibration:

VERNIER TECH BOX 1: CALIBRATION

1. Connect & Warm Up

- Connect the **SpectroVis Plus** spectrometer to the USB port on the **LabQuest 3**.
- Launch the LabQuest App. Allow the light source to warm up for 2–3 minutes.

2. Calibration (The "Blank")

- Insert the blank cuvette (deionized water) into the sample slot. **Align the clear sides with the light path arrow.**
- Tap the red **Absorbance** meter box on the screen → Select **Calibrate**.
- Select **Finish Calibration** → Tap **OK** once calibration completes.

Part C: Determine Optimal Wavelength ($\lambda_{\max}$)

6. Remove the blank cuvette from the sample slot and empty its contents.
7. Rinse the cuvette twice with small portions of the 0.10 M standard solution (Test Tube 5), then fill it $\frac{3}{4}$ full.
8. Wipe the clear sides with a lint-free tissue and place the cuvette into the spectrometer.
9. Initiate data collection to display the full absorbance spectrum. Identify and record the wavelength of maximum absorbance ($\lambda_{\max}$).
10. Export or store the spectrum to attach to your final laboratory report.

Part D: Collect Absorbance Data for Standards

11. Configure the spectrometer to record absorbance at your determined $\lambda_{\max}$:

VERNIER TECH BOX 2: BEER'S LAW MODE

Setup for Events with Entry

- Tap **Mode: Full Spectrum** → Change mode to **Events with Entry**.
- Set **Name:** Concentration
- Set **Units:** M
- Select **OK**. The instrument will measure Absorbance automatically while prompting you to manually input Concentration values.

12. Start data collection.
13. For each standard solution (Test Tubes 1–5):
 - a. Rinse the cuvette twice with the solution to be tested, then fill it $\frac{3}{4}$ full.
 - b. Wipe the optical sides clean and place the cuvette into the spectrometer.
 - c. Allow the absorbance reading to stabilize, record the value, and enter the corresponding concentration.

DATA SKILL: VISUAL REALITY CHECK

Evaluate your raw data before constructing your graph:

- **The Trend:** Absorbance must increase proportionally with concentration (e.g., doubling concentration from 0.020 M to 0.040 M should roughly double the absorbance).

- **The Check:** If Test Tube 3 (0.060 M) displays a lower absorbance than Test Tube 2 (0.040 M), an operational error has occurred (e.g., dirty optical surface or unmixed solution).
- **The Action:** Re-prepare or re-measure that specific standard immediately before applying linear regression.

Part E: Analyze Standard Data (Beer's Law Curve)

14. Display the data plot with Absorbance on the y-axis and Concentration (M) on the x-axis.
15. Apply a linear regression model ($y = mx + b$):
 - a. Record the slope (m), y-intercept (b), and correlation coefficient (r).
 - b. Verify that the regression line passes close to the origin (0, 0) and accurately models all five standard data points.

DATA SKILL: INTERPRETING REGRESSION STATISTICS

Physical interpretation of linear regression parameters ($y = mx + b$):

- **Slope (m):** Represents analytical sensitivity ($\epsilon \times b$). Higher slope values indicate greater absorbance response per unit concentration.
- **Intercept (b):** Theoretically equals zero (zero concentration = zero light absorbed).
 - If $b > 0.05$: Fingerprints or dirty cuvette walls added stray light scattering.
 - If $b < -0.05$: Baseline calibration was performed incorrectly.
- **Correlation (r):** An $r < 0.98$ reflects poor dilution technique or inconsistent cuvette orientation.

16. Record slope (m), y-intercept (b), and correlation coefficient (r) in your report data table.

Part F: Determine Unknown Concentration

17. Obtain approximately 5 mL of the unknown CuSO_4 solution in a clean test tube and record the Unknown ID.
18. Rinse the cuvette twice with the unknown solution, fill it $\frac{3}{4}$ full, wipe the optical walls, and insert it into the spectrometer.
19. Record the stabilized absorbance value in your data table.
20. Calculate the exact concentration of the unknown using your regression equation:

DATA SKILL: CALCULATING UNKNOWN CONCENTRATION

Rearrange your linear regression equation ($y = mx + b$) to solve for Concentration (x):

$$x = \frac{y - b}{m} \quad (1.1.1)$$

Substitute your measured unknown absorbance (y), slope (m), and intercept (b) into the equation to calculate the molarity.

21. Record the calculated unknown concentration in your Data Table.

CHEMICAL DISPOSAL

- Dispose of all copper(II) sulfate solutions in the designated **Aqueous Heavy Metal Waste** container. Do not pour heavy metal solutions down the drain.

1.2: Beer's Law - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. A student pipettes 3.00 mL of a 0.150 M CuSO_4 stock solution and dilutes it to a final volume of 50.0 mL in a volumetric flask. Calculate the final concentration (M) of the diluted solution.

2. State Beer's Law in equation form and define each variable in the equation. Explain how Beer's Law relates solution absorbance to concentration.

3. Why is it necessary to calibrate the spectrometer with a blank cuvette (deionized water) prior to recording absorbance values for the standard solutions?

4. Explain step-by-step how the linear regression equation ($y = mx + b$) obtained from a Beer's Law plot is used to determine the exact concentration of an unknown CuSO_4 solution.

5. **Data Logic Check:** A student records the following absorbance data for their calibration standards:

- 0.020 M: 0.150
- 0.040 M: 0.305
- 0.060 M: 0.280 (!)

Without performing calculations, identify the outlier data point. What physical handling or experimental error likely caused the absorbance to drop at higher concentration?

This page titled [1.2: Beer's Law - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

1.3: Beer's Law - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Unknown ID: _____

Wavelength of Maximum Absorbance ($\lambda_{\max}$): _____ nm

Table 1.3.1: Standard Solutions Absorbance Data

Test Tube	Concentration (M)	Absorbance
1	0.020	_____
2	0.040	_____
3	0.060	_____
4	0.080	_____
5	0.100	_____

Table 1.3.2: Beer's Law Regression Parameters and Unknown Analysis

Parameter	Value
Slope, m (M^{-1})	_____
Y-intercept, b	_____
Correlation coefficient, r	_____
Unknown Absorbance	_____
Calculated Unknown Concentration (M)	_____

Attach an image of the full absorbance spectrum for the standard solution:

Attach an image of the Beer's Law Plot (A vs. Concentration):

POST-LAB QUESTIONS

1. Explain why it is critical to measure the absorbance of standard and unknown solutions at λ_{max} rather than at an arbitrary wavelength. What physical advantage does measuring at λ_{max} provide when determining unknown concentration?

2. Evaluating regression statistics:

A student obtains a low correlation coefficient ($r = 0.85$) or a large y-intercept ($b = 0.10$) on their Beer's Law calibration plot.

DATA ANALYSIS HINT

Consider the physical source of error:

- A low **r value** indicates random error (e.g., inconsistent pipetting volume).
- A high **y-intercept (b)** indicates systematic error (e.g., dirty blank, fingerprint on cuvette, or uncalibrated instrument).

Explain specifically what experimental mistakes would cause each of these statistical outcomes.

3. Select two cuvette handling guidelines (wiping with Kimwipes, handling by ribbed sides only, dislodging bubbles, or orienting clear sides along light path) and detail how ignoring them would alter your absorbance readings and calculated unknown concentration.

This page titled [1.3: Beer's Law - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

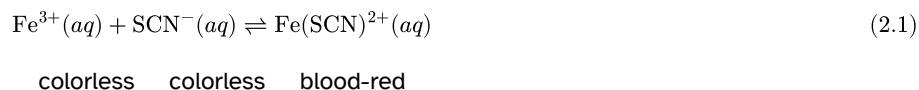
2: DETERMINATION OF AN EQUILIBRIUM CONSTANT

PURPOSE

- To apply Le Châtelier's Principle to shift equilibrium conditions and drive the formation of $\text{Fe}(\text{SCN})^{2+}$.
- To use spectrophotometry and Beer's Law to determine the concentration of $\text{Fe}(\text{SCN})^{2+}$ and calculate the proportionality constant k .
- To construct and analyze a calibration curve by plotting absorbance versus concentration and interpret the slope for quantitative analysis.
- To calculate the equilibrium concentrations of Fe^{3+} , SCN^- , and $\text{Fe}(\text{SCN})^{2+}$ in order to determine the equilibrium constant K_c .

INTRODUCTION

This experiment allows you to determine the equilibrium constant for the formation of the iron(III) thiocyanate complex ion, $\text{Fe}(\text{SCN})^{2+}$, and to apply Le Châtelier's Principle to control the position of equilibrium. When Fe^{3+} and SCN^- are mixed in aqueous solution, they form the blood-red complex ion $\text{Fe}(\text{SCN})^{2+}$:



The complex ion is highly colored and easily detected by spectrophotometry. Because the reactants are essentially colorless, absorbance measurements directly correspond to the concentration of $\text{Fe}(\text{SCN})^{2+}$ without interference.

The equilibrium constant for the reaction is given by:

$$K_c = \frac{[\text{Fe}(\text{SCN})^{2+}]_{\text{eq}}}{[\text{Fe}^{3+}]_{\text{eq}} [\text{SCN}^-]_{\text{eq}}} \quad (2.2)$$

To determine K_c , the equilibrium concentrations must be known. Beer's Law provides a direct method for determining $\text{Fe}(\text{SCN})^{2+}$ concentration from absorbance measurements:

$$A = kc \quad (2.3)$$

However, the proportionality constant k must first be determined using solutions of *known* $\text{Fe}(\text{SCN})^{2+}$ concentration. This is complicated by the fact that the complex ion participates in the equilibrium itself. Stoichiometric mixing does not guarantee a known concentration of product.

This difficulty is resolved by Le Châtelier's Principle. By adding a large excess of Fe^{3+} , the equilibrium shifts so far to the right that essentially all SCN^- is converted into $\text{Fe}(\text{SCN})^{2+}$. Under these conditions, the concentration of the complex ion is effectively equal to the initial concentration of the limiting reagent, SCN^- .

In the experiment, you will prepare solutions containing known amounts of $\text{Fe}(\text{SCN})^{2+}$ and measure their absorbances at approximately 450 nm, the wavelength of maximum absorbance (λ_{max}). Plotting absorbance versus concentration produces a straight line whose slope is the Beer's Law constant k .

After establishing k , you will measure the absorbance of equilibrium mixtures prepared with various starting concentrations of Fe^{3+} and SCN^- . From these results, equilibrium concentrations will be determined and used to calculate the equilibrium constant K_c for each trial.

DATA PREP: PRECISION VS. ACCURACY

You will calculate K_c five times. You might expect to get the exact same number every time, but experimental noise prevents this.

The Coefficient of Variation (CV): To determine if your equilibrium constant is truly "constant," we calculate the CV (Standard Deviation divided by the Average). A low CV (below 5%) proves that your technique was precise and the reaction follows the equilibrium model well.

KEY EQUATIONS

Beer's Law:

$$A = \epsilon bc \quad \text{or} \quad A = kc \quad (\text{for constant } \epsilon \text{ and path length } b) \quad (2.4)$$

Dilution Equation:

$$M_1 V_1 = M_2 V_2 \quad (2.5)$$

[2.1: Determination of an Equilibrium Constant - Experiment](#)

[2.2: Determination of an Equilibrium Constant - Pre-lab](#)

[2.3: Determination of an Equilibrium Constant - Data and Report](#)

2.1: Determination of an Equilibrium Constant - Experiment

SAFETY PRECAUTIONS

- Chemical splash goggles must be worn at all times in the laboratory.
- Iron(III) nitrate and potassium thiocyanate solutions are mild irritants. Avoid contact with skin and eyes. Wash hands thoroughly after handling.

EQUIPMENT AND CHEMICALS NEEDED

Table 2.1.1: Materials Needed

Equipment	Chemicals
UV-Vis spectrophotometer (SpectroVis Plus) & LabQuest interface	0.200 M $\text{Fe}(\text{NO}_3)_3$ solution (in 0.1 M HNO_3)
Optical cuvettes & lint-free wipes (Kimwipes)	0.00020 M KSCN solution
Ten large test tubes (20 × 150 mm) & test tube rack	0.0020 M $\text{Fe}(\text{NO}_3)_3$ solution
pipets (2 × 10-mL graduated and 1 × 5-mL volumetric) & pipet pump/bulb	0.0020 M KSCN solution
Glass stirring rod & small beakers	Deionized water & Aqueous Inorganic Waste container

EXPERIMENTAL PROCEDURE

Part A: Determine Beer's Law Constant

- Obtain five large test tubes. Wash, rinse, and dry the test tubes, then label them 1 through 5.
- Prepare the five reference solution test tubes listed in Table 2.1.2. Use a separate graduated pipet (± 0.02 mL) for each reagent to transfer the exact volumes.
- Mix each solution thoroughly using a glass stirring rod. Rinse and dry the stirring rod between solutions.

Table 2.1.2: Solution volumes for Beer's Law determination

Test Tube	0.200 M $\text{Fe}(\text{NO}_3)_3$ (mL)	0.00020 M KSCN (mL)
1	4.0	6.0
2	5.0	5.0
3	6.0	4.0
4	7.0	3.0
5	8.0	2.0

VERNIER TECH BOX: SPECTROMETER SETUP

1. Calibration (The Blank)

- Fill a cuvette $\frac{3}{4}$ full with deionized water.
- Tap the red **Absorbance** readout → **Calibrate** → **Finish Calibration**.
- Verify that the baseline reading drops to 0.0000 Abs.

2. Determining $\lambda_{\max}$

- Place **Test Tube 1** (the darkest reference solution) in the sample holder.
- Tap **Collect**. When the full spectrum appears, tap **Stop**.
- Tap the peak of the spectrum curve (near 450 nm) and record this wavelength as $\lambda_{\max}$.

3. Measuring Your Samples

- Tap the **Meter Tab** (gauge icon).
- Change the mode to **Events with Entry** to stabilize values during recording:
 - Tap **Mode: Full Spectrum** → Change mode to **Events with Entry**.
 - Set **Name:** Concentration
 - Set **Units:** M
 - Select **OK**. The instrument will measure Absorbance automatically while prompting you to manually input Concentration values.

4. Use deionized water as the blank to calibrate the spectrophotometer.
5. Obtain a full spectrum of the solution in Test Tube 1.
6. Find the wavelength at the highest point of the peak near 450 nm. Record this value as your $\lambda_{\max}$. Use this exact wavelength for all subsequent absorbance measurements.
7. Save your spectrum to include in your final lab report.
8. Measure and record the absorbance of each standard solution (Test Tubes 1–5) at $\lambda_{\max}$.
9. Display the data plot with Absorbance on the y-axis and Concentration (M) on the x-axis.
10. Apply a linear regression model ($y = mx + b$):
 - a. Record the slope (m), y-intercept (b), and correlation coefficient (r).
 - b. Verify that the regression line passes close to the origin (0, 0) and accurately models all five standard data points.

Part B: Determine the Equilibrium Constant

9. Wash and dry five additional test tubes. Label them 6 through 10.
10. Read reagent labels carefully before use! Prepare solutions according to Table 2.1.3. Use a 5-mL volumetric pipet for the iron (III) nitrate solution and separate graduated pipets for the potassium thiocyanate and deionized water.
11. Mix each solution thoroughly using a clean glass stirring rod. Rinse and dry the stirring rod between solutions.

Table 2.1.3: Solution volumes for Equilibrium Constant determination

Test Tube	0.0020 M Fe(NO ₃) ₃ (mL)	0.0020 M KSCN (mL)	Deionized Water (mL)
6	5.0	1.0	4.0
7	5.0	2.0	3.0
8	5.0	3.0	2.0
9	5.0	4.0	1.0
10	5.0	5.0	0.0

12. Measure and record the absorbances for Test Tubes 6–10 at your determined λ_{max} .

DATA SKILL: REALITY CHECK

Review your data table before leaving the instrument:

- **The Chemical Trend:** In Test Tubes 6 through 10, you add progressively more SCN^- reactant. According to Le Châtelier's Principle, this must shift the equilibrium to the right, producing more red complex.
- **The Check:** Absorbance values **must increase** sequentially from Test Tube 6 to Test Tube 10.
- **The Action:** If Test Tube 8 displays lower absorbance than Test Tube 7, a pipetting error occurred. Remake and re-measure that specific solution immediately.

CHEMICAL DISPOSAL

- Collect all waste containing thiocyanate complexes in the designated **Aqueous Inorganic Waste** container. Do not pour heavy metal or thiocyanate solutions down the drain.

2.2: Determination of an Equilibrium Constant - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. "The equilibrium concentration of $\text{Fe}(\text{SCN})^{2+}$ ions in each reference solution is essentially equal to the concentration of SCN^- ions in solution before any reaction occurs." Use Le Châtelier's Principle to explain why this statement is true.

2. The five reference solutions in Part A are prepared by mixing 0.200 M $\text{Fe}(\text{NO}_3)_3$ solution and 0.00020 M KSCN solution in the amounts listed in Part A Table 2.2.2.

The concentration of SCN^- ions in the first reference solution (M_2) before any reaction occurs is calculated using the dilution equation:

$$M_1V_1 = M_2V_2 \quad (2.2.1)$$

$$(0.00020 \text{ M})(2.0 \text{ mL}) = M_2(10.0 \text{ mL}) \rightarrow M_2 = 4.0 \times 10^{-5} \text{ M} \quad (2.2.2)$$

Use the dilution equation to calculate the initial concentration of SCN^- ions in all five reference solutions before reaction occurs. These represent the $[\text{Fe}(\text{SCN})^{2+}]$ calibration values for Test Tubes 1–5. Enter these values below and copy them into the Part A Data Table.

Table 2.2.1: Reference Solution Dilutions for Beer's Law

Test Tube	0.200 M Fe^{3+} (mL)	0.00020 M SCN^- (mL)	$[\text{SCN}^-]$ after dilution (M)
1	4.0	6.0	_____
2	5.0	5.0	_____
3	6.0	4.0	_____
4	7.0	3.0	_____
5	8.0	2.0	_____

Show a sample calculation here:

3. Calculate the initial concentrations of ions ($[\text{Fe}^{3+}]_0$ and $[\text{SCN}^-]_0$) in Test Tubes 6–10 after dilution but before equilibration. Enter the results below and copy them into the Part B Data Table. (Total volume $V_2 = 10.0 \text{ mL}$).

Table 2.2.2: Initial Ion Concentrations for Part B Test Solutions

Test Tube	$[\text{Fe}^{3+}]_0 \text{ (M)}$	$[\text{SCN}^-]_0 \text{ (M)}$
6	_____	_____
7	_____	_____
8	_____	_____
9	_____	_____
10	_____	_____

Show a sample calculation here:

4. **Trend Prediction:** In Part B, Test Tube 10 contains five times as much SCN^- reactant as Test Tube 6.
- According to Le Châtelier's Principle, should the solution in Tube 10 be darker or lighter red than Tube 6?
 - If you measure the absorbance of Tube 10 and find it is *lower* than Tube 6, what physical error occurred during solution preparation?

2.3: Determination of an Equilibrium Constant - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A: Determining Beer's Law Constant

Wavelength of Maximum Absorbance (λ_{max}): _____ nm

Table 2.3.1: Standard Concentrations and Absorbances

Test Tube	$[\text{Fe}(\text{SCN})^{2+}] \text{ (M)}$	Absorbance
1	_____	_____
2	_____	_____
3	_____	_____
4	_____	_____
5	_____	_____

Part B: Determining the Equilibrium Constant

Table 2.3.2: Initial Ion Concentrations and Absorbances for Part B

Test Tube	$[\text{Fe}^{3+}]_0 \text{ (M)}$	$[\text{SCN}^-]_0 \text{ (M)}$	Absorbance
6	_____	_____	_____
7	_____	_____	_____
8	_____	_____	_____
9	_____	_____	_____
10	_____	_____	_____

POST-LAB QUESTIONS

- Plot absorbance (A) vs. $[\text{Fe}(\text{SCN})^{2+}]$ for Test Tubes 1–5. Determine the slope of the line (k , the Beer's Law constant). Attach an image of your Beer's Law Plot here:

2. For Test Tubes 6–10, calculate equilibrium concentrations and K_c values to populate the Results Table:

- $[\text{Fe}(\text{SCN})^{2+}]_{\text{eq}}$ from Beer's Law: $[\text{Fe}(\text{SCN})^{2+}]_{\text{eq}} = \frac{A}{k}$
- $[\text{Fe}^{3+}]_{\text{eq}} = [\text{Fe}^{3+}]_0 - [\text{Fe}(\text{SCN})^{2+}]_{\text{eq}}$
- $[\text{SCN}^-]_{\text{eq}} = [\text{SCN}^-]_0 - [\text{Fe}(\text{SCN})^{2+}]_{\text{eq}}$
- $K_c = \frac{[\text{Fe}(\text{SCN})^{2+}]_{\text{eq}}}{[\text{Fe}^{3+}]_{\text{eq}} [\text{SCN}^-]_{\text{eq}}}$

DATA SKILL: LOGICAL BOUNDARIES

Apply these sanity checks when evaluating equilibrium concentrations:

- **Ceiling:** $[\text{Fe}^{3+}]_{\text{eq}}$ must be strictly less than $[\text{Fe}^{3+}]_0$.
- **Floor:** Concentrations cannot be negative. A negative value indicates an erroneous Beer's Law constant or inaccurate absorbance reading.

Show sample calculations for Test Tube 6 here:

3. Calculate the average K_c , standard deviation, and Coefficient of Variation (CV) to complete Table 2.3.3:

Table 2.3.3: Equilibrium Ion Concentrations and Calculated Constants

Test Tube	$[\text{Fe}(\text{SCN})^{2+}]_{\text{eq}} \text{ (M)}$	$[\text{Fe}^{3+}]_{\text{eq}} \text{ (M)}$	$[\text{SCN}^-]_{\text{eq}} \text{ (M)}$	K_c
6	_____	_____	_____	_____
7	_____	_____	_____	_____
8	_____	_____	_____	_____
9	_____	_____	_____	_____
10	_____	_____	_____	_____
Average K_c :				_____
Standard Deviation of K_c :				_____

4. Evaluate whether your experimental results indicate that K_c is truly "constant."

DATA SKILL: EVALUATING CONSTANCY

Calculate your Coefficient of Variation (CV) to support your claim:

$$CV = \left(\frac{\text{Standard Deviation}}{\text{Average}} \right) \times 100\% \quad (2.3.1)$$

- **CV < 5%:** Excellent precision. The system strictly follows the equilibrium model.
- **CV 5%–15%:** Acceptable experimental variance.
- **CV > 20%:** System drift or significant dilution/measurement error occurred.

Cite your calculated CV percentage to justify your conclusion:

5. Identify two major experimental sources of error and explain their specific effect on your calculated K_c values:

3: ACID DISSOCIATION CONSTANT

PURPOSE

- To titrate an unknown weak acid using a pH probe.
- To use the titration curve to determine the pK_a (and therefore the K_a) of the acid and identify candidate weak acids.
- To determine the molar concentration of the unknown weak acid solution from the titration equivalence point.

INTRODUCTION

This lab experiment focuses on titrating an unknown weak acid using a pH probe. The primary goal is to determine the acid dissociation constant (K_a) of the unknown weak acid.

The procedure involves titrating the acid with a standardized 0.10 M NaOH titrant solution while continuously monitoring pH with a pH sensor under magnetic stirring.

Data collection begins by recording the initial pH of a known volume of unknown weak acid diluted with deionized water. The NaOH titrant is added slowly while monitoring pH until the solution pH exceeds 11.

The K_a value is determined using the Henderson-Hasselbalch equation. Specifically, at the half-equivalence point of the titration, the pH equals the pK_a of the weak acid. From the pK_a , the K_a can be calculated directly. The titration curve is also analyzed to determine the exact concentration of the unknown acid from the volume of titrant required to reach the equivalence point.

Calculating the K_a

For a monoprotic weak acid, HA, the Henderson-Hasselbalch equation is:

$$pH = pK_a + \log \left(\frac{[A^-]}{[HA]} \right) \quad (3.1)$$

where $[HA]$ and $[A^-]$ represent the equilibrium molar concentrations of the weak acid and its conjugate base, respectively.

At the half-equivalence point, $V_{1/2} = \frac{1}{2}V_{eq}$:

$$[HA] = [A^-], \quad \text{so} \quad \frac{[A^-]}{[HA]} = 1, \quad \text{and} \quad \log(1) = 0 \quad (3.2)$$

Therefore, at the half-equivalence point:

$$pH = pK_a \quad (3.3)$$

The K_a value is then calculated from the definition of pK_a :

$$pK_a = -\log K_a \quad \longrightarrow \quad K_a = 10^{-pK_a} \quad (3.4)$$

DATA PREP: READING THE INFLECTION POINT

You will determine the equivalence point (V_{eq}) visually from the inflection point of your titration curve.

- **Steepness Matters:** The sharper the steep rise in pH, the more precisely you can identify V_{eq} .
- **Data Density:** Adding large increments of base (e.g., 1.0 mL) near the equivalence point will jump over the inflection point. You must add small dropwise increments near the equivalence point to locate the exact center of the steep region.

KEY EQUATIONS**Henderson-Hasselbalch Equation:**

$$\text{pH} = \text{p}K_{\text{a}} + \log \left(\frac{[\text{A}^{-}]}{[\text{HA}]} \right) \quad (3.5)$$

Definition of $\text{p}K_{\text{a}}$:

$$\text{p}K_{\text{a}} = -\log K_{\text{a}} \quad \text{or} \quad K_{\text{a}} = 10^{-\text{p}K_{\text{a}}} \quad (3.6)$$

[3.1: Acid Dissociation Constant - Experiment](#)

[3.2: Acid Dissociation Constant - Pre-lab](#)

[3.3: Acid Dissociation Constant - Data and Report](#)

3.1: Acid Dissociation Constant - Experiment

SAFETY PRECAUTIONS

- Chemical splash goggles must be worn at all times in the laboratory.
- Sodium hydroxide titrant and the unknown weak acid solutions are corrosive and caustic. Avoid skin and eye contact. Wash affected areas immediately with water.

EQUIPMENT AND CHEMICALS NEEDED

Table 3.1.1: Materials Needed

Equipment	Chemicals
Vernier pH Sensor & LabQuest interface	Unknown monoprotic weak acid solution
50-mL buret, buret clamp, & ring stand	0.10 M NaOH titrant solution
Magnetic stirrer, stir bar, & utility clamp (Note: a Vernier microstirrer and drop-counter/probe holder can also be used)	pH 4.00 & pH 7.00 buffer solutions
10.00-mL volumetric pipet & pipet pump/bulb	Deionized water
150-mL beaker, 250-mL beaker, & wash bottle	Aqueous Acid-Base Waste container

EXPERIMENTAL PROCEDURE

Part A: Set Up the Titration

VERNIER TECH BOX: pH TITRATION SETUP

1. Sensor Connections & pH Calibration (All Users)

- Connect the **pH Sensor** to Channel 1 of the LabQuest interface.
- Tap the red **pH** readout → Select **Calibrate** → Tap **Calibrate Now**.
- **Point 1:** Rinse probe, submerge in **pH 4.00 Buffer**. Enter "4.00" and tap **Keep**.
- **Point 2:** Rinse probe, submerge in **pH 7.00 Buffer**. Enter "7.00" and tap **Keep**.
- Tap **OK**.

2. Method A (No Drop Counter): Manual Titration

- Tap **Mode**. Select **Events with Entry**.
- Set **Name:** Volume | **Units:** mL.
- In this mode, add titrant, wait for pH to stabilize, tap **Keep**, and manually type in the total buret volume reading.

3. Method B (With Drop Counter): Automated Titration

- Connect the **Drop Counter** to DIG 1. Set **Mode** to **Drop Counting**.
- **CRITICAL STEP: Calibrate Drops**
 - Tap red **Drop Counting** box → **Calibrate** → **Calibrate Now**.
 - Place a 10-mL graduated cylinder below the drop tip.
 - Open the valve to establish steady drops until collecting 9–10 mL. Close the valve.
 - Enter the exact cylinder volume (e.g., 9.40 mL) and tap **Keep**. Record this calibration factor on your data sheet.

1. Measure 30 mL of deionized water into a clean 150-mL beaker. Use a 10.00-mL volumetric pipette and bulb/pump to transfer 10.00 mL of the unknown weak acid into the beaker.
2. Place the magnetic stirrer on the base of the ring stand. Carefully place a magnetic stir bar into the beaker and turn on the stirrer to maintain a gentle mixing vortex.
3. Rinse the pH sensor thoroughly with deionized water and place it into the acid solution. Ensure the spinning stir bar does not strike the glass bulb of the pH probe. Clamp the probe securely with a utility clamp.
4. Obtain approximately 100 mL of 0.10 M NaOH solution in a clean 250-mL beaker. Record the exact NaOH concentration in your data table.
5. Rinse and fill the 50-mL buret with NaOH solution. Clamp the buret above the reaction beaker.

Part B: Titrate the Unknown Weak Acid

6. Confirm the initial pH is acidic.
7. Titration Method A (No Drop Counter)
 - a. Record the initial buret volume reading (± 0.02 mL) and initial pH.
 - b. Add NaOH titrant in increments of no more than 0.50 mL, recording total buret volume and pH after each addition.
 - c. Continue adding NaOH titrant until pH reaches at least 11.0. **CRITICAL:** If you are not using a drop-counter, as you approach the steep equivalence region, reduce titrant additions to small dropwise increments (0.10–0.20 mL) to capture the inflection point accurately.
8. Titration Method B (With Drop Counter)
 - a. Press the green start triangle on the LabQuest. The system will wait until it sees the drops start.
 - b. Open the stopcock to start the LabQuest collecting data.
 - c. Monitor the data collection making sure that the drop rate is not too fast (about one drop every 1.5–2.0 seconds is best).
 - d. Stop the acquisition when the pH reaches at least 11.0.

Part C: Repeat the Titration

9. Perform a second titration trial using a fresh 10.00-mL sample of unknown weak acid following Steps 1–7/8. There is no need to re-calibrate the pH probe or drop counter.
10. Rinse the pH sensor thoroughly with deionized water and return it to its storage solution bottle.

CHEMICAL DISPOSAL

- Dispose of all titrated solutions and excess chemicals in the designated **Aqueous Acid-Base Waste** container. Neutralize residual acid or base prior to cleaning glassware.

3.2: Acid Dissociation Constant - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Write the chemical equation for the acid dissociation of a generic monoprotic weak acid, HA, in aqueous solution.

2. Write the equilibrium constant expression (K_a) for the acid dissociation of HA.

3. Which two chemical species have equal molar concentrations at the half-equivalence point of a weak acid–strong base titration?

4. A student titrates 20.00 mL of a 0.100 M weak acid solution with 0.100 M NaOH. Calculate the volume of NaOH required to reach the **half-equivalence point**.

5. **Data Strategy:** Why does the procedure instruct you to slow down titrant additions to dropwise increments (0.10–0.20 mL) near the equivalence point? Explain how this affects the precision of your determined V_{eq} .

This page titled [3.2: Acid Dissociation Constant - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

3.3: Acid Dissociation Constant - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Unknown Acid ID: _____

Exact NaOH Titrant Concentration: _____ M

Drop Counter Calibration Factor (if used): _____ mL/drop

Table 3.3.1: Titration Volume and pH Data. Record your data here if you are not using a drop counter.

Trial 1 Buret (mL)	Trial 1 pH	Trial 2 Buret (mL)	Trial 2 pH
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

POST-LAB QUESTIONS

- Construct a titration curve (pH vs. Volume of NaOH) for each trial using spreadsheet or graphing software. Attach your graphs here and comment on the key features (initial pH, buffer region, inflection steepness, and repeatability between trials):

- Determine V_{eq} , calculate $V_{1/2}$, estimate pK_a , and calculate K_a for each trial to populate Table 3.3.2:

Table 3.3.2: Titration Analysis and K_a Determination

Parameter	Trial 1	Trial 2
Volume at equivalence point, V_{eq} (mL)	_____	_____
Volume at half-equivalence point, $V_{1/2}$ (mL)	_____	_____
pK_a (= pH at $V_{1/2}$)	_____	_____
Acid dissociation constant, K_a	_____	_____
Calculated unknown acid concentration (M)	_____	_____

3. Calculate the molarity of the unknown weak acid solution for both trials and report the average value. Evaluate experimental precision using Percent Difference:

DATA SKILL: QUANTIFYING PRECISION (N=2)

Calculate the Percent Difference between your two trials:

$$\% \text{ Difference} = \frac{|\text{Trial 1} - \text{Trial 2}|}{\text{Average}} \times 100\% \quad (3.3.1)$$

- **Analysis:** A Percent Difference > 5% indicates inconsistent titrant delivery or difficulty locating the inflection point.

Show your calculations and discuss your precision below:

4. Propose a candidate identity for your unknown weak acid using a reference table of literature K_a values. Quantify accuracy using Percent Error:

DATA SKILL: QUANTIFYING ACCURACY

Compare your experimental average K_a to the literature value:

$$\% \text{ Error} = \frac{(\text{Experimental Average } K_a - \text{Literature } K_a)}{\text{Literature } K_a} \times 100\% \quad (3.3.2)$$

Note: Percent errors for K_a often range from 10–20% because K_a relates exponentially to pK_a (10^{-pK_a}). Small pH uncertainties propagate into larger K_a errors.

State your proposed acid identity, literature K_a , and calculated Percent Error below:

This page titled [3.3: Acid Dissociation Constant - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

4: TITRATION OF A DIPROTIC ACID

PRE-LAB PREPARATION

1. Prerequisite Math & Theory

- **Polyprotic Stoichiometry:** A diprotic acid (H_2A) releases two protons sequentially in separate dissociation steps:
 - **First Equivalence Point (V_{eq1}):** All H_2A is converted to HA^- .
 - **Second Equivalence Point (V_{eq2}):** All HA^- is converted to A^{2-} .
 - **The Rule:** It takes the same stoichiometric amount of base to remove the second proton as it does the first. Therefore, $V_{eq2} \approx 2 \times V_{eq1}$.
- **Finding pK_a Values:**
 - pK_{a1} is the pH at $0.5 \times V_{eq1}$.
 - pK_{a2} is the pH at $V_{eq1} + 0.5 \times (V_{eq2} - V_{eq1})$, which simplifies to $1.5 \times V_{eq1}$.

2. Required Technical Skills

- **pH Probe Calibration:** Calibrating a pH sensor using standard buffer solutions.
- **Titration Data Collection:** Recording precise incremental volume and pH readings using manual entry or an automated drop counter.
- **Finding Inflection Points:** Because the first inflection "step" on a diprotic curve can be subtle, locating the maximum slope ($\frac{\Delta pH}{\Delta V}$) yields the most accurate equivalence volume.

3. Critical Safety

- **Chemical Handling:** Sodium hydroxide and unknown acid solutions are corrosive. Wear chemical splash goggles at all times in the laboratory.

PURPOSE

- To titrate an unknown diprotic weak acid using a pH probe.
- To use the titration curve to determine pK_{a1} and pK_{a2} (and calculate K_{a1} and K_{a2}) to identify a candidate weak acid.
- To calculate the molar concentration of the unknown diprotic acid solution from the titration equivalence volumes.

INTRODUCTION

This laboratory experiment investigates the potentiometric titration of an unknown diprotic weak acid, represented as H_2A . Unlike monoprotic weak acids, a diprotic acid contains two acidic protons that dissociate in a stepwise manner, each defined by its own acid dissociation equilibrium and constant (K_{a1} and K_{a2}). The primary objective is to evaluate these stepwise dissociations and determine the quantitative constants defining the acid's chemical behavior.

A potentiometric titration is performed by continuously monitoring solution pH as a standardized strong base, 0.10 M NaOH, is added incrementally from a buret.

The collected titrant volume and pH data are plotted to produce a diprotic titration curve featuring two distinct inflection regions. From this curve, the first (V_{eq1}) and second (V_{eq2}) equivalence points are identified, corresponding to the complete neutralization of the first and second protons, respectively. The pK_{a1} and pK_{a2} values are estimated directly from the pH at the corresponding half-equivalence points. From these values, K_{a1} and K_{a2} are calculated. Finally, the molarity of the unknown acid is calculated from the equivalence points, and a candidate identity is proposed by comparing experimental pK_a values to a reference table of literature constants.

DATA PREP: THE "DOUBLE VOLUME" RULE

Diprotic acids feature two equivalence points. Because the stoichiometry is 1:1 for each proton removal step, a built-in mathematical relationship exists:

- The total volume of titrant required to reach the **second** equivalence point ($V_{\text{eq}2}$) must be approximately **twice** the volume required to reach the **first** equivalence point ($V_{\text{eq}1}$).
- **Data Check:** If your experimental $V_{\text{eq}2}$ is not approximately $2 \times V_{\text{eq}1}$, re-examine your titration curve to ensure an endpoint was not misidentified.

KEY EQUATIONS

Henderson-Hasselbalch Equation:

$$\text{pH} = \text{p}K_{\text{a}} + \log \left(\frac{[\text{Base}]}{[\text{Acid}]} \right) \quad (4.1)$$

Definition of $\text{p}K_{\text{a}}$:

$$\text{p}K_{\text{a}} = -\log K_{\text{a}} \quad \text{or} \quad K_{\text{a}} = 10^{-\text{p}K_{\text{a}}} \quad (4.2)$$

[4.1: Titration of a Diprotic Acid - Experiment](#)

[4.2: Titration of a Diprotic Acid - Pre-lab](#)

[4.3: Titration of a Diprotic Acid - Data and Report](#)

4.1: Titration of a Diprotic Acid - Experiment

SAFETY PRECAUTIONS

- Chemical splash goggles must be worn at all times in the laboratory.
- Sodium hydroxide titrant and unknown weak acid solutions are caustic and corrosive. Avoid contact with skin and eyes. Flush affected areas with water immediately.

