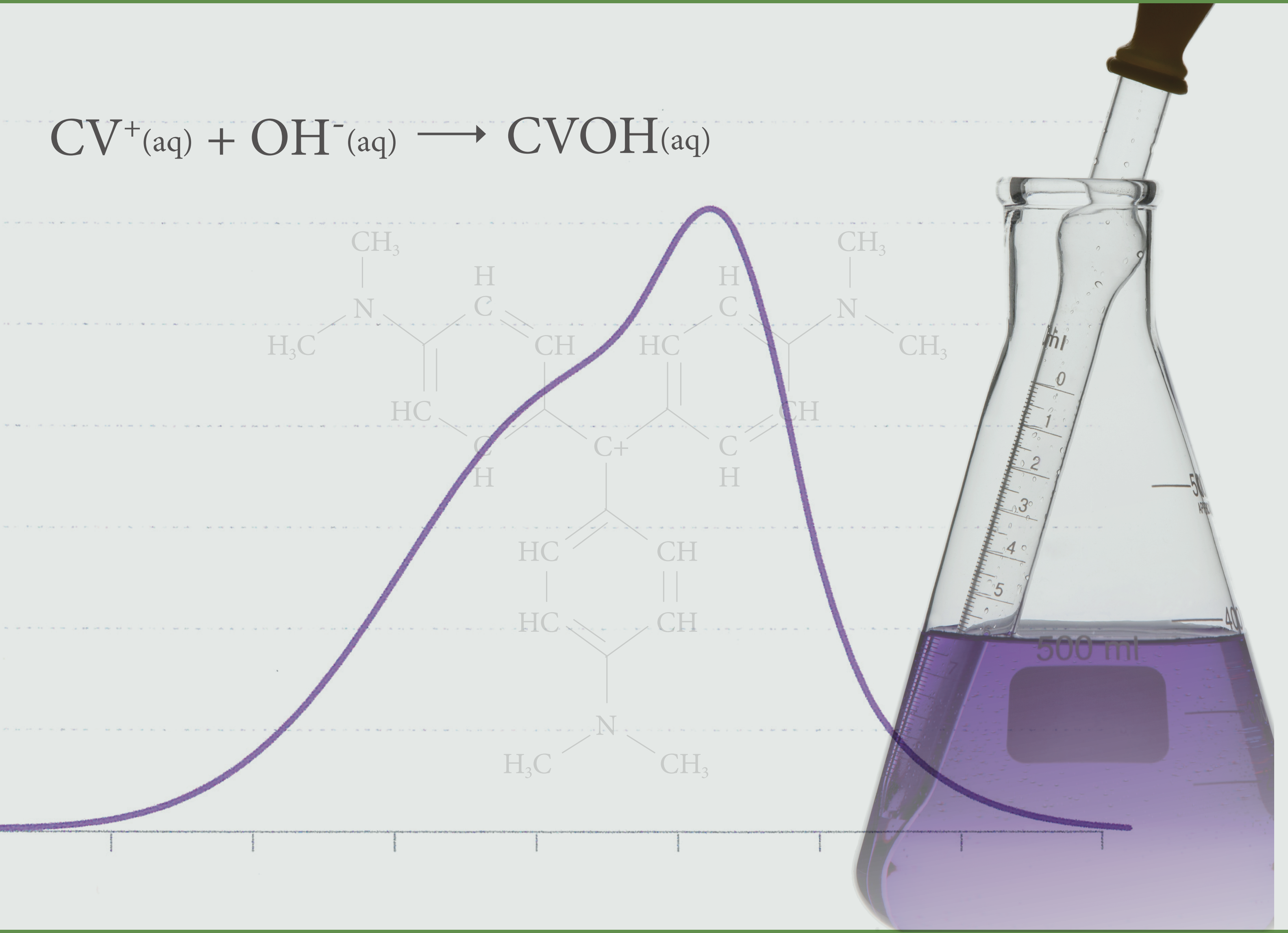
CollegeBoard

AP® CHEMISTRY

GUIDED-INQUIRY EXPERIMENTS:

Applying the Science Practices



$$CV^{+}(aq) + OH^{-}(aq) \longrightarrow CVOH(aq)$$

Student Manual

UPDATED FOR FALL 2019

CollegeBoard

AP® Chemistry Guided-Inquiry Experiments:

Applying the Science Practices

The College Board
New York, NY

Student Manual

ABOUT THE COLLEGE BOARD

The College Board is a mission-driven not-for-profit organization that connects students to college success and opportunity. Founded in 1900, the College Board was created to expand access to higher education. Today, the membership association is made up of over 6,000 of the world's leading educational institutions and is dedicated to promoting excellence and equity in education. Each year, the College Board helps more than seven million students prepare for a successful transition to college through programs and services in college readiness and college success — including the SAT® and the Advanced Placement Program®. The organization also serves the education community through research and advocacy on behalf of students, educators and schools.

For further information, visit www.collegeboard.org.

AP® EQUITY AND ACCESS POLICY

The College Board strongly encourages educators to make equitable access a guiding principle for their AP programs by giving all willing and academically prepared students the opportunity to participate in AP. We encourage the elimination of barriers that restrict access to AP for students from ethnic, racial and socioeconomic groups that have been traditionally underserved. Schools should make every effort to ensure their AP classes reflect the diversity of their student population. The College Board also believes that all students should have access to academically challenging course work before they enroll in AP classes, which can prepare them for AP success. It is only through a commitment to equitable preparation and access that true equity and excellence can be achieved.

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MAJOR CONTRIBUTORS

- Kristen Cacciatore, East Boston High School, Boston, MA; lead writer
- Thomas Holme, Iowa State University, Ames, IA; lead editor
- Jim Spencer, Franklin & Marshall College, Lancaster, PA; lead consultant

AUTHORS

- Jill Barker, Millbrook High School in Winchester, VA
- Jeff Bigler, Lynn English High School, Lynn, MA
- Brenda Brockland, Prairie Ridge High School, Crystal Lake, IL
- Mark Case, Emmaus High School, Emmaus, PA
- Bill Cumming, Gilmour Academy, Gates Mills, OH
- Marian DeWane, University of California, Irvine, CA
- Adrian Dingle, The Westminster Schools, Atlanta, GA
- Gary Freebury, Flathead Valley Community College, Kalispell, MT
- Brenda Lee Fredette, Medaille College, Buffalo, NY
- Cara Hale-Hanes, LB Polytechnic High School, Long Beach, CA
- Annis Hapkiewicz, Okemos High School, Okemos, MI
- Jeff Hepburn, Central Academy, Des Moines, IA
- David Hostage, The Taft School, Watertown, CT
- Carrie Jacobus, River Dell High School, Oradell, NJ
- Bryan Marten, Lowell High School, San Francisco, CA
- Paul Matsumoto, Galileo Academy of Science & Technology, San Francisco, CA
- Sally Mitchell, East Syracuse Minoa Central High School, East Syracuse, NY
- Joyce Nash Nadir, Roberts Wesleyan College, Rochester, NY
- Debbie Sayers, Hardin Valley Academy, Knoxville, TN
- Hannah Sevian, University of Massachusetts, Boston, MA
- Cheri Smith, Yale Secondary, Abbotsford, BC, Canada

AP LAB MANUAL VISION TEAM

In 2010, the College Board convened a group of subject matter and laboratory investigation experts to provide a model of excellence for what the investigative labs should be in AP science courses. These individuals worked diligently to create a vision for exemplary AP science labs that would serve to assist teachers in facilitating inquiry-based and student-directed investigative work. This vision also serves as the input for professional development and resource materials that will support the new course and lab investigations.

- Arthur Eisenkraft, University of Massachusetts, Boston, MA
- Eugenia Etkina, Rutgers University, Piscataway, NJ
- Nivedita Ganguly, University of Tennessee, Knoxville, TN
- Dolores Gende, Parish Episcopal School, Dallas, TX
- Peggy O'Neill Skinner, The Bush School, Seattle, WA
- Susan Singer, Carleton College, Northfield, MN
- Angelica Stacy, University of California, Berkeley, CA
- Peter Strom, Rutgers University, New Brunswick, NJ

OTHER CONTRIBUTORS

- AP Chemistry Consultants
- AP Chemistry Curriculum Development and Assessment Committee
- AP Chemistry Development Committee
- AP Chemistry Focus Group Participants
- AP Chemistry Lab Manual Reviewers
- AP Chemistry Lab Pilot Teachers
- AP Chemistry Readers
- College Board Science Academic Advisory Committee

CHAPTER 1

How to Use This Lab Manual

GOALS OF THE LABORATORY INVESTIGATIONS

Studying chemistry is not merely a matter of memorizing a collection of facts. Learning chemistry involves investigating matter in all kinds of ways in order to solve problems and answer questions of interest. So knowing a collection of facts about chemistry is useful only if you (the student) can use that information to understand and investigate a particular aspect of the world. AP® Chemistry laboratory experiments (also called “investigations” or “labs”) allow you to explore different types of chemical systems using different tools and techniques.

The labs in this manual help you learn and truly understand chemical concepts and the scientific evidence that supports them. Carrying out these labs will support and complement the work you do in other parts of your AP Chemistry course. The investigations allow you to develop and apply practices and skills used by scientists in real-world settings through a process known as inquiry. You make observations, ask questions, and then design plans for experiments, data collection, and analysis in order to answer your questions, communicate your findings to others, and demonstrate your chemical knowledge.

These labs all employ a guided-inquiry format, which is likely different from the labs you have done previously. Guided-inquiry labs do not prescribe exactly what you should do in order to accomplish the objective of the lab; rather they challenge you to generate hypotheses, pose questions, design experiments, interpret results, and decide how to communicate about your work with others. You may find this approach foreign or intimidating at first, but rest assured that your teacher will support you as you gain confidence and skills in the laboratory. Learning how to do inquiry is exciting and challenging, and evidence shows that this type of laboratory activity will lead you to learn and remember more chemistry than more traditional labs.

The following list details the sections you will find for each investigation and what those sections do.

- **Central Challenge:** describes the problem you will investigate during the lab
- **Context for the Investigation:** explains how the experiment relates to real-world applications and the lab setting
- **Prelab Guiding Questions and/or Simulations:** helps you prepare for the lab by structuring questions to help you come to conclusions about concepts or to strengthen your mental models of the particulate nature of matter
- **Explanation to Strengthen Student Understanding:** contains chemistry you need to learn and use in the lab. Some labs will not have this component if the guided-inquiry component is for you to discover the explanation or concept for yourself.

- **Preparation**

  Materials: details the equipment and chemicals you will use

  Safety and Disposal: outlines the safety concerns you need to be aware of and directions for how to safely dispose of waste
- **Practice with Instrumentation and Procedure:** includes (in most labs) activities to practice using lab equipment and techniques
- **Investigation**

  Procedure: provides instructions for how to design and carry out your investigation

  Data Collection and Computation: in most labs, this component will describe ways you could display and analyze your data

  Argumentation and Documentation: provides instructions for how to communicate your findings
- **Postlab Assessment:** gives you questions to answer based on your investigation to evaluate your understanding of the chemistry concepts you learned by designing and/or performing the investigation.
- **Supplemental Resources:** lists resources for more information on the topics in the lab

SAFETY

The chemistry laboratory is a unique environment. The equipment and chemicals used in the laboratory could cause harm if not used appropriately. Your teacher will provide safety information and training and describe his or her particular rules, and it is very important that you adhere to the rules at all times.

Some general safety guidelines are listed below.

1. Before every lab ensure you understand the potential hazards involved in the investigation and the precautions you will take and protective equipment you will need to wear in order to stay safe.
2. Know where the basic safety equipment is found in your classroom: a fire extinguisher, fire blanket, eyewash, chemical spill kit, safety goggles or glasses, and a first-aid kit.
3. Protective eyewear (safety goggles) must be worn at all times in the lab, including during set up and cleanup.
4. Follow your teacher’s safety rules at all times.
5. If you ever have any questions about safety, or are not sure if it is safe to do something, check with your teacher before proceeding.

EQUIPMENT

Chemistry laboratory equipment involves using many different types of glassware and other tools. You may find it useful to search for and review online resources to help you become familiar with the names and uses of various equipment and pieces of glassware you will be using throughout the investigations of this lab manual.

CHAPTER 2

The Science Practices: Developing Skills

Science is not a body of facts, but rather a system of understanding observations that allows us to make sense of the world around us. Scientists explore ideas and try to answer questions through experimentation, just as you will do in your AP Chemistry labs. The labs you do in this class will enable you to use the same skills that practicing scientists use in approaching questions and performing activities in a laboratory environment; these skills are known as science practices.

SCIENCE PRACTICES

Science practices (SPs) *are an assortment of laboratory and critical thinking skills used when conducting inquiry and guided inquiry laboratories.* The science practices are:

- important aspects of the work of scientists;
- a vitally important part of learning science;
- integral to laboratory experiences; and
- a foundation of scientific reasoning skills that you should master by completion of the class.

Below are descriptions of each practice and explanations of how each is used in the investigative labs in this manual.

Science Practices 1, 3 & 4:

Science Practice 1: Models and Representations
Describe models and representations, including across scales

Science Practice 3: Representing Data and Phenomena
Create representations or models of chemical phenomena

Science Practice 4: Model Analysis
Analyze and interpret models and representations on a single scale or across multiple scales

In each laboratory, you will create, use, and revise models to explain and discuss what happens in the lab. Through this practice you will develop your modeling skills.

Science Practice 2:

Science Practice 2: Question and Method
Determine scientific questions and methods

You will have many opportunities in the lab to ask questions and try to answer them through experimentation.

Science Practices 2 & 6:

Science Practice 2: Question and Method
Determine scientific questions and methods

Science Practice 6: Argumentation
Develop an explanation or scientific argument

You will have the opportunity to design your own procedure and decide what materials to use, what data to collect, and how to analyze the data to achieve your goals in many of the labs.

Science Practices 2, 5, & 6:

Science Practice 2: Question and Method
Determine scientific questions and methods

Science Practice 5: Mathematical Routines
Solve problems using mathematical relationships

Science Practice 6: Argumentation
Develop an explanation or scientific argument

You will identify trends or patterns in data, as well as possible lab errors, in order to answer questions.

Science Practices 3 & 4:

Science Practice 3: Representing Data and Phenomena
Create representations or models of chemical phenomena

Science Practice 4: Model Analysis
Analyze and interpret models and representations on a single scale or across multiple scales

As your chemistry models become more sophisticated you will apply them to different situations and contexts, such as biology and medicine, in different labs.

Together, these science practices will help you develop the range of skills used throughout the laboratories in this manual. The ultimate goal is for you to become a more sophisticated scientific thinker, prepared to continue studying and experimenting at the college level.

Science Practice 5:

Science Practice 5: Mathematical Routines
Solve problems using mathematical relationships

In each lab, you will collect data and perform calculations in order to answer questions. You will practice recording data accurately and precisely, deciding what calculations are appropriate, and carrying out calculations correctly. Part of using mathematics appropriately is also to be careful with significant figures and units as you record data and do calculations.

Science Practice 6:

Science Practice 6: Argumentation
Develop an explanation or scientific argument

Each lab requires you to apply your knowledge of scientific explanations and ideas to describe, analyze, and explain your data.

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CHAPTER 3

Quantitative Skills

Experimental chemistry relies heavily on quantitative (number-based) skills and analysis. Successful data collection and analysis in the AP Chemistry laboratory requires many skills, including accurate and precise measurement of various quantities using a wide range of instruments, unit conversions, estimation, algebraic and statistical calculations, and several types of graphing. Quantitative information obtained using these skills must then be integrated with reasoning and higher-order thinking skills in order for you to successfully analyze and interpret data and formulate and communicate conclusions. A discussion of some of the most important skills needed in AP Chemistry is presented here, followed by information about how this manual supports development and successful application of these skills.

QUANTITATIVE SKILLS IN THIS MANUAL

You have already developed some quantitative skills during earlier course work and experiences. The 16 experiments in this manual provide opportunities for you to practice and improve existing skills and to develop new skills. Each experiment demands a different combination of skills.

The labs are designed to promote new skill development and to challenge you to apply skills to solve problems using inquiry. As such, each lab contains a "Practice" component in order to introduce new techniques and equipment and to reinforce fundamental skills before you undertake the "Investigation" portion of the experiment in which you will apply your quantitative and other skills to address the central challenge.