EQUIPMENT AND CHEMICALS NEEDED

Table 4.1.1: Materials Needed

Equipment	Chemicals
Vernier pH Sensor & LabQuest interface	Unknown diprotic weak acid (H_2A) solution
50-mL buret, buret clamp, & ring stand	0.10 M NaOH titrant solution
Magnetic stirrer, stir bar, & utility clamp	pH 4.00 & pH 7.00 buffer solutions
20.00-mL volumetric pipette & pipette pump/bulb	Deionized water
150-mL beaker, 250-mL beaker, & wash bottle	Aqueous Acid-Base Waste container

EXPERIMENTAL PROCEDURE

Part A: Set Up the Titration

VERNIER TECH BOX: DIPROTIC TITRATION SETUP

1. Sensor Connections & pH Calibration (All Users)

- Connect the **pH Sensor** to CH 1 of the LabQuest interface.
- Tap the red **pH** readout → Select **Calibrate** → Tap **Calibrate Now**.
- **Point 1:** Rinse probe, submerge in **pH 4.00 Buffer**. Enter "4.00" and tap **Keep**.
- **Point 2:** Rinse probe, submerge in **pH 7.00 Buffer**. Enter "7.00" and tap **Keep**.
- Tap **OK**.

2. Method A (No Drop Counter): Manual Titration

- Tap **Mode**. Select **Events with Entry**.
- Set **Name:** Volume | **Units:** mL.
- In this mode, add titrant, wait for pH to stabilize, tap **Keep**, and manually type in the total buret volume reading.
- *Tip: Take small additions (0.10 mL) around the expected first equivalence point to capture the first inflection step clearly.*

3. Method B (With Drop Counter): Automated Titration

- Connect the **Drop Counter** to DIG 1. Set **Mode** to **Drop Counting**.

• CRITICAL STEP: Calibrate Drops

- Tap red **Drop Counting** box → **Calibrate** → **Calibrate Now**.
- Place a 10-mL graduated cylinder below the drop tip. Open valve to collect 9–10 mL. Close valve.
- Enter exact cylinder volume (e.g., 9.40 mL) and tap **Keep**. Record calibration factor on your data sheet.
- Open the stopcock valve to maintain a slow, steady drip (1–2 drops/sec). The system will plot the double-inflection curve automatically.

1. Measure 30 mL of deionized water into a clean 150-mL beaker. Use a 20.00-mL volumetric pipette and bulb/pump to transfer 20.00 mL of the unknown diprotic acid into the beaker.
2. Place the magnetic stirrer on the base of the ring stand. Carefully place a magnetic stir bar into the beaker and turn on the stirrer to maintain gentle mixing.
3. Rinse the pH sensor thoroughly with deionized water and place it into the acid solution. Ensure the spinning stir bar does not strike the glass bulb of the pH probe. Clamp the probe securely with a utility clamp.
4. Obtain approximately 100 mL of 0.10 M NaOH solution in a clean 250-mL beaker. Record the exact NaOH concentration in your data table.
5. Rinse and fill the 50-mL buret with NaOH solution. Clamp the buret above the reaction beaker.

Part B: Titrate the Unknown Diprotic Acid

6. Confirm the initial pH is acidic.
7. Method A (No Drop Counter)
 - a. Record the initial buret volume reading (± 0.02 mL) and initial pH.
 - b. Add NaOH titrant in increments of no more than 0.50 mL, recording total buret volume and pH after each addition.
 - c. Continue adding NaOH titrant until pH reaches at least 11.0. **CRITICAL:** As you approach each equivalence region, reduce titrant additions to small dropwise increments (0.10–0.20 mL) to capture both inflection points accurately.
8. Method B (With Drop Counter).
 - a. Press the green start triangle on the LabQuest. The system will wait until it sees the drops start.
 - b. Open the stopcock to start the LabQuest collecting data.
 - c. Monitor the data collection making sure that the drop rate is not too fast (about one drop every 1.5–2.0 seconds is best).
 - d. Stop the acquisition when the pH reaches at least 11.0.

DATA SKILL: RESOLUTION CHECK

Before emptying your beaker: Inspect your titration curve on the screen.

- **Can you observe two distinct inflection steps?**
- The first equivalence point of a diprotic acid is often subtle (a minor rise in slope) compared to the second steep cliff. If base was added too rapidly early on, the first step may be obscured, preventing accurate determination of pK_{a1} .
- **Decision Rule:** If the first step is unresolvable, repeat the titration using smaller, dropwise additions (0.10 mL) throughout the initial buffer and equivalence region.

Part C: Repeat the Titration

9. Perform a second titration trial using a fresh 20.00-mL sample of unknown diprotic acid following Steps 1–7/8.
10. Rinse the pH sensor thoroughly with deionized water and return it to its storage solution bottle.

CHEMICAL DISPOSAL

- Dispose of all titrated solutions and excess chemicals in the designated **Aqueous Acid-Base Waste** container. Clean all glassware before returning equipment.

This page titled [4.1: Titration of a Diprotic Acid - Experiment](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

4.2: Titration of a Diprotic Acid - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Write the two chemical equations for the stepwise dissociation of a generic diprotic weak acid, H_2A , in aqueous solution.

2. Write the equilibrium constant expressions for K_{a1} and K_{a2} corresponding to the acid dissociations in Question 1.

3. Which two chemical species have equal molar concentrations at the first half-equivalence point ($V_{1/2}^{(1)}$)? Which two chemical species have equal molar concentrations at the second half-equivalence point ($V_{1/2}^{(2)}$)?

4. **Prediction:** A student titrates 20.00 mL of a 0.100 M H_2A weak acid solution with 0.100 M NaOH.
- Calculate the volume of NaOH required to reach the **first** equivalence point (V_{eq1}).
 - Calculate the volume of NaOH required to reach the **second** equivalence point (V_{eq2}).
 - Sketch the expected shape of the titration curve, clearly labeling the half-equivalence points and equivalence points at their corresponding titrant volumes (10 mL, 20 mL, and 40 mL).

This page titled [4.2: Titration of a Diprotic Acid - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

4.3: Titration of a Diprotic Acid - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Unknown Diprotic Acid ID: _____

Exact NaOH Titrant Concentration: _____ M

Drop Counter Calibration Factor (if used): _____ mL/drop

Table 4.3.1: Titration Volume and pH Data. Record your data here if are NOT using a Drop Counter.

Trial 1 Buret (mL)	Trial 1 pH	Trial 2 Buret (mL)	Trial 2 pH
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

POST-LAB QUESTIONS

- Construct a titration curve (pH vs. Volume of NaOH) for each trial using spreadsheet or graphing software. Attach your graphs here and comment on the overall shape, presence of two distinct inflection steps, and repeatability between trials:

- Estimate V_{eq1} and V_{eq2} from your titration curves. Determine $V_{1/2}^{(1)}$ and $V_{1/2}^{(2)}$, estimate pK_{a1} and pK_{a2} , and calculate K_{a1} and K_{a2} for both trials to populate Table 4.3.2:

Table 4.3.2: Diprotic Titration Analysis Results

Parameter	Calculation Method	Trial 1	Trial 2
Volume at 1 st equivalence point, V_{eq1} (mL)	V_{eq1}	_____	_____
Volume at 1 st half-equivalence point, $V_{1/2}^{(1)}$ (mL)	$V_{1/2}^{(1)} = 0.5 \times V_{eq1}$	_____	_____
pK_{a1}	pH @ $V_{1/2}^{(1)}$	_____	_____
First acid dissociation constant, K_{a1}	$K_{a1} = 10^{-pK_{a1}}$	_____	_____
Volume at 2 nd equivalence point, V_{eq2} (mL)	V_{eq2}	_____	_____
Volume at 2 nd half-equivalence point, $V_{1/2}^{(2)}$ (mL)	$V_{1/2}^{(2)} = (V_{eq1} + V_{eq2}) / 2$	_____	_____
pK_{a2}	pH @ $V_{1/2}^{(2)}$	_____	_____
Second acid dissociation constant, K_{a2}	$K_{a2} = 10^{-pK_{a2}}$	_____	_____

3. Calculate the molar concentration of the unknown diprotic acid using both equivalence volumes (V_{eq1} and V_{eq2}) for both trials (yielding four concentration values). Report the individual molarities, calculated average molarity, and discuss precision:

DATA SKILL: INTERNAL CONSISTENCY

Calculate molarity using both endpoints:

- From V_{eq1} : $M_{acid} = \frac{M_{NaOH} \times V_{eq1}}{V_{acid}}$
- From V_{eq2} : $M_{acid} = \frac{M_{NaOH} \times V_{eq2}}{2 \times V_{acid}}$
- Analysis:** Evaluate the Percent Difference between concentrations derived from V_{eq1} versus V_{eq2} . If the difference exceeds 5%, explain which endpoint is more reliable (typically V_{eq2} due to a sharper slope).

Show your calculations and discuss internal consistency below:

4. Using your estimated pK_{a1} and pK_{a2} values, consult a reference table of acid ionization constants to propose a candidate identity for your unknown diprotic acid. Justify your choice:

This page titled [4.3: Titration of a Diprotic Acid - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

5: INVESTIGATING BUFFERS

PRE-LAB PREPARATION

1. Prerequisite Math & Theory

- Henderson-Hasselbalch Equation: You must be able to calculate the mass of sodium acetate needed to achieve a specific target pH:

$$\text{pH} = \text{p}K_{\text{a}} + \log \left(\frac{[\text{Conjugate Base}]}{[\text{Weak Acid}]} \right) \quad (5.1)$$

- Rearrangement Skill: Given a target pH and $\text{p}K_{\text{a}}$, solve for the ratio $\text{C}_2\text{H}_3\text{O}_2^- / \text{HC}_2\text{H}_3\text{O}_2$.
 - Mass Calculation: Calculate the required molarity of conjugate base, then use $\text{Mass (g)} = \text{Molarity (mol/L)} \times \text{Volume (L)} \times \text{Molar Mass (g/mol)}$ to determine grams of $\text{NaC}_2\text{H}_3\text{O}_2$.
- Buffer Capacity: A buffer prepared with 1.0 M components resists pH changes 10 times more effectively than a buffer prepared with 0.1 M components, even though both buffers possess the exact same initial pH.

2. Required Technical Skills

- pH Probe Calibration: Calibrating a pH sensor using standard buffer solutions.
- Data Logging (Events with Entry): Recording titrant volume added versus pH during buffer capacity testing.

3. Critical Safety

- Acids and Bases: Concentrated acetic acid, 0.5 M NaOH, and 0.5 M HCl are CAUSTIC and CORROSIVE. Wear chemical splash goggles at all times in the laboratory.

PURPOSE

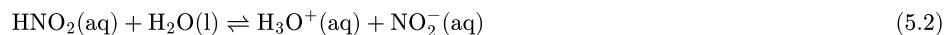
- To use the Henderson-Hasselbalch equation to calculate the required mass of sodium acetate to prepare two acidic acetate buffers (Buffer A and Buffer B).
- To experimentally prepare Buffer A (0.1 M acetic acid system) and Buffer B (1.0 M acetic acid system).
- To determine and compare the buffer capacities of Buffer A and Buffer B when challenged with strong base (0.5 M NaOH) and strong acid (0.5 M HCl).

INTRODUCTION

A **buffer solution** is an aqueous mixture of a weak acid and its conjugate base (or a weak base and its conjugate acid). The primary function of a buffer is to resist significant changes in pH upon the addition of small amounts of strong acid (H^+ or H_3O^+) or strong base (OH^-).

Buffers are vital in biological and industrial systems. For example, human blood plasma is maintained within a strict pH range of 7.35 to 7.45 by a carbonic acid/bicarbonate ($\text{H}_2\text{CO}_3 / \text{HCO}_3^-$) buffer system.

At the molecular level, a buffer is a chemical system at equilibrium. Consider a nitrous acid (HNO_2) buffer system. The weak acid establishes the following aqueous ionization equilibrium:



The corresponding acid dissociation equilibrium constant expression is:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} \quad (5.3)$$

To prepare a buffer, a soluble salt containing the conjugate base (such as sodium nitrite, NaNO_2) is added to the weak acid solution. Added OH^- ions are neutralized by reacting with the weak acid (HNO_2), while added H_3O^+ ions are neutralized by reacting with the conjugate base (NO_2^-).

Taking the negative logarithm of the equilibrium expression yields the **Henderson-Hasselbalch equation**:

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{Conjugate Base}]}{[\text{Weak Acid}]} \right) = \text{p}K_a + \log \left(\frac{[\text{NO}_2^-]}{[\text{HNO}_2]} \right) \quad (5.4)$$

A buffer is generally considered effective within a working range of **pH = pK_a ± 1**.

In this experiment, you will calculate the mass of solid sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) required to prepare two acetic acid/acetate buffer solutions (pH 4.0). You will prepare both buffers and quantitatively measure their buffer capacities by titrating with 0.5 M NaOH and 0.5 M HCl.

DATA PREP: THE LOGARITHMIC RULE

The Henderson-Hasselbalch equation demonstrates that pH depends directly on the logarithm of the conjugate base-to-acid mole ratio:

- **Optimal Capacity Ratio:** When $[\text{Base}] = [\text{Acid}]$, the ratio is 1. Since $\log(1) = 0$, $\text{pH} = \text{p}K_a$. This represents maximum buffer capacity.
- **Buffer Range Limits:** If the base-to-acid ratio drops below 0.1 or exceeds 10, the logarithmic term shifts pH by more than ±1 unit, causing buffer capacity to rapidly diminish.

KEY EQUATIONS

Henderson-Hasselbalch Equation:

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right) \quad (5.5)$$

[5.1: Investigating Buffers - Experiment](#)

[5.2: Investigating Buffers - Pre-lab](#)

[5.3: Investigating Buffers - Data and Report](#)

5.1: Investigating Buffers - Experiment

SAFETY PRECAUTIONS

- Chemical splash goggles and laboratory coats must be worn at all times.
- Sodium hydroxide (NaOH), hydrochloric acid (HCl), and glacial acetic acid solutions are CORROSIVE and CAUSTIC. Avoid skin and eye contact. Flush spills immediately with copious water.

EQUIPMENT AND CHEMICALS NEEDED

Table 5.1.1: Materials Needed

Equipment & Hardware	Chemicals & Solutions
50-mL burets (2), buret clamp, & ring stand	Sodium acetate anhydrous, $\text{NaC}_2\text{H}_3\text{O}_2(\text{s})$
Vernier pH Sensor & LabQuest interface	0.1 M Acetic acid, $\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$
Magnetic stirrer & magnetic stir bar	1.0 M Acetic acid, $\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$
Electronic balance ($\pm 0.001 \text{ g}$)	0.5 M Sodium hydroxide, $\text{NaOH}(\text{aq})$ titrant
50-mL graduated cylinder & 150-mL beakers (2)	0.5 M Hydrochloric acid, $\text{HCl}(\text{aq})$ titrant
Utility clamp & wash bottle with distilled water	pH 4.00 Buffer calibration solution

EXPERIMENTAL PROCEDURE

Part A: Prepare and Test Buffer Solution A

1. Use your Pre-Lab calculations to prepare 50.0 mL of Buffer A. Weigh the calculated mass of solid sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) on an electronic balance. Dissolve it in exactly 50.0 mL of 0.1 M acetic acid solution using a graduated cylinder and stir thoroughly.
2. Set up two 50-mL burets on a buret clamp and ring stand (Figure 5.1.1). Rinse and fill Buret 1 with 0.5 M NaOH solution. Rinse and fill Buret 2 with 0.5 M HCl solution. Record initial buret volumes.

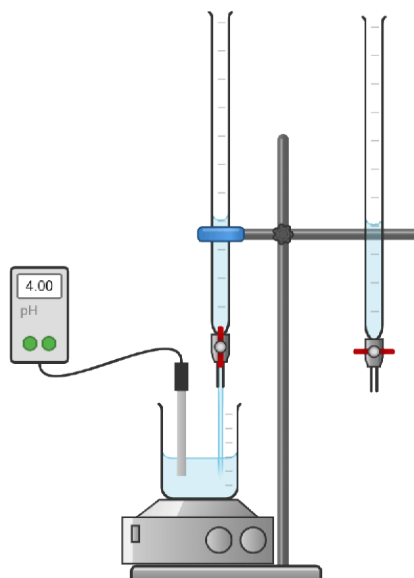


Figure 5.1.1: Buffer titration experimental apparatus setup.

3. Measure 10.0 mL of Buffer A into a 150-mL beaker and add 15 mL of distilled water. Place the beaker on a magnetic stirrer beneath the NaOH buret and add a stir bar.
4. Rinse the pH sensor with distilled water. Check calibration in pH 4.00 buffer.

VERNIER TECH BOX: BUFFER TESTING

1. Calibration Check

- Submerge probe in pH 4.00 buffer. If reading is between 3.90 and 4.10, proceed. Otherwise, recalibrate using pH 4.00 and pH 7.00 standards.

2. Data Collection Mode

- Set Mode to Events with Entry.
- Set Column Name to "Volume Added" and Units to "mL".
- Tap Collect. Tap Keep after each titrant addition, typing in cumulative added volume (0.0, 0.5, 1.0, 1.5...).

5. Base Titration of Buffer A:

- a. Record initial buret reading and initial pH of Buffer A before adding base.
- b. Add 0.50 M NaOH in small increments (0.50 mL). After each addition, allow pH to stabilize, then record buret reading and pH.
- c. Continue adding NaOH until the solution pH increases by **precisely 2.00 pH units** above the initial buffer pH.
- d. Once pH has increased by 2 units, continue adding 0.50 M NaOH in 0.50-mL increments past the endpoint until a sharp pH rise is observed.

6. Acid Titration of Buffer A:

- a. Transfer a fresh 10.0-mL sample of Buffer A into a clean 150-mL beaker and add 15 mL of distilled water.
- b. Position beaker under Buret 2 (0.5 M HCl). Record initial pH and initial buret reading.
- c. Add 0.50 M HCl in 0.50-mL increments until the solution pH decreases by **precisely 2.00 pH units** below initial buffer pH. Record all buret readings and pH values in Data Table 1.

Part B: Prepare and Test Buffer Solution B

7. Use your Pre-Lab calculations to prepare 50.0 mL of Buffer B. Weigh the calculated mass of solid sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) and dissolve it in 50.0 mL of 1.0 M acetic acid solution. Refill burets with 0.5 M NaOH and 0.5 M HCl if needed.
8. Measure 10.0 mL of Buffer B into a 150-mL beaker and add 15 mL of distilled water. Repeat the base titration (NaOH) and acid titration (HCl) procedures from Part A, testing until pH changes by ± 2.00 units. Record all data.

DATA SKILL: CAPACITY CHECK

Immediately compare Buffer A and Buffer B capacity data:

- Buffer B (1.0 M acid system) should require significantly greater titrant volume of NaOH or HCl to shift pH by 2 units compared to Buffer A (0.1 M acid system).
- Quality Check: If Buffer B resists pH change no better than Buffer A, re-verify stock acid bottle concentrations. You may have accidentally used 0.1 M acid for both buffers.

CHEMICAL DISPOSAL

- Dispose of all titrated buffer mixtures in the designated AQUEOUS ACID-BASE WASTE container.
- Rinse the pH sensor with distilled water and store in electrode storage solution. Clean all glassware.

5.2: Investigating Buffers - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Use the Henderson-Hasselbalch equation to perform the following calculations (K_a for acetic acid = 1.8×10^{-5} , $pK_a = 4.74$). Show all work clearly:

- a. **Buffer A Preparation:** Calculate the mass of solid sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$, molar mass = 82.03 g/mol) required to prepare a pH 4.00 buffer when dissolved in 50.0 mL of 0.10 M acetic acid solution.

- b. **Buffer B Preparation:** Calculate the mass of solid sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) required to prepare a pH 4.00 buffer when dissolved in 50.0 mL of 1.0 M acetic acid solution.

2. A chemist needs to prepare a buffer solution with a target pH of 6.10. Using a table of weak acid ionization constants, select the best weak acid/conjugate base pair for this buffer and justify your choice.

3. **Capacity Prediction:** Buffer B contains 1.0 M acetic acid, while Buffer A contains 0.10 M acetic acid.
- Because Buffer B is 10 times more concentrated, how many times more NaOH or HCl should Buffer B be capable of neutralizing before experiencing a 2-unit pH shift compared to Buffer A?
 - Explain why both Buffer A and Buffer B exhibit the exact same initial pH despite their 10-fold concentration difference.

This page titled [5.2: Investigating Buffers - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

5.3: Investigating Buffers - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 5.3.1: Buffer A Titration Data (0.1 M Acetic Acid System)

NaOH Titration of Buffer A		HCl Titration of Buffer A	
Buret Reading (mL)	pH	Buret Reading (mL)	pH
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

Table 5.3.2: Buffer B Titration Data (1.0 M Acetic Acid System)

NaOH Titration of Buffer B		HCl Titration of Buffer B	
Buret Reading (mL)	pH	Buret Reading (mL)	pH
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

POST-LAB ANALYSIS & RESULTS

- Construct titration curves (pH vs. Volume of Titrant) for Buffer A and Buffer B using spreadsheet or graphing software. Attach your curves here.

- Analyze your data and titration curves to complete Table 5.3.3:

Table 5.3.3: Buffer Capacity Summary Results

Experimental Parameter	Buffer A (0.1 M)	Buffer B (1.0 M)
Mass of $\text{NaC}_2\text{H}_3\text{O}_2$ used to prepare buffer (g)	_____	_____
Volume of buffer solution prepared (mL)	50.0	50.0
Initial concentration of $\text{HC}_2\text{H}_3\text{O}_2$ (M)	0.10	1.00
Base Titration (0.5 M NaOH)		
Initial pH of buffer	_____	_____
Volume of 0.5 M NaOH to raise pH by +2.00 units (mL)	_____	_____
Volume of 0.5 M NaOH at equivalence point (mL)	_____	_____
Acid Titration (0.5 M HCl)		
Initial pH of buffer	_____	_____
Volume of 0.5 M HCl to lower pH by -2.00 units (mL)	_____	_____

POST-LAB QUESTIONS

1. Write balanced net ionic equations showing how the acetic acid/acetate buffer system neutralizes added strong acid (H_3O^+) and added strong base (OH^-).

2. Calculate the quantitative **Buffer Capacity (β)** for both Buffer A and Buffer B using your experimental data:

DATA SKILL: CALCULATING BUFFER CAPACITY (β)

$$\beta = \frac{\text{Moles of strong acid or base added}}{(\Delta\text{pH}) \times (\text{Volume of buffer tested in L})} \quad (5.3.1)$$

Calculate β for Buffer A and Buffer B using your $\Delta\text{pH} = 2.00$ data. Is the capacity of Buffer B approximately 10 times greater than Buffer A? Calculate the percent difference between your experimental ratio and the theoretical 10-fold ratio.

3. Consider a Buffer C prepared by mixing 8.203 g of solid sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) with 100.0 mL of 1.0 M acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$).
 - a. Calculate the initial pH of Buffer C. Show your work.

- b. If 5.00 mL of 0.50 M NaOH solution is added to 20.0 mL of Buffer B and 20.0 mL of Buffer C, which buffer will experience a smaller pH change? Explain clearly.

This page titled [5.3: Investigating Buffers - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

6: SOLUBILITY PRODUCT

PRE-LAB PREPARATION

1. Prerequisite Math & Theory

- The 1:6 Stoichiometry Rule: In this iodometric titration, the mole ratio between the analyte (IO_3^-) and titrant ($\text{S}_2\text{O}_3^{2-}$) is 1 to 6:
 - Reaction 1: 1 mol IO_3^- reacts with excess I^- in acid to produce 3 mol I_2 .
 - Reaction 2: 3 mol I_2 reacts with 6 mol $\text{S}_2\text{O}_3^{2-}$ titrant.
 - Calculation Rule: Moles of $\text{IO}_3^- = \diamond \times \text{Moles of } \text{S}_2\text{O}_3^{2-}$. If you omit this factor of 6, your calculated K_{sp} will be erroneous by a factor of 216 (6^3)!
- K_{sp} from Molar Solubility (S): If the molar solubility of $\text{Ca}(\text{IO}_3)_2$ is S, then $[\text{Ca}^{2+}] = S$ and $[\text{IO}_3^-] = 2S$. The solubility product constant becomes:

$$K_{\text{sp}} = [\text{Ca}^{2+}][\text{IO}_3^-]^2 = (S)(2S)^2 = 4S^3 \quad (6.1)$$

2. Required Technical Skills

- Gravity Filtration: The saturated equilibrium solution must be filtered through dry filter paper into a clean, dry receiving beaker. Residual water drops will dilute the saturated solution, causing falsely low solubility and K_{sp} values.
- Starch Indicator Endpoint: Starch forms a dark blue-black complex with triiodide (I_3^-). The endpoint is signaled by the sharp disappearance of the blue-black color, yielding a colorless solution.

3. Critical Safety

- STRONG ACID & ELEMENTAL IODINE: Acidification with 3 M HCl generates molecular iodine (I_2), which can stain and irritate skin. Wear chemical splash goggles and protective gloves at all times.

PURPOSE

- To prepare a saturated aqueous solution of calcium iodate, $\text{Ca}(\text{IO}_3)_2$.
- To experimentally determine the concentration of iodate ions (IO_3^-) in a saturated solution using an iodometric redox titration with standardized sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$).
- To calculate the molar solubility (s) and solubility product constant (K_{sp}) of calcium iodate and compare the experimental result with literature values.

INTRODUCTION

The **solubility** of a substance defines the maximum amount of solute that dissolves in a given quantity of solvent at a specific temperature to establish dynamic equilibrium, yielding a saturated solution. For slightly soluble ionic salts, this dissolution equilibrium is quantitatively expressed by the **solubility product constant (K_{sp})**.

For calcium iodate, $\text{Ca}(\text{IO}_3)_2$, the heterogeneous dissolution equilibrium in water is represented by:

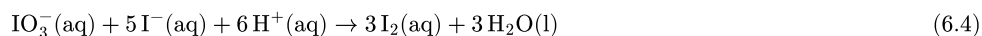


The corresponding solubility product expression is given by:

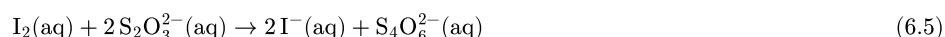
$$K_{\text{sp}} = [\text{Ca}^{2+}][\text{IO}_3^-]^2 \quad (6.3)$$

where $[Ca^{2+}]$ and $[IO_3^-]$ represent the molar equilibrium concentrations of the dissociated ions in a saturated solution.

To determine K_{sp} experimentally, the equilibrium concentration of iodate ions is measured using a two-step iodometric redox titration. An aliquot of saturated $Ca(IO_3)_2$ filtrate is treated with excess potassium iodide (KI) and acidified with hydrochloric acid (HCl). The iodate ions oxidize iodide to molecular iodine (I_2):



The liberated iodine (I_2) is immediately titrated with a standardized solution of sodium thiosulfate ($Na_2S_2O_3$), reducing iodine back to iodide while thiosulfate is oxidized to tetrathionate ($S_4O_6^{2-}$):



Starch indicator is added near the endpoint to form a deep blue-black complex with remaining iodine. The titration is complete when the blue-black color permanently disappears.

DATA PREP: THE 1:6 RATIO

Tracing electron transfer through both reactions establishes the overall stoichiometry:

- Step 1: 1 mole of IO_3^- generates 3 moles of I_2 .
- Step 2: Each mole of I_2 consumes 2 moles of $S_2O_3^{2-}$.
- Net Relationship: It requires 6 moles of thiosulfate to react with the iodine produced by 1 mole of iodate:

$$\text{Moles of } IO_3^- = \frac{1}{6} \times \text{Moles of } S_2O_3^{2-} \quad (6.6)$$

KEY EQUATIONS

Solubility Product Expression:

$$K_{sp} = [Ca^{2+}][IO_3^-]^2 \quad (6.7)$$

Titration Stoichiometry (1:6 Ratio):

$$\text{Moles of } IO_3^- = \frac{\text{Moles of } S_2O_3^{2-}}{6} = \frac{M_{S_2O_3^{2-}} \times V_{S_2O_3^{2-}}}{6} \quad (6.8)$$

Calculating K_{sp} from Molar Solubility (s):

$$S = [Ca^{2+}] = \frac{1}{2} [IO_3^-] \quad (6.9)$$

$$K_{sp} = (S)(2S)^2 = 4S^3 \quad (6.10)$$

[6.1: Solubility Product - Experiment](#)

[6.2: Solubility Product - Pre-lab](#)

[6.3: Solubility Product - Data and Report](#)

6.1: Solubility Product - Experiment

SAFETY PRECAUTIONS

- Chemical splash goggles, laboratory coats, and protective gloves must be worn at all times.
- Hydrochloric acid (3 M HCl) is corrosive; avoid skin contact and inhalation of vapors.
- Elemental iodine (I_2) generated during the titration is an irritant and causes skin staining. Handle solutions with care.

EQUIPMENT AND CHEMICALS NEEDED

Table 6.1.1: Materials Needed

Equipment & Glassware	Chemicals & Reagents
150-mL & 250-mL Beakers	1 M calcium nitrate solution, $Ca(NO_3)_2$ (aq)
125-mL Erlenmeyer flasks (2)	0.2 M potassium iodate solution, KIO_3 (aq)
10.00-mL Volumetric pipette & pipette pump	Standardized 0.025 M $Na_2S_2O_3$ titrant
50-mL Buret, buret clamp, & ring stand	potassium iodide, KI (s)
Glass funnels & quantitative filter paper	3 M hydrochloric acid, HCl (aq)
Electronic balance (± 0.001 g) & wash bottle	1% Starch indicator solution & Deionized water
Solid Chemical Waste & Redox Waste containers	

EXPERIMENTAL PROCEDURE

Part A: Prepare Saturated Calcium Iodate Solution

1. Obtain approximately 20 mL of 1 M $Ca(NO_3)_2$ solution in a small beaker. In a separate 150-mL beaker, obtain approximately 50 mL of 0.2 M KIO_3 solution.
2. Slowly pour the calcium nitrate solution into the potassium iodate solution while stirring continuously to precipitate white calcium iodate, $Ca(IO_3)_2(s)$.
3. Allow the precipitate to settle for 5 minutes.
4. Filter the solid precipitate using gravity filtration through filter paper into a waste beaker.

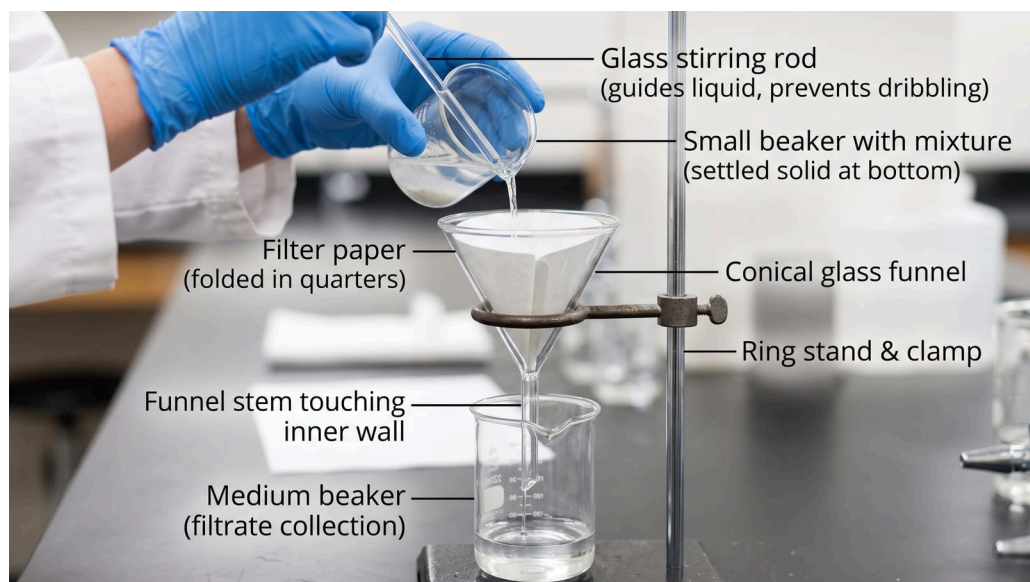


Figure 6.1.1: Gravity Filtration Setup.

5. CRITICAL TECHNIQUE ALERT: PREVENTING CHEMICAL CARRYOVER

The liquid remaining on the wet precipitate is mother liquor packed with excess IO_3^- and NO_3^- ions. If not thoroughly washed, these residual ions will cause massive titration errors or an infinite redox loop.

Precipitate Washing Protocol:

- Using a graduated cylinder filled with 30 mL *ice-cold deionized water*, gently rinse the solid precipitate on the filter paper with a 10-mL portion of water.
 - Allow the liquid to drain **completely** through the filter paper into the waste beaker before adding the next rinse.
 - Repeat this process for a total of **three 10-mL washes**.
- Transfer approximately half of the wet $\text{Ca}(\text{IO}_3)_2$ precipitate into a clean 150-mL beaker. (Reserve the remaining precipitate in case a second preparation is required).
 - Add 60 mL of deionized water to the precipitate. Stir thoroughly for **5 minutes** and allow the mixture to stand for **at least 30 minutes** at room temperature to establish saturation equilibrium.
 - Rinse and fill a 50-mL buret with standardized 0.025 M $\text{Na}_2\text{S}_2\text{O}_3$ solution. Record the exact molarity of the sodium thiosulfate titrant in your data table.

Part B: Analyze the Saturated Solution by Redox Titration

- Carefully decant and filter the saturated $\text{Ca}(\text{IO}_3)_2$ mixture into a **completely dry** receiving beaker using a **dry funnel** and **dry filter paper**. Pour only the liquid supernatant through the filter, taking care not to disturb or transfer the settled solid crystals at the bottom of the beaker.

TECHNIQUE ALERT: THE "DRY" RULE

Why all filtration apparatus must be completely DRY:

- You are isolating a Saturated Solution containing maximum equilibrium ion concentrations.
- If the funnel, filter paper, or receiving beaker contains residual water droplets, the saturated solution becomes diluted (unsaturated).
- Consequence: Dilution leads to lower measured $[\text{IO}_3^-]$, resulting in a calculated K_{sp} that is significantly too low. Thoroughly dry all glassware before filtering.

- Rinse a 10.00-mL volumetric pipette first with deionized water, then with a small 2-mL portion of the clear $\text{Ca}(\text{IO}_3)_2$ filtrate. Discard the rinse.
- Pipette exactly 10.00 mL of the saturated filtrate into a 125-mL Erlenmeyer flask.
- Add 20 mL of deionized water, 1.5 g of solid KI, and 15 drops of 3 M HCl. Swirl gently to dissolve the KI; the solution will turn dark reddish-brown due to I_2 formation.

13. Perform a **Trial Titration**: Record initial buret volume. Titrate rapidly with 0.025 M $\text{Na}_2\text{S}_2\text{O}_3$ in 1-mL increments until the dark brown solution fades to light yellow. Add 10 drops of 1% starch indicator (solution turns dark blue-black). Continue adding $\text{Na}_2\text{S}_2\text{O}_3$ dropwise until the blue-black color permanently disappears. Record final buret volume to estimate titrant usage.

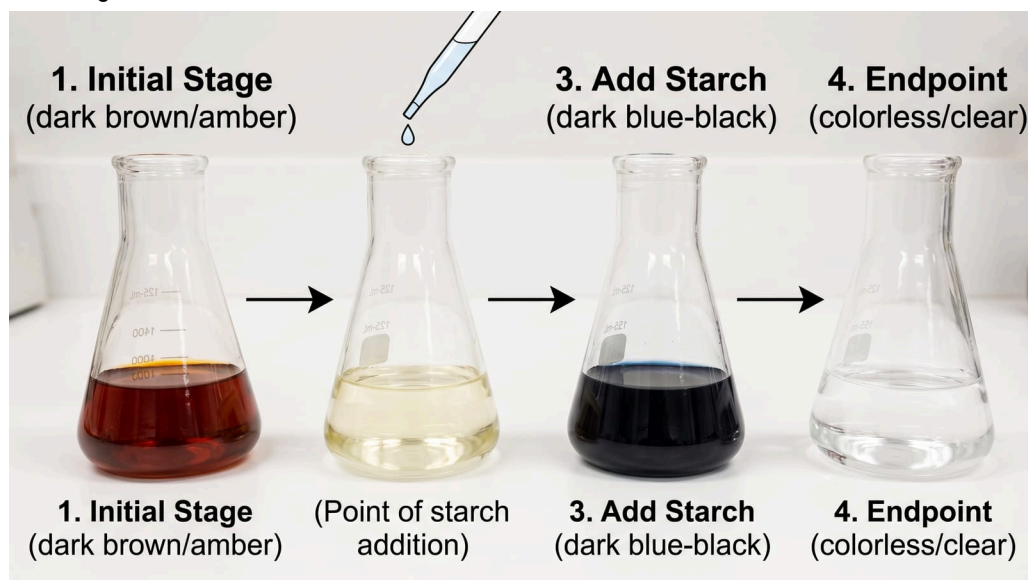


Figure 6.1.2: Iodometric titration color changes.

14. Perform **Exact Titration Trial 1**: Pipette a fresh 10.00-mL aliquot of saturated filtrate into a clean Erlenmeyer flask, add water, KI, and acid as in Step 12. Titrate rapidly until light yellow, add starch indicator, then titrate dropwise near the estimated endpoint until the blue-black color disappears. Record initial and final buret readings (± 0.02 mL).
15. Perform **Exact Titration Trial 2**: Repeat the exact titration with a third 10.00-mL sample of saturated filtrate. Record all data.