KEY QUANTITATIVE SKILLS IN AP® CHEMISTRY LABS

The most important quantitative skills in the AP Chemistry laboratory can be roughly classified into four types: **measuring**, **calculating**, **creating tables and graphs**, and **quantitative analysis**. This classification scheme is useful for discussion purposes; in practice these skills overlap and in the laboratory setting they must be used in conjunction with each other to accomplish experimental goals. You already have some skills and abilities in each of these four areas, and you will develop others through your work in the AP Chemistry laboratory.

Measuring

Measuring skills consist of using a wide range of tools including rulers, balances, thermometers, timers, and volumetric glassware to make accurate and precise measurements with appropriate significant figures; calibrating and using electronic measuring devices such as pH meters and spectrophotometers; and making decisions about appropriate measuring tools for different tasks. In AP Chemistry you will learn how to use measuring tools and techniques that you haven't used before. Careful and repeated measurements made with proper technique are needed in order for you to obtain sufficient and accurate data, which is essential to be able to solve problems and answer questions. Remember, no amount of analysis or calculation can make up for poor data!

Calculating

Most labs require several types of calculations. These include unit conversions, solving for unknowns in algebraic calculations, and statistical calculations, such as percent error and standard deviation. Many calculations involve using logarithms and scientific notation. Most of these calculations mirror those required in other parts of the AP Chemistry course, so you will find it helpful to think about connections between the material you are studying and calculations you are doing as part of other tasks, like homework problems. For example, an experiment may require you to calculate the percent composition of a substance, a calculation you would also use to solve stoichiometry problems outside of the lab. Sometimes you may conduct a lab that addresses content that you have not yet studied in other parts of the course, in which case you will have a chance to practice new calculations in the lab. You may also have the opportunity to use spreadsheets developed in programs such as Microsoft Excel to perform calculations and share data. Facility with spreadsheets is an extremely valuable skill both within and beyond AP Chemistry and it is well worth the time for you to practice using them. References related to spreadsheets in chemistry are included at the end of this chapter.

Creating Tables and Graphs

Collecting, tracking, and organizing data in the laboratory is often facilitated by use of data tables. Learning to construct data tables involves skills such as identifying dependent and independent and control variables, choosing appropriate quantities and units for measurement, and use of computer programs for word processing or spreadsheet creation. Analysis and presentation of data in the chemistry laboratory often calls for the use of different types of graphs. Generating and using these graphs requires a number of skills including plotting coordinates, determining independent and dependent variables, determining slope, and choosing appropriate scales. Higher-level skills include linear regression or best-fit lines, and determining points of inflection or derivatives; these skills often call for use of a graphing calculator or computer spreadsheet. Using these tools to generate and interpret graphs is not essential, but it is highly desirable because they allow for relatively fast analysis, generally are more accurate and precise, and are more professional in appearance than hand-drawn graphs. Use of spreadsheets for graphing and analysis is also a transferable skill that is useful in many other courses you may take. As with measuring and calculation skills discussed earlier, you will likely have experience with some of these graphing skills coming in to the course, while others will be new to you.

Quantitative Analysis

Quantitative analysis is making meaning out of quantitative information generated through measuring, calculating, and graphing. It is often the most difficult part of the quantitative work in the laboratory, but it is essential in order to answer experimental questions successfully. The only way you can develop and strengthen these skills is through practice. Three quantitative analysis tasks that are common parts of laboratory work in chemistry are interpreting data, assessing accuracy and precision using statistical measures, and error analysis.

WEB RESOURCES

- "Introduction to Statistics in Chemistry": *http://chemlab.truman.edu/data-analysis/introduction-to-statistics-in-chemistry/*
- Tutorial on the use of Excel in chemistry: *http://www.wellesley.edu/Chemistry/stats/form1.html*

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CHAPTER 4

Written, Verbal, and Visual Communication

Making an observation or measuring a quantity is never the last step you make in science. You must be ready to communicate your experimental results to others in the chemistry community, including your peers. Every laboratory experience should include small group or classwide discussions of the questions posed by the lab and possible ways to address those questions. Your discussion, guided by your teacher or conducted independently with your classmates, should help you to understand the relationships among the hypothesis, the ensuing protocol, and the results gained by following that protocol. In chemistry, your awareness of potential error and its effect on the quality of the answers gained in the process is vitally important, and effective communication is crucial in uncovering these potential errors. Furthermore, you must be able to critique in an informed way the quality of your designed or implemented protocol and identify how it might be changed for the better if you were to repeat the experiment.

COMMUNICATING YOUR RESULTS

There are many ways to communicate your laboratory results and your teacher will provide specific guidance about how to do so in your AP Chemistry class. Below is a list of some ways that scientists and students share their work with others.

- *Lab reports* The traditional laboratory report is one method of communication that you may already be familiar with. Lab reports can take a number of formats and styles, but typically include sections about the purpose, hypothesis, procedure and materials, data and observations, calculations, graphs, results analysis, and conclusion, which are tailored to the details of the particular experiment. Typically lab reports are written in third person. Preparing an individual lab report is one way to practice written communication — and becoming adept at writing up your experiments should be a major goal for you in AP Chemistry.
- *Lab notebooks* A lab notebook is a real-time record of your lab work, in which you write questions, hypotheses, experimental plans, data, analysis, and written reflections on your lab work. Many teachers use a series of questions students answer as they do the lab to guide students' work in their notebooks. A common series of questions is: What is the beginning question? How will I conduct the investigation? How will I stay safe? What observations and measurements did I make? What can I claim? What evidence do I have to support the claim? How do my ideas and results compare with others and to the literature? How have my ideas changed?

- *Group lab reports* Many articles written in scientific research journals, where scientists report their work, have several authors listed for a paper. In this AP course you might be encouraged to prepare your "publications," the lab reports, in pairs, trios, or even larger groups. Your teacher can guide you in presenting a lab report as a group effort, perhaps pointing to Google Docs as an effective way to share data and facilitate peer review.
- *Poster presentations* It is common practice for college students, both undergraduate and graduate, to give poster presentations of their work, just as scientists do at research meetings. Posters enable you to describe experiments using both written and visual components, such as by incorporating graphs of experimental data. During a presentation, you would explain the information on the posters to others, guiding them through the work and the results.
- *Multimedia projects* Another way to communicate is through multimedia such as PowerPoint presentations or videos. Some teachers may ask you to use these tools to communicate about your lab work. Like posters, multimedia projects lend themselves to visual representations of your work. Usually these projects involve an oral presentation component in which you explain the work to your peers and teacher.
- *Journal articles* Scientists generally present their work to others in scientific journal articles, which differ somewhat in format and writing conventions from a typical student lab report. Your teacher might have you prepare your lab reports in article style. Some students have actually published articles in scientific journals about their science fairs or independent research projects.
- *Peer review* Working scientists review and critique each other's work before it can be published in scientific journals in a process called peer review. As a student, evaluation of your individual work by the instructor is essential, but peer review of others' work has become an increasingly common aspect of education. Written analyses of your own work and the work of your fellow students can lead you to understand the "big picture" better and prepare you for further work in science in college and beyond.
- *Portfolios* Individual and joint lab reports and projects of various types might then be grouped by you into a portfolio, a collection of all of your lab work in AP Chemistry throughout the year. Some colleges and universities require the presentation of a portfolio or complete lab notebook to document your college-level lab experiences. This technique of developing chemistry portfolios might be required by your teacher, but, even if it is not, you should save all of your lab work for college.

QUALITY OF WRITING

While high school students write in most of their courses, in chemistry the value of clear and concise writing is particularly important. An incomplete or out-of-order description of a lab procedure might lead to an ineffective or even dangerous outcome for the lab. Dr. Lisa Zuraw of The Citadel, former chair of the AP Chemistry Test Development Committee, wrote to students and teachers about "Improving Performance on the AP Chemistry Examination." In it, she points to specific ways to improve your writing:

- Be specific in your answers
- Answer the question asked
- Understand the difference between a justification and a definition
- Realize trends are not justifications
- Use and understand chemical vocabulary
- Write in sentences
- Know the proper use of antecedents
- Review your writing

Dr. Zuraw's suggestions also apply to effective writing about laboratory topics, so in an important way, your efforts to improve your communication skills in the laboratory are directly helping you prepare to do well on the AP Chemistry Exam.

Some Writing Resources

There are a number of sources that address more generally the notion of quality writing in chemistry. Oregon State's "Writing Guide for Chemistry" (*http://www.chem.orst.edu/writing/WritingGuide2000.htm*) contains many tips for better writing in chemistry. Your teacher can direct you to a variety of other colleges and universities that also offer tips about good writing in science and chemistry.

Graphs

- Because a graph is a convenient way to present data visually, you should follow certain conventions. Each graph should fill the space available and should contain at a minimum: Title, in the form of dependent variable versus independent variable
- Properly chosen axis: convention is x-axis independent variable, y-axis dependent variable
- Axes that must be labeled, including units
- A line or curve rather than connected data points

WEB RESOURCES

- Writing to improve performance in AP Chemistry:
  http://apcentral.collegeboard.com/apc/members/courses/teachers_corner/45234.html?type=print
- A writing guide for chemistry:
  http://www.chem.orst.edu/writing/WritingGuide2000.htm
- A brief guide to writing in chemistry:
  http://chemistry.kenyon.edu/getzler/08F-CourseFiles/BriefGuideWritingChemistry.pdf

- Writing a traditional formal lab report:
  http://www.clarion.edu/academics/student-success-center/writing-center/formal-chemistry-lab-reports.pdf
- LabWrite, an online system for guiding students' lab reports:
  http://labwrite.ncsu.edu/index_forprof.html
- Lab report in journal format:
  http://chemlab.truman.edu/the-laboratory-report/
- Google for Educators, including Google Docs:
  https://edu.google.com/?modal_active=none
- Poster presentations:
  http://www.ncsu.edu/project/posters
- Alternative methods of presenting and assessing lab work:
  http://www.nj ea.org/news-and-publications/njea-review/december-2008/alternative-science-lab-reports
- Graphing data from the chemistry lab:
  https://projects.ncsu.edu/labwrite/res/gt/gt-reg-home.html

CHAPTER 5

Connecting Lab Experiences to the Real World

Most professional chemists think of chemistry as an amazing way to understand the world around them. It is not uncommon to hear them say "chemistry is everywhere," because all matter is made of chemicals. One goal of any chemistry course is to help you begin to see how knowledge of chemicals and chemical reactions can help you better understand the world around you. If your interest is in cars, you may be interested in alloys, new fuels, and engine efficiency. If you are fascinated by computers, you may become interested in materials used for video screens or computer memories. Or, if you love art, the components of different pigments in paintings, clay, or pottery glazes may pique your interest. You might enjoy TV crime programs depicting the use of forensic science, in which chemistry plays a big part. Ultimately, chemistry plays a part in all the sciences, so your chemistry class will be able to include applications covering a variety of areas that interest you.

REAL-WORLD CONNECTIONS TO INVESTIGATIVE LABS

You are probably curious about how what you are learning in school, including AP Chemistry, is connected to the world outside of school. Each of the labs in this manual is related to many real-world ideas, activities, and materials. The table below contains two such ideas for each lab, but that is just a start. If you are interested, research one or more of these topics. Remember, chemistry is, in fact, everywhere.

Lab	Connections to Real World
1. Spectroscopy: *What Is the Relationship Between the Concentration of a Solution and the Amount of Transmitted Light Through the Solution?*	■ Measuring emissions from car exhaust ■ Determining if artwork is fake or real
2. Spectrophotometry: *How Can Color Be Used to Determine the Mass Percent of Copper in Brass?*	■ Measuring blood alcohol levels in suspected drunk drivers ■ Determining concentration of lead in drinking water
3. Gravimetric Analysis: *What Makes Hard Water Hard?*	■ Elemental analysis of ores (rocks) ■ Water softening methods
4. Titration: *How Much Acid Is in Fruit Juices and Soft Drinks?*	■ Uses of citric acid ■ Soft drinks as rust removers

5. Chromatography: *Sticky Question: How Do You Separate Molecules That Are Attracted to One Another?*	■ Testing urine for illegal drugs ■ Forensic analysis
6. Bonding in Solids: *What's in That Bottle?*	■ Accidents caused by mislabeled chemicals at factories ■ Mislabeled prescription drug accidents
7. Stoichiometry: *Using the Principle That Each Substance Has Unique Properties to Purify a Mixture: An Experiment in Applying Green Chemistry to Purification*	■ Presidential Green Chemistry Challenge ■ The National High School Journal of Science (peer-reviewed journal with many articles by high school students that connects stoichiometry to real-world examples)
8. Redox Titration: *How Can We Determine the Actual Percentage of* H_2O_2 *in a Drugstore Bottle of Hydrogen Peroxide?*	■ Measuring iron content of foods or drugs ■ Wastewater treatment with hydrogen peroxide
9. Physical and Chemical Changes: *Can the Individual Components of Quick Ache Relief Be Used to Resolve Consumer Complaints?*	■ Health risks of expired drugs ■ Painkiller synthesis — how aspirin, Tylenol, Advil are manufactured
10. Kinetics: Rate of Reaction: *How Long Will That Marble Statue Last?*	■ Rate of bacterial growth on food (food spoilage) ■ Effects of different pollutants on air quality and smog
11. Kinetics: Rate Laws: *What Is the Rate Law of the Fading of Crystal Violet Using Beer's Law?*	■ Designing biodegradable containers and plastics ■ Predicting radioactivity levels and health risks of nuclear waste
12. Calorimetry: *The Hand Warmer Design Challenge: Where Does the Heat Come From?*	■ The energy content (calories) in food ■ Comparing different types of biodiesel and alternative fuels
13. Equilibrium: *Can We Make the Colors of the Rainbow? An Application of LeChâtelier's Principle*	■ The Haber process — applications in industry and manufacturing ■ Chlorination of wastewater
14. Acid-Base Titration: *How Do the Structure and the Initial Concentration of an Acid and a Base Influence the pH of the Resultant Solution During a Titration?*	■ Swimming pool and aquarium maintenance ■ Acidosis/alkalosis in humans — causes and symptoms
15. Buffering Activity: *To What Extent Do Common Household Products Have Buffering Activity?*	■ Soil management for farming and gardening ■ Monitoring pollution and environmental effects in seawater
16. Buffer Design: *The Preparation and Testing of an Effective Buffer: How Do Components Influence a Buffer's pH and Capacity?*	■ Enzymes and buffers in the stomach and small intestine ■ Buffered aspirin vs. unbuffered (regular) aspirin

SUPPLEMENTAL RESOURCES

Links

The following is a list of just a few of the resources that help give real-world contexts to chemistry.

- Carnegie Mellon and the Chem Collective: *http://www.chemcollective.org/find.php*
- ChemMatters (an ACS publication for high school students): *https://www.acs.org/content/acs/en/education/resources/highschool/chemmatters.html*
- Environmental Protection Agency (EPA): *http://www.epa.gov*
- Color in Art: *http://www.webexhibits.org/pigments/intro/look.html*
- American Chemistry: *http://www.americanchemistry.com/news*
- Science Daily: *http://www.sciencedaily.com/news/matter_energy/chemistry/*
- Science Magazine: *http://www.sciencemag.org/*
- Chemistry News at PhysOrg: *http://phys.org/chemistry-news/*
- *The New York Times* science pages: *http://www.nytimes.com/pages/science/*
- BBC News on science and environment: *http://www.bbc.co.uk/news/science_and_environment/*

INVESTIGATION 1

What Is the Relationship Between the Concentration of a Solution and the Amount of Transmitted Light Through the Solution?

CENTRAL CHALLENGE

You will be guided through an investigation to formulate your own answer to the question, “How can light be used to determine concentrations of chemical species in solutions?”

CONTEXT FOR THIS INVESTIGATION

Measuring how much of which wavelengths of light are absorbed by a substance, and getting useful information about that substance from the results, is the scientific discipline of spectroscopy. The visible spectrum is the only part of the electromagnetic spectrum that we can access with the equipment found in a typical school chemistry laboratory. The basic principles of spectral analysis that you learn in high school can also be applied to the more sophisticated instrumentation required to access the ultraviolet, infrared, and x-ray regions. What you learn by performing this lab will help you to understand more sophisticated instruments that you may encounter in a college or university.

In a visible spectrophotometer, we shine a beam of light into a solution containing the sample, and detect how much of it comes out of the other side of the solution. By comparing the amount of light transmitted by the pure solvent to the amount transmitted when the sample is dissolved in it, we can calculate a quantity called the *absorbance*. Spectrophotometers can report measurements as percent transmittance (%T) or directly as absorbance. In this investigation, you will be guided to discover the relationship between transmittance and concentration and, ultimately, the relationship between transmittance, absorbance, and concentration of solution.