CHEMICAL DISPOSAL

- Dispose of all titrated dark blue/colorless liquid mixtures in the designated INORGANIC WASTE container.
- Place excess solid calcium iodate and used filter paper into the SOLID CHEMICAL WASTE container.
- Clean all glassware and return equipment to its proper storage station.

6.2: Solubility Product - Pre-lab

Name: _____ Partner(s): _____ Date: _____

PRE-LAB QUESTIONS

1. Explain the chemical roles of potassium iodide (KI) and hydrochloric acid (HCl) in this experiment. Why are both reagents added in excess to the saturated $\text{Ca}(\text{IO}_3)_2$ sample before titrating?

2. Why is starch used as an indicator in this titration? Describe the specific chemical complex formed, state when starch should be added during the titration, and identify the exact color change at the endpoint.

3. **Experimental Error Analysis:** In Part B, the saturated $\text{Ca}(\text{IO}_3)_2$ solution is filtered into a **dry** receiving beaker using **dry** filter paper.
- If the receiving beaker contained 1.0 mL of residual water prior to filtration, would your measured concentration of IO_3^- be too high, too low, or unaffected?
 - How would this procedural error impact your experimentally calculated K_{sp} value? Explain clearly.

This page titled [6.2: Solubility Product - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

6.3: Solubility Product - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Exact Molar Concentration of Standardized $\text{Na}_2\text{S}_2\text{O}_3$ Titrant: _____ M

Volume of Saturated $\text{Ca}(\text{IO}_3)_2$ Filtrate Sample Analyzed: 10.00 mL

Table 6.3.1: Trial Titration Data

Measurement Description	Trial Run
Initial buret reading (mL)	_____
Final buret reading at endpoint (mL)	_____
Volume of $\text{Na}_2\text{S}_2\text{O}_3$ used (mL)	_____

Table 6.3.2: Exact Titration Data and Calculations

Experimental Parameter	Exact Trial 1	Exact Trial 2
Initial buret reading (mL)	_____	_____
Final buret reading at endpoint (mL)	_____	_____
Volume of $\text{Na}_2\text{S}_2\text{O}_3$ titrant used ($V_{\text{thiosulfate}}$, mL)	_____	_____
Moles of $\text{Na}_2\text{S}_2\text{O}_3$ consumed (mol)	_____	_____
Moles of IO_3^- in 10.00-mL sample ($\text{mol S}_2\text{O}_3^{2-} / 6$)	_____	_____
Concentration of IO_3^- , $[\text{IO}_3^-]$ (M)	_____	_____
Moles of Ca^{2+} in 10.00-mL sample ($\text{mol IO}_3^- / 2$)	_____	_____
Molar Solubility of $\text{Ca}(\text{IO}_3)_2$, $S = [\text{Ca}^{2+}]$ (M)	_____	_____
Solubility Product Constant, K_{sp}	_____	_____

Average Molar Solubility (S_{avg}): _____ M

Average Solubility Product Constant ($K_{\text{sp, avg}}$): _____

POST-LAB QUESTIONS & CALCULATIONS

Show a complete sample calculation for each step below using your Trial 1 data:

1. Calculate the moles of $\text{S}_2\text{O}_3^{2-}$ titrant used, moles of IO_3^- present (applying the 1:6 mole ratio), and molar concentration of IO_3^- . Show all work:

2. Calculate the moles and concentration of Ca^{2+} ions in the saturated solution. (This value represents the molar solubility, S , of $\text{Ca}(\text{IO}_3)_2$):

3. Evaluate the **precision** of your two exact titration trials by calculating the Percent Difference between your two experimental solubility values:

$$\text{Percent Difference} = \frac{|S_1 - S_2|}{S_{\text{avg}}} \times 100\% \quad (6.3.1)$$

4. The literature value for the molar solubility of calcium iodate at 25°C is $7.90 \times 10^{-3} \text{ M}$. Compare your experimental average solubility (S_{avg}) to the literature value by calculating the **Percent Error**:

$$\text{Percent Error} = \frac{S_{\text{avg}} - 7.90 \times 10^{-3} \text{ M}}{7.90 \times 10^{-3} \text{ M}} \times 100\% \quad (6.3.2)$$

5. Calculate the experimental Solubility Product Constant (K_{sp}) using your average molar solubility (S_{avg}) and compare it to the literature K_{sp} value (1.97×10^{-6}):

6. Discuss the specific effects on calculated molar solubility and K_{sp} if extraneous water was introduced into the system from wet filter paper, wet funnels, or by washing the precipitate during the final filtration step:
- _____

This page titled [6.3: Solubility Product - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

7: THERMODYNAMICS OF BORAX

PRE-LAB PREPARATION

1. Prerequisite Math & Theory

- THE VAN 'T HOFF EQUATION: You will graph your data to find enthalpy and entropy:

$$\ln K = -\frac{\Delta H^\circ}{R} \left(\frac{1}{T} \right) + \frac{\Delta S^\circ}{R} \quad (7.1)$$

- SLOPE: $m = -\Delta H^\circ / R$
- INTERCEPT: $b = \Delta S^\circ / R$
- THE TRAP: R is in Joules (8.314 J/mol·K), but ΔH° is usually reported in kJ/mol. Don't forget the factor of 1000.
- STOICHIOMETRY: The tetraborate anion ($\text{B}_4\text{O}_5(\text{OH})_4^{2-}$) accepts 2 PROTONS from HCl. Therefore:

$$\text{moles borax} = 0.5 \times \text{moles HCl} \quad (7.2)$$

2. Required Technical Skills

- Measuring temperature accurately without starting automated data logging prematurely.
- Titration data collection using a 50-mL buret.
- DATA SKILL: Linearizing data in Excel or Google Sheets (converting $T \rightarrow 1/T$ and $K_{sp} \rightarrow \ln K_{sp}$).

3. Critical Safety

- THERMAL HAZARD: You are heating solutions up to 60 °C. Glassware will be hot. Use beaker tongs or insulated gloves.
- CHEMICAL HAZARD: 0.20 M HCl is a skin and eye irritant. Wear safety goggles and lab coat.

PURPOSE

- To determine the thermodynamic properties (ΔG° , ΔH° , and ΔS°) of borax dissolution by measuring its molar solubility at various temperatures.
- To calculate the solubility product constant (K_{sp}) and $\ln K_{sp}$ of borax at different temperatures using acid-base titration.
- To apply the van 't Hoff relationship between $\ln K_{sp}$ and $1/T$ to extract enthalpy and entropy changes for a sparingly soluble salt.
- To gain hands-on experience in solution preparation, chemical equilibrium analysis, and quantitative titration techniques.

INTRODUCTION

The standard free energy change of a chemical process is related to its equilibrium constant according to the equation:

$$\Delta G^\circ = -RT \ln K \quad (7.3)$$

where R is the ideal gas constant (8.314 J/mol·K) and T is the absolute temperature in kelvins. For a sparingly soluble salt in aqueous solution, the equilibrium constant corresponds to the solubility product constant, K_{sp} .

For example, consider the dissolution equilibrium of silver chromate:



The mass action expression set equal to the solubility product constant is:

$$K_{sp} = [\text{Ag}^+]^2 [\text{CrO}_4^{2-}] \quad (7.5)$$

and the standard free energy change for this equilibrium is:

$$\Delta G^\circ = -RT \ln K_{sp} = -RT \ln ([Ag^+]^2 [CrO_4^{2-}]) \quad (7.6)$$

The free energy change is also defined by the fundamental thermodynamic equation:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \quad (7.7)$$

Setting the two free energy expressions equal yields:

$$-RT \ln K_{sp} = \Delta H^\circ - T\Delta S^\circ \quad (7.8)$$

Dividing both sides by $-RT$ and rearranging gives the van 't Hoff equation in standard linear form ($y = mx + b$):

$$\ln K_{sp} = \left(-\frac{\Delta H^\circ}{R} \right) \left(\frac{1}{T} \right) + \frac{\Delta S^\circ}{R} \quad (7.9)$$

A plot of $\ln K_{sp}$ on the y-axis versus reciprocal absolute temperature ($1/T$ in K^{-1}) on the x-axis yields a straight line with a slope $m = -\Delta H^\circ/R$ and a y-intercept $b = \Delta S^\circ/R$. From these graphical parameters, ΔH° and ΔS° for dissolution can be readily determined.

The Borax System

Borax is commonly designated as sodium tetraborate decahydrate, $Na_2B_4O_7 \cdot 10H_2O$. However, based on its crystal structure and chemical reactivity, its true chemical formula is $Na_2B_4O_5(OH)_4 \cdot 8H_2O$. The structure of the tetraborate anion, $B_4O_5(OH)_4^{2-}$, is shown in Figure 1.

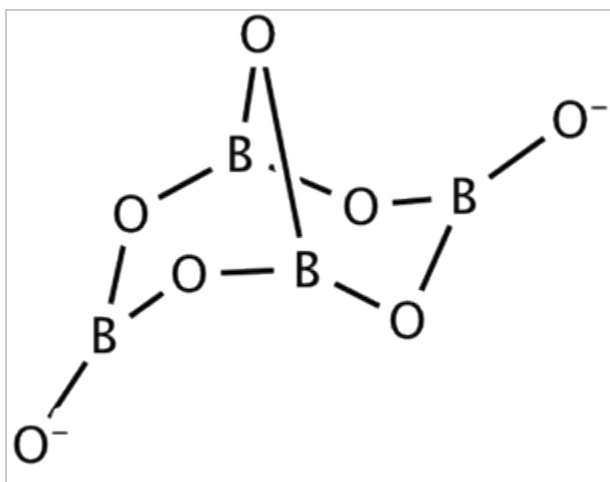
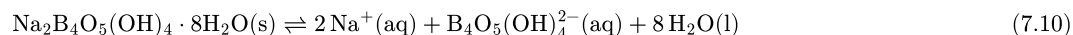


Figure 1: The tetraborate anion, $B_4O_5(OH)_4^{2-}$.

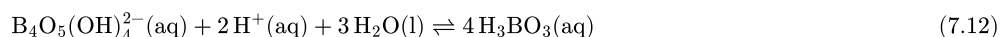
Borax dissolves and dissociates in water according to the reaction equation:



The solubility product expression for borax is:

$$K_{sp} = [Na^+]^2 [B_4O_5(OH)_4^{2-}] \quad (7.11)$$

The tetraborate anion is the conjugate base of the weak acid boric acid and accepts two protons when titrated with a strong acid:



Because one mole of tetraborate anion is produced for every mole of borax that dissolves, the concentration of $B_4O_5(OH)_4^{2-}$ measured by titration represents the molar solubility (S) of borax. Furthermore, stoichiometric stoichiometry dictates that $[Na^+] =$

$2 \times [\text{B}_4\text{O}_5(\text{OH})_4^{2-}] = 2S$. Substituting into the K_{sp} expression yields:

$$K_{sp} = (2S)^2(S) = 4S^3 \quad (7.13)$$

DATA PREP: LINEARIZING THE DATA

The relationship between K_{sp} and temperature is exponential. Plotting $\ln K_{sp}$ versus $1/T$ linearizes the data into a straight line.

- **X-AXIS ORIENTATION:** Because you plot $1/T$, higher temperatures (larger T) correspond to smaller $1/T$ values and appear on the LEFT side of the graph.
- **UNITS:** Convert all temperatures to Kelvin ($K = ^\circ\text{C} + 273.15$) before taking reciprocals.

KEY EQUATIONS

GIBBS FREE ENERGY:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ \quad (7.14)$$

FREE ENERGY AND EQUILIBRIUM CONSTANT:

$$\Delta G^\circ = -RT \ln K \quad (7.15)$$

VAN 'T HOFF EQUATION:

$$\ln K_{sp} = \left(-\frac{\Delta H^\circ}{R} \right) \left(\frac{1}{T} \right) + \frac{\Delta S^\circ}{R} \quad (7.16)$$

[7.1: Thermodynamics of Borax - Experiment](#)

[7.2: Thermodynamics of Borax - Pre-lab](#)

[7.3: Thermodynamics of Borax - Data and Report](#)

7.1: Thermodynamics of Borax - Experiment

SAFETY PRECAUTIONS

- Hydrochloric acid (HCl) is corrosive and causes skin burns and eye damage. Avoid skin contact and inhalation of vapors.
- Hot water baths and heated glassware present thermal burn hazards. Handle with care.
- Wear safety goggles throughout this procedure.

EQUIPMENT AND CHEMICALS NEEDED

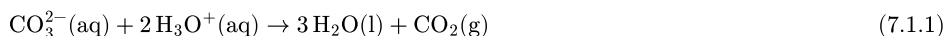
Table 1: Materials List

Equipment & Hardware	Chemicals & Reagents
Beakers (100-mL, 250-mL, 400-mL)	Borax (sodium tetraborate decahydrate)
50-mL buret & buret clamp	Anhydrous sodium carbonate, Na ₂ CO ₃
5-mL volumetric pipet & pipet bulb	~0.20 M Hydrochloric acid, HCl(aq)
125-mL Erlenmeyer flasks & stoppers	Methyl orange indicator solution
Temperature baths (ice bath, hot plate, thermometer)	Bromocresol green indicator solution
Fine-tip permanent marker	

EXPERIMENTAL PROCEDURE

Part A: Standardize the HCl Solution

A standardized HCl solution is prepared using primary standard anhydrous sodium carbonate (Na₂CO₃). Samples of known mass are titrated to a methyl orange endpoint (pH 3.1–4.4) with hydrochloric acid:



Near the endpoint, the solution is boiled to expel dissolved CO₂ gas, which shifts the carbonic acid equilibrium to ensure sharp color changes.

- Obtain approximately 1 g of anhydrous Na₂CO₃ from the desiccator and 200 mL of ~0.20 M HCl.
- Weigh approximately 0.180 g of Na₂CO₃ into a clean 125-mL Erlenmeyer flask (record mass to ±0.001 g). Calculate the estimated volume of 0.20 M HCl required for neutralization. Prepare three separate samples.
- Clean a 50-mL buret, rinse with 3–5 mL portions of the HCl solution, and fill with HCl. Read and record the initial buret volume to ±0.02 mL.
- Dissolve each Na₂CO₃ sample in 50 mL of deionized water and add 2–3 drops of methyl orange indicator. Titrate with HCl to the first faint orange color. Heat the solution to a gentle boil for 1 minute to expel CO₂ (color will revert to yellow). Cool to room temperature and complete the titration dropwise to a permanent pink-orange endpoint.

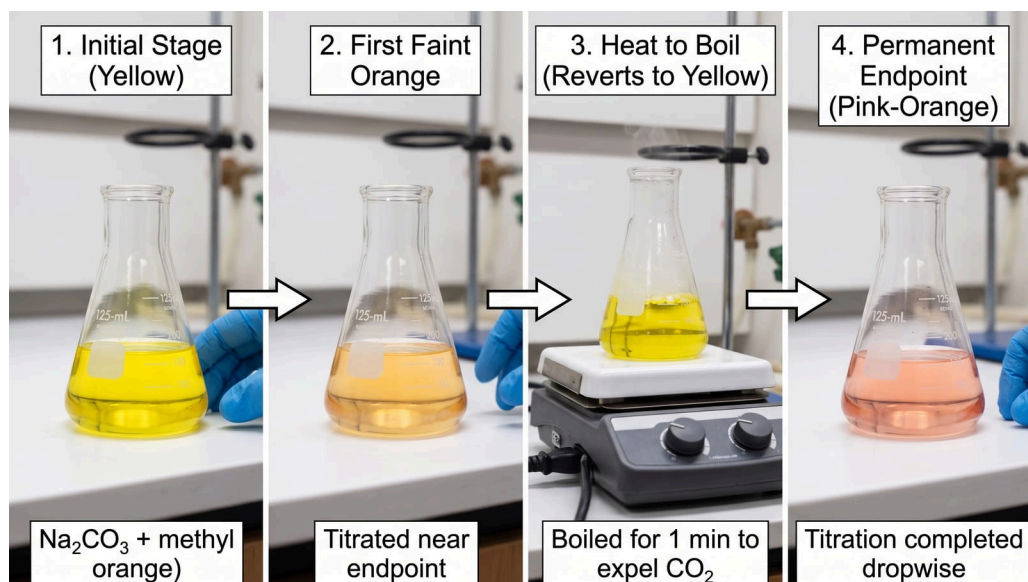


Figure 7.1.1: Methyl orange / boiling titration sequence.

5. Calculate the exact molar concentration of HCl for each trial. The average of three trials agreeing within $\pm 2\%$ will be used for Part C.

Part B: Prepare the Borax Test Solutions

6. Calibrate six test tubes (labeled 1–6) by pipetting exactly 5.00 mL of water, marking the bottom of the meniscus with a permanent marker, and pouring out the water.
7. Prepare a saturated borax stock solution by mixing 30–35 g of borax with 100 mL of deionized water in an Erlenmeyer flask. Stopper and agitate vigorously. **Undissolved solid borax must remain present to ensure saturation.**
8. Fill six additional clean test tubes half-full with the saturated borax supernatant solution containing solid crystals.
9. Place test tubes into temperature baths set to approximately 0 °C (ice bath), room temperature (~22 °C), 32 °C, 40 °C, 48 °C, and 56 °C (do not exceed 60 °C).
10. Equilibrate test tubes in their respective baths for 10–15 minutes, stirring gently every 2–3 minutes. Allow solids to settle for 1 minute prior to sampling.
11. Carefully decant exactly 5.00 mL of clear supernatant into the calibrated test tubes without transferring any solid crystals. Record the exact bath temperature at the moment of transfer.

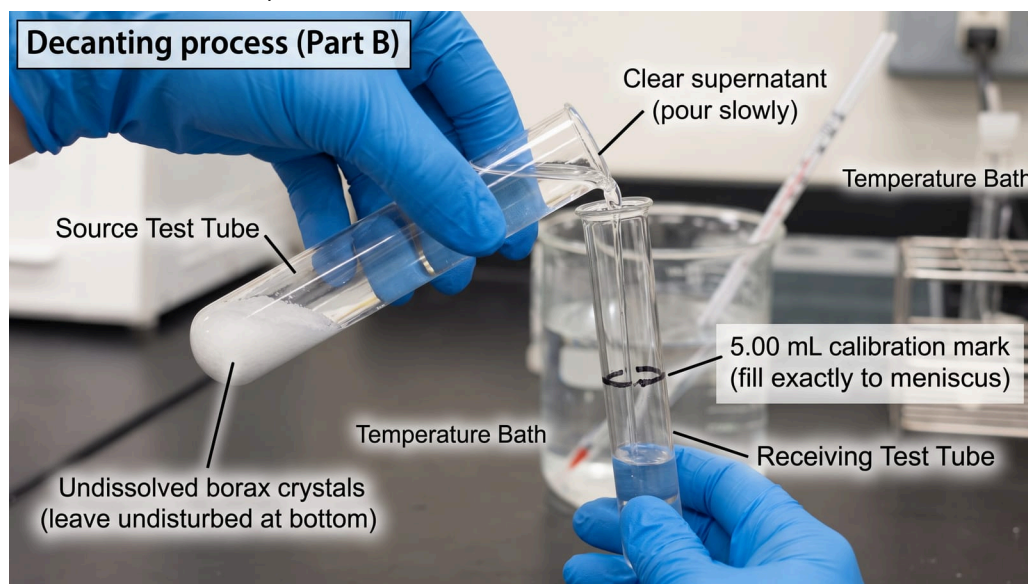


Figure 7.1.2: Saturated borax solution transfer.

Part C: Analyze the Borax Test Solutions

12. Transfer each 5.00 mL supernatant sample into a clean 125-mL Erlenmeyer flask. Rinse the calibrated sampling tube with warm deionized water and add the rinsings to the flask to ensure quantitative transfer.
13. Dilute the solution to approximately 25 mL with deionized water, add 3–4 drops of bromocresol green indicator, and titrate with your standardized HCl solution to a yellow endpoint.

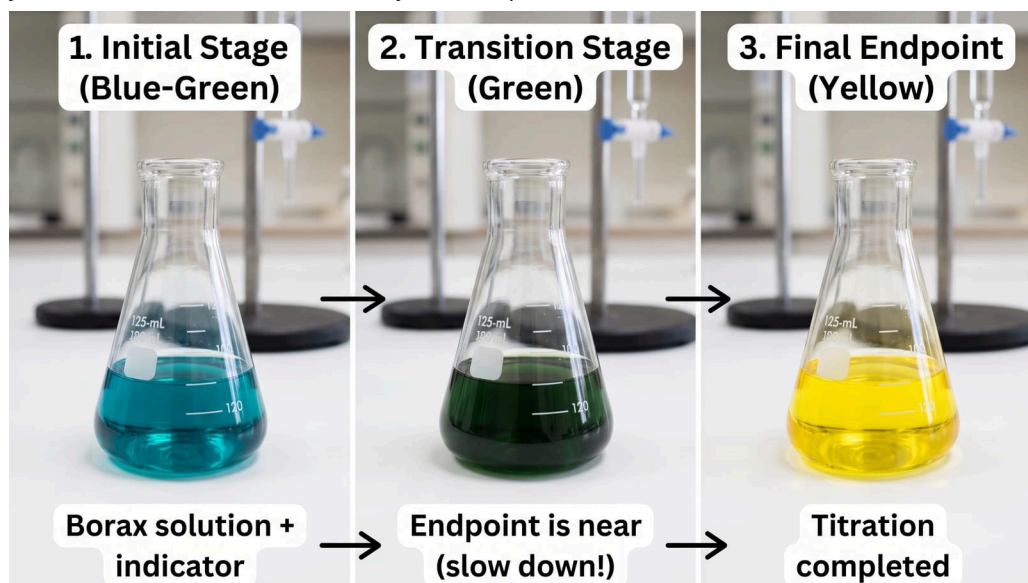


Figure 7.1.3: Bromocresol green titration sequence.

DATA SKILL: DETECTING OUTLIERS

PERFORM A VISUAL REALITY CHECK ON YOUR RAW DATA BEFORE CLEANING GLASSWARE:

- Because borax dissolution is endothermic, solubility decreases as temperature drops. The required volume of HCl MUST DECREASE monotonically with decreasing sampling temperature.
- If a lower-temperature sample requires a larger volume of HCl than a higher-temperature sample, solid crystals were transferred or a titration error occurred. Repeat that trial immediately.

CHEMICAL DISPOSAL

- Dispose of all titrated liquid waste into the designated INORGANIC ACID-BASE WASTE container.
- Unreacted solid borax residues should be returned to the stock recycling container or disposed of as solid waste as directed by your instructor.

7.2: Thermodynamics of Borax - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Why is primary standard sodium carbonate (Na_2CO_3) used to standardize the hydrochloric acid solution rather than preparing the HCl solution directly by volumetric dilution?

2. What exact mass of sodium carbonate (Na_2CO_3 , molar mass = 105.99 g/mol) is required to react completely with 17.0 mL of 0.20 M HCl?

3. Why is it essential to maintain undissolved solid borax at the bottom of each test tube while equilibrating in the temperature baths?

4. **Slope Prediction:** The dissolution of borax in water is an endothermic process ($\Delta H^\circ > 0$).

- a. Examine the van 't Hoff equation: $\ln K_{\text{sp}} = (-\Delta H^\circ/R)(1/T) + \Delta S^\circ/R$. Will the slope ($m = -\Delta H^\circ/R$) of a plot of $\ln K_{\text{sp}}$ vs. $1/T$ be positive or negative?

- b. If your Excel plot yields a slope with the opposite sign, what common mathematical error was likely made?

This page titled [7.2: Thermodynamics of Borax - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

7.3: Thermodynamics of Borax - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A: Standardization of Hydrochloric Acid

Table 1: Standardization of HCl Data

Titration Parameter	Trial 1	Trial 2	Trial 3	Trial 4 (if needed)
1. Mass of Na_2CO_3 (g)	_____	_____	_____	_____
2. Initial buret reading (mL)	_____	_____	_____	_____
3. Final buret reading (mL)	_____	_____	_____	_____
4. Volume of HCl delivered (mL)	_____	_____	_____	_____
5. Calculated Molarity of HCl (M)	_____	_____	_____	_____

Average Standardized HCl Molarity: _____ M

Sample HCl Molarity Calculation (Show work):

Part C: Analysis of Borax Test Solutions

Table 2: Borax Titration Measurements

Measurement Parameter	Tube 1	Tube 2	Tube 3	Tube 4	Tube 5	Tube 6
Approx. Temp ($^{\circ}\text{C}$)	0	Room	32	40	48	56
Actual Temp ($^{\circ}\text{C}$)	_____	_____	_____	_____	_____	_____
Initial buret (mL)	_____	_____	_____	_____	_____	_____
Final buret (mL)	_____	_____	_____	_____	_____	_____
Vol. of HCl (mL)	_____	_____	_____	_____	_____	_____

POST-LAB CALCULATIONS AND ANALYSIS

1. Calculate the molar solubility (S) of borax at each measured temperature from your titration volumes. Show a sample calculation setup below:

2. Calculate the solubility product constant ($K_{sp} = 4S^3$) for each temperature. Show a sample calculation setup below:

Table 3: Calculated Thermodynamic Data Summary

Calculated Quantity	Tube 1	Tube 2	Tube 3	Tube 4	Tube 5	Tube 6
Temperature (K)	_____	_____	_____	_____	_____	_____
1/T (K^{-1})	_____	_____	_____	_____	_____	_____
Molar Solubility, S (M)	_____	_____	_____	_____	_____	_____
K_{sp}	_____	_____	_____	_____	_____	_____
$\ln K_{sp}$	_____	_____	_____	_____	_____	_____

DATA SKILL: LOGARITHMS IN EXCEL

When calculating $\ln K_{sp}$ in Excel, keep these points in mind:

1. **NEGATIVE VALUES ARE EXPECTED:** Because $K_{sp} < 1$, $\ln K_{sp}$ will be negative. Do not drop the negative sign.
2. **FUNCTION SELECTION:** Use the Excel formula **=LN()** for the natural log. Using **=LOG()** (base-10) will distort your calculated slope by a factor of 2.303.

3. Plot $\ln K_{sp}$ (y-axis) versus $1/T$ in Kelvin (x-axis) in Excel. Add a linear trendline and display the equation ($y = mx + b$) and Coefficient of Determination (R^2) on the graph. Paste a copy of your graph below.

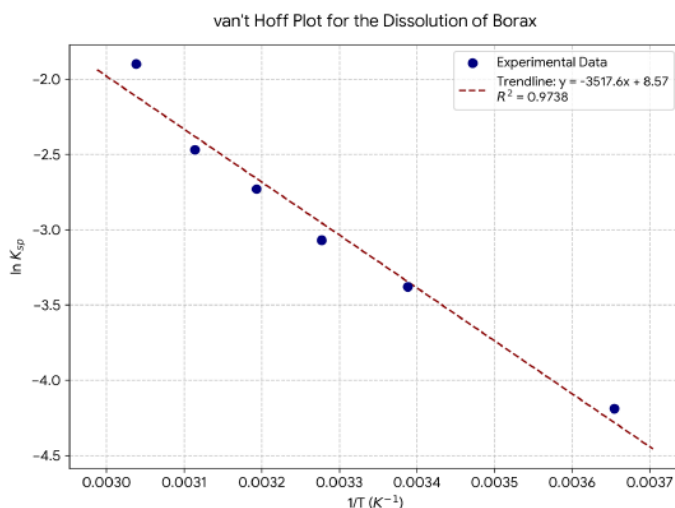


Figure 7.3.1: Example van't Hoff Plot.

4. Evaluate your Coefficient of Determination (R^2). Explain how well your data conforms to the linear van 't Hoff model and discuss potential sources of error if $R^2 < 0.95$.

5. Calculate the standard enthalpy of solution (ΔH°) in kJ/mol from the slope (m) of your trendline ($\Delta H^\circ = -m \cdot R$). Show complete calculations:

6. Calculate the standard entropy of solution (ΔS°) in J/mol·K from the y-intercept (b) of your trendline ($\Delta S^\circ = b \cdot R$). Show complete calculations:

7. Calculate the standard Gibbs free energy change (ΔG°) for borax dissolution at 25.0 °C (298.15 K) using $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$. Is the process spontaneous at 25.0 °C? Explain your reasoning.

This page titled [7.3: Thermodynamics of Borax - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

8: REDOX REACTIONS

PRE-LAB PREPARATION

1. Prerequisite Math & Theory

- **Oxidation States:** You must be able to calculate the oxidation state of manganese in various species:
 - MnO_4^- : Mn = +7 (Dark purple)
 - MnO_4^{2-} : Mn = +6 (Green)
 - MnO_2 : Mn = +4 (Dark brown solid)
 - Mn^{2+} : Mn = +2 (Colorless to pale pink)
- **Redox Definitions:** Remember "OIL RIG" (Oxidation Is Loss, Reduction Is Gain of electrons):
 - **Oxidation:** Loss of electrons (oxidation number increases).
 - **Reduction:** Gain of electrons (oxidation number decreases).
 - **Oxidizing Agent:** The reactant that accepts electrons and gets reduced.
 - **Reducing Agent:** The reactant that donates electrons and gets oxidized.

2. Required Technical Skills

- **Microscale Techniques:** Reactions are performed in a 6-well plate. Do not overfill wells. Mix thoroughly by tapping the plate or using clean toothpicks.
- **Observation Protocols:** Distinguish between "clear" (transparent) and "colorless" (water-like with no hue). Record exact colors and precipitate formation.

3. Critical Safety

- **Strong Oxidizers:** Potassium permanganate (KMnO_4) and hydrogen peroxide (H_2O_2) are strong oxidizing agents that stain skin and clothing and cause severe eye irritation.
- **Strong Caustics:** 50% NaOH solution is extremely corrosive. Wear safety goggles, gloves, and a lab coat at all times.

PURPOSE

- To investigate oxidation-reduction principles by observing the oxidation of Fe^{2+} to Fe^{3+} using various oxidizing agents (H_2O_2 , KMnO_4 , NaClO) and the reduction of Fe^{3+} to Fe^{2+} using reducing agents (Na_2SO_3 , NaBr , NaI).
- To utilize selective chemical indicators (KSCN for Fe^{3+} and $\text{K}_3[\text{Fe}(\text{CN})_6]$ for Fe^{2+}) to confirm reaction outcomes.
- To document the distinct, colorful oxidation states of manganese (+2 through +7) and use visual evidence to deduce reaction products and balance net ionic redox equations.

INTRODUCTION

Oxidation-reduction (redox) reactions involve the transfer of electrons between chemical species. In this experiment, you will explore electron transfer mechanisms through three distinct microscale procedures:

1. **Iron Ion Redox Chemistry (Parts A and B):** Part A examines the oxidation of Fe^{2+} to Fe^{3+} by strong oxidizing agents including hydrogen peroxide (H_2O_2), potassium permanganate (KMnO_4), and sodium hypochlorite (NaClO). Part B investigates the reverse process—the reduction of Fe^{3+} to Fe^{2+} —using sodium sulfite (Na_2SO_3), sodium bromide (NaBr), and sodium iodide (NaI). Qualitative indicators confirm product formation: potassium thiocyanate (KSCN) forms a blood-red complex ($[\text{Fe}(\text{SCN})]^{2+}$) with Fe^{3+} , while potassium ferricyanide ($\text{K}_3[\text{Fe}(\text{CN})_6]$) forms a deep Prussian blue precipitate ($\text{KFe}[\text{Fe}(\text{CN})_6]$) with Fe^{2+} .
2. **Colorful States of Manganese (Part C):** Manganese exhibits a remarkable range of oxidation states from +2 to +7, each displaying a characteristic color in aqueous solution as summarized in Table 8.1. By observing color changes during reactions of Mn^{2+} and MnO_4^- under acidic, neutral, and basic conditions, you can identify the reduced or oxidized manganese product, determine the number of electrons transferred, and write balanced half-reactions.

Oxidation States of Manganese

Table 8.1: Characteristic Colors of Manganese Oxidation States

Oxidation State	Manganese Species	Characteristic Appearance
+2	$\text{Mn}^{2+} (\text{aq})$	Colorless or pale pink
+3	$\text{Mn}^{3+} (\text{aq})$	Rose-red / reddish-brown
+4	$\text{MnO}_2 (\text{s})$	Dark brown precipitate
+5	$\text{MnO}_4^{3-} (\text{aq})$	Blue (hypomanganate)
+6	$\text{MnO}_4^{2-} (\text{aq})$	Bright green (manganate)
+7	$\text{MnO}_4^{-} (\text{aq})$	Deep purple (permanganate)

DATA PREP: PREDICTING COLORS

In manganese redox chemistry, solution color directly indicates the oxidation state:

- **1-Electron Reduction (+7 → +6):** Deep purple MnO_4^{-} converts to bright green MnO_4^{2-} .
- **3-Electron Reduction (+7 → +4):** Deep purple MnO_4^{-} forms a brown solid precipitate (MnO_2).
- **5-Electron Reduction (+7 → +2):** Deep purple MnO_4^{-} becomes a colorless Mn^{2+} solution.

KEY CONCEPTS

Oxidation vs. Reduction:

- **Oxidation:** Loss of electrons; increase in oxidation number.
- **Reduction:** Gain of electrons; decrease in oxidation number.

Oxidizing vs. Reducing Agents:

- **Oxidizing Agent:** Accepts electrons from another species; becomes reduced.
- **Reducing Agent:** Donates electrons to another species; becomes oxidized.

[8.1: Redox Reactions - Experiment](#)

[8.2: Redox Reactions - Pre-lab](#)

[8.3: Redox Reactions - Data and Report](#)

8.1: Redox Reactions - Experiment

SAFETY PRECAUTIONS

- Hydrochloric acid (HCl), sulfuric acid (H₂SO₄), and sodium hydroxide (NaOH) are highly corrosive. Avoid contact with skin and eyes.
- 50% NaOH is extremely caustic and generates significant heat upon mixing. Wear heavy-duty protective gloves and chemical splash goggles.
- Potassium permanganate (KMnO₄) stains skin and clothing permanently; handle dropper bottles carefully over well plates.

EQUIPMENT AND CHEMICALS NEEDED

Table 8.1.1: Equipment and Reagents List

Equipment & Supplies	Inorganic Reagents	Acids, Bases & Salts
6-well reaction plate	0.1 M (NH ₄) ₂ Fe(SO ₄) ₂	0.2 M NaBr
10-mL graduated cylinder	3% H ₂ O ₂ (hydrogen peroxide)	5% NaOCl (commercial bleach)
Mixing toothpicks & pipets	0.1 M FeCl ₃	0.2 M NaI
White paper background sheet	0.1 M K ₃ [Fe(CN) ₆]	0.2 M Na ₂ SO ₃
	0.1 M KSCN	3 M HCl
	0.006 M KMnO ₄	6 M H ₂ SO ₄
	0.025 M KMnO ₄	6 M NaOH
	0.12 M MnSO ₄	50% NaOH (caustic)

EXPERIMENTAL PROCEDURE

Place a 6-well reaction plate on a white sheet of paper to facilitate accurate color observation. Label the top three wells A1, A2, and A3, and the bottom three wells B1, B2, and B3 as illustrated in Figure 8.1.1.



Figure 8.1.1: 6-well plate labeling layout.

Part A: Reaction of Iron(II) Ions with Oxidizing Agents

1. Add 5 mL of 0.1 M ammonium iron(II) sulfate, $(\text{NH}_4)_2\text{Fe}(\text{SO}_4)_2$, to well A1.
2. Add 5 mL of 0.1 M iron(III) chloride, FeCl_3 , to well A2.
3. Record the initial colors of solutions in wells A1 and A2 in Data Table A.
4. Add 5 drops of 0.1 M potassium thiocyanate, KSCN, to both wells A1 and A2.
5. Record the final colors in Data Table A. (Well A2 serves as the positive control for Fe^{3+}).
6. Add 5 mL of 0.1 M ammonium iron(II) sulfate to wells B1, B2, and B3.
7. Add 20 drops of 3 M HCl to acidify wells B1 and B2.
8. Add 20 drops of 3% H_2O_2 to well B1.
9. Add 3 mL of 0.025 M KMnO_4 to well B2 and stir with a clean toothpick.
10. Add 3 mL of 5% NaOCl (bleach) to well B3 and stir.
11. Record the initial solution colors in wells B1, B2, and B3 in Data Table A.
12. Test for the presence of Fe^{3+} in wells B1, B2, and B3 by adding 5 drops of 0.1 M KSCN to each well.
13. Record the final test colors in Data Table A.
14. Rinse the well plate thoroughly with deionized water before proceeding to Part B.

Part B: Reaction of Iron(III) Ions with Reducing Agents

15. Add 5 mL of 0.1 M ammonium iron(II) sulfate to well A1 and 5 mL of 0.1 M FeCl_3 to well A2.
16. Record the initial solution colors in Data Table B.
17. Add 5 drops of 0.1 M potassium ferricyanide, $\text{K}_3[\text{Fe}(\text{CN})_6]$, to both wells A1 and A2.
18. Record the final colors in Data Table B. (Well A1 serves as the positive control for Fe^{2+}).
19. Add 5 mL of 0.1 M FeCl_3 to wells B1, B2, and B3.
20. Add 20 drops of 3 M HCl and 20 drops of 0.2 M Na_2SO_3 solution to well B1.
21. Add 20 drops of 0.2 M NaBr to well B2.
22. Add 20 drops of 0.2 M NaI to well B3.
23. Record the initial solution colors in wells B1, B2, and B3 in Data Table B.
24. Test for the presence of Fe^{2+} in wells B1, B2, and B3 by adding 5 drops of 0.1 M $\text{K}_3[\text{Fe}(\text{CN})_6]$ to each well.
25. Record the final test colors in Data Table B.
26. Rinse the well plate thoroughly with deionized water before proceeding to Part C.

Part C: Observing Colorful Oxidation States of Manganese

27. Add 5 mL of 0.12 M manganese(II) sulfate, MnSO_4 , to wells A1, A2, and A3. Record initial appearances in Data Table C.
28. Add 5 mL of 6 M H_2SO_4 to well A2 to acidify the solution.
29. Add 2–3 drops of 0.006 M KMnO_4 to well A2 and stir thoroughly with a toothpick.
30. Add 2–3 drops of 0.006 M KMnO_4 to well A3 (unacidified) without stirring.
31. Record the final colors and appearances of wells A1, A2, and A3 in Data Table C.

32. Add 5 mL of 0.006 M KMnO_4 solution to wells B1, B2, and B3. Record initial appearances in Data Table C.
33. Add 5 mL of 6 M NaOH and 2 mL of 0.20 M Na_2SO_3 to well B2 and stir.
34. Add 5 mL of 50% NaOH (CAUSTIC) to well B3 and stir rapidly until a distinct color change occurs.
35. Record the final colors in wells B1, B2, and B3 in Data Table C.

CHEMICAL DISPOSAL

- Dispose of all waste solutions in the designated **Heavy Metal / Redox Chemical Waste** container.
- Rinse the well plate with copious amounts of tap water, followed by a final rinse with deionized water.

8.2: Redox Reactions - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. How does potassium thiocyanate (KSCN) act as a selective indicator to distinguish between Fe^{2+} and Fe^{3+} ions in aqueous solution? State the observed color change for a positive test.

2. Identify at least two distinct oxidizing agents used in Part A and write their chemical formulas.

3. How does potassium ferricyanide ($\text{K}_3[\text{Fe}(\text{CN})_6]$) confirm that Fe^{3+} ions have been reduced to Fe^{2+} ? Describe the characteristic positive reaction product and color.

4. In Part B, the sulfite ion (SO_3^{2-}) serves as a reducing agent and is oxidized to the sulfate ion (SO_4^{2-}). Assign the oxidation number of sulfur in both ions:

- a. Oxidation number of S in SO_3^{2-} : _____
b. Oxidation number of S in SO_4^{2-} : _____

5. **Redox Strength Prediction:** Consider elemental chlorine (Cl_2), bromine (Br_2), and iodine (I_2).
- If Cl_2 oxidizes Br^- to Br_2 ...
 - And Br_2 oxidizes I^- to I_2 ...
 - Will Cl_2 be capable of oxidizing I^- to I_2 ? Explain your reasoning based on standard reduction potentials and halogen reactivity trends.

This page titled [8.2: Redox Reactions - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

8.3: Redox Reactions - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Part A: Reactions of Iron(II) with Oxidizing Agents

Table 8.3.1: Part A Observations (Fe^{2+} Oxidation)

Well	Reactants Added	Initial Color	Final Color with KSCN
A1	Fe^{2+} (blank)	_____	_____
A2	Fe^{3+} (control)	_____	_____
B1	$\text{Fe}^{2+} + \text{H}^+ + \text{H}_2\text{O}_2$	_____	_____
B2	$\text{Fe}^{2+} + \text{H}^+ + \text{MnO}_4^-$	_____	_____
B3	$\text{Fe}^{2+} + \text{ClO}^-$	_____	_____

Part B: Reactions of Iron(III) with Reducing Agents

Table 8.3.2: Part B Observations (Fe^{3+} Reduction)

Well	Reactants Added	Initial Color	Final Color with $\text{K}_3[\text{Fe}(\text{CN})_6]$
A1	Fe^{2+} (control)	_____	_____
A2	Fe^{3+} (blank)	_____	_____
B1	$\text{Fe}^{3+} + \text{H}^+ + \text{SO}_3^{2-}$	_____	_____
B2	$\text{Fe}^{3+} + \text{Br}^-$	_____	_____
B3	$\text{Fe}^{3+} + \text{I}^-$	_____	_____

Part C: Colorful Oxidation States of Manganese

Table 8.3.3: Part C Observations (Manganese Redox)

Well	Reactants Added	Initial Color	Final Color / Appearance
A1	Mn^{2+} (blank)	_____	_____
A2	$\text{Mn}^{2+} + \text{H}^+ + \text{MnO}_4^-$ (stirred)	_____	_____

Well	Reactants Added	Initial Color	Final Color / Appearance
A3	$\text{Mn}^{2+} + \text{MnO}_4^-$ (unstirred)	_____	_____
B1	MnO_4^- (blank)	_____	_____
B2	$\text{MnO}_4^- + \text{SO}_3^{2-} + \text{OH}^-$	_____	_____
B3	$\text{MnO}_4^- + 50\% \text{OH}^-$	_____	_____

POST-LAB QUESTIONS

Questions for Parts A and B

- Assign oxidation numbers to the specified atom in each reactant oxidizing agent and its corresponding product:

Table 8.3.4: Part A Oxidation Number Summary

Atom	Oxidizing Agent	Initial State	Reduction Product	Final State
Mn	MnO_4^-	_____	Mn^{2+}	_____
O	H_2O_2	_____	H_2O	_____
Cl	ClO^-	_____	Cl^-	_____

- Balance the reduction half-reactions by inserting the required number of electrons:
 - $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + \text{_____} \text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$
 - $\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{_____} \text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$
 - $\text{ClO}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{_____} \text{e}^- \rightarrow \text{Cl}^-(\text{aq}) + 2\text{OH}^-(\text{aq})$
- Combine the oxidation half-reaction for Fe^{2+} ($\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$) with each reduction half-reaction above to write overall balanced net ionic equations:
 - Reaction with MnO_4^- :**

 - Reaction with H_2O_2 :**

 - Reaction with ClO^- :**

- Sulfite ion reduction analysis:
 - Balance the oxidation half-reaction: $\text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{_____} \text{e}^-$

- b. Write the overall balanced net ionic equation for the reduction of Fe^{3+} by SO_3^{2-} :

5. Based on your experimental observations in Part B, which halide ion— Br^- or I^- —is the stronger reducing agent? Justify your choice using indicator evidence.

Questions for Part C (Manganese Redox Reactions)

For each well below: (a) explain the product color, (b) balance both half-reactions, and (c) write the net balanced overall equation.

Well A2: Potassium permanganate added to acidified manganese(II) sulfate ($\text{Mn}^{2+} + \text{MnO}_4^- + \text{H}^+$).



Well A3: Potassium permanganate added to unacidified manganese(II) sulfate ($\text{Mn}^{2+} + \text{MnO}_4^-$).



Well B2: Sodium sulfite added to basic potassium permanganate ($\text{MnO}_4^- + \text{SO}_3^{2-} + \text{OH}^-$).



Well B3: Concentrated sodium hydroxide (50% NaOH) added to potassium permanganate ($\text{MnO}_4^- + \text{OH}^-$).



This page titled [8.3: Redox Reactions - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

9: VOLTAIC CELLS

PRE-LAB PREPARATION

1. Prerequisite Math & Theory

- **Cell Potential:** $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$. At standard states (1.0 M ion concentrations, 25 °C):

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \quad (9.1)$$

- **Spontaneous Direction:** The metal species with the **more positive (higher)** standard reduction potential (E°) spontaneously undergoes reduction and serves as the **cathode**.
- **Concentration Cells (Nernst Equation):** When both half-cells utilize the same metal electrode and electrolyte species, $E_{\text{cell}}^{\circ} = 0$ V. The overall cell potential is driven solely by concentration differences:

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592 \text{ V}}{n} \log Q \quad (9.2)$$

where $Q = \frac{[\text{dilute}]}{[\text{concentrated}]}$.

2. Required Technical Skills

- Connecting and zeroing Vernier Voltage Probes for low-voltage electrochemical measurements.
- Polishing metal electrodes (copper, zinc, and unknown strips) with steel wool to remove oxide layers prior to immersion.

3. Critical Safety

- Heavy metal waste (copper, zinc, and unknown metal salt solutions) must be collected in designated waste containers. Do not pour down the drain.

PURPOSE

- To construct a semi-microscale Cu/Zn voltaic cell and measure its potential (E_{cell}) using a voltage probe.
- To measure the cell potential of two voltaic cells constructed with unknown metal electrodes and identify the unknown metals using standard reduction potential tables.
- To construct a copper concentration cell, observe its spontaneous potential, and verify the Nernst equation model.

INTRODUCTION

In electrochemistry, a **voltaic (galvanic) cell** is an electrochemical system in which a spontaneous oxidation-reduction reaction generates electrical energy. A voltaic cell consists of two separate half-cells: an oxidation half-cell occurring at the **anode** (negative terminal) and a reduction half-cell occurring at the **cathode** (positive terminal). Electrical neutrality between the two solutions is maintained by a **salt bridge** containing an inert electrolyte such as KNO_3 .

In Parts A and B of this experiment, you will construct semi-microscale voltaic cells in a 24-well plate using metal strips immersed in 0.10 M aqueous solutions of their respective metal nitrates. You will measure the cell potential of a standard Cu/Zn cell and two cells pairing copper with unknown metal electrodes (X and Y). By observing lead polarities and measured cell potentials, you will determine the reduction potentials of the unknown metals and identify them from an activity series.

In Part C, you will explore a **concentration cell**. In a concentration cell, both half-cells utilize identical copper electrodes and Cu^{2+} ions, but at different concentrations (0.050 M vs. 1.0 M). The system generates voltage as it spontaneously works to equalize Cu^{2+} concentrations across the half-cells.

DATA PREP: CONCENTRATION AS A BATTERY

Concentration cells demonstrate that concentration gradients can drive spontaneous current flow:

- **Drive Mechanism:** Thermodynamics favors equalizing concentrations. Electrons flow from the dilute half-cell (anode, where $\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$) to the concentrated half-cell (cathode, where $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$).
- **Nernst Prediction:** Because $E_{\text{cell}}^{\circ} = 0$ V, the theoretical potential is given by $E_{\text{cell}} = -\frac{0.0592}{2} \log\left(\frac{0.050}{1.0}\right) \approx +0.038$ V.

KEY EQUATIONS

Cell Potential:

$$E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}} \quad (9.3)$$

Standard Cell Potential:

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \quad (9.4)$$

Nernst Equation at 25 °C (298.15 K):

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0592 \text{ V}}{n} \log Q \quad (9.5)$$

[9.1: Voltaic Cells - Experiment](#)

[9.2: Voltaic Cells - Pre-lab](#)

[9.3: Voltaic Cells - Data and Report](#)

9.1: Voltaic Cells - Experiment

SAFETY PRECAUTIONS

- Wear chemical splash goggles and a protective lab coat throughout this experiment.
- Metal salt solutions ($\text{Cu}(\text{NO}_3)_2$, $\text{Zn}(\text{NO}_3)_2$, and unknown metal nitrates) are toxic and irritating to skin and eyes. Avoid contact.
- Dispose of all metal-containing solutions and salt-bridge strings in the designated **Heavy Metal Chemical Waste** container.

EQUIPMENT AND CHEMICALS NEEDED

Table 9.1.1: Equipment and Chemical Supplies

Hardware & Equipment	Chemicals & Electrodes
Vernier LabQuest & Voltage Probe leads	0.10 M $\text{Cu}(\text{NO}_3)_2$ solution & copper strips
24-well test plate	0.10 M $\text{Zn}(\text{NO}_3)_2$ solution & zinc strips
Cotton string (for salt bridges)	1.0 M KNO_3 (salt bridge solution)
Steel wool pads (for electrode polishing)	0.10 M Unknown X solution & Unknown X strips
Disposable transfer pipets	0.10 M Unknown Y solution & Unknown Y strips
100-mL beaker (for salt bridge preparation)	1.0 M CuSO_4 stock solution

EXPERIMENTAL PROCEDURE

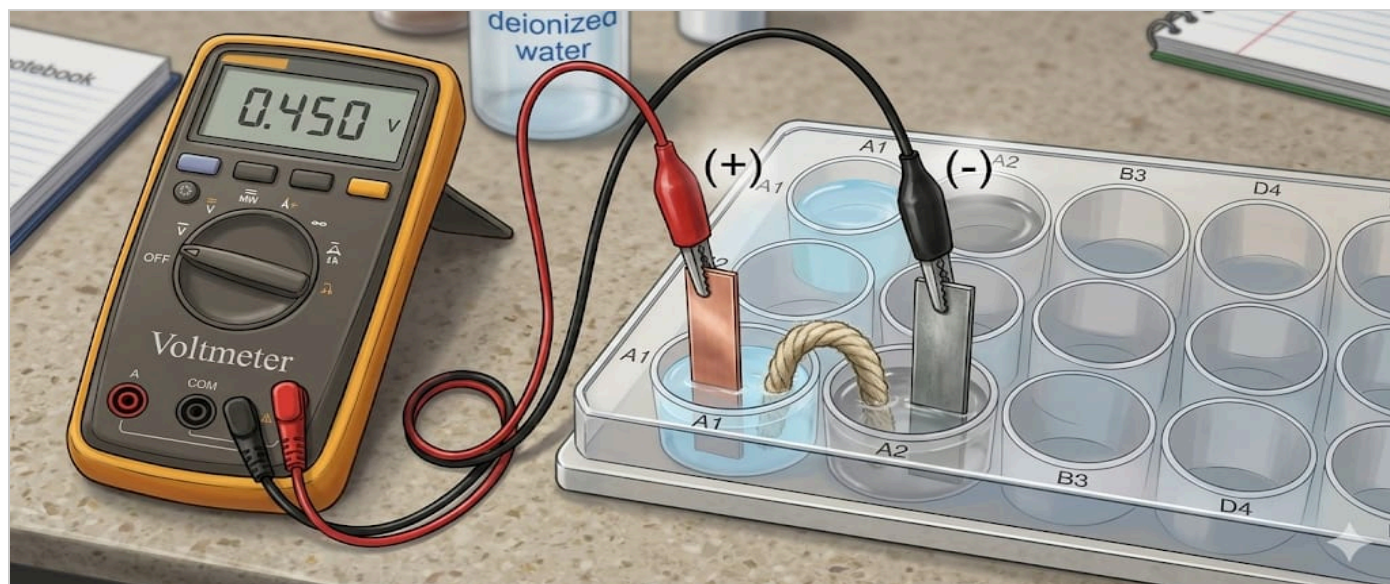


Figure 9.1.1: Experimental setup for measuring semi-microscale cell potentials.

Part A: Determine the Cell Potential (E_{cell}) for a Cu/Zn Voltaic Cell

1. Obtain a clean 24-well test plate. Pipet approximately 1 mL of 0.10 M $\text{Cu}(\text{NO}_3)_2$ into Well 1 and 1 mL of 0.10 M $\text{Zn}(\text{NO}_3)_2$ into adjacent Well 2.
2. Obtain copper and zinc metal strips. Thoroughly polish both strips with steel wool until bright and free of oxide coating.

VERNIER TECH BOX: MEASURING VOLTAGE

1. Hardware Connection

- Connect the Voltage Probe leads to **CH 1** on the LabQuest interface.

2. Lead Polarity Rules

- **RED Lead (+):** Connect to the electrode expected to be the **Cathode** (reduction site / higher reduction potential).
- **BLACK Lead (-):** Connect to the electrode expected to be the **Anode** (oxidation site / lower reduction potential).

3. Potential Sign Interpretation

- **Positive Potential (+):** The assumed cathode is correct; electrons flow spontaneously from anode to cathode.
- **Negative Potential (-):** Reversed polarity. Swap the lead positions to record a positive potential and correctly identify the cathode.

3. Prepare a salt bridge by soaking a short length of cotton string in 1.0 M KNO_3 solution. Place the string bridge so that one end dips into Well 1 and the other dips into Well 2.
4. Connect the voltage probe leads (Red to Cu, Black to Zn) and insert the metal electrodes into their respective wells.
5. Record the stabilized cell potential (E_{cell}) to the nearest 0.001 V.
6. Remove, clean, and re-polish both metal strips with steel wool. Re-insert and record a second potential reading. Repeat once more for a third trial. Calculate the average cell potential.

Part B: Cell Potentials of Copper paired with Unknown Metals X and Y

7. Pipet 1 mL of 0.10 M Unknown X solution into Well 3 adjacent to a well containing 0.10 M $\text{Cu}(\text{NO}_3)_2$.
8. Polish an Unknown X metal strip with steel wool. Connect a new KNO_3 -soaked salt bridge between the two wells.
9. Measure the cell potential (X/Cu) three times with cleaning/polishing between trials. Note which electrode acts as the cathode (Red lead). Record all data and the average potential.
10. Repeat the complete procedure for Unknown Y pairing 0.10 M Unknown Y solution and an Unknown Y metal strip against 0.10 M $\text{Cu}(\text{NO}_3)_2$. Record all data and the average potential.

DATA SKILL: THE LOGIC MATRIX

Identifying Unknown Metals:

- **If Red Lead is on Copper:** Copper is the cathode ($E_{\text{Cu}}^\circ = +0.34 \text{ V}$). The unknown is the anode with a lower reduction potential ($E_{\text{unknown}}^\circ = +0.34 \text{ V} - E_{\text{cell}}$).
- **If Red Lead is on Unknown:** Unknown is the cathode with a higher reduction potential ($E_{\text{unknown}}^\circ = +0.34 \text{ V} + E_{\text{cell}}$).

Part C: Prepare and Test a Copper Concentration Cell

11. Prepare 20 mL of 0.050 M CuSO_4 solution by mixing 1.0 mL of 1.0 M CuSO_4 stock solution with 19.0 mL of deionized water in a small beaker.
12. Pipet 1 mL of 0.050 M CuSO_4 into Well 5 and 1 mL of 1.0 M CuSO_4 into Well 6.
13. Insert polished copper metal strips into both wells and connect them with a fresh KNO_3 salt bridge.
14. Connect the Red lead to the electrode in the concentrated solution (1.0 M) and the Black lead to the dilute solution (

0.050 M). Measure and record the cell potential three times. Calculate the average.

CHEMICAL DISPOSAL

- Transfer all waste solutions from the well plate into the designated **Heavy Metal Chemical Waste** bottle.
- Rinse the 24-well plate thoroughly with tap water followed by deionized water and air-dry.

9.2: Voltaic Cells - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

- Write the balanced overall cell reaction and calculate the theoretical standard cell potential (E_{cell}°) for the following cell shorthand notation:



- Complete Table 9.2.1 using a standard table of reduction potentials at 25 °C. Identify the anode for each cell pairing by circling it.

Table 9.2.1: Theoretical Galvanic Cell Potentials

Electrodes Pair (Circle Anode)	Reduction Half-Reactions	Standard Potential, E° (V)	Standard Cell Potential, E_{cell}° (V)
(Zn) / Cu	$\text{Zn}^{2+} + 2\text{e}^{-} \rightarrow \text{Zn(s)}$ $\text{Cu}^{2+} + 2\text{e}^{-} \rightarrow \text{Cu(s)}$	-0.76 $+0.34$	$+1.10$
Cu / Ag	_____ _____	_____ _____	_____
Cu / Mg	_____ _____	_____ _____	_____
Zn / Sn	_____ _____	_____ _____	_____
Cu / Sn	_____ _____	_____ _____	_____

- For a Zn/Cu cell, if $[\text{Zn}^{2+}] = 0.010 \text{ M}$ and $[\text{Cu}^{2+}] = 1.0 \text{ M}$, will the measured cell potential (E_{cell}) be greater than, less than, or equal to the standard cell potential ($E_{\text{cell}}^{\circ} = 1.10 \text{ V}$)? Explain using the Nernst equation.

- Logic Check:** You measure a potential of $+0.46 \text{ V}$ for a cell constructed from Unknown Metal X and Copper ($E_{\text{Cu}}^{\circ} = +0.34 \text{ V}$).

- a. If Copper is the cathode (Red lead), what is the reduction potential of Unknown X?

- b. If Copper is the anode (Black lead), what is the reduction potential of Unknown X?

- c. Consult a standard reduction table ($\text{Ag}^+/\text{Ag} = +0.80 \text{ V}$, $\text{Sn}^{2+}/\text{Sn} = -0.14 \text{ V}$, $\text{Zn}^{2+}/\text{Zn} = -0.76 \text{ V}$, $\text{Mg}^{2+}/\text{Mg} = -2.37 \text{ V}$). Which scenario corresponds to a real metal electrode?

This page titled [9.2: Voltaic Cells - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

9.3: Voltaic Cells - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 9.3.1: Experimental Cell Potential Measurements

Experimental Parameter	Part A: Cu / Zn	Part B: X / Cu	Part B: Y / Cu	Part C: Conc. Cell
1. Identified Anode Electrode	_____	_____	_____	_____
2. E_{cell} , Trial 1 (V)	_____	_____	_____	_____
3. E_{cell} , Trial 2 (V)	_____	_____	_____	_____
4. E_{cell} , Trial 3 (V)	_____	_____	_____	_____
5. Average E_{cell} (V)	_____	_____	_____	_____

POST-LAB CALCULATIONS AND ANALYSIS

- Part A Analysis:** Compare your experimental average cell potential for the Cu/Zn cell with the theoretical standard potential ($E_{\text{cell}}^{\circ} = 1.10 \text{ V}$). Calculate your percent error and discuss experimental factors (e.g., non-standard 0.10 M concentrations or oxide films) that contribute to the difference.

- Part B Analysis:** Unknown metals X and Y were selected from magnesium (Mg), silver (Ag), tin (Sn), or zinc (Zn). Calculate the reduction potential for each unknown and state their identities. Show work:

- Identity of Unknown X:** _____

- Identity of Unknown Y:** _____

- Part C Analysis:** Use the Nernst equation to calculate the theoretical cell potential (E_{cell}) for the copper concentration cell (0.050 M vs. 1.0 M Cu^{2+} at 298 K). Compare this value to your measured average potential.

DATA SKILL: VALIDATING THE MODEL

If your measured potential is significantly lower than theoretical ($+0.038\text{ V}$), evaluate potential volumetric dilution errors in preparing the 0.050 M solution or incomplete salt bridge equilibration.

This page titled [9.3: Voltaic Cells - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

10: DECOMPOSITION OF HYDROGEN PEROXIDE

PRE-LAB PREPARATION

1. Prerequisite Math & Theory

- **Method of Initial Rates:** You will determine the reaction order (n) by comparing two runs where the reactant concentration changes while other variables are held constant:

$$\frac{\text{Rate}_2}{\text{Rate}_1} = \left(\frac{[\text{KI}]_2}{[\text{KI}]_1} \right)^n \quad (10.1)$$

- *Example:* If doubling the concentration doubles the reaction rate ($2^1 = 2$), the reaction is first-order with respect to that reactant.
- **Arrhenius Equation:** To determine activation energy (E_a), compare rate constants at two different temperatures (Part A vs. Part D):

$$\ln \left(\frac{k_2}{k_1} \right) = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right) \quad (10.2)$$

2. Required Technical Skills

- Configuring time profiles on data loggers (300-second run duration at 0.1 samples/second).
- Linear regression analysis: selecting only the initial linear region (first 60 seconds) to avoid error from reactant depletion.

3. Critical Safety

- **Pressure Hazard:** The reaction produces oxygen gas. If the stopper is stuck, pressure can build up rapidly. **Never** point the reaction vessel at yourself or others.

PURPOSE

- To conduct the catalyzed decomposition of hydrogen peroxide under varied reactant concentrations and temperatures.
- To determine the reaction orders and write the differential rate law expression.
- To calculate the rate constant (k) at room temperature and elevated temperature.
- To determine the activation energy (E_a) for the reaction using the two-point Arrhenius equation.

INTRODUCTION

The uncatalyzed decomposition of hydrogen peroxide, H_2O_2 , in aqueous solution proceeds very slowly at room temperature. A standard commercial 3% bottle remains stable for months. The decomposition is described by the balanced equation:



To accelerate this process, catalysts such as iodide ions (I^- from KI), manganese(IV) oxide (MnO_2), or the enzyme catalase are added. In this experiment, potassium iodide (KI) serves as a homogeneous catalyst.

By conducting the reaction in a sealed vessel attached to a Gas Pressure Sensor, the reaction rate is determined by monitoring the rate of pressure increase ($\Delta P/\Delta t$) caused by oxygen gas generation. This approach allows key kinetic parameters to be evaluated:

- **Reaction Orders & Rate Law:** Varying initial concentrations of H_2O_2 and I^- (Parts A, B, and C) yields the reaction orders m and n .
- **Rate Constant (k):** Using the rate law, k is calculated for each experimental trial.

- **Activation Energy (E_a):** Comparing rate constants at $\sim 20^\circ\text{C}$ and $\sim 30^\circ\text{C}$ (Parts A and D) allows calculation of E_a .

DATA PREP: THE INITIAL RATE METHOD

As H_2O_2 is consumed, the reaction slows down, causing the Pressure vs. Time curve to level off:

- **Objective:** Measure the rate at the *very beginning* of the reaction ($t = 0$), where initial concentrations are precisely known.
- **Method:** Apply a linear fit strictly to the **first 60 seconds** of data following initial mixing. Including curved data points will artificially depress your calculated rate.

KEY EQUATIONS

Rate Law Expression:

$$\text{Rate} = k[\text{H}_2\text{O}_2]^m[\text{I}^-]^n \quad (10.4)$$

Reaction Order (Logarithmic Solution):

$$n = \frac{\log(\text{Rate}_2/\text{Rate}_1)}{\log([\text{Reactant}]_2/[\text{Reactant}]_1)} \quad (10.5)$$

Two-Point Arrhenius Equation:

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2}\right) \quad (10.6)$$

[10.1: Decomposition of Hydrogen Peroxide - Experiment](#)

[10.2: Decomposition of Hydrogen Peroxide - Pre-lab](#)

[10.3: Decomposition of Hydrogen Peroxide - Data and Report](#)

10.1: Decomposition of Hydrogen Peroxide - Experiment

SAFETY PRECAUTIONS

- Gas pressure builds continuously inside the test tube during data collection. Loosen and remove the rubber stopper promptly after each trial to safely vent the system.
- Eye protection (goggles) must be worn at all times. Potassium iodide and hydrogen peroxide solutions are mild skin and eye irritants.

EQUIPMENT AND CHEMICALS NEEDED

Table 10.1.1: Equipment and Materials List

Equipment & Hardware	Reagents & Solutions
Vernier Gas Pressure Sensor & interface	3% Hydrogen peroxide (H_2O_2) solution
Plastic tubing with quick-connect valve	0.50 M Potassium iodide (KI) solution
Digital thermometer or temperature probe	Deionized water
#1 One-hole rubber stopper with stem	
#1 Solid rubber stopper	
Medium test tube (~ 20 mL)	
Two 10-mL graduated cylinders	
Graduated plastic transfer pipettes	
1-L beaker (water bath container)	

EXPERIMENTAL PROCEDURE

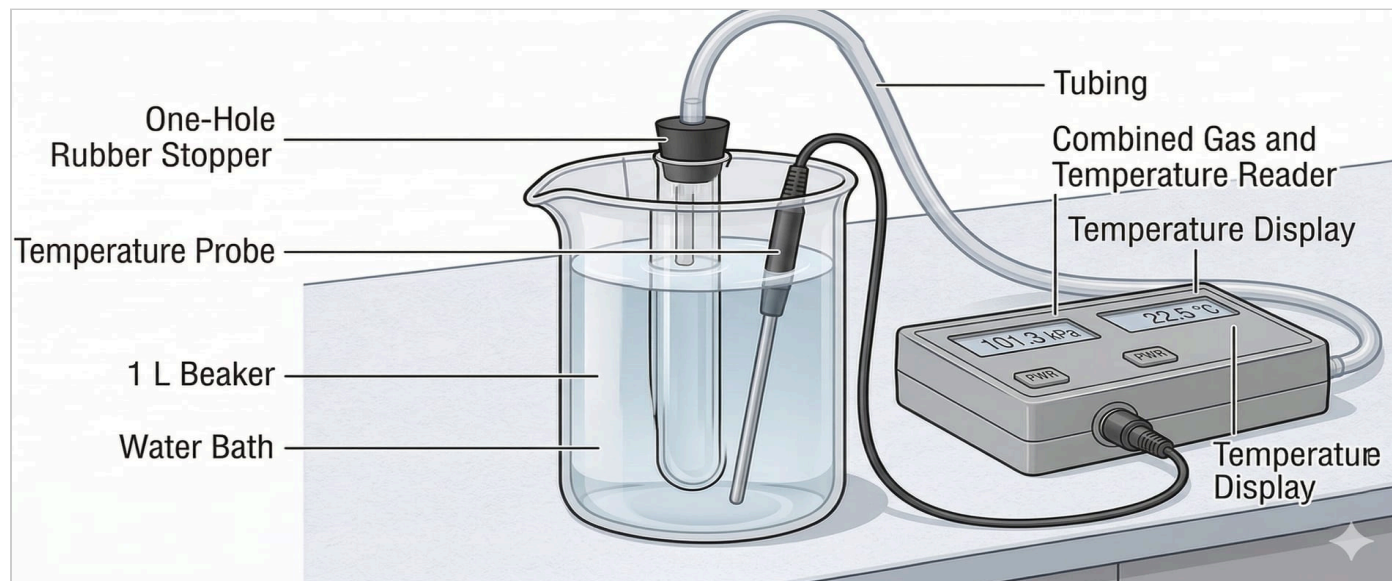


Figure 10.1.1: Reaction vessel apparatus with Gas Pressure Sensor setup.

Part A: Decomposition of 3% H_2O_2 with 0.50 M KI at $\sim 20^\circ\text{C}$

- Prepare the water bath and thermal equilibration setup:
 - Fill a 1-L beaker with room-temperature water ($\sim 20^\circ\text{C}$). Record the exact water bath temperature in Data Table 1. Presume this temperature remains constant for Parts A, B, and C.
 - Pipet 4.0 mL of 3% H_2O_2 into a clean, dry test tube. Seal with the solid stopper and submerge in the water bath.
 - Measure 2.0 mL of 0.50 M KI solution in a graduated cylinder. Draw 1.0 mL into a transfer pipette, invert the pipette, and immerse the bulb end into the water bath to thermally equilibrate.
- Connect the one-hole stopper and tubing assembly to Channel 1 on the interface.

VERNIER TECH BOX: KINETICS SETUP

1. Data Logger Configuration

- Set collection mode to **Time Based**.
- Duration:** 300 seconds (5 minutes).
- Rate / Interval:** 0.1 samples/second (10 seconds/sample).

2. Sensor Calibration Note

- Do NOT zero the sensor.** The initial baseline represents atmospheric pressure ($\sim 101.3\text{ kPa}$). Kinetics calculations depend strictly on the initial slope ($\Delta P/\Delta t$).

- Start data collection on the Vernier software.
- Rapidly initiate the reaction:
 - Remove the test tube and solid stopper from the bath.
 - Dispense the 1.0 mL of thermally equilibrated KI solution quickly into the test tube. Swirl once to mix.
 - Firmly insert the one-hole stopper connected to the pressure sensor into the test tube.
 - Submerge the test tube back into the water bath and hold firmly in place.
- When data collection completes at 300 seconds, carefully loosen the stopper to vent pressure.
- Determine the initial reaction rate ($\Delta P/\Delta t$):

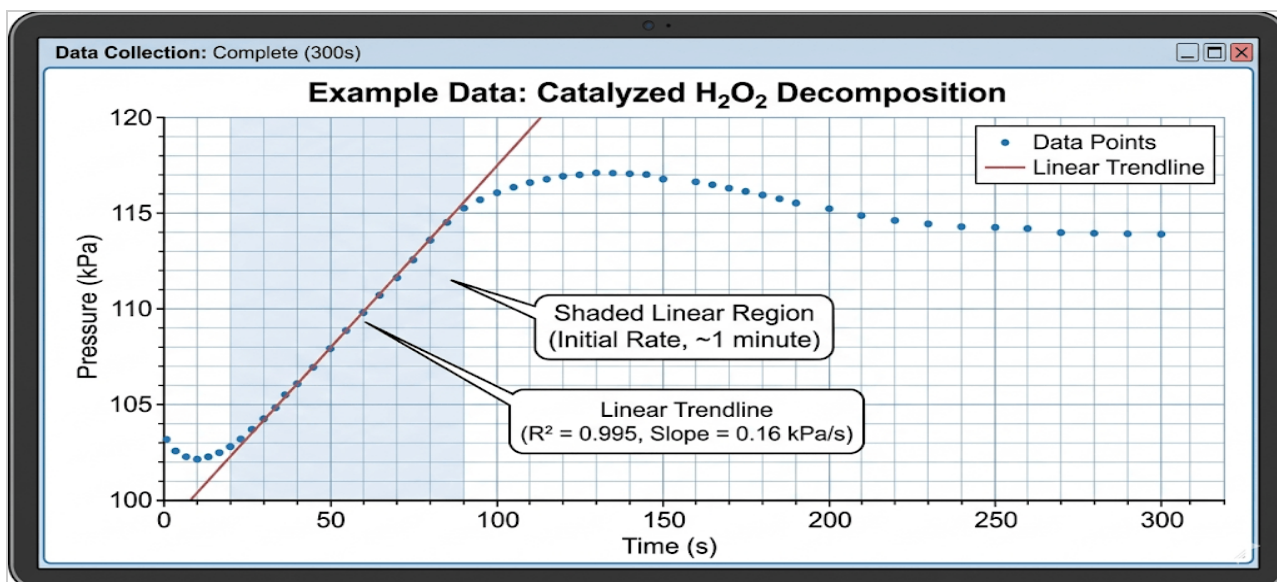


Figure 10.1.2: Selection of initial 60-second linear region for regression analysis.

- Highlight approximately 60 seconds of data in the linear region immediately following initial mixing.
 - Apply a linear regression model ($y = mx + b$). Ensure $R^2 > 0.98$.
 - Record the linear slope (kPa/s) as the initial reaction rate in Data Table 1.
7. Rinse the test tube thoroughly with deionized water and dry before Part B.

Part B: Decomposition of 3% H_2O_2 with 0.25 M KI at $\sim 20^\circ\text{C}$

- Add 4.0 mL of 3% H_2O_2 to the test tube, seal with the solid stopper, and submerge in the room-temperature bath.
- Dilute the catalyst: mix 1.0 mL of 0.50 M KI with 1.0 mL of deionized water to prepare 0.25 M KI.
- Draw 1.0 mL of 0.25 M KI into a pipette and place in the water bath for 2 minutes to equilibrate.
- Repeat Steps 3–7 to collect data and record the initial rate (kPa/s).

Part C: Decomposition of 1.5% H_2O_2 with 0.50 M KI at $\sim 20^\circ\text{C}$

- Dilute the substrate: combine 2.0 mL of 3% H_2O_2 with 2.0 mL of deionized water to yield 4.0 mL of 1.5% H_2O_2 . Transfer into the test tube, seal, and submerge in the water bath.
- Draw 1.0 mL of fresh 0.50 M KI into a pipette and equilibrate in the bath.
- Repeat Steps 3–7 to collect data and record the initial rate (kPa/s).

Part D: Decomposition of 3% H_2O_2 with 0.50 M KI at $\sim 30^\circ\text{C}$

- Prepare a warm water bath ($\sim 30^\circ\text{C}$) in the 1-L beaker using warm tap water. Record the exact temperature in Data Table 1.
- Conduct Part D identically to Part A (4.0 mL 3% H_2O_2 + 1.0 mL 0.50 M KI) using the 30°C bath for thermal equilibration and monitoring. Record the initial rate (kPa/s).

CHEMICAL DISPOSAL

- Dispose of all reaction mixtures in the labeled **Aqueous Inorganic Waste** bottle.

10.2: Decomposition of Hydrogen Peroxide - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. Explain the role of a catalyst in a chemical reaction. How does adding a catalyst alter the **activation energy** (E_a) and the **initial reaction rate**?

2. The reaction equation is: $2 \text{H}_2\text{O}_2(aq) \rightarrow 2 \text{H}_2\text{O}(l) + \text{O}_2(g)$. Why does gas pressure increase inside the sealed test tube during the reaction?

3. Commercial hydrogen peroxide is sold as a 3.0% (m/v) solution (3.0 g H_2O_2 /100 mL solution). Calculate the molarity of 3.0% (m/v) H_2O_2 (Parts A, B, D) and 1.5% (m/v) H_2O_2 (Part C). Record these values in Table 10.2.1 using 2 significant figures (Molar mass of $\text{H}_2\text{O}_2 = 34.01 \text{ g/mol}$).

Table 10.2.1: Stock Solution Volumes and Initial Molarities

Part	Vol. H_2O_2 (mL)	$[\text{H}_2\text{O}_2]$ Stock (M)	Vol. KI (mL)	$[\text{KI}]$ Stock (M)
A	4.0	_____	1.0	0.50 M
B	4.0	_____	1.0	0.25 M
C	4.0	_____	1.0	0.50 M
D	4.0	_____	1.0	0.50 M

Show calculation work for stock molarities below:

4. **Unit Conversion Derivation:** The pressure sensor measures rate in kPa/s, but kinetics rate laws require M/s ($\text{mol} \cdot \text{L}^{-1} \cdot \text{s}^{-1}$). Rearrange the Ideal Gas Law ($PV = nRT$) to show that:

$$\text{Rate [M/s]} = \frac{\text{Rate [kPa/s]}}{R \times T} \quad (10.2.1)$$

State the correct value and units for R when pressure is expressed in kPa.