PRELAB GUIDING QUESTIONS/SIMULATIONS

Your teacher will guide you through a series of questions to identify the relationship (if any) between the transmittance of light through a solution and the concentration of the solution.

Step 1: Your group will be given a known concentration or stock solution of the blue #1 dye (Brilliant Blue) molecule dissolved in water (see Figure 1). The blue #1 dye molecule is the only chemical species in solution. You will be measuring the percent transmittance of light through a solution containing only this chemical species.

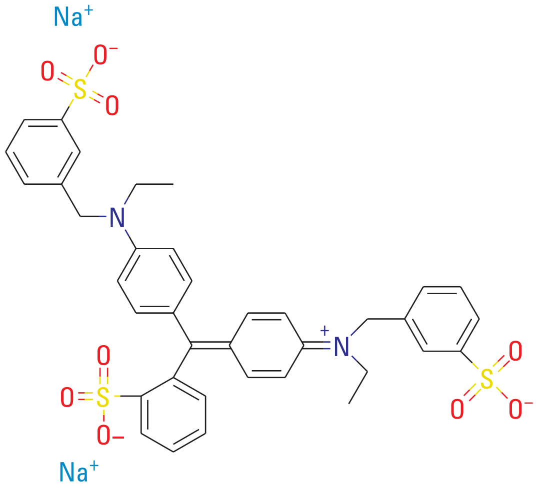


Figure 1. FD&C blue #1 (Brilliant Blue) molecule

Step 2: Your teacher will tell you the appropriate wavelength at which to take transmittance measurements using the SPEC 20 (or other spectrophotometer/colorimeter).

Step 3: Your teacher will assign your group to gather data as to the percent transmittance of the molecule in solution at various concentrations. Your group will need to make a specific dilution or set of dilutions according to this assignment. Circle and then make the dilution(s) to which you are assigned below:

Collect percent transmittance data for the following dilutions of stock/water:

10 mL stock/0 mL water

8 mL/2 mL

6 mL/4 mL

4 mL/6 mL

3 mL/7 mL

2 mL/8 mL

1 mL/9 mL

0 mL/10 mL

Step 4: Collect the percent transmittance values of your group's assigned diluted solutions using the SPEC 20 (or other spectrophotometer). Calculate the molar concentration of each of your assigned diluted solutions and record both pieces of data in a data table in a central location in the lab. Use the original stock solution molarity and $M_1V_1 = M_2V_2$ to determine the concentrations of each of your dilutions.

Note: You will need to convert %T to a decimal and record that information in the data table as well.

Table 1

Solution	Dilution Ratio mL stock/mL water	Molar Concentration (μM)	Measured percent transmittance	Measured %T in Decimal Form
1. (stock solution)	10 mL/0 mL			
2.	8 mL/2 mL			
3.	6 mL/4 mL			
4.	4 mL/6 mL			
5.	3 mL/7 mL			
6.	2 mL/8 mL			
7.	1 mL/9 mL			
8.	0 mL/10 mL			

Step 5: Determine the relationship between transmittance and molarity of the solution by graphing the data in the data table. Plot transmittance (as a decimal) on the y-axis and concentration in μM on the x-axis. Discuss your proposed relationship between transmittance and molarity of the solution with other groups and with your teacher.

Step 6: Work with your group to make a series of plots to determine if there is a linear relationship (which has a positive slope and goes through zero) between transmittance and the molarity of the solution. Your teacher may assign you one type of plot to make from the following list:

- 1/T versus [dye]
- 1×10^T versus [dye]
- logT versus [dye]
- –logT versus [dye]

If there is a linear relationship between two values, that goes through zero and has a positive slope, then an equation can be constructed and the scientist can predict values of one variable given the other. Discuss with the other groups and with your teacher which treatment of the data (using the plots above) showed such a linear relationship between transmittance and concentration. Your teacher will further explain to you the relationship between transmittance, absorbance, and concentration.

Record that relationship in your lab notebook.

PREPARATION

Materials

Spectrophotometer SPEC 20, SPEC 200, or colorimeter	Cuvettes	Disposable pipettes
Stock solution of known concentration of blue dye #1 liquid food dye.	Blue-colored sports drink containing blue #1 dye	Distilled water
5–10 Test tubes (20 mL)	Graduated cylinders (10 mL)	Test tube rack

Safety and Disposal

The food dye and sports drink can be flushed down the sink with plenty of water.

INVESTIGATION

Many common sports drinks contain blue #1 dye. Use the relationship between transmittance, absorbance, and concentration (as well as your plot or calibration line and slope from your prelab) to determine the concentration of this dye in the sports drink.

Procedure

Obtain a sample of the blue-colored drink. Design a data-collection and data-analysis procedure to determine the molarity or concentration of blue #1 dye in the sports drink.

Data Collection and Computation

1. Determine the molar concentration of blue #1 dye in the sports drink. Show all work.
2. Determine the mass of blue #1 dye found in 500 mL of the drink. Show all work.

Argumentation and Documentation

In the conclusion of your lab, justify the procedure you chose, the instrumentation you used, and the selection of the kind of data needed to determine the concentration of blue #1 dye in a blue-colored sports drink containing only this dye.

POSTLAB ASSESSMENT

1. Suppose a solution was too concentrated for an accurate reading with the spectrophotometer. The concentrated solution was diluted by placing 1.00 mL of the concentrated solution in 4.00 mL of water. The solution was then placed in the spectrophotometer, an absorbance was obtained, and after a few calculations the molar

concentration was calculated to be 3.5 × 10^{-6} M. What was the concentration of the original stock solution before dilution?

2. If a 0.10 *M* solution of a colored substance has a maximum absorbance at 500 nm and an absorbance of 0.26 at this wavelength, what will be the measured absorbance of a 0.20 *M* solution at 500 nm?
3. The spectrophotometer really measures the percent of light that is transmitted through the solution. The instrument then converts %T (transmittance) into absorbance by using the equation you determined in the prelab section. If the absorbance of a sample is 0.85, what is the percent light transmitted through the colored sample at this collected wavelength?

SUPPLEMENTAL RESOURCE

Link

"USGS Spectroscopy Lab." United States Geological Survey. Accessed July 31, 2012. *http://speclab.cr.usgs.gov/*

INVESTIGATION 2

How Can Color Be Used to Determine the Mass Percent of Copper in Brass?

CENTRAL CHALLENGE

What are the relationships between color, wavelength, absorbance, and concentration? You will need to understand these relationships in order to design an experiment that can quantitatively measure the absorption of light by a colored solution in order to determine the concentration of the absorbing species in that solution.

CONTEXT FOR THIS INVESTIGATION

Spectrophotometry is an extremely important tool used in forensic science to determine the detailed chemical composition of evidence obtained from a crime scene. It can be used to determine the concentration of either a single chemical species in solution or even the concentration of a species within a mixture of species in solution. For example, it can be used to determine the mass percent of copper in brass shell casings collected by the crime scene investigator (CSI), and then match the brass composition to a particular manufacturer.

PRELAB GUIDING QUESTIONS/SIMULATIONS

Step 1: You will be collecting data (in your group) to determine the relationship between wavelength (lambda, λ), concentration, and absorbance. You will use the interaction of light with solutions to determine the concentration of an absorbing species. However, it is important to identify what is doing the absorbing prior to determining the concentration of the absorbing species. Such information can be gathered by collecting absorbance data for a solution at various wavelengths.

Step 2: Your group will be given two different salt solutions of the exact same concentration. You are to measure the absorbance for each solution at every 20 nanometers from 400–700 nm to generate a spectrum and to determine the best wavelength at which to measure the absorbance of the two solutions. Make sure you use the appropriate spectrophotometer or colorimeter to collect such data. You will be sharing your absorption spectra with the rest of the class to identify what species in solution is actually absorbing light.

Step 3: Circle the salt solutions below to which your group is assigned:

salt 1 — copper (II) sulfate
salt 2 — copper (II) nitrate
salt 3 — zinc nitrate
salt 4 — zinc sulfate
salt 5 — iron (III) nitrate
salt 6 — iron (III) sulfate

Step 4: Look at the different plots or spectra for the six different salts. Can you identify the chemical species doing the absorbing (the metal ion, the polyatomic anion, or both)? How do you identify it?

Step 5: Engage in whole-class discussion to answer the following questions:

Extension Question 1. Why do we want to use a particular wavelength when determining the absorption of a particular chemical species? Is it important to measure absorbance for all wavelengths from 400–700 nm?

Extension Question 2. If you know the molar absorptivity of copper sulfate at 630 nm, explain how you could determine a wavelength where the molar absorptivity is half of that simply by examining the absorption spectrum.

Step 6: Make sure to write down your understanding of the relationship between concentration, absorbance, and wavelength to help inform your decisions throughout the remaining components of the lab.

PREPARATION

Materials

Spectrophotometer (or colorimeter)	Concentrated 15.8 *M* Nitric acid (HNO_3)
	10 mL of 0.400 *M* Copper(II) nitrate trihydrate ($Cu(NO_3)_2 \cdot 3H_2O$)
Various copper, zinc, and iron salt solutions (0.1 M)	1–2 g brass sample
3 thin-stem Beral pipettes	Digital balance, ±0.001 g
50 mL beaker with watch glass	Distilled water
5 mL and 10 mL graduated pipettes (or cylinders)	Pipette pump or pipette bulb
516 × 150 mm test tubes	1 cuvette
Test-tube rack	Tissues (preferably lint free)

Safety and Disposal

Concentrated nitric acid is corrosive and will attack and destroy metals, proteins, and most plastics. Avoid skin contact and neutralize any spills with baking soda, then rinse with copious amounts of water. The acid will discolor the skin for days after contact, so be sure to wear rubber gloves. The NO gas that forms quickly oxidizes in air to produce a toxic, reddish-brown gas of NO_2. For more information, read the Material Safety Data Sheet (MSDS) for nitric acid found at *http://www.ehso.com/msds.php*

Perform this reaction under a fume hood. Take normal laboratory precautions, including wearing splash-proof goggles and chemical-resistant gloves and apron at all times. The remaining brass solution should be neutralized by adding small amounts of solid baking soda until the bubbling has subsided and the pH is 5–9. Then the waste solution can be safely disposed of by following standard procedures as directed by your instructor.

INVESTIGATION

Procedure

1. Determine the mass of a 1–2 gram sample of brass to ± 0.001 g. Place the sample in a small beaker.
2. Assuming your brass sample is 100 percent copper by mass, calculate the minimum volume of concentrated 15.8 M $HNO_3(aq)$ that needs to be added to react completely with the brass. Under the fume hood, have your teacher add approximately 2 mL more than this volume of 15.8 M $HNO_3(aq)$ so that the acid is in excess, and then your teacher will cover the beaker with a watch glass.
3. After the metal dissolves completely, add 50 mL of distilled water to the beaker **(again, your teacher will perform this part of the investigation)**. Then you will remove the beaker from the fume hood and transfer the solution to a 100 mL volumetric flask. Rinse the beaker 3–4 times with 5 mL of distilled water and add the washings to the flask. Dilute to a final volume of 100.0 mL. The excess nitric acid will dissolve the zinc and copper metals in the brass.
4. Obtain 10.0 mL of 0.400 M $Cu(NO_3)_2(aq)$ stock solution in a 10 mL graduated cylinder. Determine what volume is required to make 10.00 mL of 0.200 M $Cu(NO_3)_2(aq)$. Use a volumetric pipette to transfer this volume of the stock solution into a clean test tube. Then add a sufficient amount of distilled water to reach 10.00 mL. Thoroughly mix the solution. Repeat the dilution process to make 10.0 mL each of three more additional dilute solutions that are 0.100 M, 0.0500 M, and 0.0250 M, respectively. Have your instructor verify your dilution calculations!
5. Based upon the results of the prelab experimentation, set the wavelength of the SPEC 20 to that which is strongly absorbed by the blue-colored Cu^{2+} solutions. Ideally, the maximum absorbance value should be ≤ 1. Since absorbance is a logarithmic function of the percent transmittance, the instrument is in a nonlinear region to measure the light passing through when the absorbance value is at a range of 1–2. When selecting a wavelength for measurement, keep in mind that a wavelength at maximum absorbance provides maximum sensitivity but the smallest concentration range, while a wavelength with a smaller absorbance would provide less sensitivity but a larger concentration

range to be measured in the experiment. Have your teacher approve of your selected wavelength before you continue with the next step.

6. Calibrate the SPEC 20 to read 0% transmittance with no cuvette in the instrument, and 100% transmittance with a blank inside (cuvette filled with distilled water).
7. Empty the water from the "blank" cuvette. Using the most dilute $Cu(NO_3)_2$ standard solution, rinse the cuvette twice with ~1 mL amounts and then fill it three-quarters full. Wipe the outside with a tissue, place it in the SPEC 20, and close the lid. Read and record the absorbance value. Discard the cuvette contents back into the original test tube.
8. Continue testing the other solutions, starting with the most dilute $Cu(NO_3)_2$ to the most concentrated (0.400 *M*). Finally, determine the absorbance value of the unknown $Cu(NO_3)_2$ solution from your brass sample. Using the absorbance and concentration values for the five standard solutions, prepare a graph of the absorbance (y) versus the concentration (x) values. Draw a best-fit straight line for your data and calculate the slope and y-intercept for Beer's plot. Then determine the concentration of your unknown brass solution. This is a good opportunity to use a spreadsheet program to plot the data and perform a linear regression.
9. Use a visual comparison test to also determine the concentration of the unknown solution. The amount of light absorbed by the solution is directly related to its concentration and the depth of the solution through which the light passes. By comparing the depth of two solutions with the same color intensity, the molarity of an unknown solution can be determined based upon the known concentration of the other solution.

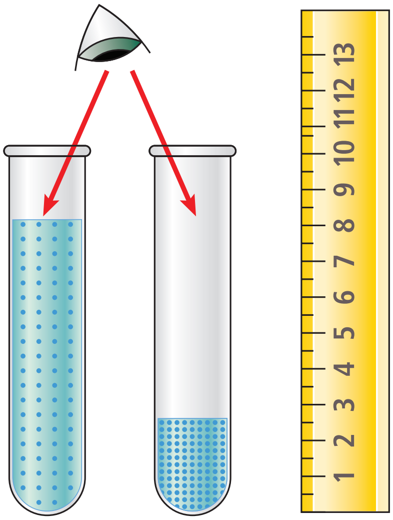


Figure 1. Visual comparison method

The intensity of the color of a solution in a test tube depends upon both the concentration of the solution and thd depth of solution in the tube. A higher concentration of the colored solution has a darker, more intense color than a solution of lower concentration. The greater the depth of a solution in a tube, the more particles there are to absorb the light beween the light source and they eye. This makes the color appear to be more intense.

To determine the concentration (molarity) of an unknown solution, place it in a test tube on top of a piece of white paper, next to a test tube filled with a solution of known concentration (molarity). While visually comparing the color intensity of each tube, use a pipette to remove liquid from the more concentrated solution until the intensities appear to be the same. At this point, the following relationship applies:

$$(\text{Molarity}_1)\ (\text{Depth}_1) = (\text{Molarity}_2)(\text{Depth}_2)$$

10. Fill a 16 × 150 mm test tube with the unknown brass solution. Use your standard 0.400 *M* $Cu(NO_3)_2$ to fill a second 16 × 150 mm test tube to a depth where the color intensity appears to match that of the unknown. Calculate the molarity of the unknown based upon this visual comparison method and compare to the molarity determined by the SPEC 20.
11. Dispose of all solutions as directed by your teacher. Rinse the cuvettes with distilled water. Return all equipment to your teacher for final approval.

Data Collection and Computation

1. Prepare a data table to record all of your measured data and calculated values.
2. Show the calculations used to prepare the $Cu(NO_3)_2(aq)$ with known molarities by diluting the 0.400 *M* $Cu(NO_3)_2(aq)$ standard solution.
3. Determine the molarity of the $Cu(NO_3)_2(aq)$ found in 100.0 mL of the brass solution using both the SPEC 20 (or colorimeter) analysis AND the visual comparison method. Support your answers with the appropriate calculations.
4. Determine the mass of Cu dissolved in the brass solution based upon the molarities calculated in Step 3, and use these values to calculate the mass percent of Cu in the brass sample using both experimental techniques.
5. Compare the mass percent of Cu values that were determined by the two different experimental methods for finding the molarity of the brass solution. Which method do you think will provide a more accurate value? Justify your answer.

Argumentation and Documentation

Use a shared Google Docs Excel spreadsheet to analyze the precision of the class data. Enter your group's values for the calculated mass percent of copper in brass based upon the spectrophotometer (colorimeter) in one column and the visual comparison method in another column. Calculate the averages and standard deviations for both methods. Which method provided a more precise value? Justify your answer.

POSTLAB ASSESSMENT

This lab involves several key skills and concepts that are part of the AP Chemistry curriculum, making it ideally suited for a formal lab report as a postlab assessment. Special emphasis should be placed on the two methods of analysis used to determine the molarity of the brass solution, and your comparison of the precision and probable accuracy of the class average values for the mass percent of copper in brass. However, your teacher may assign a different postlab assessment such as the following lab practical.

The following lab practical can be used to assess your ability to interpret a Beer's law plot and apply it to a new situation.

Problem: Based upon the Beer's law plot shown in Figure 2 for the absorbance of $CuSO_4(aq)$ versus its concentration, what mass of copper could be produced by a complete reaction of excess zinc metal with 50.0 mL of a $CuSO_4(aq)$ that has a measured absorbance of 0.685?

x Concentration (mol/L)	y Absorbance
0.08	0.186
0.16	0.372
0.24	0.587
0.32	0.753
0.40	0.955

Linear Fit for Latest/Absorbance
y = mx + b
m (Slope): 2.399
b (Y-Intercept): –0.005494
Correlation: 0.9995
RMSE: 0.01151

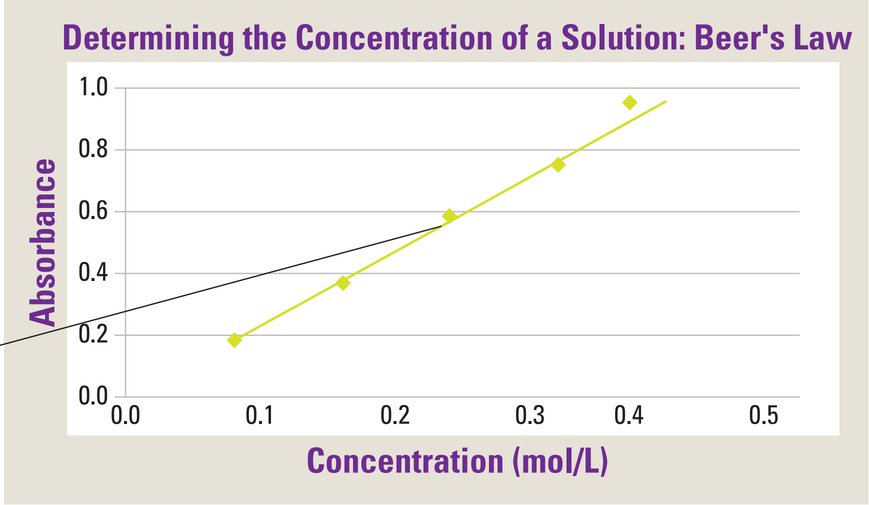


Figure 2. $CuSO_4$ **Beer's law sample data and graph**

SUPPLEMENTAL RESOURCES

Links

"Color Vision Interactive Simulation." University of Colorado at Boulder, PhET Interactive Simulations. Accessed July 31, 2012. *http://phet.colorado.edu/en/simulation/color-vision*

"Light Waves and Color — Lesson 2." The Physics Classroom. Accessed July 31, 2012. *http://www.physicsclassroom.com/class/light/u12l2c.cfm*

"The Franklin Institute Resources for Learning — Light and Color." The Franklin Institute. Accessed July 31, 2012. *https://www.fi.edu/color-science*

Reference

Brown, Theodore L. "Chemistry of Coordination Compounds." In *Chemistry: The Central Science*, Theodore L. Brown, H. Eugene LeMay Jr., Bruce E. Bursten, and Catherine J. Murphy, 1031–1033. Upper Saddle River, NJ: Pearson Education, 2009.

INVESTIGATION 3

What Makes Hard Water Hard?

CENTRAL CHALLENGE

Six samples of water will be analyzed for their quantities of water hardness through principles of metal ion precipitation and separation. The samples will then be ranked in order of increasing water hardness.

CONTEXT FOR THIS INVESTIGATION

A client seeks to move into your community. While there are multiple factors that will help him decide where to purchase a house, he is hoping to find a location where a water softener will *not* be needed. Several samples of water were collected from around your community. Help this client identify which of the areas around your community have the lowest hard-water quantities by ranking the hardness of the water in the areas.

PRELAB GUIDING QUESTIONS/SIMULATIONS

1. Questions a–g relate to the interactive simulation, a PhET simulation designed by the University of Colorado. Go to *http://phet.colorado.edu/en/simulation/soluble-salts* to open the interactive simulation. Click *Run Now*. After obtaining access to the simulation, note the three tabs at the top: *Table Salt*, *Slightly Soluble Salts*, and *Design a Salt*.
   - **a.** Under the *Table Salt* tab, shake the salt shaker. Describe what happens to the solid table salt, NaCl.
   - **b.** Click *Reset All*. Shake the salt shaker until some of the particles are designated as *Bound*. How many sodium ions are designated as *Dissolved*? How many sodium ions are designated as *Bound*? Use the simulation to describe what *bound* means.
   - **c.** Click the *Slightly Soluble Salts* tab. Using the pull down menu, select *Mercury(II) Bromide*. Slowly shake the salt shaker until some of the ions are designated as *Bound.* How many shakes did it take? Compare how this mercury(II) bromide is different from table salt.
   - **d.** Shake a large amount of mercury(II) bromide into the container. How do the number of dissolved ions change as more mercury(II) bromide is added to the container?
   - **e.** Slowly drain some of the mixture out of the container and stop. Where do the dissolved ions go as the solution is drained? What else do you notice as the mixture is drained? If the mixture left the container through a long pipe as it was drained, how might problems arise inside the pipe?

   f. How might the simulation look different if the mercury(II) bromide was created from two salts, such as mercury(II) nitrate and sodium bromide, rather than added directly?
   g. Predict an appropriate experimental method to collect the bound mercury(II) bromide.
2. Watch the following video: *http://www.youtube.com/watch?v=YcZSNcaHHN8&feature=youtube_gdata_player*

After watching the video, describe what you believe to be "hard water."

3. Describe some ways that water can be softened. Helpful sites include, but are not limited to: *http://www.chem1.com/CQ/hardwater.html*
   https://www.ag.ndsu.edu/publications/home-farm/water-softening-ionexchange
4. A reaction occurs between solutions of strontium bromide and silver nitrate, as shown in the equation below:

$$SrBr_2\ (aq) + AgNO_3\ (aq) \rightarrow Sr(NO_3)_2\ (aq) + AgBr\ (s)$$

   a. If 3.491 grams of the precipitate is formed, how many moles of strontium bromide were reacted?
   b. If 45.61 mL of strontium bromide were reacted in part (a), what is the molarity of the strontium bromide solution that was used?
   c. In collecting the precipitate, why would it be inappropriate to heat the reacted mixture and evaporate off the water?
5. Below is a table of K_{sp} (solubility product constant) values. Consider how the values in this table may help in deciding how to remove one of these ions by selective precipitation. You have been assigned one of the cations — **Mg** $^{2+}$, **Ca** $^{2+}$, **Fe** $^{2+}$. How would you remove the ion from hard water? Include which anion you would use to remove the cation and explain why you chose that anion.

Table 1. K_{sp} Values for Salts

Mg^{2+} Ion	K_{sp} Value	Ca^{2+} Ion	K_{sp} Value	Fe^{2+} Ion	K_{sp} Value
$MgCO_3$	3.5×10^{-8}	$CaCO_3$	2.8×10^{-9}	$FeCO_3$	3.2×10^{-11}
MgC_2O_4	1.0×10^{-8}	CaC_2O_4	4.0×10^{-9}	FeC_2O_4	3.2×10^{-7}
MgF_2	6.5×10^{-9}	CaF_2	4.0×10^{-11}	FeF_2	2.4×10^{-6}
$Mg(OH)_2$	1.8×10^{-11}	$Ca(OH)_2$	5.0×10^{-6}	$Fe(OH)_2$	8.0×10^{-16}

Note: Solubility product constants are extremely difficult to obtain experimentally because of the necessity to identify all chemical species and processes present in the chemical system used to obtain their values. Literature K_{sp} values may disagree widely, even by several orders of magnitude.

6. Go to *http://www.ehso.com/msds.php*

Identify the potential health risks and the appropriate measures for first aid for the following chemicals:

- a. calcium chloride
- b. sodium chloride
- c. sodium carbonate, anhydrous

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

Have you ever had difficulty lathering soap or find that the scum in your shower constantly needs to be removed? These are signs of "hard water." Soap doesn't lather well in hard water because metal ions, such as Ca^{2+}, form precipitates, creating "soap scum." A *precipitate* is an insoluble compound that forms when soluble ions in separate solutions are mixed together. Because this happens, soap is a less effective cleanser in hard water. Even laundry can appear dingy or feel rough when washed in hard water.

While these metal ions are generally harmless, hard water has other disadvantages, such as "boiler scale." Boiler scale is a scaly buildup of calcium carbonate, $CaCO_3$, which occurs when the calcium ions in hard water have precipitated with dissolved carbonate ions, CO_3^{2-}. This scale can build up on the inside of water pipes and coffee makers. One of the biggest problems boiler scale creates is that it reduces the operating efficiency of water heaters. Even a thin layer of scale inside a water heater can reduce the energy efficiency by 10 percent or more. Scale can also result in the failure of boiler tubes as they become clogged. Once these insoluble salts form a deposit, other metal ions present in the water can become bound to the deposit, increasing the thickness of the boiler scale layer.

The best way to control the formation of boiler scale is through water pretreatment, such as installing a water softener. Water softeners typically replace the Ca^{2+} ions with soluble Na^+ salts. Other water softeners cause the calcium carbonate to form before the water is circulated to the water heater. Because metal ions enter water when it travels through rocks and soil en route to a home, the amount of water hardness varies. When the hard water within your home is 120–150 mg/L as $CaCO_3$, it is recommended that a water softener is installed. If the hardness falls between 60–120 mg/L as $CaCO_3$, the hardness is considered acceptable (moderately hard) and no water softener is needed.

Hard water can contain various metal ions, including Ca^{2+}, Mg^{2+}, and Fe^{3+}. In order to fully determine the hardness of the water, each ion must be isolated separately. When a particular ion species, an *analyte*, needs to be isolated, it is possible to use the tendency of that ion to form an insoluble compound by a precipitation reaction. When the analyte ion is formed into a precipitate, it can be collected through a process called *gravimetric analysis*, where the precipitate is isolated, purified, dried, and weighed. From the mass and the known composition of the precipitate, the amount of the analyte in the original solution can be calculated stoichiometrically. When done properly, gravimetric analysis provides an extremely precise quantitative analysis of the analyte.

Since hard water is commonly expressed as the milligrams of $CaCO_3$ per liter of solution, the quantity of Ca^{2+} in the water identifies how hard the water is. This analyte can be isolated by mixing it with a solution of Na_2CO_3 to form the slightly soluble salt $CaCO_3$.

$$Ca^{2+}\,(aq) + Na_2CO_3\,(aq) \rightarrow 2\,Na^{+}\,(aq) + CaCO_3\,(s)$$

When completing a gravimetric analysis, an important consideration is that the analyte is completely precipitated. This can be accomplished by ensuring that the analyte acts as the limiting reactant in the precipitation reaction.

Once the salt has precipitated, it can be collected through filtration. All of the impurities should be removed from the precipitate through washing and drying.

PREPARATION

Materials

100 mL beakers	250 mL beakers
100 mL graduated cylinders	Analytical balances (±0.001 or ± 0.0001g)
Drying oven	Watch glass
Sodium carbonate, anhydrous (Na_2CO_3)	Calcium chloride, anhydrous ($CaCl_2$)
Vacuum filtration apparatus	Filter papers
Permanent markers	Rubbing alcohol
0.50 *M* sodium carbonate (Na_2CO_3)	Water samples of unknown hardness
Distilled water	Deionized water
Metal scoop	

Safety and Disposal

All solutions can be flushed down the drain. Normal laboratory precautions, including wearing goggles at all times, should be taken as these solutions may be harmful if swallowed and can irritate the eyes.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

Complete the following Procedure section in your lab group. Then answer the questions individually. Be ready to discuss your responses with the class.

Procedure

In this part of the lab, you will make a solution of sodium carbonate and mix it with a solution of calcium chloride. Read the following procedure. Create a data table for the data you will need to collect.

1. Weigh about 2 g of sodium carbonate in a clean, dry beaker.
2. Weigh about 2 g of calcium chloride in a second clean, dry beaker.
3. To each beaker, add about 20 mL of distilled water. Stir each until each solid is dissolved.
4. Pour a small amount of the sodium carbonate solution into the beaker containing the calcium chloride solution. Stir and observe the mixture. Add some more of the sodium carbonate solution, while stirring and observing. After all of the sodium carbonate solution has been added, continue to stir this mixture for another couple of minutes.
5. Weigh a piece of filter paper.
6. Set up the filtering apparatus.

Your teacher might demonstrate the proper procedure for filtration before sending you into lab.

7. Insert the filter paper into the filtering apparatus. Wet the filter paper with deionized water.
8. Pour the contents of your beaker slowly into the funnel. Be careful as you pour so that none of the mixture flows out of the filter paper or the funnel. Use a wash bottle of deionized water to rinse the precipitate out of the beaker with small quantities of water. Use a little more deionized water to wash the precipitate that is now collected in the filter paper.
9. Using a permanent marker, label a clean, dry watch glass with your name. Weigh the watch glass.
10. Carefully remove the filter paper with the precipitate from the filtering apparatus and set it on the watch glass. The filtrate can be poured down the drain.
11. Place the watch glass with the filter paper on it into the drying oven, which should be set between 110°C and 120°C.
12. Allow the filter paper to dry for 10–15 minutes. Answer Questions 1–2 on the next page while you wait.
13. Carefully remove the warm watch glass. Use a metal scoop to break the precipitate into small pieces.
14. Return the watch glass to the drying oven for another 5 minutes.

15. Carefully remove the watch glass from the oven and set aside to cool. Once cool, weigh the watch glass, filter paper, and precipitate. Record this mass.
16. The precipitate and filter paper may be discarded in the wastebasket. Wipe the permanent marker ink off of the watch glass with some rubbing alcohol.

Practice with Instrumentation and Procedure Questions

1. Use the masses of sodium carbonate and calcium chloride to predict the mass of calcium carbonate that will form in your experiment.
2. If one more gram of sodium carbonate were used, how would it affect the amount of calcium carbonate that you *calculated* would form?* You may wish to test your answer by running the procedure again.
3. Which mass of the precipitate, the first or second, better represents the amount of dry precipitate collected? What mass of precipitate did you collect?
4. Is the mass you measured close to the expected mass you calculated based on stoichiometry in Question 1? What may be the reason(s) for any differences?
5. Would the mass of precipitate that you measured be larger or smaller if you did not wash the precipitate before drying it?
6. If the precipitate were weighed without drying, would you believe that you had started with 2 grams of calcium chloride? Explain.
7. Do you feel that the second weighing of your precipitate was dry? What experimental changes could be made to improve this portion of the procedure?

INVESTIGATION

In order to help the client, six samples of water have been collected from areas around your community. You need to analyze two of the samples, determining the hardness of each in mg of $CaCO_3$ per L of solution. The results of your analysis will be shared with the other groups in order to rank the community areas in increasing order of water hardness.

The samples have been boiled to concentrate the metal ions. Since water hardness measures $CaCO_3$ in mg/L, saying that a sample of water is 60 mg/L as $CaCO_3$ means that there are 60 grams of $CaCO_3$ for every 1×10^6 g of solution (ppm). Therefore, the volume of the water samples has been reduced by 200 times. This means that 240 mg (0.24 g) of $CaCO_3$ can be collected from a 20 mL water sample that has been boiled.

Procedure

You will design your own procedure.

Design a procedure that will determine the unknown concentration of Ca^{2+} ions as mg of $CaCO_3$ per L of solution, in a 20 mL aliquot of a water sample.

All of the materials listed are available to use in your experiment. If there are other materials you would like to use, check with your teacher before deciding to use them, as they may not be available.