This page titled [10.2: Decomposition of Hydrogen Peroxide - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

10.3: Decomposition of Hydrogen Peroxide - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 10.3.1: Experimental Data Measurements

Part	Reactants Used	Temp. (°C)	Initial Rate (kPa/s)
A	4 mL 3.0% H ₂ O ₂ + 1 mL 0.50 M KI	_____	_____
B	4 mL 3.0% H ₂ O ₂ + 1 mL 0.25 M KI	_____	_____
C	4 mL 1.5% H ₂ O ₂ + 1 mL 0.50 M KI	_____	_____
D	4 mL 3.0% H ₂ O ₂ + 1 mL 0.50 M KI (Warm)	_____	_____

Attach printed graphs for all four pressure vs. time runs, clearly showing linear regression fits for the initial rates.

POST-LAB CALCULATIONS AND ANALYSIS

Table 10.3.2: Results Summary

Part	Initial Rate (kPa/s)	Initial Rate (M/s)	[H ₂ O ₂] Mixed (M)	[I ⁻] Mixed (M)	Rate Const. <i>k</i> (units)
A	_____	_____	_____	_____	_____
B	_____	_____	_____	_____	_____
C	_____	_____	_____	_____	_____
D	_____	_____	_____	_____	_____

- Show sample conversion of initial rate from kPa/s to M/s for Part A using $R = 8.314 \text{ L} \cdot \text{kPa} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$:

- Calculate initial concentrations after mixing ($V_{\text{total}} = 5.0 \text{ mL}$) for Part A using $M_1V_1 = M_2V_2$:

3. Determine reaction orders m and n using the logarithmic method:

$$n = \frac{\log(\text{Rate}_2/\text{Rate}_1)}{\log([\text{Reactant}]_2/[\text{Reactant}]_1)} \quad (10.3.1)$$

Round calculated values to the nearest integer (0, 1, or 2) and write the final rate law:

4. Calculate the rate constant (k) for each trial. Show calculation work and unit derivation for Part A:

5. Calculate activation energy (E_a) in kJ/mol using Part A (T_1, k_1) and Part D (T_2, k_2) with $R = 8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$:

This page titled [10.3: Decomposition of Hydrogen Peroxide - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

11: DETERMINATION OF AN ACTIVATION ENERGY

PRE-LAB PREPARATION

1. Prerequisite Math & Theory

- **The Modified Arrhenius Equation:** Because we measure reaction time (t) instead of rate constant (k), the Arrhenius slope equation inverts sign:

$$\ln t = \frac{E_a}{R} \left(\frac{1}{T} \right) + \text{constant} \quad (11.1)$$

- **Slope:** $m = +\frac{E_a}{R}$. A positive slope reflects longer reaction times at lower absolute temperatures.
- **Calculation:** $E_a = \text{Slope} \times 8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$.
- **Catalyst Effect:** A catalyst lowers the activation energy of a reaction. Consequently, the calculated E_a for Part B (Catalyzed) must be **lower** than that for Part A (Uncatalyzed).

2. Required Technical Skills

- Accurately monitoring solution temperatures using a temperature probe.
- Linearizing kinetic data ($\ln t$ vs. $\frac{1}{T}$) using spreadsheet software.

3. Critical Safety

- **Oxidizer Hazard:** Ammonium peroxydisulfate ($(\text{NH}_4)_2\text{S}_2\text{O}_8$) is a strong oxidizing agent. Avoid contact with skin and eyes.
- **Heavy Metal Waste:** Reaction mixtures containing copper catalyst and iodine products must be disposed of in the designated **Heavy Metal Waste** container.

PURPOSE

- To investigate the effects of temperature and a homogeneous catalyst (Cu^{2+}) on the rate of the iodide-peroxydisulfate clock reaction.
- To calculate the activation energy (E_a) for both uncatalyzed and catalyzed pathways by measuring reaction times across a range of absolute temperatures.
- To apply the logarithmic Arrhenius equation ($\ln t = \frac{E_a}{R} \left(\frac{1}{T} \right) + \text{constant}$) via graphical analysis of $\ln t$ versus $\frac{1}{T}$.

INTRODUCTION

Reaction rates depend strongly on **temperature** and the presence of a **catalyst**. From kinetic theory, only molecular collisions exceeding a minimum threshold energy—known as the **activation energy** (E_a)—result in reaction, forming an activated complex. The Maxwell-Boltzmann distribution dictates that the fraction of molecules possessing this minimum energy increases exponentially with absolute temperature (T).

A **catalyst** accelerates a reaction by providing an alternative reaction pathway with a lower activation energy, allowing a larger fraction of molecules to react at any given temperature. The temperature dependence of the rate constant k is described by the Arrhenius equation:

$$k = A \exp \left(-\frac{E_a}{RT} \right) \quad (11.2)$$

where R is the gas constant ($8.314 \text{ J/mol} \cdot \text{K}$) and A is the frequency factor. Taking the natural logarithm yields the linear form:

$$\ln k = -\frac{E_a}{R} \left(\frac{1}{T} \right) + \ln A \quad (11.3)$$

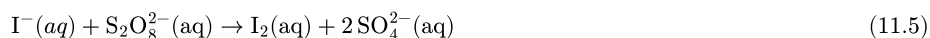
Because the reaction rate constant k is inversely proportional to reaction time (t required to consume a fixed amount of reactant), substituting $k \propto 1/t$ transforms the linear equation into:

$$\ln t = \frac{E_a}{R} \left(\frac{1}{T} \right) + \text{constant} \quad (11.4)$$

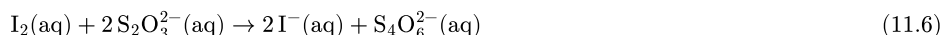
Thus, graphing $\ln t$ **on the y-axis against $\frac{1}{T}$ (in K^{-1}) on the x-axis** yields a straight line with a **positive slope** $m = \frac{E_a}{R}$.

The Iodine Clock Reaction

This experiment examines the oxidation of iodide (I^-) by peroxydisulfate ($\text{S}_2\text{O}_8^{2-}$):



To measure the reaction time conveniently, a small, precisely measured amount of sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$) and starch indicator are added. Thiosulfate instantly consumes iodine as it is formed:



As long as thiosulfate remains in solution, no free iodine accumulates. The instant thiosulfate is completely exhausted, the next drop of generated iodine reacts with starch, producing an abrupt, deep blue-black color change that marks the exact reaction time t .

DATA PREP: TIME VS. RATE CONSTANT

Because reaction time t is inversely proportional to rate ($k \propto 1/t$), substituting into logarithmic form gives $\ln(1/t) = -\ln t$.

- **Slope Sign:** Consequently, plotting $\ln t$ vs. $\frac{1}{T}$ produces a **positive slope** ($+\frac{E_a}{R}$), unlike a standard Arrhenius plot of $\ln k$ vs. $\frac{1}{T}$ which has a negative slope.

KEY EQUATIONS

Arrhenius Activation Energy:

$$E_a = \text{Slope} \times R \quad (11.7)$$

[11.1: Determination of an Activation Energy - Experiment](#)

[11.2: Determination of an Activation Energy - Pre-lab](#)

[11.3: Determination of an Activation Energy - Data and Report](#)

11.1: Determination of an Activation Energy - Experiment

SAFETY PRECAUTIONS

- **Chemical Hazards:** Ammonium peroxydisulfate ($(\text{NH}_4)_2\text{S}_2\text{O}_8$) is a strong oxidizing agent. Copper(II) sulfate is toxic. Wear splash goggles, nitrile gloves, and a lab coat.
- **Thermal Hazards:** Exercise caution when handling hot water baths ($\sim 50^\circ\text{C}$) and ice-water baths to avoid thermal burns or frostbite.

EQUIPMENT AND CHEMICALS NEEDED

Table 11.1.1: Materials List

Equipment & Hardware	Chemicals & Solutions
Burets (3) & buret clamps	0.200 M Potassium iodide, KI
50-mL Erlenmeyer flasks (2)	0.100 M Ammonium peroxydisulfate, $(\text{NH}_4)_2\text{S}_2\text{O}_8$
250-mL beaker (reaction vessel)	0.0050 M Sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$
400-mL beaker (temperature bath)	Starch indicator solution
Digital thermometer or temperature probe	0.1 M Copper(II) sulfate, CuSO_4 (catalyst)
Stopwatch / timer	

EXPERIMENTAL PROCEDURE

Part A: Determine the Effect of Temperature (Uncatalyzed)

For all reaction mixtures in Part A, maintain identical initial reactant volumes:

- 20.00 mL of 0.200 M potassium iodide (KI)
- 20.00 mL of 0.100 M ammonium peroxydisulfate ($(\text{NH}_4)_2\text{S}_2\text{O}_8$)
- 10.00 mL of 0.0050 M sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$)
- 3 drops of starch indicator solution

Table 11.1.2: Part A Experimental Procedures

Mixture	Procedure & Temperature Control
Mixture 1 (Room Temp)	- Fill three burets with 0.200 M KI, 0.100 M $(\text{NH}_4)_2\text{S}_2\text{O}_8$, and 0.0050 M $\text{Na}_2\text{S}_2\text{O}_3$. - Measure exactly 20.00 mL of KI into one clean 50-mL Erlenmeyer flask and 20.00 mL of $(\text{NH}_4)_2\text{S}_2\text{O}_8$ into a second flask. - In a 250-mL reaction beaker, add 10.00 mL of 0.0050 M $\text{Na}_2\text{S}_2\text{O}_3$ and 3 drops of starch indicator. Insert a thermometer. - Quickly pour both reactant solutions into the reaction beaker, swirl to mix, and start the timer immediately. - Record the initial temperature. Stop the timer the exact instant the solution turns blue-black. Record the final temperature. Use the average temperature for calculations. - Repeat to check duplicate times within ± 1 second.
Mixture 2 (Warm, $\sim 30\text{--}35^\circ\text{C}$)	Warm reactant flasks and the reaction beaker in a warm water bath until final temperature reaches $30\text{--}35^\circ\text{C}$. Mix, time, and record temperatures.
Mixture 3 (Hot, $\sim 45\text{--}50^\circ\text{C}$)	Warm reactant flasks and reaction beaker in a hot water bath until final temperature reaches $45\text{--}50^\circ\text{C}$. Mix, time, and record temperatures.
Mixture 4 (Cool, $\sim 10\text{--}15^\circ\text{C}$)	Cool reactant flasks and reaction beaker in an ice-water bath until temperature reaches $10\text{--}15^\circ\text{C}$. Mix, time, and record temperatures.
Mixture 5 (Cold, $\sim 3\text{--}7^\circ\text{C}$)	Cool reactant flasks and reaction beaker further in an ice-water bath to reach $3\text{--}7^\circ\text{C}$. Mix, time, and record temperatures.

DATA SKILL: REALITY CHECK

Reaction rates increase with temperature, meaning reaction times must **decrease** as temperature increases. If a higher temperature trial takes longer than a cooler one, repeat the measurement.

Part B: Determine the Effect of a Catalyst

- Prepare 150 mL of catalyzed peroxydisulfate solution by adding 3 drops of 0.1 M CuSO_4 (Cu^{2+}) to stock peroxydisulfate in a beaker. Mix thoroughly and refill the buret.
- Perform rate measurements (Mixtures 6–10) using this catalyzed peroxydisulfate solution at five temperatures spanning the same range ($\sim 5^\circ\text{C}$ to $\sim 48^\circ\text{C}$) as in Part A.

Table 11.1.3: Summary of Experimental Conditions

Reaction Series	Mixture Numbers	Catalyst Status	Target Temperature Range
Part A: Uncatalyzed	1 – 5	None	Ambient to $3\text{--}50^\circ\text{C}$
Part B: Catalyzed	6 – 10	Added Cu^{2+}	Ambient to $3\text{--}50^\circ\text{C}$

CHEMICAL DISPOSAL

- Dispose of all chemical waste into the designated **Heavy Metal / Inorganic Waste** container.

This page titled [11.1: Determination of an Activation Energy - Experiment](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

11.2: Determination of an Activation Energy - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

1. To determine activation energy (E_a), what two mathematical variables are plotted against each other, and how is the slope of the resulting straight line used with the gas constant (R) to calculate E_a ? Provide the governing equation.

2. If the slope of an Arrhenius plot ($\ln t$ vs. $\frac{1}{T}$) is $+6.4 \times 10^3 \text{ K}$, calculate the activation energy (E_a) in kJ/mol ($R = 8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$).

3. In this iodine clock reaction, describe the specific chemical roles of sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$) and starch indicator in making reaction rates visually measurable.

4. What specific visual signal marks the exact endpoint of the timed reaction?

5. **Graphical Prediction:** You will plot both uncatalyzed and catalyzed data sets on the same graph:
- Which line should exhibit a steeper slope? (Recall: Steeper slope = Higher E_a).
 - Which line should appear lower on the y-axis (representing faster reaction times t) at any given temperature?

This page titled [11.2: Determination of an Activation Energy - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

11.3: Determination of an Activation Energy - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Table 11.3.1: Experimental Temperature and Time Data

Reaction Series	Mixture No.	Average Temp, T (°C)	Time, t (s)
A. Uncatalyzed	1	_____	_____
	2	_____	_____
	3	_____	_____
	4	_____	_____
	5	_____	_____
B. Catalyzed	6	_____	_____
	7	_____	_____
	8	_____	_____
	9	_____	_____
	10	_____	_____

POST-LAB CALCULATIONS AND ANALYSIS

Use a spreadsheet program (Excel or Google Sheets) to perform the following analysis:

1. Convert all Celsius temperatures to Kelvin ($T = ^\circ\text{C} + 273.15$).
2. Calculate $\ln t$ and reciprocal temperature $\frac{1}{T}$ (in K^{-1}) for each mixture.
3. Plot $\ln t$ (y-axis) versus $\frac{1}{T}$ (x-axis) for both uncatalyzed and catalyzed mixtures on the same graph. Include linear trendlines, equations, and R^2 values.
4. Calculate E_a in kJ/mol for both series from the slopes ($E_a = \text{Slope} \times 8.314/1000$).

Table 11.3.2: Arrhenius Analysis Results Summary

Calculated Parameter	Uncatalyzed Series	Catalyzed Series
Trendline Slope (K)	_____	_____
Trendline y-Intercept	_____	_____
Activation Energy, E_a (kJ/mol)	_____	_____

Insert your combined Arrhenius plot here:

5. **Rule of Thumb Evaluation:** Using your data for uncatalyzed mixtures 1–5, test the common rule of thumb that "reaction rate doubles for every 10 °C temperature increase." Calculate reaction time ratios ($t_{\text{lower}}/t_{\text{higher}}$) across 10 °C intervals and comment on its validity for this specific reaction.

6. **Catalyst Mechanism:** Based on your experimental activation energies, explain specifically how the Cu^{2+} catalyst alters the reaction rate on a molecular level.

This page titled [11.3: Determination of an Activation Energy - Data and Report](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

12: RADIOACTIVITY SIMULATION

PRE-LAB PREPARATION

1. Prerequisite Math & Theory

- **First-Order Decay:** Radioactive decay follows the same mathematical formulation as first-order kinetics:

$$\ln N_t = -\lambda t + \ln N_0 \quad (12.1)$$

- **y-axis:** $\ln(\text{Number of Dice Remaining})$
- **x-axis:** Round Number (t)
- **Slope:** $-\lambda$ (Decay constant)
- **Half-Life Math:** $t_{1/2} = \frac{\ln 2}{\lambda}$. Knowing the decay constant λ allows calculation of the half-life.

2. Required Technical Skills

- Simulation logic: Recognizing that repeated trials with a small population (e.g., rolling 10 dice 10 times) statistically mirrors rolling a large population (100 dice) once.
- Graphical analysis: Linearizing decay data ($\ln N$ vs. Round Number) in spreadsheet software to extract the decay constant.

3. Critical Safety

- **Choking Hazard:** Dice are small parts. Keep them secure and never ingest them.
- **Slip Hazard:** Contain dice within trays or boxes during rolling to prevent floor spills and tripping hazards.

PURPOSE

- To simulate radioactive decay using a random probabilistic process (rolling dice).
- To investigate the concept of half-life ($t_{1/2}$) by observing the exponential decrease of a population over successive rounds.
- To determine the experimental decay constant (λ) and half-life via graphical and statistical analysis.

INTRODUCTION

Radioactive decay is a spontaneous and completely random process. It is impossible to predict the exact moment an individual radioactive nucleus will disintegrate; however, the probability of decay over a given time interval can be modeled using statistical and probabilistic simulations such as coin tosses or rolling dice.

Radioactive isotopes decay at widely varying rates, characterized by their **half-life** ($t_{1/2}$)—the time required for exactly one-half of the radioactive nuclei in a given sample to decay. For example, Polonium-218 has a half-life of about 3 minutes, whereas Uranium-238 has a half-life exceeding 4 billion years. Regardless of initial sample size, the fraction remaining always decreases by 50% per half-life.

In this experiment, radioactive decay is simulated by rolling dice. Designating a specific outcome (e.g., landing on a '6') as "decay" models the probabilistic nature of nuclear disintegration.

Table 12.1: Simulated "Radioactive Decay" of Coins

Round (Coin Toss)	1	2	3	4	5	6
Initial Coins	100	50	25	12	6	3

Round (Coin Toss)	1	2	3	4	5	6
Coins Decayed	50	25	13	6	3	2
Coins Remaining	50	25	12	6	3	1

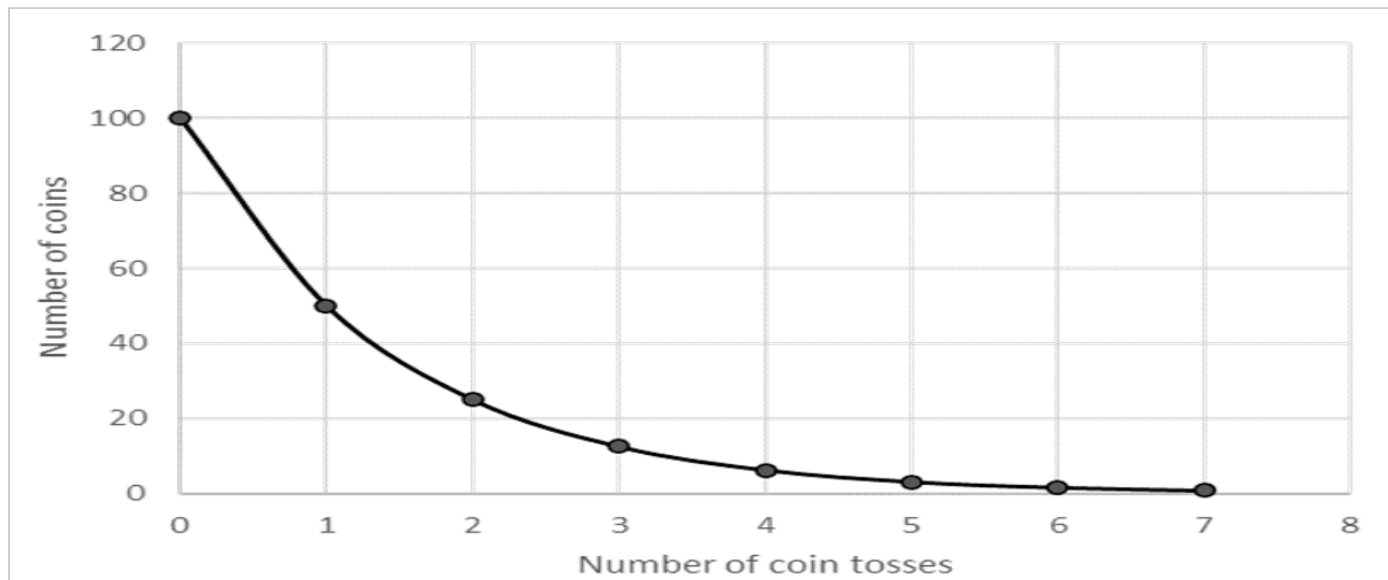


Figure 12.1: Simulated radioactive decay curve of coins.

DATA PREP: KINETICS REVISITED

Radioactive decay is mathematically identical to **first-order kinetics**:

- **Rate Law:** $\text{Rate} = \lambda N$, where the rate of decay depends directly on the number of active nuclei (dice) present.
- **Linearization:** Plotting the natural logarithm of remaining dice ($\ln N_t$) versus time (t) yields a straight line with slope $-\lambda$.

KEY EQUATIONS

First-Order Decay Law:

$$N_t = N_0 e^{-\lambda t} \quad \text{or} \quad t = -\frac{1}{\lambda} \ln \left(\frac{N_t}{N_0} \right) \quad (12.2)$$

Half-Life Relationship:

$$t_{1/2} = \frac{\ln 2}{\lambda} \quad (12.3)$$

[12.1: Radioactivity Simulation - Experiment](#)

[12.2: Radioactivity Simulation - Pre-lab](#)

[12.3: Radioactivity Simulation - Data and Report](#)

12.1: Radioactivity Simulation - Experiment

SAFETY PRECAUTIONS

- Dice are small objects and represent a choking hazard. Never put them in your mouth.
- Perform all rolls inside a confined box or tray to prevent dice from falling onto the floor and creating slipping hazards.

EQUIPMENT AND CHEMICALS NEEDED

Table 12.1.1: Materials List

Equipment & Supplies
10 multi-sided dice (4-sided, 6-sided, 8-sided, 10-sided, 12-sided, or 20-sided)
Shaking box or small container for randomized rolling

EXPERIMENTAL PROCEDURE

Part A: Simulating Radioactive Decay with Dice

1. Obtain your assigned set of 10 multi-sided dice and note the designated "decay number" assigned by your instructor. Record the number of sides and the decay number in Data Table 1.
2. Place all 10 dice in your container, shake vigorously for 3 seconds, and roll them onto a flat surface. Ensure all dice land flat. Re-roll any cocked dice leaning against objects.

TECHNIQUE ALERT: RANDOMNESS IS KEY

Ensuring valid data collection:

- **Vigorous Shaking:** Always shake the container thoroughly to ensure true randomization before every roll.
- **Rigorous Counting:** Carefully tally every die showing the decay number. Missing a decayed die artificially inflates remaining populations (N_t) and skews half-life calculations.

3. Identify and count any dice displaying the assigned decay number as "decayed." Record this count.
4. Repeat rolling the 10 dice a total of 10 times (representing a combined population of 100 rolls for Round 1). Record the number of decayed dice for each individual roll.
5. Sum the total number of decayed dice across all 10 rolls to establish the total number of decayed dice for Round 1.
6. Subtract the total decayed dice from the starting population (100 for Round 1) to determine the number of dice remaining. This remainder serves as the starting population for Round 2.
7. For subsequent rounds, roll the 10-dice set as many full sets of 10 as needed, plus a final partial roll, to match your current starting population. For example, if 79 dice remain, roll the 10-dice set 7 times and roll 9 dice once ($79 = 10 \times 7 + 9$).

DATA SKILL: SAMPLE SIZE NOISE

The Law of Large Numbers: Probabilistic models perform best with massive sample sizes.

- **Early Rounds:** With large populations (near 100), data closely hugs the smooth theoretical decay curve.
- **Late Rounds:** As remaining dice drop below 10, statistical noise increases significantly because a single random roll exerts a massive percentage impact.

8. Continue this procedure until no dice remain or until 20 rounds of decay have been completed.

12.2: Radioactivity Simulation - Pre-lab

Name: _____ Date: _____

PRE-LAB QUESTIONS

- Radioactive decay follows first-order kinetics. State the mathematical relationship connecting half-life ($t_{1/2}$) and the decay constant (λ).

- Strontium-90 is a radioactive isotope with a half-life of 29 years. Assume an initial sample of 10,000 atoms of Sr-90.
 - Complete Table 12.2.1 to determine the elapsed time when 1250 atoms of Sr-90 remain.

Table 12.2.1: Strontium-90 Radioactive Decay Data

Number of Half-Lives	0	1	2	3	4	5
Number of Years	0	29	_____	_____	_____	_____
Number of Sr-90 Atoms	10,000	_____	_____	_____	_____	_____

- Graph the Sr-90 decay data and fit a smooth exponential curve in spreadsheet software. Insert your graph here:

- Describe the geometric shape of the radioactive decay curve.

- What is the exact probability that a **specific Sr-90 atom** will decay after (a) 29 years and (b) 58 years?

- Theoretical Prediction (Dice Simulation):** Assume standard 6-sided dice.
 - If 1 side represents "decay" and 5 sides represent "safe," what is the probability P of decay per roll?

- Calculate the theoretical decay constant: $\lambda_{\text{theo}} = -\ln(1 - P) = -\ln\left(\frac{\text{Sides}-1}{\text{Sides}}\right)$.

- Calculate the theoretical half-life: $t_{1/2} = \frac{\ln 2}{\lambda_{\text{theo}}}$.

This page titled [12.2: Radioactivity Simulation - Pre-lab](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

- [2.3: Determination of an Equilibrium Constant - Data and Report](#) by Vince Hradil is licensed [CC BY 4.0](#).

12.3: Radioactivity Simulation - Data and Report

Name: _____ Partner(s): _____ Date: _____

DATA AND OBSERVATIONS

Number of sides on the dice: _____

Assigned "decay number": _____

Table 12.3.1: Dice Decay Experimental Data

Round	Initial Dice	Total Decayed	Dice Remaining	Roll Decays (10 Individual Rolls)
0	—	—	100	—
1	100	_____	_____	_____
2	_____	_____	_____	_____
3	_____	_____	_____	_____
4	_____	_____	_____	_____
5	_____	_____	_____	_____
6	_____	_____	_____	_____
7	_____	_____	_____	_____
8	_____	_____	_____	_____
9	_____	_____	_____	_____
10	_____	_____	_____	_____
11	_____	_____	_____	_____
12	_____	_____	_____	_____
13	_____	_____	_____	_____
14	_____	_____	_____	_____
15	_____	_____	_____	_____
16	_____	_____	_____	_____
17	_____	_____	_____	_____
18	_____	_____	_____	_____

Round	Initial Dice	Total Decayed	Dice Remaining	Roll Decays (10 Individual Rolls)
19	_____	_____	_____	_____
20	_____	_____	_____	_____

POST-LAB QUESTIONS

- Plot the experimental dice decay data in spreadsheet software as a scatterplot with symbols only (include Round 0 at $N_0 = 100$). Insert your graph here:

DATA SKILL: THE LAW OF LARGE NUMBERS

Notice any "bumpiness" in your decay curve? Small sample sizes (100 dice) generate statistical noise compared to massive nuclear populations, reflecting a fundamental property of probability rather than human error.

- Determine the experimental decay constant (λ): Add an exponential trendline to your graph, display the equation, display the R^2 value, and set the intercept fixed at $N_0 = 100$.

DATA SKILL: EXPERIMENT VS. THEORY

Compare your experimental λ to the theoretical probability value: $\lambda_{\text{theo}} = -\ln\left(\frac{\text{Sides}-1}{\text{Sides}}\right)$. Calculate percent error.

- Use the fitted decay constant to calculate the experimental half-life ($t_{1/2}$).

- Does the half-life depend on the chosen decay number? Explain based on probability.

- Predict the remaining number of dice after 15 rounds using your decay constant. Compare this prediction with your actual experimental data for round 15 and discuss potential sources of variance.

- Compare the decay curve of the dice simulation with the radioactive decay curve of Sr-90. Would the curve shape differ if you started with 1000 dice instead of 100?

13: APPENDICES

13.1: Vernier LabQuest 3 Instructions

13.2: Tables

13.2.1: The Periodic Table

13.2.2: Fundamental Physical Constants

13.2.3: Water Properties

13.2.4: Composition of Commercial Acids and Bases

13.2.5: Standard Thermodynamic Properties for Selected Substances

13.2.6: Ionization Constants of Weak Acids

13.2.7: Ionization Constants of Weak Bases

13.2.8: Solubility Products

13.2.9: Formation Constants for Complex Ions

13.2.10: Standard Electrode (Half-Cell) Potentials

13.2.11: Half-Lives for Several Radioactive Isotopes

13.1: Vernier LabQuest 3 Instructions

This appendix contains step-by-step instructions for the common tasks used in General Chemistry II. Refer to these sections whenever a lab procedure directs you to "Calibrate" or "Setup" a sensor.

A. Temperature Probe

Measuring Temperature

The Stainless Steel Temperature Probe is auto-identified by the LabQuest.

1. Plug the probe into **CH 1**.
 2. The screen will display the live temperature in °C.
 3. **Note:** You typically do *not* need to press "Start/Collect" to read a temperature. Just use the live meter readout.
 4. **Calibration:** These probes are factory-calibrated and rarely need adjustment. If the reading seems wrong (e.g., room temperature is 80°C), check that the probe is not set to °F or K in the units menu.
-

B. pH Sensor

Two-Point pH Calibration

pH probes drift over time. For accurate K_a or titration data, you must calibrate before use.

1. Connect the pH Sensor to **CH 1**.
 2. Tap the red **pH** box on the meter screen → Select **Calibrate**.
 3. Tap **Calibrate Now**.
 4. **First Point (Acidic):**
 - Rinse the probe with distilled water and place it in the **pH 4.00** buffer solution.
 - Wait for the voltage reading to stabilize.
 - Enter **4.00** in the "Value 1" field.
 - Tap **Keep**.
 5. **Second Point (Neutral):**
 - Rinse the probe and place it in the **pH 7.00** buffer solution.
 - Wait for the voltage to stabilize.
 - Enter **7.00** in the "Value 2" field.
 - Tap **Keep**.
 6. Tap **OK**. The probe is now calibrated.
-

C. Drop Counter (Automated Titration)

Calibrating Drops/mL

The "Drop Counter" counts drops, but you need to know Volume (mL). Because every plastic reagent reservoir tip is slightly different, you must tell the device exactly how large your drops are.

1. Connect the Drop Counter to **DIG 1**.

2. Tap the **Mode** box → Select **Drop Counting**.
3. Tap the red **Drop Count** box → Select **Calibrate** → **Calibrate Now**.
4. **The Procedure:**
 - Place a clean **10-mL graduated cylinder** directly under the drop tip.
 - Open the bottom valve fully to establish a steady, slow drop rate (about 1-2 drops per second).
 - Allow the drops to fall until the cylinder is nearly full (between 9 and 10 mL).
 - **Close the valve completely.**
5. **The Calculation:**
 - Read the exact volume in the graduated cylinder (e.g., 9.45 mL).
 - Enter this value into the LabQuest.
 - Tap **Keep**.
6. The device now calculates a precise "Drops/mL" ratio for your specific setup.

D. Manual Data Entry

"Events with Entry" Mode

Use this mode when you want to type in one variable (like Concentration or Volume) while the sensor records the other (like Absorbance or pH).

1. Tap the **Mode** box (default is usually "Time Based").
2. Select **Events with Entry** from the dropdown menu.
3. **Configure the X-Axis:**
 - **Name:** Enter the variable you will type (e.g., "Concentration").
 - **Units:** Enter the units (e.g., "mol/L" or "M").
4. Tap **OK**.
5. **How to Collect Data:**
 - Tap **Start/Collect**.
 - When the sensor reading stabilizes, tap **Keep**.
 - A keyboard will appear. Type in the value for your X-axis (e.g., "0.05") and tap **OK**.
 - Repeat for all samples. Tap **Stop** when finished.

E. Spectrometer (SpectroVis Plus)

Calibration & Setup

Step 1: Calibration (Blanking)

1. Allow the lamp to warm up for 2-3 minutes.
2. Fill a cuvette 3/4 full with your blank solvent (usually water).
3. Insert the cuvette so the clear sides face the arrow/light path.
4. Tap the red **Absorbance** box → **Calibrate**.
5. Tap **Finish Calibration** → **OK**. (The reading should be 0.000).

Step 2: Finding $\lambda_{\max}$ (Peak Wavelength)

1. Insert a cuvette containing your most concentrated colored solution.
2. Tap **Start/Collect**. A spectrum graph will appear.
3. Tap **Stop**.
4. Tap the highest point on the curve. The wavelength (nm) and absorbance will appear on the right. Note this wavelength.

Step 3: Setting Up for Beer's Law

1. Tap the **Meter** tab (gauge icon).
2. Tap the **Mode** box.
3. Select **Events with Entry** (See Appendix A.5).
4. The device will now lock onto the $\lambda_{\max}$ you found in Step 2.

F. Linear Fit (Regression)

Finding Slope and Intercept

To find the best-fit line for your data (e.g., Beer's Law or Kinetics):

1. Tap the **Graph** tab.
2. Tap **Analyze** → **Curve Fit**.
3. Select the checkbox for your data series (usually "Absorbance").
4. From the "Fit Equation" dropdown, select **Linear**.
5. The screen will display the equation form $y = mx + b$:
 - **m (Slope):** The rate of change or molar absorptivity.
 - **b (Intercept):** The y-value at zero.
 - **r (Correlation):** A value close to 1.000 indicates a good fit.
6. Tap **OK** to display the line on the graph.

G. Time-Based Settings (Kinetics)

Changing Duration & Rate

For Kinetics or Temperature experiments, you often need to adjust how long the sensor records.

1. Tap the **Meter** tab.
2. Tap the **Mode** box (it usually says "Time Based").
3. **Duration:** Change this to match your reaction time (e.g., change "180 s" to "600 s").
4. **Rate/Interval:**
 - **Rate:** Samples per second (e.g., 2 samples/s).
 - **Interval:** Time between samples (e.g., 0.5 s/sample).
5. Tap **OK**.

H. Managing Data Runs

Storing & Hiding Runs

If you want to do multiple trials (e.g., Trial 1 and Trial 2) without losing your first graph:

1. **Store a Run:** Tap the **File Cabinet Icon** (next to the "Run 1" label). The screen will clear and prepare for "Run 2." Your old data is safe.
2. **Viewing Multiple Runs:**
 - Tap the **Graph** tab.
 - Tap the **Run Number** button (e.g., "Run 2").
 - Select **All Runs** to see them overlayed.
3. **Strike Through Data (Outliers):**

- Tap the **Table** tab.
- Tap and drag to select the "bad" data cells.
- Tap **Table** (menu) → **Strike Through Data**.
- The graph will ignore these points, but the data remains visible with a line through it.

I. Graph Analysis Tools

Statistics & Tangents

1. Statistics (Mean/Max/Min)

1. On the **Graph** tab, drag across the region of interest.
2. Tap **Analyze** → **Statistics**.
3. Check the box for your data column.
4. A box will appear showing the Mean (Average), Min, and Max for that selection.

2. Tangent (Finding Slope at a Point)

Use this for Titrations to find the steepest part of the curve (Equivalence Point).

1. Tap **Analyze** → **Tangent**.
2. Tap anywhere on the graph curve.
3. A crosshair will appear. The number displayed is the **Slope** at that exact instant.
4. For titrations, the Equivalence Point is where the slope is **highest** (maximum value).

J. Exporting Data

Emailing Files for Excel

To analyze your data in Excel or Google Sheets, you must email the text file to yourself.

1. Ensure the LabQuest is connected to WiFi (Connections icon).
2. Tap **File** → **Email**.
3. **Option A: Text File (Recommended for Excel)**
 - Tap the **Text** icon.
 - Enter your email address and subject.
 - Tap **Send**. (You can open the .txt file in Excel).
4. **Option B: Screenshot (For Lab Reports)**
 - Tap the **Screenshot** icon (Camera).
 - Enter your email and send.

13.2: TABLES

13.2.1: The Periodic Table

13.2.2: Fundamental Physical Constants

13.2.3: Water Properties

13.2.4: Composition of Commercial Acids and Bases

13.2.5: Standard Thermodynamic Properties for Selected Substances

13.2.6: Ionization Constants of Weak Acids

13.2.7: Ionization Constants of Weak Bases

13.2.8: Solubility Products

13.2.9: Formation Constants for Complex Ions

13.2.10: Standard Electrode (Half-Cell) Potentials

13.2.11: Half-Lives for Several Radioactive Isotopes

13.2.1: The Periodic Table

Period

Group

1

2

3

4

5

6

7

8

9

10

11

12

13

14

15

16

17

18

1

H

1.008

hydrogen

2

He

4.003

helium

3

Li

6.94

lithium

4

Be

9.012

beryllium

5

B

10.81

boron

6

C

12.01

carbon

7

N

14.01

nitrogen

8

O

16.00

oxygen

9

F

18.99

fluorine

10

Ne

20.18

neon

11

Na

22.99

sodium

12

Mg

24.31

magnesium

13

Al

26.98

aluminum

14

Si

28.09

silicon

15

P

30.97

phosphorus

16

S

32.06

sulfur

17

Cl

35.45

chlorine

18

Ar

39.95

argon

19

K

39.10

potassium

20

Ca

40.08

calcium

21

Sc

44.96

scandium

22

Ti

47.87

titanium

23

V

50.94

vanadium

24

Cr

52.00

chromium

25

Mn

54.94

manganese

26

Fe

55.85

iron

27

Co

58.93

cobalt

28

Ni

58.69

nickel

29

Cu

63.55

copper

30

Zn

65.38

zinc

31

Ga

69.72

gallium

32

Ge

72.63

germanium

33

As

74.92

arsenic

34

Se

78.97

selenium

35

Br

79.90

bromine

36

Kr

83.80

krypton

37

Rb

85.47

rubidium

38

Sr

87.62

strontium

39

Y

88.91

yttrium

40

Zr

91.22

zirconium

41

Nb

92.91

niobium

42

Mo

95.95

molybdenum

43

Tc

[97]

technetium

44

Ru

101.1

ruthenium

45

Rh

102.9

rhodium

46

Pd

106.4

palladium

47

Ag

107.9

silver

48

Cd

112.4

cadmium

49

In

114.8

indium

50

Sn

118.7

tin

51

Sb

121.8

antimony

52

Te

127.6

tellurium

53

I

126.9

iodine

54

Xe

131.3

xenon

55

Cs

132.9

cesium

56

Ba

137.3

barium

57-71

La-Lu

lanthanum series

72

Hf

178.5

hafnium

73

Ta

180.9

tantalum

74

W

183.8

tungsten

75

Re

186.2

rhenium

76

Os

190.2

osmium

77

Ir

192.2

iridium

78

Pt

195.1

platinum

79

Au

197.0

gold

80

Hg

200.6

mercury

81

Tl

204.4

thallium

82

Pb

207.2

lead

83

Bi

209.0

bismuth

84

Po

[209]

polonium

85

At

[210]

astatine

86

Rn

[222]

radon

87

Fr

[223]

francium

88

Ra

[226]

radium

89-103

Ac-Lr

actinide series

104

Rf

[261]

rutherfordium

105

Db

[262]

dubnium

106

Sg

[266]

seaborgium

107

Bh

[270]

bohrium

108

Hs

[277]

hassium

109

Mt

[276]

meitnerium

110

Ds

[281]

darmstadtium

111

Rg

[282]

roentgenium

112

Cn

[285]

copernicium

113

Uut

[288]

ununtrium

114

Fl

[289]

flerovium

115

Uup

[294]

ununpentium

116

Lv

[293]

livermorium

117

Uus

[294]

ununseptium

118

Uuo

[294]

ununoctium

57

La

138.9

lanthanum

58

Ce

140.1

cerium

59

Pr

140.9

praseodymium

60

Nd

144.2

neodymium

61

Pm

[145]

promethium

62

Sm

150.4

samarium

63

Eu

152.0

europium

64

Gd

157.3

gadolinium

65

Tb

158.9

terbium

66

Dy

162.5

dysprosium

67

Ho

164.9

holmium

68

Er

167.3

erbium

69

Tm

168.9

thulium

70

Yb

173.1

ytterbium

71

Lu

175.0

lutetium

89

Ac

[227]

actinium

90

Th

232.0

thorium

91

Pa

231.0

protactinium

92

U

238.0

uranium

93

Np

[237]

neptunium

94

Pu

[244]

plutonium

95

Am

[243]

americium

96

Cm

[247]

curium

97

Bk

[247]

berkelium

98

Cf

[251]

californium

99

Es

[252]

einsteinium

100

Fm

[257]

fermium

101

Md

[258]

mendelevium

102

No

[259]

nobelium

103

Lr

[262]

lawrencium

Atomic number

1

Symbol

H

Atomic mass

1.008

Name

hydrogen

Color Code

Metal

Solid

Metalloid

Liquid

Nonmetal

Gas

This page titled 13.2.1: The Periodic Table is shared under a CC BY 4.0 license and was authored, remixed, and/or curated by Vince Hradil.