Write out a detailed, step-by-step procedure, a list of materials needed, and a data table for measurements and observations. Consider what data is needed to obtain the concentration of the Ca^{2+} ions. Show your proposed procedure to your teacher before attempting the experiment yourself.

Once your teacher has reviewed your proposed procedure, you may begin the data collection. Keep a detailed record of your data and observations.

Data Collection and Computation

Answer the following questions after you have completed collecting all of your data.

1. How many grams of each precipitate were collected?
2. What is the hardness, in mg/L as $CaCO_3$, of each water sample?
3. The water softener discussed here relies on precipitation softening, also known as ion-exchange. Based on the reactants used (Na_2CO_3 and $CaCl_2$), what ions would remain in the softened water that would be consumed by the homeowners? What could be some negative aspects of consuming these ions?
4. What other types of water softeners are available, besides the precipitation softening discussed here? Are there any advantages or disadvantages to using one of these, rather than an ion-exchange softener?

Argumentation and Documentation

After your group has completed the Data Collection and Computation, share the hardness of your samples in the class table. Be prepared to defend your laboratory techniques and calculated water hardness of each of your water samples.

After collecting and comparing the data, draft a letter directed to the client. Outline what your group did to determine the water hardness in the community, what was discovered, and which area you would recommend as having the lowest water hardness. Although you do not know the client's financial situation or health history, include which type of water softener you might recommend.

POSTLAB ASSESSMENT

Answer the questions below.

1. Excess Na_2SO_4 (*aq*) is added to a 42.53 mL sample of $Ba(NO_3)_2$ (*aq*).
   a. What is the formula of the precipitate?
   b. If 3.046 g of precipitate was formed, what was the molarity of the $Ba(NO_3)_2$ (*aq*)?

2. A 5.000 gram mixture contains strontium nitrate and potassium bromide. Excess lead(II) nitrate solution, $Pb(NO_3)_2$ (*aq*), is added to precipitate out 0.7822 grams of $PbBr_2$ (*s*).
   a. What is the percent by mass of potassium bromide in the mixture?
   b. What is the percent by mass of strontium nitrate in the mixture?

SUPPLEMENTAL RESOURCES

Links

"Gordon Bros./Kinetico Hard Water Test." YouTube. Accessed July 25, 2012. *http://www.youtube.com/watch?v=YcZSNcaHHN8&feature=youtube_gdata_player*

"Hard Water and Water Softening." Stephen Lower. Accessed July, 25, 2012. *http://www.chem1.com/CQ/hardwater.html*

"Salts and Solubility." University of Colorado at Boulder, PhET Interactive Simulations. Accessed July, 25, 2012. *http://phet.colorado.edu/en/simulation/soluble-salts*

"Treatment Systems for Household Water Supplies: Softening." NDSU. Accessed July 25, 2012. *https://www.ag.ndsu.edu/publications/home-farm/water-softening-ionexchange*

INVESTIGATION 4

How Much Acid Is in Fruit Juice and Soft Drinks?

CENTRAL CHALLENGE

Imagine that you have just been told the pain you feel in your chest after eating is caused by small amounts of stomach acid coming in contact with the lower part of your esophagus and irritating the lining of your esophagus over time. It is recommended that you avoid spicy foods, carbonated beverages, and orange juice. You wonder why. This laboratory investigation gives you an opportunity to explore that question as you design your own experimental procedure to determine how much acid is in fruit juice and soft drinks.

CONTEXT FOR THIS INVESTIGATION

Have you ever wondered how doctors determine the ratios for IV drips during surgery or hospital stays? They use a common laboratory procedure called a titration to calculate specific ratios of different substances using volume measurements. A similar process occurs when someone uses a machine to monitor blood glucose levels, to analyze urine samples, or to conduct a pregnancy test. Pharmacists use titrations when compounding drugs, which allows them to more precisely match a person's drug prescription to their body weight, size, or medical condition. While it may sound like titrations are common practice only in the chemistry lab or medical community, titrations have great practical use. Food scientists use them when testing for levels of salt, sugar, and vitamins in different foods, and for deciding if wine and cheese are ready for consumption. Others use titrations to test water quality or hardness and in neutralizing the free fatty acids in waste vegetable oil before refining it as biodiesel.

In all cases, titrations are used to quantitatively analyze the unknown concentration of a solution or the amount of a substance by comparing it to a solution of known concentration.

PRELAB GUIDING QUESTIONS/SIMULATIONS

In your group, answer the following nine questions and be prepared to discuss your answers as a class.

1. Write the complete chemical equation for the reaction of a solution of sodium hydroxide (NaOH) with hydrochloric acid (HCl).
2. How many mL of 0.1 *M* HCl are required to react completely with 5 mL of 0.1 *M* NaOH?

3. If equal molar amounts of NaOH and HCl are mixed, when the reaction is complete what will be the chemical species in the resulting solution?
4. Will the pH of the mixture in question 3 be acidic, neutral, or basic? Explain.
5. Write the complete chemical equation for the reaction of a 0.1 *M* solution of acetic acid (CH_3COOH) with a 0.1 *M* solution of NaOH.
6. How many mL of the 0.1 *M* NaOH solution will required to react completely with 5 mL of a 0.1 *M* acetic acid (CH_3COOH) solution? Explain.
7. When the reaction is complete what will be the pH (acidic, neutral, basic) of the solution in question 6? Explain.
8. How is it possible to determine when an acid-base reaction is complete when one of the reactant's concentrations is unknown?
9. Using the table below, explain how indicators are chosen and used during titrations.

Indicator	pH range	Color Change
Methyl orange	3.1–4.4	Orange to yellow
Methyl red	4.2–6.2	Red to yellow
Bromthymol blue	6.0–7.6	Yellow to green-blue
Phenolphthalein	8.3–10.0	Colorless to pink
Thymol blue	1.2–2.8	Red to yellow

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

One of the first titrations you perform is the reaction between a hydrochloric acid solution of unknown concentration and sodium hydroxide. Since we are analyzing HCl to determine its concentration, we call it the *analyte*. The hydrochloric acid solution would be "titrated" by adding a standard solution of sodium hydroxide dropwise. The sodium hydroxide solution is called the *titrant*, and is a standard solution because its concentration is accurately known. In this example, the titrant (NaOH) reacts with the analyte (HCl) in the neutralization reaction shown below:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

The point at which enough titrant has been added to react exactly with the analyte is called the equivalence point. Here, the number of moles of OH^- (*aq*) added is equal to the number of moles of H^+ (*aq*) consumed. Since this point is not visible to the eye, we use a pH indicator to help us detect it. The point at which the indicator changes color is called the endpoint, and assuming it has been selected appropriately, indicates that the equivalence point has been reached. The best indicator for the titration is one in which the endpoint and equivalence point are as close together as possible.

In this experiment, you will design your own acid-base titration to determine the acid concentration of fruit juice or carbonated beverage by using a standardized solution of sodium hydroxide, and make a prediction about how knowing the acid content of a certain beverage would be of benefit to them. Nutritionists cite the dangers of regular consumption of carbonated beverages on bone density and recommend a low-acid diet. Dentists caution patients about the relationship between acid and oral health. Is there cause for concern? Might the knowledge gained in this lab help us better understand their recommendations, and make reasonable decisions about our beverage of choice?

PREPARATION

Materials

Your teacher will instruct you on equipment and materials which you may select from when developing your lab procedure. Be sure to indicate in your lab reports which pieces of equipment you will be using.

Beakers of varying sizes (100 mL, 150 mL, 250 mL, and 400 mL)	Various acid/base indicators — methyl orange, methyl red, bromothymol blue, phenolphthalein, thymol blue (or others you may have on hand)
50 mL burets	pH meters or pH paper (if meters are unavailable)
Erlenmeyer flasks of varying sizes (125 and 250 mL)	0.10 *M* acetic acid (CH_3COOH) (500 mL total)
Graduated cylinders of varying sizes (10 mL, 25 mL, and 100 mL)	Various light-colored or clear fruit juices and sodas
Volumetric pipettes of varying sizes (10 mL and 25 mL)	100 mL bottles of standardized 0.10 *M* and 0.25 *M* sodium hydroxide (NaOH) solutions (4.0 L total of each solution)
Droppers	Wash bottles with distilled water

Safety and Disposal

Make sure you consult the relevant MSDS or websites for information regarding safe handling of any chemicals you have chosen to use during your experiment. This information can be found at the following website:
http://www.ehso.com/msds.php

In addition, you will want to follow all safety precautions and lab protocols as described by your instructor.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

If you do not know how to perform a titration or use a volumetric pipette or buret, it is suggested that you follow the steps below before performing this investigation.

Step 1: Put 5.0 mL of 0.10 *M* HCl in a test tube that contains a few drops of phenolphthalein indicator. Add dropwise a 0.10 *M* solution of NaOH until a color change is observed. Count your drops. Be prepared to justify why phenolphthalein was chosen as an indicator.

Step 2: Put 5.0 mL of a 0.10 *M* acetic acid solution in a test tube. Select an appropriate indicator from the list given and add a few drops to the acetic acid solution. Add dropwise 0.10 *M* NaOH until a color change is observed. Count your drops. Be prepared to discuss what the color change of the indicator signifies.

Step 3: Identify the relationship of the volume added to the molarity of the reactants.

Step 4: Some bottled fruit juices contain $C_3H_5O(COOH)_3$ (citric acid). Write a chemical equation for the reaction between citric acid and NaOH (*aq*). Select an appropriate indicator to use in a titration of a citric acid solution (of unknown concentration) with a solution of known concentration of NaOH to the equivalence point.

Step 5: Some colas or sodas contain H_3PO_4 (phosphoric acid). Write the equation for the reaction between phosphoric acid and NaOH. Select an appropriate indicator to use in the titration of phosphoric acid (of unknown concentration) using a solution of known concentration of NaOH.

Step 6: If your group needs additional information on performing titrations, access the animations and simulations below:

1. If a particulate view of the neutralization reaction between a weak acid and a strong base is preferred to help you form a mental image of the interaction between hydronium and hydroxide ions as the reaction proceeds, one can be found at: *http://www.chembio.uoguelph.ca/educmat/chm19104/chemtoons/chemtoons.htm*
    a. Click on "Animation 8 — Titration of a Weak Acid, HA, with a Strong Base."

INVESTIGATION

Procedure

Your instructor will provide you with various fruit juices and sodas, two standardized solutions of sodium hydroxide (NaOH), a phenolphthalein indicator, and a variety of laboratory equipment that may or may not need to be used (see Materials section).

1. Identify a research question pertaining to acid concentration in fruit juices and sodas. Propose a hypothesis that answers your question, and then outline a tentative procedure in your lab notebook or on another sheet of paper. Be sure to clearly identify which pieces of equipment (listed in the Materials section) you plan to use during your investigation.

2. Consult MSDS or websites for information regarding safe handling of any chemicals you have chosen to use during your experiment. Describe at least three safety precautions you need to take as you perform your experiment. MSDS information can be found at the following website:
   http://www.ehso.com/msds.php
3. Make a simple sketch showing how you plan to set up your equipment, labeling the following: analyte, titrant, buret, Erlenmeyer flask, volumetric pipette. Show where the indicator solution will be added, and comment on your choice of indicator: why did you choose this one and not another?
4. Consult with your instructor for approval before beginning your experiment.
5. As you conduct your experiment, keep detailed written records. Be sure to list all steps taken as you perform your experiment, and all measurements and observations made during the experiment.
6. Write the balanced chemical equation for the reaction between sodium hydroxide and the primary acid in the fruit juices or carbonated beverages your lab team selected. If you identified different acids in the drinks, write balanced chemical equations that represent each.
7. When complete, graph pH versus volume of base added for all beverages tested. Make a note of where the indicator chosen changes color. Share your results with your instructor so that experimental data from the entire class may be reviewed (if requested by your instructor).

Data Collection and Computation

Record all measurements taken in a data table of your own design. Show your work as you calculate the acid concentration of each juice or soda tested. Give this information to your instructor so that class data can be pooled and discussed. Then answer the following questions after all data is collected and results are compiled.

1. Calculate the acid concentration(s) of the fruit juice or sodas you tested, being sure to show all work. Include units.
2. Share your results with three other lab groups, and then explain how your answers compare. What similarities and differences do you notice? Is this what you expected? Why or why not?
3. Calculate the pH of the fruit juices or carbonated beverages you tested. Give this information to you instructor so they can pool class data.
4. Graph pH versus volume of titrant (NaOH) added. In general, as volume of titrant increases, what happens to pH? Why does this occur?

Argumentation and Documentation

Pool class data and compare primary acid, acid concentration, and associated pH of the various soft drinks and juices tested. Provide a general statement and explanation for any trends you notice. In what ways do they surprise you and why?

If you were to perform this experiment again, what changes would you make to your procedure and why? How do your answers compare to those of other student groups?

POSTLAB ASSESSMENT

1. Suppose a student chose to measure solution volumes using the beakers or graduated cylinders provided. What effect would this have had on the calculated acid concentration? How might this affect the number of significant figures in your final answer? Explain your answers.
2. A student rinsed the buret with water, but neglected to rinse the buret with titrant before conducting the experiment. What effect would this have on the calculated acid concentration in the juices or sodas? Why?
3. If you did not titrate a carbonated beverage, find another lab group that did so, and ask them to discuss their procedure and results. What step was necessary to determine the endpoint when titrating a cola with a standard sodium hydroxide solution, and why did this step matter?
4. Imagine a lab team that consistently added base past the first appearance of a pale pink color. What would happen to the average calculated acid concentration of the juices or sodas? Explain your answer.
5. Create a diagram that shows the molecular interaction between the acid and base as the titration proceeds. Display this at four points along a titration curve: (a) when 0 mL of base is added, (b) when 5 mL of base is added, (c) at the equivalence point, and (d) when an excess of base has been added, and provide an explanation for what is happening at the molecular level.
6. Review your graph of pH versus volume of base added. Where did the indicator change color for each trial? What was the pH associated with the color change? What assumptions can you make regarding the pH at this point and the shape of your titration curve? Use a diagram to show what you think might be happening at the molecular level.
7. After conducting this experiment, what recommendation(s) might you make for a person with acid reflux disease or tooth decay despite the fact that they drink juice or diet sodas? Justify your answer.

SUPPLEMENTAL RESOURCES

Links

“Acid-Base Titration.” Mark Bishop. Accessed July 27, 2012.
http://preparatorychemistry.com/Bishop_Titration.htm

“Chemical Analysis by Acid-Base Titration.” W. G. Rushin. Accessed July 27, 2012.
http://chem.lapeer.org/Chem2Docs/AcidBaseTitration.html

References

Burns, Andrew. “Idea Bank: Fruits and Soft Drinks in the Lab.” *The Science Teacher* 78, no. 8 (2011): 68–69.

Nelson, John H., Kenneth C. Kemp, and Theodore L. Brown. *Laboratory Experiments [to Accompany] Chemistry: the Central Science*, 11th ed. Upper Saddle River, NJ: Pearson-Prentice Hall, 2009.

Slowinski, Emil J., Wayne C. Wolsey, and William L. Masterton, “Experiment 43: Analysis for Vitamin C.” *Chemical Principles in the Laboratory*, 7th ed., 197–203. Harcourt College Publishers, Inc., Fort Worth, 2001.

INVESTIGATION 5

Sticky Question: How Do You Separate Molecules That Are Attracted to One Another?

CENTRAL CHALLENGE

The central challenge in this laboratory is to develop a method to separate three similar molecules.

CONTEXT FOR THIS INVESTIGATION

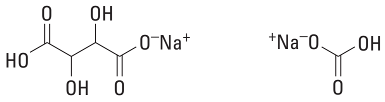
You are working for a crime lab and a chemical residue has been turned in for analysis. To identify the chemicals in the residue, you will need to separate them from the mixture and identify them individually. Another lab technologist has made an attempt to separate the molecules but was not as successful as the boss would like. There was only one molecule separated from the mixture, but the boss suspects that there are at least three different molecules. Science is often a process, in which a method is tried and then modified for a second attempt. Your job will be to propose a modification and attempt to improve the separation attained.

PRELAB GUIDING QUESTIONS/SIMULATIONS

Part I

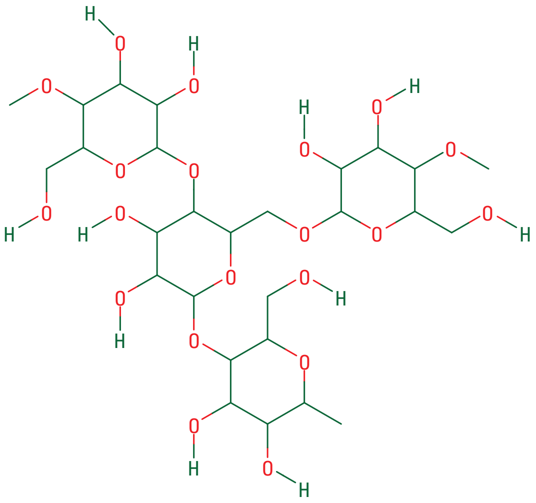
Your teacher is going to show you a demonstration.

1. Why does the water creep up the paper? Explain this. Draw the molecular interactions associated with the macroscopic observation.



Sodium bitartrate

Sodium bicarbonate



Cornstarch polymer

Figure 1. Components of baking powder

2. In your kitchen there is a mixture that is usually listed as a single ingredient in recipes: baking powder. Baking powder is actually a mixture of sodium bicarbonate, cream of tartar (sodium bitartrate), and cornstarch (see Figure 1). Looking at the three molecules in Figure 1, how are they similar and how are they different? Draw a picture of how each molecule would interact with a bunch of water molecules
3. How would you separate this baking powder mixture into the three parts?

Part II

Your teacher may or may not ask you to perform this part of the prelab.

Computer simulation at Concord Consortium: *http://mw.concord.org/modeler/*

Select "More" at the bottom of the "Selected Curriculum Modules" section and then, from the chemistry column, select "Intermolecular attractions." When doing the simulation, make sure to take snapshots and answer all the questions. On the final page, there is a button labeled "Create a report of my work" that will include all written answers and snapshots. Your teacher may also assign the following questions for you to answer after performing the simulations.

1. How can molecules attract each other when they are in a mixture? Predict how ethanol would interact with those molecules. Draw a picture illustrating the interactions between the components of the mixture and the solvent, ethanol.
2. What does the R_f value describe on a microscopic level? Why is this important?
3. If the molecule had a very high affinity for the stationary phase, how would this affect the R_f value? Explain.

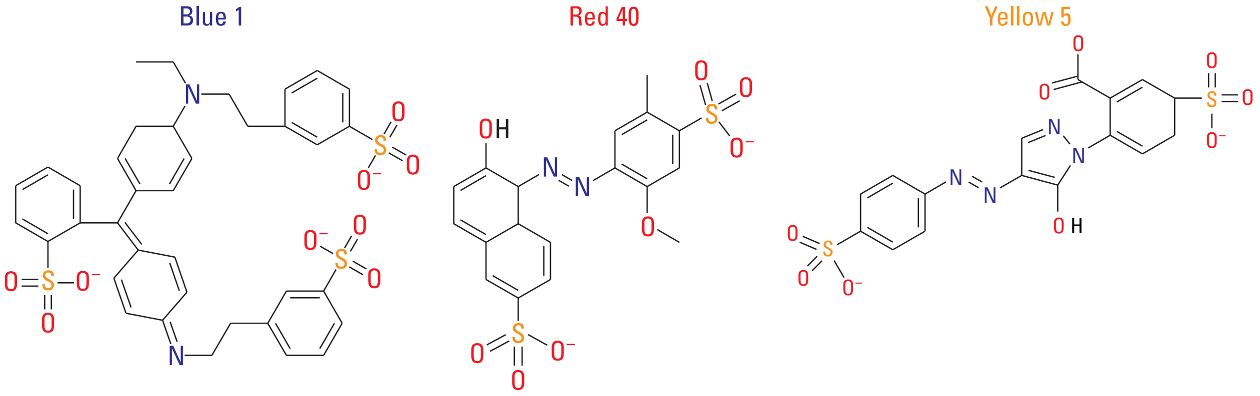
4. What role does the mobile phase play in the distance a molecule travels in chromatography? What does the mobile phase describe?
5. If you combined a polar solvent with a molecule that has a carbonyl group (carbon with a double-bonded oxygen) would it have a high or low R_f value? Justify your answer with what you understand of intermolecular forces.

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

There are two phases in paper chromatography, a stationary phase (the paper) and a mobile phase (the solvent). A molecule can have a greater affinity for either the paper or for the solvent. The filter paper is made of cellulose, a polymer. Cellulose will attract water molecules to the exposed hydroxyl groups along the polymer. This interaction makes a thin layer of water on the paper that competes for the attraction of the molecules being separated. Alternately, the molecule can be attracted to the solvent and travel with the solvent up the paper. When doing chromatography, a small amount of solvent is placed in a sealed container. The mixture being separated is put on a piece of paper, the starting point is marked, and the paper is put into the solvent. The container must be sealed so the solvent saturates the paper and does not evaporate first. The level of separation is measured by a ratio that compares the distance that the molecule travels to the distance the solvent travels. This ratio is called the R_f value. To get the R_f value, the experimenter must identify the distance that the solvent traveled on the paper and measure the distance. Secondly, the experimenter must identify the distance that each molecule traveled and measure that distance. It is best to run the test more than once to reach the best separation values possible. The R_f value is a ratio of the distance of the molecule divided by the distance of the solvent. The greater the distance the molecule travels, the greater its affinity for the solvent and the greater the R_f value.

In this experiment you will have a choice of different solvents to use. When you propose the best solvent for separation of the mixture you will also need to evaluate it in terms of “greenness.” In modern chemistry, chemists use principles of green chemistry to evaluate solvents that are used in a chemical process for their level of toxicity to humans and the environment. Solvents are also evaluated in terms of their life cycle or how long the molecule remains in the environment and if the molecule breaks down to become more benign or more toxic. The overall focus of green chemistry is to be more efficient in chemical production, producing less waste, using fewer toxic molecules, and producing waste that biodegrades and does not pose a risk to the environment. See Tables 1 and 2 in the GSK solvent selection guide, available at *http://pubs.rsc.org/en/content/articlelanding/2011/gc/c0gc00918k* under “supplementary information,” for an evaluation of solvents based on these guidelines. To understand the differences between solvents, look up the green rating for hexane (found in “chromatography solvent”) and compare it to that for 2-propanol.

The food dyes that are in the mixture have their own green chemistry issues. For example, the molecules used may have a life cycle that is longer than previously anticipated and possibly increased toxicity. While scientists evaluate the toxicity of molecules based on the experimental data, efforts to understand the origins of toxicity often look at the structure of the molecule. A key strategy for looking at molecular structure is to identify functional groups that are present. There are specific functional groups that are known to create toxic by-products when they are metabolized in the human body, such as acetaminophen that can be converted to N-acetyl-p-benzoquinoneimine. Some of the food dyes used in this lab are azo dyes, which means that they contain a double-bonded nitrogen connecting multiple aromatic carbons. While the molecules resemble each other, only red 40 has been linked to allergic reactions in some people, but the FDA has not found conclusive evidence that such a dye is unsafe. In Europe these food dyes are not used and natural pigments are used instead.

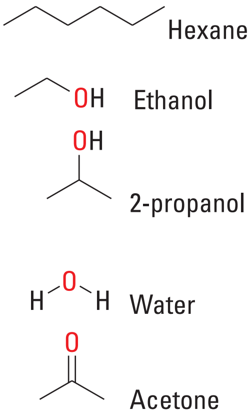


Blue 1

Red 40

Yellow 5

Figure 2. Molecular structure of food dyes



Hexane

Ethanol

2-propanol

Water

Acetone

Figure 3. Molecular structure of typical solvents

PREPARATION

Materials

Sample mixture to separate	Small (25–50 mL) cylindrical glass containers with lids	Tape to make labels
Pencils	100 mL graduated cylinders	Ethanol (C_2H_5OH)
2-propanol ($(CH_3)_2CHOH$)	Distilled water	Acetone ($(CH_3)_2CO$)
Chromatography solvent (petroleum ether and acetone mixture)	Metric rulers accurate to mm	Laboratory filter paper or chromatography paper (cut into strips narrow enough to fit into container and shorter than container)

Safety and Disposal

The mixture of food dyes is safe to dispose down the sink, as well as the distilled water. The other solvents — acetone, 2-propanol, ethanol, and chromatography solvent — should be collected and disposed of in an organic waste container. Teachers and students should take normal laboratory precautions, including wearing splash-proof goggles at all times. If solutions are spilled on skin, wash with copious amounts of water. For student-friendly MSDS, use the searchable index found at the following link: *http://www.ehso.com/msds.php*

INVESTIGATION

You will design an experiment that tests the solvents that you believe will provide the best separation of the three food dyes in Figure 2. There are five solvents listed as available for you to use in this experiment, including chromatography solvent (petroleum ether/acetone mix), acetone, ethanol, 2-propanol, and distilled water. If there are other materials you would like to use, check with your teacher to see if they are available.

Procedure

Develop a hypothesis as to what could lead to an effective separation of the three food dyes in the sample your teacher has provided. Include intermolecular attractions in your hypothesis. You must test at least two of the solvents provided by your teacher (see Figure 3). You will need to write a step-by-step procedure and design a data table to record data. The procedure needs to be clear enough that someone else could repeat the experiment and get the same results. Your teacher will have materials set out for you to choose from.

Data Collection and Computation

1. Why did you select the solvents that you tested? Did your data support your hypothesis or disprove your hypothesis?
2. What explanations can you provide for your separation of the three molecules? How was the choice of the solvent connected to the separation process?
3. What part of the chromatography setup did the molecules interact with, stationary or mobile phase? How would you explain this interaction using intermolecular forces?
4. Draw a picture of how the chromatography worked. Explain your picture using the following terms: stationary phase, mobile phase, and intermolecular forces.
5. Evaluate which solvent is the one with the best "green chemistry" rating (using the reference in the Explanation to Strengthen Student Understanding section). What intermolecular forces would this solvent form with the three molecules in the mixture?
6. Which molecule spent the most time in the stationary phase and why?
7. Calculate the R_f values for each chromatography trial that you completed and include them in the data table.

Argumentation and Documentation

Write an explanation of how intermolecular forces and molecular structure of the molecules being separated determine the ideal solvent for use in chromatographic separation. Be sure to cite specific evidence from this experiment to support your explanation.

POSTLAB ASSESSMENT

Write an explanation of what are the intermolecular forces that would explain why pyridine is soluble in water and benzene is not.

SUPPLEMENTAL RESOURCES

Links

"Chromatography Basics." Science Spot. Accessed July 28, 2012. *http://sciencespot.net/Media/FrnsScience/chromacard.pdf*

"GSK Solvent Selection Guide." RSC Publishing. Accessed July 28, 2012. *http://pubs.rsc.org/en/content/articlelanding/2011/gc/c0gc00918k* (under "supplemental resources")

"Isolation of Plant Pigments by Column Chromatography." Amrita. Accessed July 28, 2012. *http://amrita.vlab.co.in/?sub=3&brch=64&sim=160&cnt=1*

Marsden, Steve. "Paper Chromatography of Plant Pigments." Chem Topics. Accessed July 28, 2012. *http://www.chemtopics.com/unit06/pchrom.pdf*

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Fulton, April. "FDA Probes Link Between Food Dyes, Kids' Behavior." *Morning Edition*, NPR. Accessed July 28, 2012. *http://www.npr.org/2011/03/30/134962888/fda-probes-link-between-food-dyes-kids-behavior#chart*

Gilman, Victoria. "Food Coloring." *Chemical & Engineering News* 81, no. 34 (2003): 34. Accessed July 28, 2012. *http://pubs.acs.org/cen/whatstuff/stuff/8134foodcoloring.html*

Murphy, Parvathi. "Molecular Handshake: Recognition through Weak Noncovalent Interactions." *Journal of Chemical Education* 83, no. 7 (2006): 1010–1013.

INVESTIGATION 6

What's in That Bottle?

CENTRAL CHALLENGE

You will determine the type of bonding in unlabeled chemicals using physical and chemical properties of substances containing ionic, molecular (polar and nonpolar covalent), and metallic bonds.

CONTEXT FOR THIS INVESTIGATION

There is a problem in the chemical storeroom. The high humidity in the storeroom caused the labels on some of the chemical bottles to fall off. The labels are lying all over the shelves and it is your job, as a chemistry intern, to design a method that will help identify the chemicals so the labels can be put onto the correct bottles. The unlabeled chemicals are all solids but may be ionic compounds, nonpolar or polar covalent compounds, or metals. There are at least four unlabeled bottles that represent at least one of each type of bond. If the type of substance, or, even better, the identity can be determined, disposal will be less costly to the school. Once the properties of the unknown substances are determined, you will be given information that can help identify the name of each chemical within the unlabeled bottles.

PRELAB GUIDING QUESTIONS/SIMULATIONS

Answer Questions 1–2 using Table 1.

Table 1. Properties and Bond Types of Solid Substances

Compound	Observations	MP (°C)	Solubility in 25°C Water	Types of Elements Metal (M), Nonmetal(NM)	Type of Bond
Potassium chloride (KCl)	White solid	993	Yes	M/NM	Ionic
Sucrose ($C_{12}H_{22}O_{11}$)	White solid	186	Yes	NM/NM	Polar covalent
Iodine (I_2)	Dark gray solid	114	Slightly soluble	NM/NM	Nonpolar covalent
Zinc (Zn)	Gray, shiny metal	420	No	M	Metallic

1. Compare the type of bond with regard to the properties below using Table 1 and explain any relationships. HINT: Think of what is happening between the bonded atoms as well as what occurs between the particles.
   a. melting point
   b. solubility in 25°C water
2. Predict the properties of each substance below based on Table 1.

Substances	Bond Type: Nonpolar Covalent, Polar Covalent, Metallic, Ionic	Relative Melting Point (High or Low)	Solubility in Water
Hexane (C_6H_6)			
Bromobenzene (C_6H_5Br)			
Sodium chloride (NaCl)			
Iron (Fe)			

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

There are few greater potential hazards around the laboratory than that of unmarked or improperly labeled chemicals. Many schools house unused, unlabeled, and improperly stored chemicals. These chemicals can pose a risk to humans and the environment. All chemicals must have complete identification securely fastened to their containers. Chemicals of unknown stability and those that deteriorate over time should have a preparation date clearly indicated on the label.

The chemistry storerooms in schools need to be cleaned out periodically and chemicals properly disposed of. Most chemicals should not be flushed down the drain or thrown into the garbage. Proper disposal of chemicals is costly. If the identity of a substance is not known due to poor labeling or lack of a label, the cost of proper chemical disposal can increase. There are many accidents associated with chemicals that are thrown out and inadvertently mix.

Disposal of unlabeled bottles is dangerous and therefore very expensive and closely regulated by law. The purpose of proper labels is to indicate the source, supplier or manufacturer of the chemicals, the production date, CAS (Chemical Abstract Service) number of the chemical, and to warn of possible hazards. The MSDS (Material Safety Data Sheet) provides personnel with procedures for handling and cleaning up each substance in a safe manner along with details on their physical and chemical properties and toxicity.

PREPARATION

Materials

Potential unknown solids:					
Ammonium chloride (NH_4Cl)	Magnesium oxide (MgO)	Benzoic acid (C_6H_5COOH)	Aluminum	Wax/paraffin	Magnesium
Calcium carbonate ($CaCO_3$)	Potassium nitrate (KNO_3)	Urea ($(NH_2)_2CO$)	Calcium	Iodine (I_2)	Zinc
Copper (II) sulfate, anhydrous ($CuSO_4$)	Sodium carbonate (Na_2CO_3)	Sucrose ($C_{12}H_{22}O_{11}$)	Copper	Sodium acetate ($NaC_2H_3O_2$)	Sodium hydrogen carbonate ($NaHCO_3$)
Copper (II) sulfate pentahydrate ($CuSO_4{\bullet}5H_2O$)	Sodium chloride ($NaCl$)	Salicylic acid ($C_6H_4(OH)COOH$)			
Materials for testing:					
95% Ethanol, (C_2H_5OH) 30 mL dropper bottle	Ice	Magnifying lens	Capillary tube	Beaker, 100 mL	Small test tubes and rack
Hexanes, 30 mL dropper bottle	Phenolphthalein (dropper bottle)	Ring stand	Thermometer clamp	Thermometer	Conductivity meter or tester (metals and aqueous only)
Distilled Water, 30 mL dropper bottle	pH paper and glass rod	Wooden splints	Orthodontic rubber band	Canned food lid	Relevant MSDS for all knowns and unknowns used
0.1 M sodium hydroxide ($NaOH$) 30 mL dropper bottle	Corks for test tubes	Magnet	Cotton swabs	Well plate or micro well plate	Tongs
0.1 M hydrochloric acid (HCl) 30 mL dropper bottle	Hot plate	Disposable gloves	Wire gauze	Sandpaper	Toothpicks
Universal indicator (dropper bottle)					

Safety and Disposal

Safety goggles should be worn at all times in the laboratory. Be cautious of acidic and basic solutions since they can cause skin burns and eye damage. Liquids and solids are to be disposed of in properly labeled waste containers per MSDS guidelines. It is recommended that the lab be done on a small scale to minimize solvent and chemical exposure. If hexanes and iodine are used, it is advisable for the teacher to leave these substances in the hood for student use and to use and exercise appropriate safety precautions. A recommended site for MSDS information is *http://www.ehso.com/msds.php*

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

The purpose of this portion of the lab is to identify properties that allow one to determine the type of bonding in a substance and to carry out tests that allow one to characterize these properties.