13.2.2: Fundamental Physical Constants

Fundamental Physical Constants	
Name and Symbol	Value
atomic mass unit (amu)	$1.6605402 \times 10^{-27} \text{ kg}$
Avogadro's number	$6.0221367 \times 10^{23} \text{ mol}^{-1}$
Boltzmann's constant (k)	$1.380658 \times 10^{-23} \text{ J K}^{-1}$
charge-to-mass ratio for electron (e/m_e)	$1.75881962 \times 10^{11} \text{ C kg}^{-1}$
electron charge (e)	$1.60217733 \times 10^{-19} \text{ C}$
electron rest mass (m_e)	$9.1093897 \times 10^{-31} \text{ kg}$
Faraday's constant (F)	$9.6485309 \times 10^4 \text{ C mol}^{-1}$
gas constant (R)	$8.205784 \times 10^{-2} \text{ L atm mol}^{-1} \text{ K}^{-1} = 8.314510 \text{ J mol}^{-1} \text{ K}^{-1}$
molar volume of an ideal gas, 1 atm, 0 °C	$22.41409 \text{ L mol}^{-1}$
molar volume of an ideal gas, 1 bar, 0 °C	$22.71108 \text{ L mol}^{-1}$
neutron rest mass (m_n)	$1.6749274 \times 10^{-27} \text{ kg}$
Planck's constant (h)	$6.6260755 \times 10^{-34} \text{ J s}$
proton rest mass (m_p)	$1.6726231 \times 10^{-27} \text{ kg}$
Rydberg constant (R)	$1.0973731534 \times 10^7 \text{ m}^{-1} = 2.1798736 \times 10^{-18} \text{ J}$
speed of light (in vacuum) (c)	$2.99792458 \times 10^8 \text{ m s}^{-1}$

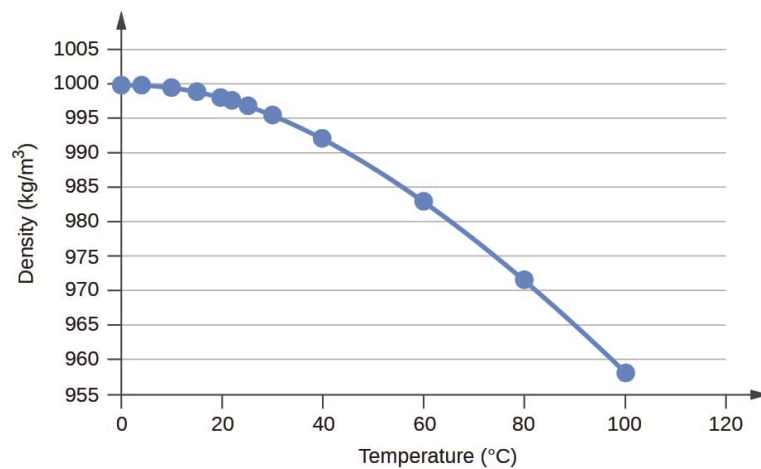
This page titled [13.2.2: Fundamental Physical Constants](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

13.2.3: Water Properties

Water Density (kg/m^3) at Different Temperatures ($^{\circ}\text{C}$)

Temperature ¹	Density
0	999.8395
4	999.9720 (density maximum)
10	999.7026
15	999.1026
20	998.2071
22	997.7735
25	997.0479
30	995.6502
40	992.2
60	983.2
80	971.8
100	958.4

Density of Water as a Function of Temperature

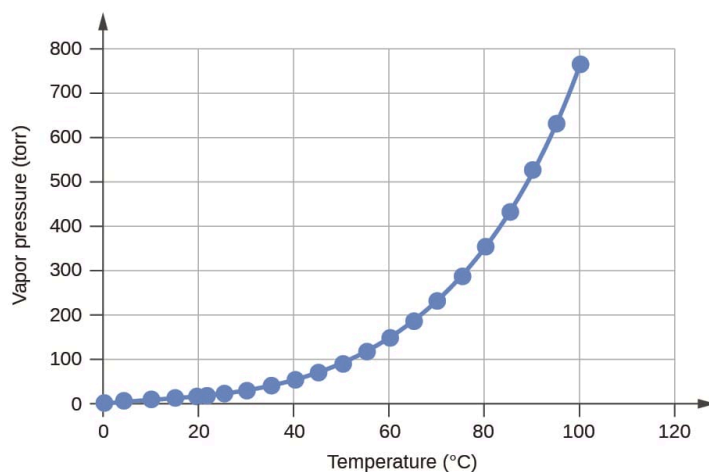


Water Vapor Pressure at Different Temperatures ($^{\circ}\text{C}$)

Temperature	Vapor Pressure (torr)	Vapor Pressure (Pa)
0	4.6	613.2812
4	6.1	813.2642
10	9.2	1226.562
15	12.8	1706.522
20	17.5	2333.135

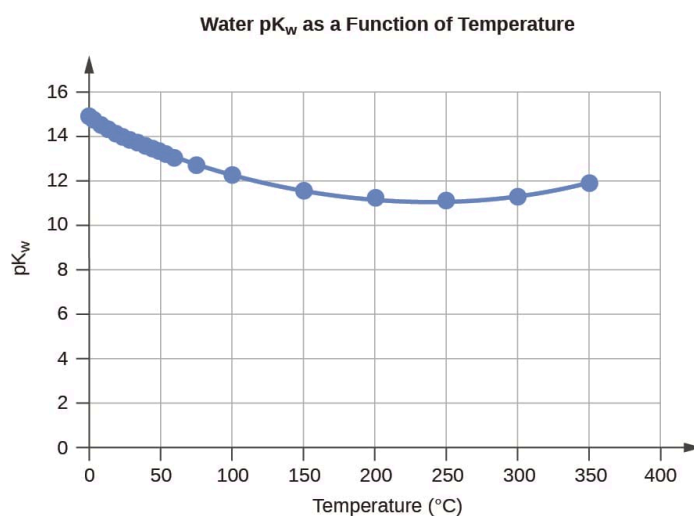
Water Vapor Pressure at Different Temperatures (°C)		
Temperature	Vapor Pressure (torr)	Vapor Pressure (Pa)
22	19.8	2639.776
25	23.8	3173.064
30	31.8	4239.64
35	42.2	5626.188
40	55.3	7372.707
45	71.9	9585.852
50	92.5	12332.29
55	118.0	15732
60	149.4	19918.31
65	187.5	24997.88
70	233.7	31157.35
75	289.1	38543.39
80	355.1	47342.64
85	433.6	57808.42
90	525.8	70100.71
95	633.9	84512.82
100	760.0	101324.7

Vapor Pressure as a Function of Temperature



Water K_w and pK_w at Different Temperatures (°C)		
Temperature	$K_w \cdot 10^{-14}$	pK_w
0	0.112	14.95
5	0.182	14.74
10	0.288	14.54

Water K_w and pK_w at Different Temperatures (°C)		
Temperature	$K_w \cdot 10^{-14}$	pK_w
15	0.465	14.33
20	0.671	14.17
25	0.991	14.00
30	1.432	13.84
35	2.042	13.69
40	2.851	13.55
45	3.917	13.41
50	5.297	13.28
55	7.080	13.15
60	9.311	13.03
75	19.95	12.70
100	56.23	12.25



Specific Heat Capacity for Water	
$C^\circ(\text{H}_2\text{O}(l)) = 4184 \text{ J}\cdot\text{K}^{-1}\cdot\text{kg}^{-1} = 4.184 \text{ J}\cdot\text{g}^{-1}\cdot^\circ\text{C}^{-1}$	
$C^\circ(\text{H}_2\text{O}(s)) = 1864 \text{ J}\cdot\text{K}^{-1}\cdot\text{kg}^{-1}$	
$C^\circ(\text{H}_2\text{O}(g)) = 2093 \text{ J}\cdot\text{K}^{-1}\cdot\text{kg}^{-1}$	

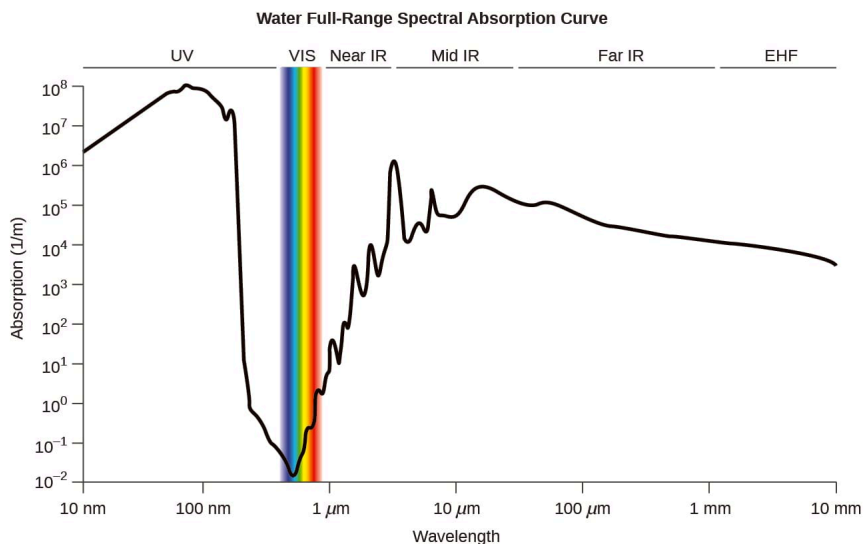
Standard Water Melting and Boiling Temperatures and Enthalpies of the Transitions		
	Temperature (K)	ΔH (kJ/mol)
melting	273.15	6.088
boiling	373.15	40.656 (44.016 at 298 K)

Water Cryoscopic (Freezing Point Depression) and Ebullioscopic (Boiling Point Elevation) Constants

$K_f = 1.86^\circ\text{C}\cdot\text{kg}\cdot\text{mol}^{-1}$ (cryoscopic constant)

$K_b = 0.51^\circ\text{C}\cdot\text{kg}\cdot\text{mol}^{-1}$ (ebullioscopic constant)

<figure class="ui-has-child-figcaption" id="CNX_Chem_00_EE_LiqWatAbso">



<figcaption>Water full-range spectral

absorption curve. This curve shows the full-range spectral absorption for water. The y -axis signifies the absorption in $1/\text{cm}$. If we divide 1 by this value, we will obtain the length of the path (in cm) after which the intensity of a light beam passing through water decays by a factor of the base of the natural logarithm e ($e = 2.718281828$).</figcaption> </figure>

Footnotes

- 1 Data for $t < 0^\circ\text{C}$ are for supercooled water
- 2 $\text{p}K_w = -\log_{10}(K_w)$

13.2.4: Composition of Commercial Acids and Bases

Composition of Commercial Acids and Bases			
Acid or Base ¹	Density (g/mL) ²	Percentage by Mass	Molarity
acetic acid, glacial	1.05	99.5%	17.4
aqueous ammonia ³	0.90	28%	14.8
hydrochloric acid	1.18	36%	11.6
nitric acid	1.42	71%	16.0
perchloric acid	1.67	70%	11.65
phosphoric acid	1.70	85%	14.7
sodium hydroxide	1.53	50%	19.1
sulfuric acid	1.84	96%	18.0

Footnotes

- 1 Acids and bases are commercially available as aqueous solutions. This table lists properties (densities and concentrations) of common acid and base solutions. Nominal values are provided in cases where the manufacturer cites a range of concentrations and densities.
- 2 This column contains specific gravity data. In the case of this table, specific gravity is the ratio of density of a substance to the density of pure water at the same conditions. Specific gravity is often cited on commercial labels.
- 3 This solution is sometimes called “ammonium hydroxide,” although this term is not chemically accurate.

This page titled [13.2.4: Composition of Commercial Acids and Bases](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

13.2.5: Standard Thermodynamic Properties for Selected Substances

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
aluminum			
Al(s)	0	0	28.3
Al(g)	324.4	285.7	164.54
Al ₂ O ₃ (s)	-1676	-1582	50.92
AlF ₃ (s)	-1510.4	-1425	66.5
AlCl ₃ (s)	-704.2	-628.8	110.67
AlCl ₃ ·6H ₂ O(s)	-2691.57	-2269.40	376.56
Al ₂ S ₃ (s)	-724.0	-492.4	116.9
Al ₂ (SO ₄) ₃ (s)	-3445.06	-3506.61	239.32
antimony			
Sb(s)	0	0	45.69
Sb(g)	262.34	222.17	180.16
Sb ₄ O ₆ (s)	-1440.55	-1268.17	220.92
SbCl ₃ (g)	-313.8	-301.2	337.80
SbCl ₅ (g)	-394.34	-334.29	401.94
Sb ₂ S ₃ (s)	-174.89	-173.64	182.00
SbCl ₃ (s)	-382.17	-323.72	184.10
SbOCl(s)	-374.0	—	—
arsenic			
As(s)	0	0	35.1
As(g)	302.5	261.0	174.21
As ₄ (g)	143.9	92.4	314
As ₄ O ₆ (s)	-1313.94	-1152.52	214.22
As ₂ O ₅ (s)	-924.87	-782.41	105.44
AsCl ₃ (g)	-261.50	-248.95	327.06
As ₂ S ₃ (s)	-169.03	-168.62	163.59
AsH ₃ (g)	66.44	68.93	222.78
H ₃ AsO ₄ (s)	-906.3	—	—
barium			
Ba(s)	0	0	62.5

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
Ba(g)	180	146	170.24
BaO(s)	-548.0	-520.3	72.1
BaCl ₂ (s)	-855.0	-806.7	123.7
BaSO ₄ (s)	-1473.2	-1362.3	132.2
beryllium			
Be(s)	0	0	9.50
Be(g)	324.3	286.6	136.27
BeO(s)	-609.4	-580.1	13.8
bismuth			
Bi(s)	0	0	56.74
Bi(g)	207.1	168.2	187.00
Bi ₂ O ₃ (s)	-573.88	-493.7	151.5
BiCl ₃ (s)	-379.07	-315.06	176.98
Bi ₂ S ₃ (s)	-143.1	-140.6	200.4
boron			
B(s)	0	0	5.86
B(g)	565.0	521.0	153.4
B ₂ O ₃ (s)	-1273.5	-1194.3	53.97
B ₂ H ₆ (g)	36.4	87.6	232.1
H ₃ BO ₃ (s)	-1094.33	-968.92	88.83
BF ₃ (g)	-1136.0	-1119.4	254.4
BCl ₃ (g)	-403.8	-388.7	290.1
B ₃ N ₃ H ₆ (l)	-540.99	-392.79	199.58
HBO ₂ (s)	-794.25	-723.41	37.66
bromine			
Br ₂ (l)	0	0	152.23
Br ₂ (g)	30.91	3.142	245.5
Br(g)	111.88	82.429	175.0
BrF ₃ (g)	-255.60	-229.45	292.42
HBr(g)	-36.3	-53.43	198.7
cadmium			
Cd(s)	0	0	51.76
Cd(g)	112.01	77.41	167.75
CdO(s)	-258.2	-228.4	54.8
CdCl ₂ (s)	-391.5	-343.9	115.3

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
CdSO ₄ (s)	-933.3	-822.7	123.0
CdS(s)	-161.9	-156.5	64.9
calcium			
Ca(s)	0	0	41.6
Ca(g)	178.2	144.3	154.88
CaO(s)	-634.9	-603.3	38.1
Ca(OH) ₂ (s)	-985.2	-897.5	83.4
CaSO ₄ (s)	-1434.5	-1322.0	106.5
CaSO ₄ ·2H ₂ O(s)	-2022.63	-1797.45	194.14
CaCO ₃ (s) (calcite)	-1220.0	-1081.4	110.0
CaSO ₃ ·H ₂ O(s)	-1752.68	-1555.19	184.10
carbon			
C(s) (graphite)	0	0	5.740
C(s) (diamond)	1.89	2.90	2.38
C(g)	716.681	671.2	158.1
CO(g)	-110.52	-137.15	197.7
CO ₂ (g)	-393.51	-394.36	213.8
CH ₄ (g)	-74.6	-50.5	186.3
CH ₃ OH(l)	-239.2	-166.6	126.8
CH ₃ OH(g)	-201.0	-162.3	239.9
CCl ₄ (l)	-128.2	-62.5	214.4
CCl ₄ (g)	-95.7	-58.2	309.7
CHCl ₃ (l)	-134.1	-73.7	201.7
CHCl ₃ (g)	-103.14	-70.34	295.71
CS ₂ (l)	89.70	65.27	151.34
CS ₂ (g)	116.9	66.8	238.0
C ₂ H ₂ (g)	227.4	209.2	200.9
C ₂ H ₄ (g)	52.4	68.4	219.3
C ₂ H ₆ (g)	-84.0	-32.0	229.2
CH ₃ CO ₂ H(l)	-484.3	-389.9	159.8
CH ₃ CO ₂ H(g)	-434.84	-376.69	282.50
C ₂ H ₅ OH(l)	-277.6	-174.8	160.7
C ₂ H ₅ OH(g)	-234.8	-167.9	281.6
C ₃ H ₈ (g)	-103.8	-23.4	270.3

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
C ₆ H ₆ (g)	82.927	129.66	269.2
C ₆ H ₆ (l)	49.1	124.50	173.4
CH ₂ Cl ₂ (l)	-124.2	-63.2	177.8
CH ₂ Cl ₂ (g)	-95.4	-65.90	270.2
CH ₃ Cl(g)	-81.9	-60.2	234.6
C ₂ H ₅ Cl(l)	-136.52	-59.31	190.79
C ₂ H ₅ Cl(g)	-112.17	-60.39	276.00
C ₂ N ₂ (g)	308.98	297.36	241.90
HCN(l)	108.9	125.0	112.8
HCN(g)	135.5	124.7	201.8
chlorine			
Cl ₂ (g)	0	0	223.1
Cl(g)	121.3	105.70	165.2
ClF(g)	-54.48	-55.94	217.78
ClF ₃ (g)	-158.99	-118.83	281.50
Cl ₂ O(g)	80.3	97.9	266.2
Cl ₂ O ₇ (l)	238.1	—	—
Cl ₂ O ₇ (g)	272.0	—	—
HCl(g)	-92.307	-95.299	186.9
HClO ₄ (l)	-40.58	—	—
chromium			
Cr(s)	0	0	23.77
Cr(g)	396.6	351.8	174.50
Cr ₂ O ₃ (s)	-1139.7	-1058.1	81.2
CrO ₃ (s)	-589.5	—	—
(NH ₄) ₂ Cr ₂ O ₇ (s)	-1806.7	—	—
cobalt			
Co(s)	0	0	30.0
CoO(s)	-237.9	-214.2	52.97
Co ₃ O ₄ (s)	-910.02	-794.98	114.22
Co(NO ₃) ₂ (s)	-420.5	—	—
copper			
Cu(s)	0	0	33.15
Cu(g)	338.32	298.58	166.38
CuO(s)	-157.3	-129.7	42.63

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
Cu ₂ O(s)	-168.6	-146.0	93.14
CuS(s)	-53.1	-53.6	66.5
Cu ₂ S(s)	-79.5	-86.2	120.9
CuSO ₄ (s)	-771.36	-662.2	109.2
Cu(NO ₃) ₂ (s)	-302.9	—	—
fluorine			
F ₂ (g)	0	0	202.8
F(g)	79.4	62.3	158.8
F ₂ O(g)	24.7	41.9	247.43
HF(g)	-273.3	-275.4	173.8
hydrogen			
H ₂ (g)	0	0	130.7
H(g)	217.97	203.26	114.7
H ₂ O(l)	-285.83	-237.1	70.0
H ₂ O(g)	-241.82	-228.59	188.8
H ₂ O ₂ (l)	-187.78	-120.35	109.6
H ₂ O ₂ (g)	-136.3	-105.6	232.7
HF(g)	-273.3	-275.4	173.8
HCl(g)	-92.307	-95.299	186.9
HBr(g)	-36.3	-53.43	198.7
HI(g)	26.48	1.70	206.59
H ₂ S(g)	-20.6	-33.4	205.8
H ₂ Se(g)	29.7	15.9	219.0
iodine			
I ₂ (s)	0	0	116.14
I ₂ (g)	62.438	19.3	260.7
I(g)	106.84	70.2	180.8
IF(g)	95.65	-118.49	236.06
ICl(g)	17.78	-5.44	247.44
IBr(g)	40.84	3.72	258.66
IF ₇ (g)	-943.91	-818.39	346.44
HI(g)	26.48	1.70	206.59
iron			
Fe(s)	0	0	27.3
Fe(g)	416.3	370.7	180.5

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
Fe ₂ O ₃ (s)	-824.2	-742.2	87.40
Fe ₃ O ₄ (s)	-1118.4	-1015.4	146.4
Fe(CO) ₅ (l)	-774.04	-705.42	338.07
Fe(CO) ₅ (g)	-733.87	-697.26	445.18
FeCl ₂ (s)	-341.79	-302.30	117.95
FeCl ₃ (s)	-399.49	-334.00	142.3
FeO(s)	-272.0	-255.2	60.75
Fe(OH) ₂ (s)	-569.0	-486.5	88.
Fe(OH) ₃ (s)	-823.0	-696.5	106.7
FeS(s)	-100.0	-100.4	60.29
Fe ₃ C(s)	25.10	20.08	104.60
lead			
Pb(s)	0	0	64.81
Pb(g)	195.2	162.	175.4
PbO(s) (yellow)	-217.32	-187.89	68.70
PbO(s) (red)	-218.99	-188.93	66.5
Pb(OH) ₂ (s)	-515.9	—	—
PbS(s)	-100.4	-98.7	91.2
Pb(NO ₃) ₂ (s)	-451.9	—	—
PbO ₂ (s)	-277.4	-217.3	68.6
PbCl ₂ (s)	-359.4	-314.1	136.0
lithium			
Li(s)	0	0	29.1
Li(g)	159.3	126.6	138.8
LiH(s)	-90.5	-68.3	20.0
Li(OH)(s)	-487.5	-441.5	42.8
LiF(s)	-616.0	-587.5	35.7
Li ₂ CO ₃ (s)	-1216.04	-1132.19	90.17
manganese			
Mn(s)	0	0	32.0
Mn(g)	280.7	238.5	173.7
MnO(s)	-385.2	-362.9	59.71
MnO ₂ (s)	-520.03	-465.1	53.05
Mn ₂ O ₃ (s)	-958.97	-881.15	110.46
Mn ₃ O ₄ (s)	-1378.83	-1283.23	155.64

Standard Thermodynamic Properties for Selected Substances

Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
mercury			
Hg(l)	0	0	75.9
Hg(g)	61.4	31.8	175.0
HgO(s) (red)	-90.83	-58.5	70.29
HgO(s) (yellow)	-90.46	-58.43	71.13
HgCl ₂ (s)	-224.3	-178.6	146.0
Hg ₂ Cl ₂ (s)	-265.4	-210.7	191.6
HgS(s) (red)	-58.16	-50.6	82.4
HgS(s) (black)	-53.56	-47.70	88.28
HgSO ₄ (s)	-707.51	-594.13	0.00
nitrogen			
N ₂ (g)	0	0	191.6
N(g)	472.704	455.5	153.3
NO(g)	90.25	87.6	210.8
NO ₂ (g)	33.2	51.30	240.1
N ₂ O(g)	81.6	103.7	220.0
N ₂ O ₃ (g)	83.72	139.41	312.17
N ₂ O ₄ (g)	11.1	99.8	304.4
N ₂ O ₅ (g)	11.3	115.1	355.7
NH ₃ (g)	-45.9	-16.5	192.8
N ₂ H ₄ (l)	50.63	149.43	121.21
N ₂ H ₄ (g)	95.4	159.4	238.5
NH ₄ NO ₃ (s)	-365.56	-183.87	151.08
NH ₄ Cl(s)	-314.43	-202.87	94.6
NH ₄ Br(s)	-270.8	-175.2	113.0
NH ₄ I(s)	-201.4	-112.5	117.0
NH ₄ NO ₂ (s)	-256.5	—	—
HNO ₃ (l)	-174.1	-80.7	155.6
HNO ₃ (g)	-133.9	-73.5	266.9
oxygen			
O ₂ (g)	0	0	205.2
O(g)	249.17	231.7	161.1
O ₃ (g)	142.7	163.2	238.9
phosphorus			
P ₄ (s)	0	0	164.4

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
P ₄ (g)	58.91	24.4	280.0
P(g)	314.64	278.25	163.19
PH ₃ (g)	5.4	13.5	210.2
PCl ₃ (g)	-287.0	-267.8	311.78
PCl ₅ (g)	-374.9	-305.0	364.4
P ₄ O ₆ (s)	-1640.1	—	—
P ₄ O ₁₀ (s)	-2984.0	-2697.0	228.86
HPO ₃ (s)	-948.5	—	—
H ₃ PO ₂ (s)	-604.6	—	—
H ₃ PO ₃ (s)	-964.4	—	—
H ₃ PO ₄ (s)	-1279.0	-1119.1	110.50
H ₃ PO ₄ (l)	-1266.9	-1124.3	110.5
H ₄ P ₂ O ₇ (s)	-2241.0	—	—
POCl ₃ (l)	-597.1	-520.8	222.5
POCl ₃ (g)	-558.5	-512.9	325.5
potassium			
K(s)	0	0	64.7
K(g)	89.0	60.5	160.3
KF(s)	-576.27	-537.75	66.57
KCl(s)	-436.5	-408.5	82.6
silicon			
Si(s)	0	0	18.8
Si(g)	450.0	405.5	168.0
SiO ₂ (s)	-910.7	-856.3	41.5
SiH ₄ (g)	34.3	56.9	204.6
H ₂ SiO ₃ (s)	-1188.67	-1092.44	133.89
H ₄ SiO ₄ (s)	-1481.14	-1333.02	192.46
SiF ₄ (g)	-1615.0	-1572.8	282.8
SiCl ₄ (l)	-687.0	-619.8	239.7
SiCl ₄ (g)	-662.75	-622.58	330.62
SiC(s, <i>beta cubic</i>)	-73.22	-70.71	16.61
SiC(s, <i>alpha hexagonal</i>)	-71.55	-69.04	16.48
silver			
Ag(s)	0	0	42.55
Ag(g)	284.9	246.0	172.89

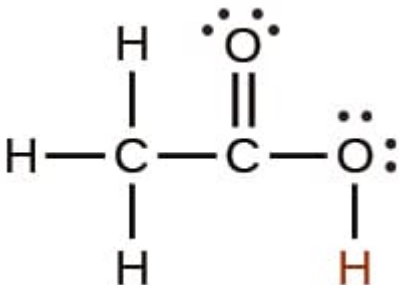
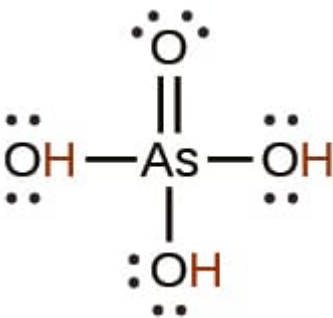
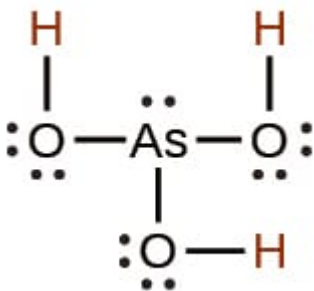
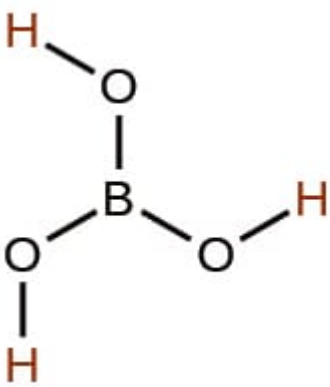
Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
Ag ₂ O(s)	-31.05	-11.20	121.3
AgCl(s)	-127.0	-109.8	96.3
Ag ₂ S(s)	-32.6	-40.7	144.0
sodium			
Na(s)	0	0	51.3
Na(g)	107.5	77.0	153.7
Na ₂ O(s)	-414.2	-375.5	75.1
NaCl(s)	-411.2	-384.1	72.1
sulfur			
S ₈ (s) (rhombic)	0	0	256.8
S(g)	278.81	238.25	167.82
SO ₂ (g)	-296.83	-300.1	248.2
SO ₃ (g)	-395.72	-371.06	256.76
H ₂ S(g)	-20.6	-33.4	205.8
H ₂ SO ₄ (l)	-813.989	690.00	156.90
H ₂ S ₂ O ₇ (s)	-1273.6	—	—
SF ₄ (g)	-728.43	-684.84	291.12
SF ₆ (g)	-1220.5	-1116.5	291.5
SCl ₂ (l)	-50	—	—
SCl ₂ (g)	-19.7	—	—
S ₂ Cl ₂ (l)	-59.4	—	—
S ₂ Cl ₂ (g)	-19.50	-29.25	319.45
SOCl ₂ (g)	-212.55	-198.32	309.66
SOCl ₂ (l)	-245.6	—	—
SO ₂ Cl ₂ (l)	-394.1	—	—
SO ₂ Cl ₂ (g)	-354.80	-310.45	311.83
tin			
Sn(s)	0	0	51.2
Sn(g)	301.2	266.2	168.5
SnO(s)	-285.8	-256.9	56.5
SnO ₂ (s)	-577.6	-515.8	49.0
SnCl ₄ (l)	-511.3	-440.1	258.6
SnCl ₄ (g)	-471.5	-432.2	365.8
titanium			
Ti(s)	0	0	30.7

Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
Ti(g)	473.0	428.4	180.3
TiO ₂ (s)	-944.0	-888.8	50.6
TiCl ₄ (l)	-804.2	-737.2	252.4
TiCl ₄ (g)	-763.2	-726.3	353.2
tungsten			
W(s)	0	0	32.6
W(g)	849.4	807.1	174.0
WO ₃ (s)	-842.9	-764.0	75.9
zinc			
Zn(s)	0	0	41.6
Zn(g)	130.73	95.14	160.98
ZnO(s)	-350.5	-320.5	43.7
ZnCl ₂ (s)	-415.1	-369.43	111.5
ZnS(s)	-206.0	-201.3	57.7
ZnSO ₄ (s)	-982.8	-871.5	110.5
ZnCO ₃ (s)	-812.78	-731.57	82.42
complexes			
[Co(NH ₃) ₄ (NO ₂) ₂]NO ₃ , <i>cis</i>	-898.7	—	—
[Co(NH ₃) ₄ (NO ₂) ₂]NO ₃ , <i>trans</i>	-896.2	—	—
NH ₄ [Co(NH ₃) ₂ (NO ₂) ₄]	-837.6	—	—
[Co(NH ₃) ₆][Co(NH ₃) ₂ (NO ₂) ₄] ₃	-2733.0	—	—
[Co(NH ₃) ₄ Cl ₂]Cl, <i>cis</i>	-874.9	—	—
[Co(NH ₃) ₄ Cl ₂]Cl, <i>trans</i>	-877.4	—	—
[Co(en) ₂ (NO ₂) ₂]NO ₃ , <i>cis</i>	-689.5	—	—
[Co(en) ₂ Cl ₂]Cl, <i>cis</i>	-681.2	—	—
[Co(en) ₂ Cl ₂]Cl, <i>trans</i>	-677.4	—	—
[Co(en) ₃](ClO ₄) ₃	-762.7	—	—
[Co(en) ₃]Br ₂	-595.8	—	—
[Co(en) ₃]I ₂	-475.3	—	—
[Co(en) ₃]I ₃	-519.2	—	—
[Co(NH ₃) ₆](ClO ₄) ₃	-1034.7	-221.1	615
[Co(NH ₃) ₅ NO ₂](NO ₃) ₂	-1088.7	-412.9	331
[Co(NH ₃) ₆](NO ₃) ₃	-1282.0	-524.5	448
[Co(NH ₃) ₅ Cl]Cl ₂	-1017.1	-582.5	366.1

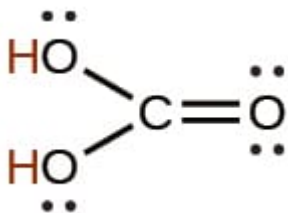
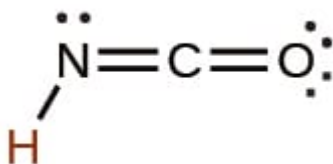
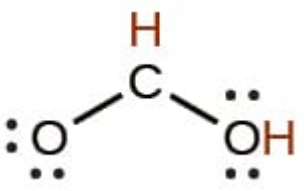
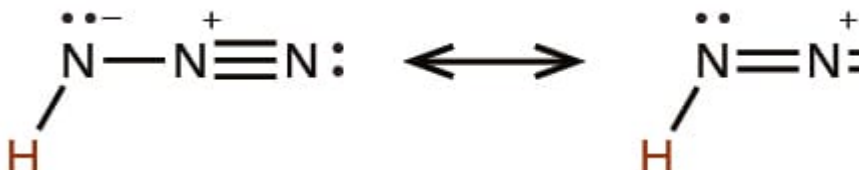
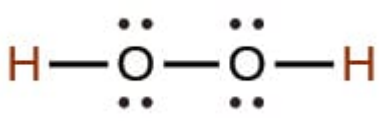
Standard Thermodynamic Properties for Selected Substances			
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S_{298}° (J K ⁻¹ mol ⁻¹)
[Pt(NH ₃) ₄]Cl ₂	-725.5	—	—
[Ni(NH ₃) ₆]Cl ₂	-994.1	—	—
[Ni(NH ₃) ₆]Br ₂	-923.8	—	—
[Ni(NH ₃) ₆]I ₂	-808.3	—	—

This page titled [13.2.5: Standard Thermodynamic Properties for Selected Substances](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

13.2.6: Ionization Constants of Weak Acids

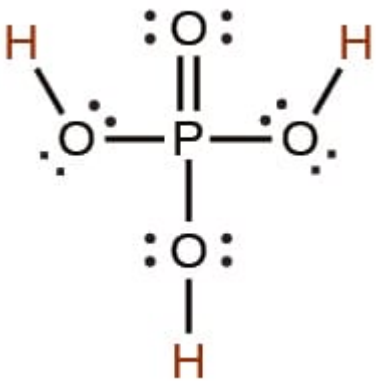
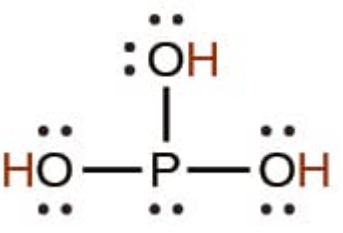
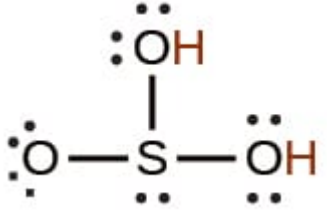
Ionization Constants of Weak Acids			
Acid	Formula	K_a at 25 °C	Lewis Structure
acetic	$\text{CH}_3\text{CO}_2\text{H}$	1.8×10^{-5}	
arsenic	H_3AsO_4	5.5×10^{-3}	
	H_2AsO_4^-	1.7×10^{-7}	
	HAsO_4^{2-}	5.1×10^{-12}	
arsenous	H_3AsO_3	5.1×10^{-10}	
boric	H_3BO_3	5.4×10^{-10}	

Ionization Constants of Weak Acids

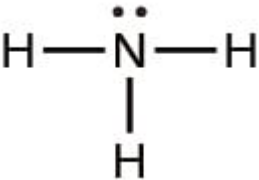
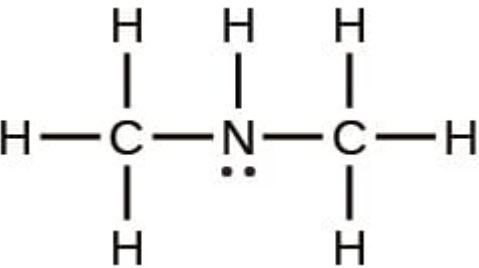
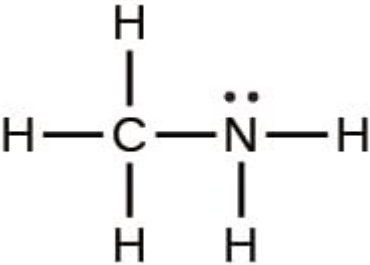
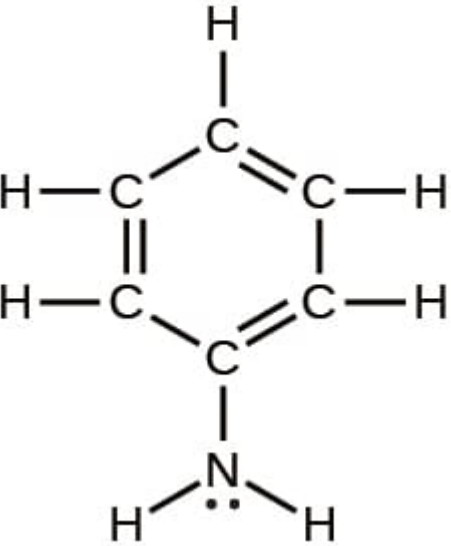
Acid	Formula	K_a at 25 °C	Lewis Structure
carbonic	H_2CO_3	4.3×10^{-7}	
	HCO_3^-	5.6×10^{-11}	
cyanic	HCNO	2×10^{-4}	
formic	HCO_2H	1.8×10^{-4}	
hydrazoic	HN_3	2.5×10^{-5}	
hydrocyanic	HCN	4.9×10^{-10}	
hydrofluoric	HF	3.5×10^{-4}	
hydrogen peroxide	H_2O_2	2.4×10^{-12}	
hydrogen selenide	H_2Se	1.29×10^{-4}	
	HSe^-	1×10^{-12}	

Ionization Constants of Weak Acids

Acid	Formula	K_a at 25 °C	Lewis Structure
hydrogen sulfate ion	HSO_4^-	1.2×10^{-2}	
hydrogen sulfide	H_2S	8.9×10^{-8}	
	HS^-	1.0×10^{-19}	
hydrogen telluride	H_2Te	2.3×10^{-3}	
	HTe^-	1.6×10^{-11}	
hypobromous	HBrO	2.8×10^{-9}	
hypochlorous	HClO	2.9×10^{-8}	
nitrous	HNO_2	4.6×10^{-4}	
oxalic	$\text{H}_2\text{C}_2\text{O}_4$	6.0×10^{-2}	
	HC_2O_4^-	6.1×10^{-5}	

Ionization Constants of Weak Acids			
Acid	Formula	K_a at 25 °C	Lewis Structure
phosphoric	H_3PO_4	7.5×10^{-3}	
	H_2PO_4^-	6.2×10^{-8}	
	HPO_4^{2-}	4.2×10^{-13}	
phosphorous	H_3PO_3	5×10^{-2}	
	H_2PO_3^-	2.0×10^{-7}	
sulfurous	H_2SO_3	1.6×10^{-2}	
	HSO_3^-	6.4×10^{-8}	

13.2.7: Ionization Constants of Weak Bases

Ionization Constants of Weak Bases		
Base	Lewis Structure	K_b at 25 °C
ammonia		1.8×10^{-5}
dimethylamine		5.9×10^{-4}
methylamine		4.4×10^{-4}
phenylamine (aniline)		4.3×10^{-10}

Ionization Constants of Weak Bases		
Base	Lewis Structure	K_b at 25 °C
trimethylamine		6.3×10^{-5}

This page titled [13.2.7: Ionization Constants of Weak Bases](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

13.2.8: Solubility Products

Solubility Products	
Substance	K_{sp} at 25 °C
aluminum	
$\text{Al}(\text{OH})_3$	2×10^{-32}
barium	
BaCO_3	1.6×10^{-9}
$\text{BaC}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$	1.1×10^{-7}
BaSO_4	2.3×10^{-8}
BaCrO_4	8.5×10^{-11}
BaF_2	2.4×10^{-5}
$\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$	5.0×10^{-3}
$\text{Ba}_3(\text{PO}_4)_2$	6×10^{-39}
$\text{Ba}_3(\text{AsO}_4)_2$	1.1×10^{-13}
bismuth	
$\text{BiO}(\text{OH})$	4×10^{-10}
BiOCl	1.8×10^{-31}
Bi_2S_3	1×10^{-97}
cadmium	
$\text{Cd}(\text{OH})_2$	5.9×10^{-15}
CdS	1.0×10^{-28}
CdCO_3	5.2×10^{-12}
calcium	
$\text{Ca}(\text{OH})_2$	1.3×10^{-6}
CaCO_3	8.7×10^{-9}
$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$	6.1×10^{-5}
$\text{CaC}_2\text{O}_4 \cdot \text{H}_2\text{O}$	1.96×10^{-8}
$\text{Ca}_3(\text{PO}_4)_2$	1.3×10^{-32}
CaHPO_4	7×10^{-7}
CaF_2	4.0×10^{-11}
chromium	
$\text{Cr}(\text{OH})_3$	6.7×10^{-31}

Solubility Products	
Substance	K_{sp} at 25 °C
cobalt	
Co(OH) ₂	2.5×10^{-16}
CoS(α)	5×10^{-22}
CoS(β)	3×10^{-26}
CoCO ₃	1.4×10^{-13}
Co(OH) ₃	2.5×10^{-43}
copper	
CuCl	1.2×10^{-6}
CuBr	6.27×10^{-9}
CuI	1.27×10^{-12}
CuSCN	1.6×10^{-11}
Cu ₂ S	2.5×10^{-48}
Cu(OH) ₂	2.2×10^{-20}
CuS	8.5×10^{-45}
CuCO ₃	2.5×10^{-10}
iron	
Fe(OH) ₂	1.8×10^{-15}
FeCO ₃	2.1×10^{-11}
FeS	3.7×10^{-19}
Fe(OH) ₃	4×10^{-38}
lead	
Pb(OH) ₂	1.2×10^{-15}
PbF ₂	4×10^{-8}
PbCl ₂	1.6×10^{-5}
PbBr ₂	4.6×10^{-6}
PbI ₂	1.4×10^{-8}
PbCO ₃	1.5×10^{-15}
PbS	7×10^{-29}
PbCrO ₄	2×10^{-16}
PbSO ₄	1.3×10^{-8}
Pb ₃ (PO ₄) ₂	1×10^{-54}

Solubility Products	
Substance	K_{sp} at 25 °C
magnesium	
Mg(OH) ₂	8.9×10^{-12}
MgCO ₃ ·3H ₂ O	$ca\ 1 \times 10^{-5}$
MgNH ₄ PO ₄	3×10^{-13}
MgF ₂	6.4×10^{-9}
MgC ₂ O ₄	7×10^{-7}
manganese	
Mn(OH) ₂	2×10^{-13}
MnCO ₃	8.8×10^{-11}
MnS	2.3×10^{-13}
mercury	
Hg ₂ O·H ₂ O	3.6×10^{-26}
Hg ₂ Cl ₂	1.1×10^{-18}
Hg ₂ Br ₂	1.3×10^{-22}
Hg ₂ I ₂	4.5×10^{-29}
Hg ₂ CO ₃	9×10^{-15}
Hg ₂ SO ₄	7.4×10^{-7}
Hg ₂ S	1.0×10^{-47}
Hg ₂ CrO ₄	2×10^{-9}
HgS	1.6×10^{-54}
nickel	
Ni(OH) ₂	1.6×10^{-16}
NiCO ₃	1.4×10^{-7}
NiS(α)	4×10^{-20}
NiS(β)	1.3×10^{-25}
potassium	
KClO ₄	1.05×10^{-2}
K ₂ PtCl ₆	7.48×10^{-6}
KHC ₄ H ₄ O ₆	3×10^{-4}
silver	
$\frac{1}{2}\text{Ag}_2\text{O}(\text{Ag}^+ + \text{OH}^-)$	2×10^{-8}
AgCl	1.6×10^{-10}

Solubility Products	
Substance	K_{sp} at 25 °C
AgBr	5.0×10^{-13}
AgI	1.5×10^{-16}
AgCN	1.2×10^{-16}
AgSCN	1.0×10^{-12}
Ag ₂ S	1.6×10^{-49}
Ag ₂ CO ₃	8.1×10^{-12}
Ag ₂ CrO ₄	9.0×10^{-12}
Ag ₄ Fe(CN) ₆	1.55×10^{-41}
Ag ₂ SO ₄	1.2×10^{-5}
Ag ₃ PO ₄	1.8×10^{-18}
strontium	
Sr(OH) ₂ ·8H ₂ O	3.2×10^{-4}
SrCO ₃	7×10^{-10}
SrCrO ₄	3.6×10^{-5}
SrSO ₄	3.2×10^{-7}
SrC ₂ O ₄ ·H ₂ O	4×10^{-7}
thallium	
TlCl	1.7×10^{-4}
TlSCN	1.6×10^{-4}
Tl ₂ S	6×10^{-22}
Tl(OH) ₃	6.3×10^{-46}
tin	
Sn(OH) ₂	3×10^{-27}
SnS	1×10^{-26}
Sn(OH) ₄	1.0×10^{-57}
zinc	
ZnCO ₃	2×10^{-10}

13.2.9: Formation Constants for Complex Ions

Formation Constants for Complex Ions	
Equilibrium	K_f
$\text{Al}^{3+} + 6 \text{F}^- \rightleftharpoons [\text{AlF}_6]^{3-}$	7×10^{19}
$\text{Cd}^{2+} + 4 \text{NH}_3 \rightleftharpoons [\text{Cd}(\text{NH}_3)_4]^{2+}$	1.3×10^7
$\text{Cd}^{2+} + 4 \text{CN}^- \rightleftharpoons [\text{Cd}(\text{CN})_4]^{2-}$	3×10^{18}
$\text{Co}^{2+} + 6 \text{NH}_3 \rightleftharpoons [\text{Co}(\text{NH}_3)_6]^{2+}$	1.3×10^5
$\text{Co}^{3+} + 6 \text{NH}_3 \rightleftharpoons [\text{Co}(\text{NH}_3)_6]^{3+}$	2.3×10^{33}
$\text{Cu}^+ + 2 \text{CN}^- \rightleftharpoons [\text{Cu}(\text{CN})_2]^-$	1.0×10^{16}
$\text{Cu}^{2+} + 4 \text{NH}_3 \rightleftharpoons \text{Cu}(\text{NH}_3)_4^{2+}$	1.7×10^{13}
$\text{Fe}^{2+} + 6 \text{CN}^- \rightleftharpoons \text{Fe}(\text{CN})_6^{4-}$	1.5×10^{35}
$\text{Fe}^{3+} + 6 \text{CN}^- \rightleftharpoons \text{Fe}(\text{CN})_6^{3-}$	2×10^{43}
$\text{Fe}^{3+} + 6 \text{SCN}^- \rightleftharpoons \text{Fe}(\text{SCN})_6^{3-}$	3.2×10^3
$\text{Hg}^{2+} + 4 \text{Cl}^- \rightleftharpoons \text{HgCl}_4^{2-}$	1.1×10^{16}
$\text{Ni}^{2+} + 6 \text{NH}_3 \rightleftharpoons \text{Ni}(\text{NH}_3)_6^{2+}$	2.0×10^8
$\text{Ag}^+ + 2 \text{Cl}^- \rightleftharpoons \text{AgCl}_2^-$	1.8×10^5
$\text{Ag}^+ + 2 \text{CN}^- \rightleftharpoons \text{Ag}(\text{CN})_2^-$	1×10^{21}
$\text{Ag}^+ + 2 \text{NH}_3 \rightleftharpoons \text{Ag}(\text{NH}_3)_2^+$	1.7×10^7
$\text{Zn}^{2+} + 4 \text{CN}^- \rightleftharpoons \text{Zn}(\text{CN})_4^{2-}$	2.1×10^{19}
$\text{Zn}^{2+} + 4 \text{OH}^- \rightleftharpoons \text{Zn}(\text{OH})_4^{2-}$	2×10^{15}
$\text{Fe}^{3+} + \text{SCN}^- \rightleftharpoons \text{Fe}(\text{SCN})_2^+$	8.9×10^2
$\text{Ag}^+ + 4 \text{SCN}^- \rightleftharpoons \text{Ag}(\text{SCN})_4^{3-}$	1.2×10^{10}
$\text{Pb}^{2+} + 4 \text{I}^- \rightleftharpoons \text{PbI}_4^{2-}$	3.0×10^4
$\text{Pt}^{2+} + 4 \text{Cl}^- \rightleftharpoons \text{PtCl}_4^{2-}$	1×10^{16}
$\text{Cu}^{2+} + 4 \text{CN}^- \rightleftharpoons \text{Cu}(\text{CN})_4^{2-}$	1.0×10^{25}
$\text{Co}^{2+} + 4 \text{SCN}^- \rightleftharpoons \text{Co}(\text{SCN})_4^{2-}$	1×10^3

This page titled [13.2.9: Formation Constants for Complex Ions](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

13.2.10: Standard Electrode (Half-Cell) Potentials

Standard Electrode (Half-Cell) Potentials	
Half-Reaction	E° (V)
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	+0.7996
$\text{AgCl} + \text{e}^- \rightarrow \text{Ag} + \text{Cl}^-$	+0.22233
$[\text{Ag}(\text{CN})_2]^- + \text{e}^- \rightarrow \text{Ag} + 2 \text{CN}^-$	-0.31
$\text{Ag}_2\text{CrO}_4 + 2 \text{e}^- \rightarrow 2 \text{Ag} + \text{CrO}_4^{2-}$	+0.45
$[\text{Ag}(\text{NH}_3)_2]^+ + \text{e}^- \rightarrow \text{Ag} + 2 \text{NH}_3$	+0.373
$[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-} + \text{e}^- \rightarrow \text{Ag} + 2 \text{S}_2\text{O}_3^{2-}$	+0.017
$[\text{AlF}_6]^{3-} + 3 \text{e}^- \rightarrow \text{Al} + 6 \text{F}^-$	-2.07
$\text{Al}^{3+} + 3 \text{e}^- \rightarrow \text{Al}$	-1.662
$\text{Am}^{3+} + 3 \text{e}^- \rightarrow \text{Am}$	-2.048
$\text{Au}^{3+} + 3 \text{e}^- \rightarrow \text{Au}$	+1.498
$\text{Au}^+ + \text{e}^- \rightarrow \text{Au}$	+1.692
$\text{Ba}^{2+} + 2 \text{e}^- \rightarrow \text{Ba}$	-2.912
$\text{Be}^{2+} + 2 \text{e}^- \rightarrow \text{Be}$	-1.847
$\text{Br}_2(\text{aq}) + 2 \text{e}^- \rightarrow 2 \text{Br}^-$	+1.0873
$\text{Ca}^{2+} + 2 \text{e}^- \rightarrow \text{Ca}$	-2.868
$\text{Ce}^{3+} + 3 \text{e}^- \rightarrow \text{Ce}$	-2.483
$\text{Ce}^{4+} + \text{e}^- \rightarrow \text{Ce}^{3+}$	+1.61
$\text{Cd}^{2+} + 2 \text{e}^- \rightarrow \text{Cd}$	-0.4030
$[\text{Cd}(\text{CN})_4]^{2-} + 2 \text{e}^- \rightarrow \text{Cd} + 4 \text{CN}^-$	-1.09
$[\text{Cd}(\text{NH}_3)_4]^{2+} + 2 \text{e}^- \rightarrow \text{Cd} + 4 \text{NH}_3$	-0.61
$\text{CdS} + 2 \text{e}^- \rightarrow \text{Cd} + \text{S}^{2-}$	-1.17
$\text{Cl}_2 + 2 \text{e}^- \rightarrow 2 \text{Cl}^-$	+1.35827
$\text{ClO}_4^- + \text{H}_2\text{O} + 2 \text{e}^- \rightarrow \text{ClO}_3^- + 2 \text{OH}^-$	+0.36
$\text{ClO}_3^- + \text{H}_2\text{O} + 2 \text{e}^- \rightarrow \text{ClO}_2^- + 2 \text{OH}^-$	+0.33
$\text{ClO}_2^- + \text{H}_2\text{O} + 2 \text{e}^- \rightarrow \text{ClO}^- + 2 \text{OH}^-$	+0.66
$\text{ClO}^- + \text{H}_2\text{O} + 2 \text{e}^- \rightarrow \text{Cl}^- + 2 \text{OH}^-$	+0.89
$\text{ClO}_4^- + 2 \text{H}_3\text{O}^+ + 2 \text{e}^- \rightarrow \text{ClO}_3^- + 3 \text{H}_2\text{O}$	+1.189
$\text{ClO}_3^- + 3 \text{H}_3\text{O}^+ + 2 \text{e}^- \rightarrow \text{HClO}_2 + 4 \text{H}_2\text{O}$	+1.21
$\text{HClO} + \text{H}_3\text{O}^+ + 2 \text{e}^- \rightarrow \text{Cl}^- + 2 \text{H}_2\text{O}$	+1.482
$\text{HClO} + \text{H}_3\text{O}^+ + \text{e}^- \rightarrow \frac{1}{2} \text{Cl}_2 + 2 \text{H}_2\text{O}$	+1.611
$\text{HClO}_2 + 2 \text{H}_3\text{O}^+ + 2 \text{e}^- \rightarrow \text{HClO} + 3 \text{H}_2\text{O}$	+1.628
$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+} \quad (2 \text{ mol } // \text{H}_2\text{SO}_4)$	+1.83

Standard Electrode (Half-Cell) Potentials	
Half-Reaction	E° (V)
$\text{Co}^{2+} + 2\text{e}^- \rightarrow \text{Co}$	-0.28
$[\text{Co}(\text{NH}_3)_6]^{3+} + \text{e}^- \xrightarrow{\text{Co}(\text{NH}_3)_6} 2+$	+0.1
$\text{Co}(\text{OH})_3 + \text{e}^- \rightarrow \text{Co}(\text{OH})_2 + \text{OH}^-$	+0.17
$\text{Cr}^3 + 3\text{e}^- \rightarrow \text{Cr}$	-0.744
$\text{Cr}^{3+} + \text{e}^- \rightarrow \text{Cr}^{2+}$	-0.407
$\text{Cr}^{2+} + 2\text{e}^- \rightarrow \text{Cr}$	-0.913
$[\text{Cu}(\text{CN})_2]^- + \text{e}^- \rightarrow \text{Cu} + 2\text{CN}^-$	-0.43
$\text{CrO}_4^{2-} + 4\text{H}_2\text{O} + 3\text{e}^- \rightarrow \text{Cr}(\text{OH})_3 + 5\text{OH}^-$	-0.13
$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}_3\text{O}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 21\text{H}_2\text{O}$	+1.232
$[\text{Cr}(\text{OH})_4]^- + 3\text{e}^- \rightarrow \text{Cr} + 4\text{OH}^-$	-1.2
$\text{Cr}(\text{OH})_3 + 3\text{e}^- \rightarrow \text{Cr} + 3\text{OH}^-$	-1.48
$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	+0.153
$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	+0.34
$\text{Cu}^+ + \text{e}^- \rightarrow \text{Cu}$	+0.521
$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	+2.866
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.447
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	+0.771
$[\text{Fe}(\text{CN})_6]^{3-} + \text{e}^- \xrightarrow{\text{Fe}(\text{CN})_6} 4-$	+0.36
$\text{Fe}(\text{OH})_2 + 2\text{e}^- \rightarrow \text{Fe} + 2\text{OH}^-$	-0.88
$\text{FeS} + 2\text{e}^- \rightarrow \text{Fe} + \text{S}^{2-}$	-1.01
$\text{Ga}^{3+} + 3\text{e}^- \rightarrow \text{Ga}$	-0.549
$\text{Gd}^{3+} + 3\text{e}^- \rightarrow \text{Gd}$	-2.279
$\frac{1}{2}\text{H}_2 + \text{e}^- \rightarrow \text{H}^-$	-2.23
$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$	-0.8277
$\text{H}_2\text{O}_2 + 2\text{H}_3\text{O}^+ + 2\text{e}^- \rightarrow 4\text{H}_2\text{O}$	+1.776
$2\text{H}_3\text{O}^+ + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{H}_2\text{O}$	0.00
$\text{HO}_2^- + \text{H}_2\text{O} + 2\text{e}^- \rightarrow 3\text{OH}^-$	+0.878
$\text{Hf}^{4+} + 4\text{e}^- \rightarrow \text{Hf}$	-1.55
$\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}$	+0.851
$2\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}$	+0.92
$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow 2\text{Hg}$	+0.7973
$[\text{HgBr}_4]^{2-} + 2\text{e}^- \rightarrow \text{Hg} + 4\text{Br}^-$	+0.21
$\text{Hg}_2\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Hg} + 2\text{Cl}^-$	+0.26808
$[\text{Hg}(\text{CN})_4]^{2-} + 2\text{e}^- \rightarrow \text{Hg} + 4\text{CN}^-$	-0.37
$[\text{HgI}_4]^{2-} + 2\text{e}^- \rightarrow \text{Hg} + 4\text{I}^-$	-0.04

Standard Electrode (Half-Cell) Potentials	
Half-Reaction	E° (V)
$\text{HgS} + 2\text{e}^- \rightarrow \text{Hg} + \text{S}^{2-}$	-0.70
$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	+0.5355
$\text{In}^{3+} + 3\text{e}^- \rightarrow \text{In}$	-0.3382
$\text{K}^+ + \text{e}^- \rightarrow \text{K}$	-2.931
$\text{La}^{3+} + 3\text{e}^- \rightarrow \text{La}$	-2.52
$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.04
$\text{Lu}^{3+} + 3\text{e}^- \rightarrow \text{Lu}$	-2.28
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.372
$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.185
$\text{MnO}_2 + 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Mn}(\text{OH})_2 + 2\text{OH}^-$	-0.05
$\text{MnO}_4^- + 2\text{H}_2\text{O} + 3\text{e}^- \rightarrow \text{MnO}_2 + 4\text{OH}^-$	+0.558
$\text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O}$	+1.23
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.507
$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71
$\text{Nd}^{3+} + 3\text{e}^- \rightarrow \text{Nd}$	-2.323
$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.257
$[\text{Ni}(\text{NH}_3)_6]^{2+} + 2\text{e}^- \rightarrow \text{Ni} + 6\text{NH}_3$	-0.49
$\text{NiO}_2 + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{Ni}^{2+} + 2\text{H}_2\text{O}$	+1.593
$\text{NiO}_2 + 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Ni}(\text{OH})_2 + 2\text{OH}^-$	+0.49
$\text{NiS} + 2\text{e}^- \rightarrow \text{Ni} + \text{S}^{2-}$	+0.76
$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$	+0.957
$\text{NO}_3^- + 3\text{H}^+ + 2\text{e}^- \rightarrow \text{HNO}_2 + \text{H}_2\text{O}$	+0.92
$\text{NO}_3^- + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{NO}_2^- + 2\text{OH}^-$	+0.10
$\text{Np}^{3+} + 3\text{e}^- \rightarrow \text{Np}$	-1.856
$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$	+0.401
$\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O}_2$	+0.695
$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.229
$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.1262
$\text{PbO}_2 + \text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^- \rightarrow \text{PbSO}_4 + 2\text{H}_2\text{O}$	+1.69
$\text{PbS} + 2\text{e}^- \rightarrow \text{Pb} + \text{S}^{2-}$	-0.95
$\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.3505
$\text{Pd}^{2+} + 2\text{e}^- \rightarrow \text{Pd}$	+0.987
$[\text{PdCl}_4]^{2-} + 2\text{e}^- \rightarrow \text{Pd} + 4\text{Cl}^-$	+0.591
$\text{Pt}^{2+} + 2\text{e}^- \rightarrow \text{Pt}$	+1.20
$[\text{PtBr}_4]^{2-} + 2\text{e}^- \rightarrow \text{Pt} + 4\text{Br}^-$	+0.58
$[\text{PtCl}_4]^{2-} + 2\text{e}^- \rightarrow \text{Pt} + 4\text{Cl}^-$	+0.755

Standard Electrode (Half-Cell) Potentials	
Half-Reaction	E° (V)
$[\text{PtCl}_6]^{2-} + 2\text{e}^- \xrightarrow{\text{PtCl}_2} 2\text{Cl}^-$	+0.68
$\text{Pu}^3 + 3\text{e}^- \rightarrow \text{Pu}$	-2.03
$\text{Ra}^{2+} + 2\text{e}^- \rightarrow \text{Ra}$	-2.92
$\text{Rb}^+ + \text{e}^- \rightarrow \text{Rb}$	-2.98
$[\text{RhCl}_6]^{3-} + 3\text{e}^- \rightarrow \text{Rh} + 6\text{Cl}^-$	+0.44
$\text{S} + 2\text{e}^- \rightarrow \text{S}^{2-}$	-0.47627
$\text{S} + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{S}$	+0.142
$\text{Sc}^{3+} + 3\text{e}^- \rightarrow \text{Sc}$	-2.09
$\text{Se} + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{Se}$	-0.399
$[\text{SiF}_6]^{2-} + 4\text{e}^- \rightarrow \text{Si} + 6\text{F}^-$	-1.2
$\text{SiO}_3^{2-} + 3\text{H}_2\text{O} + 4\text{e}^- \rightarrow \text{Si} + 6\text{OH}^-$	-1.697
$\text{SiO}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow \text{Si} + 2\text{H}_2\text{O}$	-0.86
$\text{Sm}^{3+} + 3\text{e}^- \rightarrow \text{Sm}$	-2.304
$\text{Sn}^{4+} + 2\text{e}^- \rightarrow \text{Sn}^{2+}$	+0.151
$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.1375
$[\text{SnF}_6]^{2-} + 4\text{e}^- \rightarrow \text{Sn} + 6\text{F}^-$	-0.25
$\text{SnS} + 2\text{e}^- \rightarrow \text{Sn} + \text{S}^{2-}$	-0.94
$\text{Sr}^{2+} + 2\text{e}^- \rightarrow \text{Sr}$	-2.89
$\text{TeO}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow \text{Te} + 2\text{H}_2\text{O}$	+0.593
$\text{Th}^{4+} + 4\text{e}^- \rightarrow \text{Th}$	-1.90
$\text{Ti}^{2+} + 2\text{e}^- \rightarrow \text{Ti}$	-1.630
$\text{U}^{3+} + 3\text{e}^- \rightarrow \text{U}$	-1.79
$\text{V}^{2+} + 2\text{e}^- \rightarrow \text{V}$	-1.19
$\text{Y}^{3+} + 3\text{e}^- \rightarrow \text{Y}$	-2.37
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.7618
$[\text{Zn}(\text{CN})_4]^{2-} + 2\text{e}^- \rightarrow \text{Zn} + 4\text{CN}^-$	-1.26
$[\text{Zn}(\text{NH}_3)_4]^{2+} + 2\text{e}^- \rightarrow \text{Zn} + 4\text{NH}_3$	-1.04
$\text{Zn}(\text{OH})_2 + 2\text{e}^- \rightarrow \text{Zn} + 2\text{OH}^-$	-1.245
$[\text{Zn}(\text{OH})_4]^{2-} + 2\text{e}^- \rightarrow \text{Zn} + 4\text{OH}^-$	-1.199
$\text{ZnS} + 2\text{e}^- \rightarrow \text{Zn} + \text{S}^{2-}$	-1.40
$\text{Zr}^{4+} + 4\text{e}^- \rightarrow \text{Zr}$	-1.539

This page titled [13.2.10: Standard Electrode \(Half-Cell\) Potentials](#) is shared under a [CC BY 4.0](#) license and was authored, remixed, and/or curated by Vince Hradil.

13.2.11: Half-Lives for Several Radioactive Isotopes

Isotope	Half-Life ¹	Type of Emission ²	Isotope	Half-Life ³	Type of Emission ⁴
$^{14}_6\text{C}$	5730 <i>y</i>	(β^-)	$^{210}_{83}\text{Bi}$	5.01 <i>d</i>	(β^-)
$^{13}_7\text{N}$	9.97 <i>m</i>	(β^+)	$^{212}_{83}\text{Bi}$	60.55 <i>m</i>	(α or β^-)
$^{15}_9\text{F}$	4.1×10^{-22} <i>s</i>	(<i>p</i>)	$^{210}_{84}\text{Po}$	138.4 <i>d</i>	(α)
$^{24}_{11}\text{Na}$	15.00 <i>h</i>	(β^-)	$^{212}_{84}\text{Po}$	3×10^{-7} <i>s</i>	(α)
$^{32}_{15}\text{P}$	14.29 <i>d</i>	(β^-)	$^{216}_{84}\text{Po}$	0.15 <i>s</i>	(α)
$^{40}_{19}\text{K}$	1.27×10^9 <i>y</i>	(β or <i>E.C.</i>)	$^{218}_{84}\text{Po}$	3.05 <i>m</i>	(α)
$^{49}_{26}\text{Fe}$	0.08 <i>s</i>	(β^+)	$^{215}_{85}\text{At}$	1.0×10^{-4} <i>s</i>	(α)
$^{60}_{26}\text{Fe}$	2.6×10^6 <i>y</i>	(β^-)	$^{218}_{85}\text{At}$	1.6 <i>s</i>	(α)
$^{60}_{27}\text{Co}$	5.27 <i>y</i>	(β^-)	$^{220}_{86}\text{Rn}$	55.6 <i>s</i>	(α)
$^{87}_{37}\text{Rb}$	4.7×10^{10} <i>y</i>	(β^-)	$^{222}_{86}\text{Rn}$	3.82 <i>d</i>	(α)
$^{90}_{38}\text{Sr}$	29 <i>y</i>	(β^-)	$^{224}_{88}\text{Ra}$	3.66 <i>d</i>	(α)
$^{115}_{49}\text{In}$	5.1×10^{15} <i>y</i>	(β^-)	$^{226}_{88}\text{Ra}$	1600 <i>y</i>	(α)
$^{131}_{53}\text{I}$	8.040 <i>d</i>	(β^-)	$^{228}_{88}\text{Ra}$	5.75 <i>y</i>	(β^-)
$^{142}_{58}\text{Ce}$	5×10^{15} <i>y</i>	(α)	$^{228}_{89}\text{Ac}$	6.13 <i>h</i>	(β^-)
$^{208}_{81}\text{Tl}$	3.07 <i>m</i>	(β^-)	$^{228}_{90}\text{Th}$	1.913 <i>y</i>	(α)
$^{210}_{82}\text{Pb}$	22.3 <i>y</i>	(β^-)	$^{232}_{90}\text{Th}$	1.4×10^{10} <i>y</i>	(α)
$^{212}_{82}\text{Pb}$	10.6 <i>h</i>	(β^-)	$^{233}_{90}\text{Th}$	22 <i>m</i>	(β^-)
$^{214}_{82}\text{Pb}$	26.8 <i>m</i>	(β^-)	$^{234}_{90}\text{Th}$	24.10 <i>d</i>	(β^-)
$^{206}_{83}\text{Bi}$	6.243 <i>d</i>	(<i>E.C.</i>)	$^{233}_{91}\text{Pa}$	27 <i>d</i>	(β^-)
$^{233}_{92}\text{U}$	1.59×10^5 <i>y</i>	(α)	$^{242}_{96}\text{Cm}$	162.8 <i>d</i>	(α)
$^{234}_{92}\text{U}$	2.45×10^5 <i>y</i>	(α)	$^{243}_{97}\text{Bk}$	4.5 <i>h</i>	(α or <i>E.C.</i>)
$^{235}_{92}\text{U}$	7.03×10^8 <i>y</i>	(α)	$^{253}_{99}\text{Es}$	20.47 <i>d</i>	(α)
$^{238}_{92}\text{U}$	4.47×10^9 <i>y</i>	(α)	$^{254}_{100}\text{Fm}$	3.24 <i>h</i>	(α or <i>S.F.</i>)
$^{239}_{92}\text{U}$	23.54 <i>m</i>	(β^-)	$^{255}_{100}\text{Fm}$	20.1 <i>h</i>	(α)
$^{239}_{93}\text{Np}$	2.3 <i>d</i>	(β^-)	$^{256}_{101}\text{Md}$	76 <i>m</i>	(α or <i>E.C.</i>)
$^{239}_{94}\text{Pu}$	2.407×10^4 <i>y</i>	(α)	$^{254}_{102}\text{No}$	55 <i>s</i>	(α)
$^{239}_{94}\text{Pu}$	6.54×10^3 <i>y</i>	(α)	$^{257}_{103}\text{Lr}$	0.65 <i>s</i>	(α)
$^{241}_{94}\text{Pu}$	14.4 <i>y</i>	(α or β^-)	$^{260}_{105}\text{Ha}$	1.5 <i>s</i>	(α or <i>S.F.</i>)
$^{241}_{95}\text{Am}$	432.2 <i>y</i>	(α)	$^{263}_{106}\text{Sg}$	0.8 <i>s</i>	(α or <i>S.F.</i>)

Footnotes

1. *y* = years, *d* = days, *h* = hours, *m* = minutes, *s* = seconds
2. *E.C.* = electron capture, *S.F.* = Spontaneous fission

3. y = years, d = days, h = hours, m = minutes, s = seconds
4. *E.C.* = electron capture, *S.F.* = Spontaneous fission

INDEX

A

absorbance

- 1: Beer's Law
- 1.2: Beer's Law - Pre-lab
- 1.3: Beer's Law - Data and Report
- 2: Determination of an Equilibrium Constant
- 2.1: Determination of an Equilibrium Constant - Experiment

accuracy

- 6.3: Solubility Product - Data and Report

Acetic

- 13.2.4: Composition of Commercial Acids and Bases
- 13.2.6: Ionization Constants of Weak Acids

acetic acid

- 5: Investigating Buffers
- 5.2: Investigating Buffers - Pre-lab
- 5.3: Investigating Buffers - Data and Report

acid

- 4.1: Titration of a Diprotic Acid - Experiment
- 4.2: Titration of a Diprotic Acid - Pre-lab

acid dissociation

- 3.2: Acid Dissociation Constant - Pre-lab
- 4: Titration of a Diprotic Acid

acid dissociation constant

- 3: Acid Dissociation Constant

acids

- 13: Appendices
- 13.2.4: Composition of Commercial Acids and Bases
- 13.2.6: Ionization Constants of Weak Acids

activation energy

- 10: Decomposition of Hydrogen Peroxide
- 10.2: Decomposition of Hydrogen Peroxide - Pre-lab
- 11: Determination of an Activation Energy
- 11.3: Determination of an Activation Energy - Data and Report

Ag+

- 13.2.9: Formation Constants for Complex Ions

agents

- 8: Redox Reactions

Al3+

- 13.2.9: Formation Constants for Complex Ions

alpha

- 13.2.11: Half-Lives for Several Radioactive Isotopes

aluminum

- 13.2.10: Standard Electrode (Half-Cell) Potentials
- 13.2.5: Standard Thermodynamic Properties for Selected Substances
- 13.2.8: Solubility Products

American

- 0.1: Lab Safety Videos

ammonia

- 13.2.7: Ionization Constants of Weak Bases

ammonium iron(II) sulfate

- 8.1: Redox Reactions - Experiment

analysis

- 6.2: Solubility Product - Pre-lab
- 7.1: Thermodynamics of Borax - Experiment
- 9: Voltaic Cells

analytical chemistry

- 2.1: Determination of an Equilibrium Constant - Experiment

aniline

- 13.2.7: Ionization Constants of Weak Bases

anode

- 9.3: Voltaic Cells - Data and Report

antimony

- 13.2.5: Standard Thermodynamic Properties for Selected Substances

aqueous

- 13.2.4: Composition of Commercial Acids and Bases

Arrhenius equation

- 11: Determination of an Activation Energy

arsenic

- 13.2.5: Standard Thermodynamic Properties for Selected Substances

atomic mass unit

- 13.2: Tables
- 13.2.2: Fundamental Physical Constants

Avogadro's number

- 13.2: Tables

Avogadro's number

- 13.2.2: Fundamental Physical Constants

B

Barium

- 13.2.5: Standard Thermodynamic Properties for Selected Substances
- 13.2.8: Solubility Products

bases

- 13: Appendices
- 13.2.4: Composition of Commercial Acids and Bases
- 13.2.7: Ionization Constants of Weak Bases

Beer's Law

- 1.3: Beer's Law - Data and Report
- 2.1: Determination of an Equilibrium Constant - Experiment
- 2.2: Determination of an Equilibrium Constant - Pre-lab