Procedure

Given four to six known substances, you will choose at least four different tests, qualitative or quantitative, to study the physical and chemical properties of each of the given substances. Based on your results you will develop a system that will help determine whether an unknown solid is ionic, covalent (polar or nonpolar), or metallic using these tests. Characteristics to consider testing include: color, solubility in water, conductivity of the solid, conductivity in water, pH of the solution in water, solubility in ethanol, solubility in hexanes, high/low melting point (order of melting if qualitative, or quantitative value), reaction with 0.1 *M* HCl, reaction with 0.1 *M* NaOH, and magnetism.

Select at least four tests and write a detailed procedure to carry out the tests. Refer to the Materials section for guidance on available materials. Create a data table to record results.

Practice Questions

1. Based on the Practice Instrumentation and Procedure data, list the general properties associated with each bond type (metallic, ionic, polar covalent, nonpolar covalent).
2. Go to two other groups and compare your answers. Write a brief statement about what you learned from discussing results with other groups or as a class.
3. How can your experimental procedure be improved? List what your team/group would do differently. If any additional materials are needed, please inform the teacher.
4. Design a flowchart using your experimental procedure that can help you identify unknown substances.
5. You are given a blue crystalline solid. Using your flowchart, explain how the type of bond can be determined, and what you might observe in the lab.

INVESTIGATION

Procedure

The knowledge acquired in the Practice section for known compounds will now be applied in order to determine the type of bond for four to six unknown solids. Given four to six unknown compounds, you will choose at least four different tests to study physical and chemical properties of the substances. Based on your results, you will develop a system that will help determine whether an unknown solid is ionic, covalent (polar or nonpolar), or metallic using these tests. Characteristics to consider testing include: color, solubility in water, conductivity of the solid, conductivity in water, pH of the solution in water, solubility in ethanol, solubility in hexanes, high/low melting point (order of melting if qualitative, or quantitative value), reaction with 0.1 *M* HCl, reaction with 0.1 *M* NaOH, and magnetism.

Select at least four tests, quantitative and qualitative, and write a detailed procedure to carry out the tests. Refer to the Materials section for guidance on available materials. Create a data table to record results.

Data Collection and Computation

After you have completed your procedure, identify the bond type in each unknown.

Obtain a list of all of the unlabelled bottles (the unknowns) from your instructor. Using your results and the MSDS for the unknowns, identify the four to six chemicals you tested. Check your results with your instructor.

Argumentation and Documentation

1. To what extent do you believe the classification of your unknown is reliable? Justify your claim with evidence.
2. Discuss in your group the two most significant tests done to identify each of the types of bonds.
3. Go to another group and compare your answers for Questions 1 and 2. Do you need to revise the answers?
4. Obtain the MSDS of your substances to summarize the toxicity and method of disposal for each of your "unlabeled bottles."

POSTLAB ASSESSMENT

You may find it necessary to talk to other groups to compare findings as you complete these questions.

1. How do the melting points of ionic compounds compare to those of covalent compounds? What evidence from the investigation supports your conclusion?
2. When the solids were placed in water were all the results the same? What types of solids conduct electricity in water? Use your investigations to explain what happened.
3. Explain how you were able to determine each unknown as being an ionic, metallic, or covalent (polar or nonpolar) compound.
4. If the solid is ionic, explain why you cannot make the general statement that "all ionic compounds are soluble in water." What evidence from the investigation supports your conclusions?
5. Why was it necessary to use distilled water and not tap water?
6. Metal oxides dissolved in water show a pH in what range? In contrast to these metal oxides, do nonmetal oxides produce the same pH range?
7. Wax is a saturated hydrocarbon, a covalent compound. Wax is not soluble in water yet sugar is also a covalent compound and is soluble in water. Look at the structure of both compounds and explain what could justify these results.

SUPPLEMENTAL RESOURCE

Link

Carpi, Anthony. “Chemical Bonding.” Vision Learning. Accessed July 27, 2012. *http://visionlearning.com/library/module_viewer.php?c3=&mid=55%20*

INVESTIGATION 7

Using the Principle That Each Substance Has Unique Properties to Purify a Mixture: An Experiment Applying Green Chemistry to Purification

CENTRAL CHALLENGE

In this laboratory activity, you will be a chemist who has been asked to participate in the peer-review process by an editor of a journal on green chemistry. This editor has received three different manuscripts (lab reports) that report on the same process of separating two substances.

You will design your own procedure to determine the relative amounts of two substances in a mixture using green chemistry principles. Afterward, you will receive one of the three manuscripts submitted to the journal. It is your job to assess the quality of the lab report you are reviewing and to write a review of the lab report and submit it to the editor.

CONTEXT FOR THIS INVESTIGATION

Some chemists design chemicals and chemical processes. These practices involve a few main principles. One of these is that pure substances have unique properties that can be used to distinguish these principles from one another. A second principle is that chemists consider the benefits and risks of different options when deciding which is the best chemical process. These two principles form the focus of this experiment.

PRELAB GUIDING QUESTIONS/SIMULATIONS

Prelab Part I:

Green chemistry is a set of guidelines used by chemists to address sustainability when applying both of these principles. We will focus on one particular principle of green chemistry that is relevant to this lab.

Review the 12 Principles of Green Chemistry at
https://www.epa.gov/greenchemistry/basics-green-chemistry#twelve

Principles of Green Chemistry

1. Create no waste
2. Nothing should be left over
3. No toxicity
4. Green products have to work as well as nongreen products
5. Get rid of all nonessential additives
6. Reduce energy usage
7. Use renewable materials
8. Get rid of as many steps as possible
9. Make use of a reusable method to speed up a reaction
10. Use materials that break down in the environment (biodegradable)
11. Check everything you do against the other principles
12. Safety first

Principle 2 is called the atom economy principle: the less that is left over, the better. In the most ideal case, all of the starting materials are converted into the desired product and there is nothing left over. In order to compare processes in terms of how much is left over, a calculation can be made of atom economy that compares the mass of the desired product (D) to the total mass of products (T). Atom economy = D/T. When the ratio D/T equals 100%, then there is nothing left over because there is only one product, the desired one. If the ratio is less than 100%, it is because there are products that are wastes. The lower the ratio gets, the less product is produced in comparison to the wastes produced. The atom economy is therefore a measure of how green, or efficient, a chemical process is. The higher the atom economy is, the greener the process is, and the more efficient the process is for producing the desired product.

Prelab Part I Questions

1. Select one of the principles of green chemistry. Explain in your own words how this principle shifts chemistry toward more environmentally conscious practices. How is this relevant to considering the benefits and risks when making decisions about which of two (or more) possible chemical processes is better?
2. The following two reactions are possible methods for refining copper in the final step of a smelting process, i.e., getting pure copper (Cu) from copper ores found in rocks. Calculate the theoretical atom economy for each reaction.
   a. $2\,CuO\,(s) + C\,(s) \rightarrow 2\,Cu\,(s) + CO_2\,(g)$
   b. $CuO\,(s) + CO\,(g) \rightarrow Cu\,(s) + CO_2\,(g)$

3. Use your calculations from the previous question (i.e., 2a and 2b) to answer the following questions.
    a. Which one of the methods for refining copper ore is greener according to the atom economy principle of green chemistry?
    b. Why is a calculation of atom economy helpful in comparing two chemical reactions to determine which one is greener? In other words, what does atom economy tell you about "greenness"?
    c. What is another possible consideration from the principles of green chemistry that could tell you more about comparing the "greenness" of these two reactions?
4. Peer review is a critical component of how scientists communicate what they have learned and contribute new knowledge. What is peer review? Why do you think that scientists believe it is important for published work to be peer reviewed prior to publication? Why is peer review an important part of a process of chemists considering benefits and risks? Why might peer review be important in chemists reporting results of a chemical process?

Prelab Part II Questions:

1. What is a substance? What is a mixture? How are they related?
2. What are the general characteristics of substances chemists use to separate mixtures into individual substances?
3. How are the two ways of measuring the efficiency of a reaction the same? How do they differ?

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

One can characterize a major practice of chemists as harvesting materials from nature, separating the individual substances within them, and putting the pure substances in bottles so that they can make new substances out of them. The practice of designing chemicals is fundamental to all of the sciences that build on knowledge of chemistry. Pharmacists and medical biochemists create, purify, and test the safety of medicines. Analytical chemists use and design instruments to detect the presence and properties of individual substances. Bioorganic chemists design new chemicals for specific purposes and invent and compare the synthetic processes needed to make such new chemicals. Chemical engineers manage chemical process industries to produce useful products on larger scales and minimize harm. But even if you do not go into one of these fields, understanding some of the basic principles of chemistry is relevant to your life. For the rest of your life, you will use ways of thinking as a chemist as you make educated decisions. You will use chemistry principles as you think about the safety of products you buy, energy reduction in your home, which containers are safest for cooking or storing food, and how to dispose of all sorts of things in the home and office, such as cleaning agents, unused medicines, and batteries.

This experiment involves a major practice of chemists: separating substances in a mixture by taking advantage of properties of the substances that are unique to each one, as well as determining the relative amount of each substance within the mixture. In this case, you will rely on the substances' chemical reactivity upon heating as the property that differs between them. From antiquity, two very important substances in society have been obtained from a salt mixture called *natron*. Natron has been harvested for thousands of years from dry lake beds. In ancient Egypt, and still today, natron is blended with oil and used as soap. Natron primarily consists of two substances, sodium carbonate (Na_2CO_3) and sodium bicarbonate ($NaHCO_3$). Each of these substances, when separated, also has important uses. Sodium carbonate is used in the manufacture of glass, as a water softener for laundry, as an additive in community swimming pools to raise pH, and as an additive in foods. Sodium bicarbonate has many uses, ranging from cooking and medical uses, to cleaning, pesticide, and fire extinguishing uses.

Every process that a chemist uses involves some kind of transformation and has an efficiency of that transformation associated with it. There are many ways to measure and report efficiency; two of these are discussed here. For the first way, the efficiency of a chemical process can be interpreted in terms of its financial cost or in terms of its benefit or detriment to the environment. When considering financial cost, one might compare the amount of the desired product that is actually produced and the amount that could be produced in the ideal situation where all of the starting materials are used up and converted to useful products. Percent yield is such a calculation, as the ratio of the two values.

$$Percent\ Yield = \frac{Actual\ Yield}{Theoretical\ Yield} \times 100$$

Equation 1

The number is usually reported as a percentage (out of 100%) instead of as a decimal, in order to make comparison to the ideal situation easier. Percent yield is a way of comparing actual runs in the laboratory of a chemical process, when striving for a maximum yield. If the percent yield is closer to 100%, then maximum financial value is derived because starting materials (which might be expensive) are maximally used, and the end product is maximally produced. In a second way of measuring and reporting efficiency of a process, considering environmental health, one might instead try to minimize undesired products such as wastes or byproducts that are an inevitable outcome of a chemical process. Atom economy is such a calculation, comparing the theoretical yield of the desired product and the total theoretical yield of all products as a ratio.

$$\text{Percent Atom Economy} = \frac{\text{Mass of desired product produced}}{\text{Mass of all products produced}} \times 100$$

Equation 2

Different chemical processes have different atom economies, because the desired product can range from being the only substance produced to being only of one several substances that result from the reaction. If the atom economy is closer to 100%, then the amount of desired product is maximized. Thus, different chemical processes can be compared to each other in terms of which one is more efficient in resulting in the desired product. The process with the largest atom economy is the one that most cleanly produces the substance that is the goal of the process. Today, when we recognize that disposing of unwanted byproducts is expensive and potentially harmful, this method of comparing chemical processes provides a very useful tool in making decisions that support the principles of green chemistry.

INVESTIGATION

Procedure

You have a mixture of $NaHCO_3$ and Na_2CO_3. How will you determine the relative amounts of $NaHCO_3$ and Na_2CO_3 in a mixture of the two substances using the 12 principles of green chemistry and stoichiometry? Work in your group to devise a procedure to collect data for the aforementioned question. Below is a list of materials available for you to use during your procedure.

A mixture of sodium hydrogen carbonate ($NaHCO_3$) and sodium carbonate (Na_2CO_3)	Bunsen burner lighter, also known as a striker
Transferring scoop	Ring stands
Digital balance, with at least ±0.01 g precision	Iron ring
Crucible and lid	Ceramic triangle
Crucible tong	Collection container (can be glass or plastic) for Na_2CO_3 produced, labeled “Product Made from Heating Samples”
Bunsen burner and associated tubing	Collection container for unused mixture, labeled “Unused Sample”

Safety and Disposal

Do not heat covered crucibles. Always place the lid askew on top of the crucible while heating. Completely covering the crucible is dangerous because gases can build up inside while heating and the hot lid can fly off.

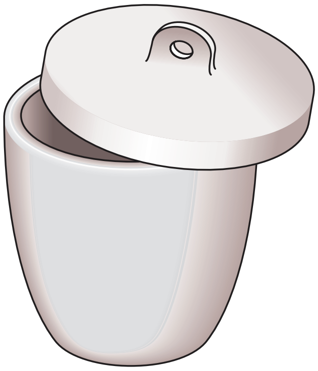

Figure 1: Crucible Lid Askew

Place the solid product that remains after heating in the container labeled “Product Made from Heating Samples.” Place any mixture that was taken but not heated in the container labeled “Unused Sample.” After the laboratory session, the contents of the “Product Made from Heating Samples” container should be heated to constant mass again, in case any partially heated sample was placed in it. This product will be saved and stored, to be used the next time this experiment is done. There are no significant chemical hazards in this experiment. Normal laboratory precautions are sufficient.

- Make sure your instructor has reviewed and approved your proposed procedure and safety precautions prior to proceeding with your lab investigation.
- Make sure you carefully record and organize your data as well as your calculations.
- Important note about use of laboratory equipment: Allow crucibles to cool before weighing. A hot crucible can damage an analytical balance. Also, convection currents caused by a hot crucible will cause the mass reading to be inaccurate by being too low.
- Use the same analytical balance every time, to reduce systematic errors that could be introduced by balances that might be calibrated differently.

Data Collection and Computation

- Remember that there are two very important principles of chemistry that you should be thinking about throughout the entire lab: 1) each substance has unique properties, which enables chemists to distinguish it from other substances, and 2) chemists consider benefits and risks in making decisions about chemical processes. Green chemistry is a set of guidelines used by chemists when applying both of these principles, and the second principle of green chemistry (the atom economy principle) uses both of these. Try to find instances where the above two principles of chemistry are relevant to what you are doing.
- Record your data and observations as you carry out your verification experiments in an organized manner of your own design.
- After you have performed your procedure, collected your data, and determined the relative amounts of $NaHCO_3$ and Na_2CO_3 in a mixture of the two substances, you will work in your group to review another scientist’s lab report pertaining to the same question that you addressed with your procedure, data, and calculations.

Argumentation and Documentation

As a complication to this situation, the very same investigation you performed has been conducted by three other scientists, who have all submitted lab reports to the same journal. You will review one of the three lab reports and make a report on its strengths and weaknesses.

You will review a lab report given to you by your teacher, and then provide a written review to your teacher, who is acting as the editor.

- All scientists in each group receive a copy of the same lab report. There are three different lab reports, each one written by a scientist who conducted a similar investigation using the same chemical sample you will use, with a report on the results of the investigation. Scientists in your group will receive only one of the three lab reports, and all scientists in your group will receive the same lab report. Each lab report has some information that is reported well, according to conventions that scientists have agreed upon, and some information that is reported poorly.
- The journal to which the lab reports have been submitted is a green chemistry journal. The signature of this journal is that all chemical reactions reported on are assessed for their "greenness." Therefore, you will calculate the atom economy of the reaction to contribute to the journal's collection of chemical processes, and you must assess the "greenness" of the chemical reaction in at least one other way, according to one of the other principles of green chemistry.
- In summary, your job as a scientist who has been asked to participate in the peer-review process is to assess the quality of the lab report you are reviewing, and to write a review of the lab report and submit it to the editor. There are three main factors you are assessing:
    1. What is the quality of the science in the lab report?
    2. What is the quality of the reporting the scientist did in the lab report?
    3. How green is the chemical reaction that is reported on?