Beer's Law

- 1: Beer's Law
- 1.2: Beer's Law - Pre-lab

BeerLaw

- 2: Determination of an Equilibrium Constant

beta

- 13.2.11: Half-Lives for Several Radioactive Isotopes

boiling temperature

- 13.2.3: Water Properties

Boltzmann's constant

- 13.2.2: Fundamental Physical Constants

broken glass disposal

- 0.3: Lab Safety Quiz

buffer

- 5: Investigating Buffers
- 5.2: Investigating Buffers - Pre-lab
- 5.3: Investigating Buffers - Data and Report

buffer capacity

- 5: Investigating Buffers
- 5.3: Investigating Buffers - Data and Report

buffer solution

- 5.1: Investigating Buffers - Experiment

buffer solutions

- 5: Investigating Buffers
- 5.2: Investigating Buffers - Pre-lab

C

cadmium

- 13.2.8: Solubility Products

calcium

- 13.2.8: Solubility Products

calcium iodate

- 6: Solubility Product

calcium ions

- 6.3: Solubility Product - Data and Report

calculations

- 5.2: Investigating Buffers - Pre-lab

calibration

- 1.1: Beer's Law - Experiment

calibration curve

- 1: Beer's Law

calibrationCurve

- 2: Determination of an Equilibrium Constant

capture

- 13.2.11: Half-Lives for Several Radioactive Isotopes

catalyst

- 10.2: Decomposition of Hydrogen Peroxide - Pre-lab

catalysts

- 11: Determination of an Activation Energy

catalyzed

- 11.3: Determination of an Activation Energy - Data and Report

catalyzed decomposition

- 10: Decomposition of Hydrogen Peroxide

cathode

- 9.3: Voltaic Cells - Data and Report

caustic

- 3.1: Acid Dissociation Constant - Experiment
- 5.1: Investigating Buffers - Experiment

cell

- 9: Voltaic Cells

cell potential

- 9.3: Voltaic Cells - Data and Report

Chemical

- 0.1: Lab Safety Videos
- 8: Redox Reactions

chemical disposal

- 10.1: Decomposition of Hydrogen Peroxide - Experiment
- 11.1: Determination of an Activation Energy - Experiment
- 6.1: Solubility Product - Experiment
- 7.1: Thermodynamics of Borax - Experiment
- 8.1: Redox Reactions - Experiment
- 9.1: Voltaic Cells - Experiment

chemical equilibrium

- 2.2: Determination of an Equilibrium Constant - Pre-lab

chemical exposure

- 0: Lab Safety Orientation

chemical hazards

11.1: Determination of an Activation Energy - Experiment

chemical kinetics

10.2: Decomposition of Hydrogen Peroxide - Pre-lab

chemical process

7: Thermodynamics of Borax

chemical reaction

10.2: Decomposition of Hydrogen Peroxide - Pre-lab
11: Determination of an Activation Energy
11.3: Determination of an Activation Energy - Data and Report

chemical reactions

2.1: Determination of an Equilibrium Constant - Experiment

chemical safety

0.2: Lab Safety Contract

chemical spills

0.2: Lab Safety Contract
0.3: Lab Safety Quiz

chemical reactions

2: Determination of an Equilibrium Constant

chemicals

1.1: Beer's Law - Experiment
10.1: Decomposition of Hydrogen Peroxide - Experiment
11.1: Determination of an Activation Energy - Experiment
12.1: Radioactivity Simulation - Experiment
3.1: Acid Dissociation Constant - Experiment
4.1: Titration of a Diprotic Acid - Experiment
5.1: Investigating Buffers - Experiment
6.1: Solubility Product - Experiment
7.1: Thermodynamics of Borax - Experiment
8.1: Redox Reactions - Experiment
9.1: Voltaic Cells - Experiment

chemistry

0.2: Lab Safety Contract
4.2: Titration of a Diprotic Acid - Pre-lab
6.2: Solubility Product - Pre-lab
8.2: Redox Reactions - Pre-lab
9: Voltaic Cells

choking hazard

12.1: Radioactivity Simulation - Experiment

college

0.1: Lab Safety Videos

college policies

0.2: Lab Safety Contract

Complex Ions

13.2.9: Formation Constants for Complex Ions

composition

13.2.4: Composition of Commercial Acids and Bases

concentration

1: Beer's Law
1.1: Beer's Law - Experiment
1.2: Beer's Law - Pre-lab
1.3: Beer's Law - Data and Report
2: Determination of an Equilibrium Constant
2.1: Determination of an Equilibrium Constant - Experiment
3: Acid Dissociation Constant
5.2: Investigating Buffers - Pre-lab
6.2: Solubility Product - Pre-lab
6.3: Solubility Product - Data and Report

concentration calculations

10.2: Decomposition of Hydrogen Peroxide - Pre-lab

concentration cell

9.1: Voltaic Cells - Experiment

concentrations

13.2.4: Composition of Commercial Acids and Bases
4.2: Titration of a Diprotic Acid - Pre-lab

concentrations calculation

2.2: Determination of an Equilibrium Constant - Pre-lab

concepts

13: Appendices

Constants

13.2.7: Ionization Constants of Weak Bases

coordination chemistry

13.2.9: Formation Constants for Complex Ions

copper

13.2.8: Solubility Products
9.3: Voltaic Cells - Data and Report

copper(II) sulfate

1: Beer's Law
1.2: Beer's Law - Pre-lab

Correlation coefficient

1.3: Beer's Law - Data and Report

corrosive

6.1: Solubility Product - Experiment
7.1: Thermodynamics of Borax - Experiment

cryoscopic constant

13.2.3: Water Properties

D

data

1.3: Beer's Law - Data and Report
11.3: Determination of an Activation Energy - Data and Report
12.3: Radioactivity Simulation - Data and Report
6.3: Solubility Product - Data and Report
9.3: Voltaic Cells - Data and Report

data analysis

1.2: Beer's Law - Pre-lab

data collection

11.1: Determination of an Activation Energy - Experiment
12: Radioactivity Simulation
2.1: Determination of an Equilibrium Constant - Experiment

data strategy

3.2: Acid Dissociation Constant - Pre-lab

decay

12: Radioactivity Simulation

decay constant

12.2: Radioactivity Simulation - Pre-lab
12.3: Radioactivity Simulation - Data and Report

decay probability

12.2: Radioactivity Simulation - Pre-lab

decomposition reaction

10.2: Decomposition of Hydrogen Peroxide - Pre-lab

deionized water

3.1: Acid Dissociation Constant - Experiment

density

13.2.3: Water Properties
13.2.4: Composition of Commercial Acids and Bases

dice

12.1: Radioactivity Simulation - Experiment

dilution

1.2: Beer's Law - Pre-lab

dilution equation

2.2: Determination of an Equilibrium Constant - Pre-lab

dimethylamine

13.2.7: Ionization Constants of Weak Bases

diprotic

4: Titration of a Diprotic Acid

diprotic acid

4.2: Titration of a Diprotic Acid - Pre-lab

disposal

4.1: Titration of a Diprotic Acid - Experiment

dissociation

4.2: Titration of a Diprotic Acid - Pre-lab

E

ebullioscopic constant

13.2.3: Water Properties

electrochemistry

13.2.10: Standard Electrode (Half-Cell) Potentials
9: Voltaic Cells
9.3: Voltaic Cells - Data and Report

electrode potentials

13.2.10: Standard Electrode (Half-Cell) Potentials

electrodes

9: Voltaic Cells
9.1: Voltaic Cells - Experiment

electron

13.2.11: Half-Lives for Several Radioactive Isotopes

electron charge

13.2.2: Fundamental Physical Constants

electron mass

13.2: Tables

emergency equipment

0: Lab Safety Orientation

emergency response

0.3: Lab Safety Quiz

emission

13.2.11: Half-Lives for Several Radioactive Isotopes

enthalpy

13.2.3: Water Properties
13.2.5: Standard Thermodynamic Properties for Selected Substances
7: Thermodynamics of Borax

entropy

13.2.5: Standard Thermodynamic Properties for Selected Substances
7: Thermodynamics of Borax

enzyme catalase

10: Decomposition of Hydrogen Peroxide

equations

8: Redox Reactions

equilibrium

13.2.9: Formation Constants for Complex Ions
2: Determination of an Equilibrium Constant

equilibrium concentration

2.2: Determination of an Equilibrium Constant - Pre-lab

equilibrium constant

2.1: Determination of an Equilibrium Constant - Experiment
6: Solubility Product
7: Thermodynamics of Borax

equipment

- 1.1: Beer's Law - Experiment
- 10.1: Decomposition of Hydrogen Peroxide - Experiment
- 11.1: Determination of an Activation Energy - Experiment
- 12.1: Radioactivity Simulation - Experiment
- 3.1: Acid Dissociation Constant - Experiment
- 4.1: Titration of a Diprotic Acid - Experiment
- 5.1: Investigating Buffers - Experiment
- 6.1: Solubility Product - Experiment
- 9.1: Voltaic Cells - Experiment

equipment list

- 11: Determination of an Activation Energy

equivalence point

- 3: Acid Dissociation Constant
- 3.2: Acid Dissociation Constant - Pre-lab
- 4: Titration of a Diprotic Acid

error

- 12.3: Radioactivity Simulation - Data and Report

experiment

- 1.3: Beer's Law - Data and Report
- 12: Radioactivity Simulation
- 12.3: Radioactivity Simulation - Data and Report
- 6.2: Solubility Product - Pre-lab
- 8: Redox Reactions
- 9: Voltaic Cells

experimental data

- 11: Determination of an Activation Energy

experimental procedure

- 11.1: Determination of an Activation Energy - Experiment
- 12.1: Radioactivity Simulation - Experiment
- 3.1: Acid Dissociation Constant - Experiment
- 6.1: Solubility Product - Experiment
- 7.1: Thermodynamics of Borax - Experiment

exponential trendline

- 12.2: Radioactivity Simulation - Pre-lab

eye protection

- 0.3: Lab Safety Quiz

F

Faraday's constant

- 13.2: Tables

Faraday's constant

- 13.2.2: Fundamental Physical Constants

first aid

- 0.3: Lab Safety Quiz

fission

- 13.2.11: Half-Lives for Several Radioactive Isotopes

formation constants

- 13.2.9: Formation Constants for Complex Ions

fundamental

- 13: Appendices

fundamental physical constants

- 13.2.2: Fundamental Physical Constants

G

gas constant

- 13.2.2: Fundamental Physical Constants

gas pressure sensor

- 10.1: Decomposition of Hydrogen Peroxide - Experiment

Gibbs Free Energy

- 13.2.5: Standard Thermodynamic Properties for Selected Substances
- 7: Thermodynamics of Borax

gloves

- 8.1: Redox Reactions - Experiment

goggles

- 1.1: Beer's Law - Experiment
- 9.1: Voltaic Cells - Experiment

gold

- 13.2.10: Standard Electrode (Half-Cell) Potentials

graph

- 12.3: Radioactivity Simulation - Data and Report

H

H2A

- 4.2: Titration of a Diprotic Acid - Pre-lab

halogens

- 8.2: Redox Reactions - Pre-lab

handling

- 1.1: Beer's Law - Experiment

handling chemicals

- 0.3: Lab Safety Quiz

hazardous materials

- 0.2: Lab Safety Contract

HCl

- 5.3: Investigating Buffers - Data and Report

hydrochloric acid

- 5.1: Investigating Buffers - Experiment
- 6.1: Solubility Product - Experiment
- 6.2: Solubility Product - Pre-lab
- 7.1: Thermodynamics of Borax - Experiment
- 8.1: Redox Reactions - Experiment

hydrocyanic

- 13.2.6: Ionization Constants of Weak Acids

hydrofluoric

- 13.2.6: Ionization Constants of Weak Acids

hydrogen

- 13.2.6: Ionization Constants of Weak Acids

hydrogen peroxide

- 10: Decomposition of Hydrogen Peroxide
- 10.1: Decomposition of Hydrogen Peroxide - Experiment
- 10.2: Decomposition of Hydrogen Peroxide - Pre-lab
- 8.1: Redox Reactions - Experiment

I

initial rate method

- 10: Decomposition of Hydrogen Peroxide

ionization

- 13.2.6: Ionization Constants of Weak Acids
- 13.2.7: Ionization Constants of Weak Bases

iron

- 13.2.10: Standard Electrode (Half-Cell) Potentials
- 13.2.8: Solubility Products

iron ion

- 8.2: Redox Reactions - Pre-lab

iron thiocyanate

- 2.1: Determination of an Equilibrium Constant - Experiment

isotope

- 13.2.11: Half-Lives for Several Radioactive Isotopes

K

Ka

- 3: Acid Dissociation Constant
- 4: Titration of a Diprotic Acid

Ka expression

- 3.2: Acid Dissociation Constant - Pre-lab

Ka1

- 4.2: Titration of a Diprotic Acid - Pre-lab

Ka2

- 4.2: Titration of a Diprotic Acid - Pre-lab

Kc

- 2: Determination of an Equilibrium Constant

Kf

- 13.2.9: Formation Constants for Complex Ions

Ksp

- 13.2.8: Solubility Products
- 6.2: Solubility Product - Pre-lab
- 6.3: Solubility Product - Data and Report

L

lab

- 0.1: Lab Safety Videos

lab attire

- 0.3: Lab Safety Quiz

lab equipment

- 0.2: Lab Safety Contract
- 7.1: Thermodynamics of Borax - Experiment

lab procedures

- 2.1: Determination of an Equilibrium Constant - Experiment

lab safety

- 0: Lab Safety Orientation
- 0.3: Lab Safety Quiz

laboratory

- 6.3: Solubility Product - Data and Report
- 9.3: Voltaic Cells - Data and Report

laboratory environment

- 0.2: Lab Safety Contract

laboratory experiment

- 3: Acid Dissociation Constant
- 7: Thermodynamics of Borax

laboratory procedure

- 10.1: Decomposition of Hydrogen Peroxide - Experiment

laboratory safety

- 0: Lab Safety Orientation
- 11.1: Determination of an Activation Energy - Experiment
- 5.1: Investigating Buffers - Experiment

law of large numbers

- 12.1: Radioactivity Simulation - Experiment

lead

- 13.2.8: Solubility Products

LeChatelier

- 2.2: Determination of an Equilibrium Constant - Pre-lab

lechateliersprinciple

- 2: Determination of an Equilibrium Constant

linear regression

- 1: Beer's Law
- 1.2: Beer's Law - Pre-lab

M

magnesium

- 13.2.8: Solubility Products

materials list

- 5.1: Investigating Buffers - Experiment
- 6.1: Solubility Product - Experiment

measurements

- 9: Voltaic Cells
- 9.3: Voltaic Cells - Data and Report

melting temperature

13.2.3: Water Properties

metal complexes

13.2.9: Formation Constants for Complex Ions

methylamine

13.2.7: Ionization Constants of Weak Bases

molarity

10.2: Decomposition of Hydrogen Peroxide - Pre-lab

13.2.4: Composition of Commercial Acids and Bases

6: Solubility Product

monoprotic acid

3.2: Acid Dissociation Constant - Pre-lab

N

NaOH

5.3: Investigating Buffers - Data and Report

NaOH titration

3.2: Acid Dissociation Constant - Pre-lab

Nernst

9: Voltaic Cells

neutralization

5.2: Investigating Buffers - Pre-lab

5.3: Investigating Buffers - Data and Report

neutron mass

13.2: Tables

nuclear reaction

12.2: Radioactivity Simulation - Pre-lab

O

observations

11.3: Determination of an Activation Energy - Data and Report

12.3: Radioactivity Simulation - Data and Report

6.3: Solubility Product - Data and Report

8: Redox Reactions

9.3: Voltaic Cells - Data and Report

order determination

10: Decomposition of Hydrogen Peroxide

orientation

0.1: Lab Safety Videos

outlier

1.2: Beer's Law - Pre-lab

oxidation

8: Redox Reactions

8.2: Redox Reactions - Pre-lab

P

Periodic

13: Appendices

periodic table

13.2: Tables

peroxide

13.2.6: Ionization Constants of Weak Acids

pH

4.1: Titration of a Diprotic Acid - Experiment

5: Investigating Buffers

5.2: Investigating Buffers - Pre-lab

5.3: Investigating Buffers - Data and Report

pH probe

3: Acid Dissociation Constant

3.1: Acid Dissociation Constant - Experiment

4: Titration of a Diprotic Acid

phenylamine

13.2.7: Ionization Constants of Weak Bases

photometry

1: Beer's Law

physical

13: Appendices

physical constants

13.2: Tables

pKa

3: Acid Dissociation Constant

4: Titration of a Diprotic Acid

Planck's constant

13.2: Tables

Planck's constant

13.2.2: Fundamental Physical Constants

potassium ferricyanide

8.2: Redox Reactions - Pre-lab

potassium iodide

10.1: Decomposition of Hydrogen Peroxide - Experiment

6: Solubility Product

6.2: Solubility Product - Pre-lab

potassium thiocyanate

8.2: Redox Reactions - Pre-lab

precautions

10.1: Decomposition of Hydrogen Peroxide - Experiment

12.1: Radioactivity Simulation - Experiment

3.1: Acid Dissociation Constant - Experiment

4.1: Titration of a Diprotic Acid - Experiment

probability

12: Radioactivity Simulation

12.1: Radioactivity Simulation - Experiment

procedure

4.1: Titration of a Diprotic Acid - Experiment

properties

13: Appendices

13.2.4: Composition of Commercial Acids and Bases

protective equipment

11.1: Determination of an Activation Energy - Experiment

proton mass

13.2: Tables

R

Radioactive

12: Radioactivity Simulation

13.2.11: Half-Lives for Several Radioactive Isotopes

radioactive decay

12.1: Radioactivity Simulation - Experiment

12.2: Radioactivity Simulation - Pre-lab

radioactive isotope

12.2: Radioactivity Simulation - Pre-lab

radioactive isotopes

12: Radioactivity Simulation

random process

12: Radioactivity Simulation

rate

11.3: Determination of an Activation Energy - Data and Report

rate constant

10: Decomposition of Hydrogen Peroxide

rate law

10: Decomposition of Hydrogen Peroxide

reaction

9.1: Voltaic Cells - Experiment

reaction kinetics

10: Decomposition of Hydrogen Peroxide

11: Determination of an Activation Energy

reaction rate

10.1: Decomposition of Hydrogen Peroxide - Experiment

11: Determination of an Activation Energy

reaction rates

11.1: Determination of an Activation Energy - Experiment

reactions

4.2: Titration of a Diprotic Acid - Pre-lab

8: Redox Reactions

redox

8: Redox Reactions

8.2: Redox Reactions - Pre-lab

9: Voltaic Cells

redox reactions

13.2.10: Standard Electrode (Half-Cell) Potentials

redox titration

6: Solubility Product

reducing agent

8.2: Redox Reactions - Pre-lab

reduction

8: Redox Reactions

reference solutions

2.2: Determination of an Equilibrium Constant - Pre-lab

S

safety

0.1: Lab Safety Videos

1.1: Beer's Law - Experiment

10.1: Decomposition of Hydrogen Peroxide - Experiment

11.1: Determination of an Activation Energy - Experiment

12.1: Radioactivity Simulation - Experiment

3.1: Acid Dissociation Constant - Experiment

4.1: Titration of a Diprotic Acid - Experiment

6.1: Solubility Product - Experiment

7.1: Thermodynamics of Borax - Experiment

8.1: Redox Reactions - Experiment

9.1: Voltaic Cells - Experiment

safety attire

0: Lab Safety Orientation

safety contract

0: Lab Safety Orientation

safety equipment

0.3: Lab Safety Quiz

safety goggles

0: Lab Safety Orientation

0.2: Lab Safety Contract

safety precautions

10: Decomposition of Hydrogen Peroxide

11: Determination of an Activation Energy

5.1: Investigating Buffers - Experiment

6: Solubility Product

safety procedures

0.2: Lab Safety Contract

0.3: Lab Safety Quiz

safety quiz

0: Lab Safety Orientation

safety rules

0: Lab Safety Orientation

salt bridge

9.1: Voltaic Cells - Experiment

saturated solution

6: Solubility Product

scatterplot

12.3: Radioactivity Simulation - Data and Report

scientific

13: Appendices

significant figures

10.2: Decomposition of Hydrogen Peroxide - Pre-lab

silver

13.2.10: Standard Electrode (Half-Cell) Potentials

simulation

12: Radioactivity Simulation

Society

0.1: Lab Safety Videos

sodium

4.1: Titration of a Diprotic Acid - Experiment

sodium acetate

5: Investigating Buffers
5.2: Investigating Buffers - Pre-lab

sodium bromide

8.1: Redox Reactions - Experiment

sodium hydroxide

4: Titration of a Diprotic Acid
5.1: Investigating Buffers - Experiment
8.1: Redox Reactions - Experiment

sodium thiosulfate

6.3: Solubility Product - Data and Report

solubility

13: Appendices
13.2.8: Solubility Products
6: Solubility Product
6.3: Solubility Product - Data and Report
7: Thermodynamics of Borax

solubility product

6: Solubility Product
7: Thermodynamics of Borax

solution preparation

2.2: Determination of an Equilibrium Constant - Pre-lab

solutionchemistry

2: Determination of an Equilibrium Constant

solutions

13.2.4: Composition of Commercial Acids and Bases

specific heat capacity

13.2.3: Water Properties

spectrometer calibration

1.2: Beer's Law - Pre-lab

spectrophotometer

1.1: Beer's Law - Experiment
1.3: Beer's Law - Data and Report
2.1: Determination of an Equilibrium Constant - Experiment

Spectrophotometry

2: Determination of an Equilibrium Constant

spectroscopy

1: Beer's Law

speed of light

13.2: Tables
13.2.2: Fundamental Physical Constants

spills

1.1: Beer's Law - Experiment

spontaneous

13.2.11: Half-Lives for Several Radioactive Isotopes

standard electrode potentials

13.2.10: Standard Electrode (Half-Cell) Potentials

standard solution

1.3: Beer's Law - Data and Report

standard solutions

1: Beer's Law
1.2: Beer's Law - Pre-lab

standard thermodynamic properties

13.2.5: Standard Thermodynamic Properties for Selected Substances

standardized solution

7.1: Thermodynamics of Borax - Experiment

starch indicator

6: Solubility Product
6.2: Solubility Product - Pre-lab

statistics

12.3: Radioactivity Simulation - Data and Report

strong acid

5: Investigating Buffers

strong acids

5.3: Investigating Buffers - Data and Report

strong base

5: Investigating Buffers

student safety

0: Lab Safety Orientation

students

0.1: Lab Safety Videos

substance

13.2.5: Standard Thermodynamic Properties for Selected Substances

sulfate

13.2.6: Ionization Constants of Weak Acids

sulfate ion

8.2: Redox Reactions - Pre-lab

sulfide

13.2.6: Ionization Constants of Weak Acids

sulfite ion

8.2: Redox Reactions - Pre-lab

sulfuric acid

8.1: Redox Reactions - Experiment

T

techniques

0.1: Lab Safety Videos

temperature

11: Determination of an Activation Energy
11.3: Determination of an Activation Energy - Data and Report
13.2.3: Water Properties

test solutions

2.2: Determination of an Equilibrium Constant - Pre-lab

theoretical prediction

12.2: Radioactivity Simulation - Pre-lab

theory

12.3: Radioactivity Simulation - Data and Report

thermodynamic

13: Appendices

thermodynamics

13.2.5: Standard Thermodynamic Properties for Selected Substances
7: Thermodynamics of Borax

time

11.3: Determination of an Activation Energy - Data and Report

titration

3: Acid Dissociation Constant
3.1: Acid Dissociation Constant - Experiment
4: Titration of a Diprotic Acid
4.1: Titration of a Diprotic Acid - Experiment
5.1: Investigating Buffers - Experiment
5.3: Investigating Buffers - Data and Report
6.1: Solubility Product - Experiment
6.3: Solubility Product - Data and Report
7.1: Thermodynamics of Borax - Experiment

titration curve

4: Titration of a Diprotic Acid

titration curves

3.2: Acid Dissociation Constant - Pre-lab

titration process

6.2: Solubility Product - Pre-lab

transition metals

13.2.9: Formation Constants for Complex Ions

trimethylamine

13.2.7: Ionization Constants of Weak Bases

Triton College

0.2: Lab Safety Contract

U

uncatalyzed

11.3: Determination of an Activation Energy - Data and Report

unit conversion

10.2: Decomposition of Hydrogen Peroxide - Pre-lab

unknown concentration

1.3: Beer's Law - Data and Report

V

vapor pressure

13.2.3: Water Properties

videos

0.1: Lab Safety Videos

voltaic

9: Voltaic Cells

voltaic cell

9.1: Voltaic Cells - Experiment

volume calculation

3.2: Acid Dissociation Constant - Pre-lab

volumes mixing

2.2: Determination of an Equilibrium Constant - Pre-lab

W

waste disposal

3.1: Acid Dissociation Constant - Experiment

waste management

6.1: Solubility Product - Experiment

water

13.2.3: Water Properties

wavelength

1.3: Beer's Law - Data and Report

weak

13.2.6: Ionization Constants of Weak Acids
13.2.7: Ionization Constants of Weak Bases

weak acid

3: Acid Dissociation Constant
4: Titration of a Diprotic Acid
5: Investigating Buffers

weak acid titration

3.2: Acid Dissociation Constant - Pre-lab

weak acids

5.2: Investigating Buffers - Pre-lab

5.3: Investigating Buffers - Data and Report

Z

zinc

9.3: Voltaic Cells - Data and Report

Zn²⁺

13.2.9: Formation Constants for Complex Ions

GLOSSARY

S

Sample Word 1 | Sample Definition 1

DETAILED LICENSING

Overview

Title: General Chemistry II Lab Manual

Webpages: 78

All licenses found:

- [CC BY 4.0](#): 97.4% (76 pages)
- [Undeclared](#): 2.6% (2 pages)

By Page

- General Chemistry II Lab Manual - [CC BY 4.0](#)
 - Front Matter - [CC BY 4.0](#)
 - TitlePage - [CC BY 4.0](#)
 - InfoPage - [CC BY 4.0](#)
 - Table of Contents - [Undeclared](#)
 - Licensing - [CC BY 4.0](#)
 - 0: Lab Safety Orientation - [CC BY 4.0](#)
 - 0.1: Lab Safety Videos - [CC BY 4.0](#)
 - 0.2: Lab Safety Contract - [CC BY 4.0](#)
 - 0.3: Lab Safety Quiz - [CC BY 4.0](#)
 - 0.4: Additional Resources - [CC BY 4.0](#)
 - 1: Beer's Law - [CC BY 4.0](#)
 - 1.1: Beer's Law - Experiment - [CC BY 4.0](#)
 - 1.2: Beer's Law - Pre-lab - [CC BY 4.0](#)
 - 1.3: Beer's Law - Data and Report - [CC BY 4.0](#)
 - 2: Determination of an Equilibrium Constant - [CC BY 4.0](#)
 - 2.1: Determination of an Equilibrium Constant - Experiment - [CC BY 4.0](#)
 - 2.2: Determination of an Equilibrium Constant - Pre-lab - [CC BY 4.0](#)
 - 2.3: Determination of an Equilibrium Constant - Data and Report - [CC BY 4.0](#)
 - 3: Acid Dissociation Constant - [CC BY 4.0](#)
 - 3.1: Acid Dissociation Constant - Experiment - [CC BY 4.0](#)
 - 3.2: Acid Dissociation Constant - Pre-lab - [CC BY 4.0](#)
 - 3.3: Acid Dissociation Constant - Data and Report - [CC BY 4.0](#)
 - 4: Titration of a Diprotic Acid - [CC BY 4.0](#)
 - 4.1: Titration of a Diprotic Acid - Experiment - [CC BY 4.0](#)
 - 4.2: Titration of a Diprotic Acid - Pre-lab - [CC BY 4.0](#)
 - 4.3: Titration of a Diprotic Acid - Data and Report - [CC BY 4.0](#)
 - 5: Investigating Buffers - [CC BY 4.0](#)
 - 5.1: Investigating Buffers - Experiment - [CC BY 4.0](#)
 - 5.2: Investigating Buffers - Pre-lab - [CC BY 4.0](#)
 - 5.3: Investigating Buffers - Data and Report - [CC BY 4.0](#)
 - 6: Solubility Product - [CC BY 4.0](#)
 - 6.1: Solubility Product - Experiment - [CC BY 4.0](#)
 - 6.2: Solubility Product - Pre-lab - [CC BY 4.0](#)
 - 6.3: Solubility Product - Data and Report - [CC BY 4.0](#)
 - 7: Thermodynamics of Borax - [CC BY 4.0](#)
 - 7.1: Thermodynamics of Borax - Experiment - [CC BY 4.0](#)
 - 7.2: Thermodynamics of Borax - Pre-lab - [CC BY 4.0](#)
 - 7.3: Thermodynamics of Borax - Data and Report - [CC BY 4.0](#)
 - 8: Redox Reactions - [CC BY 4.0](#)
 - 8.1: Redox Reactions - Experiment - [CC BY 4.0](#)
 - 8.2: Redox Reactions - Pre-lab - [CC BY 4.0](#)
 - 8.3: Redox Reactions - Data and Report - [CC BY 4.0](#)
 - 9: Voltaic Cells - [CC BY 4.0](#)
 - 9.1: Voltaic Cells - Experiment - [CC BY 4.0](#)
 - 9.2: Voltaic Cells - Pre-lab - [CC BY 4.0](#)
 - 9.3: Voltaic Cells - Data and Report - [CC BY 4.0](#)
 - 10: Decomposition of Hydrogen Peroxide - [CC BY 4.0](#)
 - 10.1: Decomposition of Hydrogen Peroxide - Experiment - [CC BY 4.0](#)
 - 10.2: Decomposition of Hydrogen Peroxide - Pre-lab - [CC BY 4.0](#)
 - 10.3: Decomposition of Hydrogen Peroxide - Data and Report - [CC BY 4.0](#)
 - 11: Determination of an Activation Energy - [CC BY 4.0](#)
 - 11.1: Determination of an Activation Energy - Experiment - [CC BY 4.0](#)
 - 11.2: Determination of an Activation Energy - Pre-lab - [CC BY 4.0](#)
 - 11.3: Determination of an Activation Energy - Data and Report - [CC BY 4.0](#)

- 12: Radioactivity Simulation - [CC BY 4.0](#)
 - 12.1: Radioactivity Simulation - Experiment - [CC BY 4.0](#)
 - 12.2: Radioactivity Simulation - Pre-lab - [CC BY 4.0](#)
 - 12.3: Radioactivity Simulation - Data and Report - [CC BY 4.0](#)
- 13: Appendices - [CC BY 4.0](#)
 - 13.1: Vernier LabQuest 3 Instructions - [Undeclared](#)
 - 13.2: Tables - [CC BY 4.0](#)
 - 13.2.1: The Periodic Table - [CC BY 4.0](#)
 - 13.2.2: Fundamental Physical Constants - [CC BY 4.0](#)
 - 13.2.3: Water Properties - [CC BY 4.0](#)
 - 13.2.4: Composition of Commercial Acids and Bases - [CC BY 4.0](#)
 - 13.2.5: Standard Thermodynamic Properties for Selected Substances - [CC BY 4.0](#)
 - 13.2.6: Ionization Constants of Weak Acids - [CC BY 4.0](#)
 - 13.2.7: Ionization Constants of Weak Bases - [CC BY 4.0](#)
 - 13.2.8: Solubility Products - [CC BY 4.0](#)
 - 13.2.9: Formation Constants for Complex Ions - [CC BY 4.0](#)
 - 13.2.10: Standard Electrode (Half-Cell) Potentials - [CC BY 4.0](#)
 - 13.2.11: Half-Lives for Several Radioactive Isotopes - [CC BY 4.0](#)
- Back Matter - [CC BY 4.0](#)
 - Index - [CC BY 4.0](#)
 - Glossary - [CC BY 4.0](#)
 - Detailed Licensing - [CC BY 4.0](#)
 - Detailed Licensing - [CC BY 4.0](#)

DETAILED LICENSING

Overview

Title: General Chemistry II Lab Manual

Webpages: 78

All licenses found:

- [CC BY 4.0](#): 97.4% (76 pages)
- [Undeclared](#): 2.6% (2 pages)

By Page

- General Chemistry II Lab Manual - [CC BY 4.0](#)
 - Front Matter - [CC BY 4.0](#)
 - TitlePage - [CC BY 4.0](#)
 - InfoPage - [CC BY 4.0](#)
 - Table of Contents - [Undeclared](#)
 - Licensing - [CC BY 4.0](#)
 - 0: Lab Safety Orientation - [CC BY 4.0](#)
 - 0.1: Lab Safety Videos - [CC BY 4.0](#)
 - 0.2: Lab Safety Contract - [CC BY 4.0](#)
 - 0.3: Lab Safety Quiz - [CC BY 4.0](#)
 - 0.4: Additional Resources - [CC BY 4.0](#)
 - 1: Beer's Law - [CC BY 4.0](#)
 - 1.1: Beer's Law - Experiment - [CC BY 4.0](#)
 - 1.2: Beer's Law - Pre-lab - [CC BY 4.0](#)
 - 1.3: Beer's Law - Data and Report - [CC BY 4.0](#)
 - 2: Determination of an Equilibrium Constant - [CC BY 4.0](#)
 - 2.1: Determination of an Equilibrium Constant - Experiment - [CC BY 4.0](#)
 - 2.2: Determination of an Equilibrium Constant - Pre-lab - [CC BY 4.0](#)
 - 2.3: Determination of an Equilibrium Constant - Data and Report - [CC BY 4.0](#)
 - 3: Acid Dissociation Constant - [CC BY 4.0](#)
 - 3.1: Acid Dissociation Constant - Experiment - [CC BY 4.0](#)
 - 3.2: Acid Dissociation Constant - Pre-lab - [CC BY 4.0](#)
 - 3.3: Acid Dissociation Constant - Data and Report - [CC BY 4.0](#)
 - 4: Titration of a Diprotic Acid - [CC BY 4.0](#)
 - 4.1: Titration of a Diprotic Acid - Experiment - [CC BY 4.0](#)
 - 4.2: Titration of a Diprotic Acid - Pre-lab - [CC BY 4.0](#)
 - 4.3: Titration of a Diprotic Acid - Data and Report - [CC BY 4.0](#)
 - 5: Investigating Buffers - [CC BY 4.0](#)
 - 5.1: Investigating Buffers - Experiment - [CC BY 4.0](#)
 - 5.2: Investigating Buffers - Pre-lab - [CC BY 4.0](#)
 - 5.3: Investigating Buffers - Data and Report - [CC BY 4.0](#)
 - 6: Solubility Product - [CC BY 4.0](#)
 - 6.1: Solubility Product - Experiment - [CC BY 4.0](#)
 - 6.2: Solubility Product - Pre-lab - [CC BY 4.0](#)
 - 6.3: Solubility Product - Data and Report - [CC BY 4.0](#)
 - 7: Thermodynamics of Borax - [CC BY 4.0](#)
 - 7.1: Thermodynamics of Borax - Experiment - [CC BY 4.0](#)
 - 7.2: Thermodynamics of Borax - Pre-lab - [CC BY 4.0](#)
 - 7.3: Thermodynamics of Borax - Data and Report - [CC BY 4.0](#)
 - 8: Redox Reactions - [CC BY 4.0](#)
 - 8.1: Redox Reactions - Experiment - [CC BY 4.0](#)
 - 8.2: Redox Reactions - Pre-lab - [CC BY 4.0](#)
 - 8.3: Redox Reactions - Data and Report - [CC BY 4.0](#)
 - 9: Voltaic Cells - [CC BY 4.0](#)
 - 9.1: Voltaic Cells - Experiment - [CC BY 4.0](#)
 - 9.2: Voltaic Cells - Pre-lab - [CC BY 4.0](#)
 - 9.3: Voltaic Cells - Data and Report - [CC BY 4.0](#)
 - 10: Decomposition of Hydrogen Peroxide - [CC BY 4.0](#)
 - 10.1: Decomposition of Hydrogen Peroxide - Experiment - [CC BY 4.0](#)
 - 10.2: Decomposition of Hydrogen Peroxide - Pre-lab - [CC BY 4.0](#)
 - 10.3: Decomposition of Hydrogen Peroxide - Data and Report - [CC BY 4.0](#)
 - 11: Determination of an Activation Energy - [CC BY 4.0](#)
 - 11.1: Determination of an Activation Energy - Experiment - [CC BY 4.0](#)
 - 11.2: Determination of an Activation Energy - Pre-lab - [CC BY 4.0](#)
 - 11.3: Determination of an Activation Energy - Data and Report - [CC BY 4.0](#)

- 12: Radioactivity Simulation - [CC BY 4.0](#)
 - 12.1: Radioactivity Simulation - Experiment - [CC BY 4.0](#)
 - 12.2: Radioactivity Simulation - Pre-lab - [CC BY 4.0](#)
 - 12.3: Radioactivity Simulation - Data and Report - [CC BY 4.0](#)
- 13: Appendices - [CC BY 4.0](#)
 - 13.1: Vernier LabQuest 3 Instructions - [Undeclared](#)
 - 13.2: Tables - [CC BY 4.0](#)
 - 13.2.1: The Periodic Table - [CC BY 4.0](#)
 - 13.2.2: Fundamental Physical Constants - [CC BY 4.0](#)
 - 13.2.3: Water Properties - [CC BY 4.0](#)
 - 13.2.4: Composition of Commercial Acids and Bases - [CC BY 4.0](#)
 - 13.2.5: Standard Thermodynamic Properties for Selected Substances - [CC BY 4.0](#)
 - 13.2.6: Ionization Constants of Weak Acids - [CC BY 4.0](#)
 - 13.2.7: Ionization Constants of Weak Bases - [CC BY 4.0](#)
 - 13.2.8: Solubility Products - [CC BY 4.0](#)
 - 13.2.9: Formation Constants for Complex Ions - [CC BY 4.0](#)
 - 13.2.10: Standard Electrode (Half-Cell) Potentials - [CC BY 4.0](#)
 - 13.2.11: Half-Lives for Several Radioactive Isotopes - [CC BY 4.0](#)
- Back Matter - [CC BY 4.0](#)
 - Index - [CC BY 4.0](#)
 - Glossary - [CC BY 4.0](#)
 - Detailed Licensing - [CC BY 4.0](#)
 - Detailed Licensing - [CC BY 4.0](#)