The review template provides more specific questions to help you assess these.

Lab Report Review

In each section of the lab report review table below provide a concise report on what is done well in the report and what is lacking. Provide suggestions for the author of the lab report on how to improve it. Write in complete sentences. You may use second person (e.g., "You should do more of …") or third person (e.g., "The author should do more of …"), but be consistent with whichever one you choose.

Lab Report Reviewed (circle one): 1 2 3	Date That Review Was Submitted:
Name of Reviewer(s):	

Reviewer Criteria and Response
1. What is the quality of the science in the lab report?
a. Was the investigation conducted well?
b. Were sufficient trials run to instill confidence in the results reported?
c. Is there evidence that the data were collected carefully?
d. Are the explanations of chemistry in the lab report correct?
e. Are the calculations done correctly?
f. Does the report account for any problems in the investigation?
2. What is the quality of the reporting the scientist did in the lab report?
a. Does the lab report communicate the investigation effectively?
b. Is the procedure detailed enough that it can be repeated to verify the results?
c. Were the materials and equipment specified sufficiently?
d. Are the calculations presented clearly?
e. Are any of the sections of the lab report weak? If so, how can they be strengthened? (For example: Is the procedure clear? Are the data reported effectively?)
3. How green is the chemical process that is reported?
a. What is the atom economy of the chemical reaction? To calculate this, you must identify the desired product. Show your calculation.
b. Provide an argument, based on a different principle of green chemistry, for or against the "greenness" of the chemical process reported in the lab report.

POSTLAB ASSESSMENT

1. Epsom salt is a strong laxative used by veterinarians to treat animals, and soaking swollen feet or skin with a rash in a solution of Epsom salt and warm water is sometimes prescribed by doctors to relieve swelling or itchiness. Epsom salt is a hydrate, which means that a specific ratio of water molecules per unit of salt formula regularly repeats in the crystalline structure of Epsom salt. The formula for Epsom salt can be written as $MgSO_4 \cdot xH_2O$, where the ratio of water molecules per unit of salt formula is x:1. When hydrated Epsom salt is heated to at least 250°C, all of the waters of hydration are lost, according to the reaction:

$$MgSO_4 \cdot xH_2O\ (s) \rightarrow MgSO_4\ (s) + x\ H_2O\ (g)$$

   a. When 3.648 g of Epsom salt were heated to constant mass at 250°C, 1.782 g of $MgSO_4$ powder remained. What is the value of x?

   b. If anhydrous $MgSO_4$ is the desired product, what is the atom economy of this reaction?

2. Propene (CH_3-CH-CH_2) is a raw material used in a vast number of chemical processes. Two-thirds of the demand for it is for manufacturing polypropylene, a material that is ubiquitous in the modern world. Propene can be synthesized by several processes. Two such processes are:

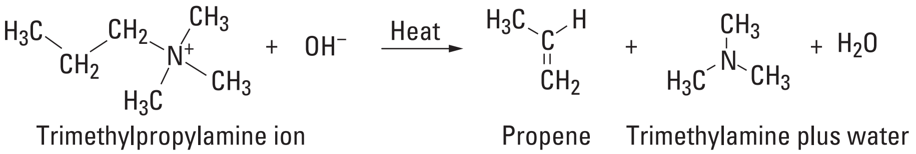

Figure 2. Process A to synthesize propene

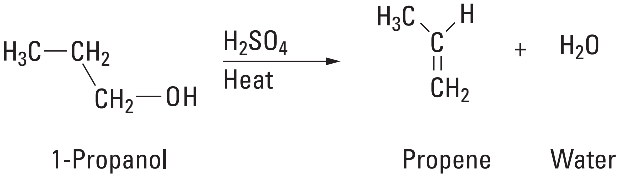

Figure 3. Process B to synthesize propene

Use atom economy to argue which of these processes may be greener.

SUPPLEMENTAL RESOURCES

Links

"Chemical Reactions and Stoichiometry." The Concord Consortium. Accessed July 27, 2012.
http://www.concord.org/activities/chemical-reactions-and-stoichiometry

“Gravimetric Determination of Arsenic Lab.” The Chem Collective. Accessed July 27, 2012. *http://chemcollective.org/activities/vlab/75*

“Unknown Silver Chloride Lab.” The Chem Collective. Accessed July 27, 2012. *http://chemcollective.org/vlab/83*

Reference

Cacciatore, Kristen L., and Hannah Sevian. “Teaching Lab Report Writing Through Inquiry: A Green Chemistry Stoichiometry Experiment for General Chemistry.” *Journal of Chemical Education* 83, no. 7 (2006): 1039–1041.

INVESTIGATION 8

How Can We Determine the Actual Percentage of H_2O_2 in a Drugstore Bottle of Hydrogen Peroxide?

CENTRAL CHALLENGE

This lab has two major tasks. The first task is to standardize the concentration of a $KMnO_4$ solution. This task is necessary in order to complete the second task, which is to evaluate how close commercial H_2O_2 solutions are to their labeled concentrations. Different groups of students will work with different brands and then share their results.

CONTEXT FOR THIS INVESTIGATION

Container labels provide detailed information about the contents present in a given container. Who determines that information? Are there mechanisms used by manufacturers to verify the information present on those labels? Also, what happens when container seals are broken? Will this have an effect on the contents present in the container? These are questions you will address in this lab.

PRELAB GUIDING QUESTIONS/SIMULATIONS

Prelab Part I

Complete the following prelab assessment independently or in small groups as you read the Explanation to Strengthen Student Understanding section. It will prepare you for developing your own procedure, analyzing your data, and performing lab calculations. After completing the prelab assessment, your teacher will direct you on how to share your responses with the entire class.

1. What measuring devices are used to obtain precise measurements of volumes? What measuring devices are used to obtain approximate measurements of volumes?
2. Write a balanced half-reaction for the reduction of permanganate ions in acidic solution. What are the oxidation states of manganese in this reaction?
3. Write a balanced half-reaction for the oxidation of hydrogen peroxide. What are the oxidation states of oxygen in this reaction?
4. What is the balanced reaction for the reduction of permanganate ions by hydrogen peroxide? How many electrons are transferred in this reaction?

5. Write a balanced half-reaction for the oxidation of iron (II) ions.
6. Write a balanced half-reaction for the reduction of permanganate ions by iron (II) ions. How many electrons are transferred in this reaction?
7. Besides iron (II) ions and hydrogen peroxide, what are one or two other species that could be used to reduce permanganate ions?
8. A sample of oxalic acid, $H_2C_2O_4$, was analyzed using a standardized solution of $KMnO_4$. 25.0 mL of oxalic acid is titrated after heating. 12.30 mL of a 0.0226 M $KMnO_4$ was added to the sample when a faint pink color was observed. The balanced equation for this reaction is shown below:

$$6\,H^+(aq) + 2\,MnO_4^-(aq) + 5\,H_2C_2O_4(aq) \rightarrow 10\,CO_2(g) + 8\,H_2O(l) + 2\,Mn^{2+}(aq)$$

   a. What is the ratio of MnO_4^- ions to $H_2C_2O_4$ molecules in this reaction?
   b. How many moles of MnO_4^- ions reacted with the given amount of oxalic acid solution?
   c. How many moles of $H_2C_2O_4$ were present?
   d. What was the molarity of the oxalic acid solution?
   e. If the density of the oxalic acid solution was 1.00 g/mL, what was the percentage by mass of oxalic acid in the solution?
   f. What does the faint pink color indicate about the reaction?

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

Hydrogen peroxide, H_2O_2, is easily oxidized. It is used in commercial bleaching processes and in wastewater treatment plants as an environmentally friendly alternative to chlorine. Dilute solutions of H_2O_2 are used to bleach hair and to clean wounds. It readily decomposes in the presence of light, heat, or metallic catalysts into water and oxygen. It is important to know the actual concentration of a solution of H_2O_2 as its effectiveness can decrease with smaller concentrations.

Oxidation-reduction (redox) reactions involve a transfer of electrons between the species being oxidized and the species being reduced. The reactions are often balanced by separating the reaction components into two half-reactions: *oxidation* (loss of electrons) and *reduction* (gain of electrons). In a redox reaction, the number of electrons lost by the species being oxidized is always equal to the number of electrons gained by the species being reduced. In the reaction being studied in this lab, solutions of hydrogen peroxide, H_2O_2, and potassium permanganate, $KMnO_4$, will be combined in acidic solution. Deep purple in solution, the Mn in $KMnO_4$ undergoes reduction very easily. In acidic solution, permanganate ions (MnO_4^-) from $KMnO_4$ reduce to nearly colorless Mn^{2+} ions. In the presence of permanganate ions in acidic solution, an aqueous solution of H_2O_2 will undergo oxidation to make oxygen gas and hydrogen ions.

Solutions of $KMnO_4$ are not easily standardized solely by preparation, as solid $KMnO_4$ often contains impurities such as chlorides, sulfates, and nitrates. $KMnO_4$ can be standardized in acidic solution with a known concentration of iron (II) ions. In this oxidation-reduction reaction, manganese (II) ions and iron (III) ions are formed.

PREPARATION

Materials

100 mL or 150 mL beaker	125 mL or 250 mL Erlenmeyer flask	400 mL or 600 mL beaker
10 mL or 25 mL graduated cylinder	50 mL or 100 mL graduated cylinder	50 mL buret
Buret clamp	Ring stand	10 mL volumetric pipette
1 mL volumetric pipette	5 mL volumetric pipette	10 mL graduated pipettes
Dropping pipette	Pipette bulb	Funnel
Stirring rod or magnetic stirrer with magnet	Waste container	0.100 *M* ammonium iron(II) sulfate $Fe(NH_4)_2(SO_4)_2{\cdot}6H_2O$ (Fe^{2+} solution)
~0.02 *M* potassium permanganate ($KMnO_4$)	6 *M* sulfuric acid (H_2SO_4)	Various 3% hydrogen peroxide (H_2O_2) solutions
Distilled water		

Safety and Disposal

Review chemical MSDS for $KMnO_4$, 3% H_2O_2, 6M H_2SO_4, and $Fe(NH_4)_2(SO_4)_2{\cdot}6H_2O$ prior to carrying out the experiment. Online MSDS for these chemicals are located at *http://www.ehso.com/msds.php*

The sulfuric acid and potassium permanganate solutions require careful handling; gloves are strongly recommended. Solid $KMnO_4$ and H_2SO_4 are not to be mixed, as an explosion could result. Proper ventilation is essential. Both solutions can cause skin burns and eye damage. Teachers and students should take normal laboratory precautions, including wearing splash-proof goggles and aprons at all times. If solutions are spilled on skin, wash those areas immediately with copious amounts of water. Review local and/or state guidelines and specific procedures regarding the disposal of laboratory chemicals and waste materials.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

Procedure

Part A. For the first 15 minutes, you will:

- develop a procedure that will allow you to determine the concentration of the $KMnO_4$ solution;
- collect qualitative and quantitative data that will allow you to determine the concentration of the $KMnO_4$ solution; and
- determine the concentration of the $KMnO_4$ solution.

You will do this using:

- 10 mL of $KMnO_4$ solution that is added in small portions to a 5–15 mL sample of an acidified 0.100 M Fe^{2+} solution;
- 10 mL or less of 6 M H_2SO_4 solution to be added in order to acidify the solution; and
- any of the equipment made available for the lab to collect your data.

You will then have 1 minute to present your description and reported concentration to the class.

Part B. After the presentation discussion, you are to complete the *first part* of the central challenge by designing a procedure in groups of two or three that will enable you to more accurately standardize the $KMnO_4$ solution. The procedures are to be teacher reviewed before being conducted. Any additional changes to the procedure also need to be teacher reviewed.

INVESTIGATION

Procedure

All of the $KMnO_4$ standardization testing is to be completed before you begin the *second part* of the central challenge: analyzing the H_2O_2 solutions. You may wish to use the procedure developed to standardize the $KMnO_4$ solution while designing the procedure to analyze the H_2O_2 solutions. Each group will be given two different H_2O_2 solutions to titrate. You should share your data with each other to obtain additional data for the various H_2O_2 solutions and to compare your results.

DATA COLLECTION AND COMPUTATION

1. Calculate the moles of $KMnO_4$ solution needed to react with all of the 0.100 M $Fe(NH_4)_2(SO_4)_2$ solution for each trial.
2. Calculate the molarity of the $KMnO_4$ solution for each trial.
3. Calculate the average molarity of the $KMnO_4$ solution.
4. Calculate the moles and mass of H_2O_2 titrated for each trial.

5. If the density of the H_2O_2 solution titrated was 1.00 g/mL, calculate the percentage of H_2O_2 in solution in each trial.
6. Calculate the average percentage of the H_2O_2 solution.

ARGUMENTATION AND DOCUMENTATION

1. Are your average values higher or lower than the reported values of the H_2O_2 you tested? What are the likely causes of any errors? Justify your explanations. Be sure to discuss any data values that are outliers.
2. Why and how did you modify your procedure and materials from Part A in the Practice section? How did these modifications impact your data and/or your calculations?
3. New dietary supplements do not undergo the same rigorous approval process as new medications. You performed a redox titration to determine the percentage of a component in a dietary supplement and in a medication. When you performed repeated trials, your percentages varied widely for the supplement but were very consistent for the medication. Why do you think this happened?

POSTLAB ASSESSMENT

1. What is a titrant? What is an analyte?
2. In this lab, did a substance serve as both a titrant and an analyte? If so, what was this substance? Support your answer.
3. What might have been the product(s) in the original solution if it had remained neutral (the solution was not acidified with H_2SO_4)? How could you determine this?
4. How would the concentrations of the $KMnO_4$ and the H_2O_2 solutions have been affected if the following observations were made about the Erlenmeyer flask? Justify your answers.
   a. When the titration was completed, the flask was colorless.
   b. When the titration was completed, the flask was dark red or purple.

SUPPLEMENTAL RESOURCES

Links

“Chemistry Tutorial: Redox.” Ausetute. Accessed July 29, 2012. *http://www.ausetute.com.au/redoxtitr.html*

Chieh, Chung. “Solutions Stoichiometry.” University of Waterloo. Accessed July 29, 2012. *http://www.science.uwaterloo.ca/~cchieh/cact/c120/sltnstoich.html*

References

Rees, Thomas. “The Stability of Potassium Permanganate Solutions.” *Journal of Chemical Education* 64, no. 12 (1987):1058.

Webb, Michael J. “Aqueous Hydrogen Peroxide: Its Household Uses and Concentration Units.” *Journal of Chemical Education* 62, no. 2 (1985):152.

Worley, John D. “Hydrogen Peroxide in Cleansing Antiseptics.” *Journal of Chemical Education* 60, no. 8 (1983): 678.

Young, J. A. “Hydrogen Peroxide, 3%.” *Journal of Chemical Education* 80, no.11, (2003): 1132.

Young, J. A. “Potassium Permanganate.” *Journal of Chemical Education* 80, no. 8 (2003): 873.

INVESTIGATION 9

Can the Individual Components of Quick Ache Relief Be Used to Resolve Consumer Complaints?

CENTRAL CHALLENGE

Akin Omore works for a large pharmaceutical company. Recently there have been complaints of chipped or broken pills in bottles of the company's popular Quick Ache Relief product. Quick Ache Relief is a generic pain relief medication that contains acetaminophen. One analytical team at the company, Team A, suspects that the broken pills are the result of insufficient amounts of sucrose. Sucrose is used as a binder to prevent the components of the pills from crumbling. Other examples of binders include starch, cellulose, and silica gel. A second team at the company, Team B, disagrees with this conclusion. This second team believes that while binder ratios appear to be inaccurate, the chipped residue is the result of a manufacturing problem in which tablets from a different product may have been retained in the packaging machinery. They believe that the chipped residue is most likely aspirin residue that was inadvertently carried over into the Quick Ache Relief packaging process. Akin would like your group to test a sample of the product and provide evidence supporting one of two hypotheses.

CONTEXT FOR THIS INVESTIGATION

Many drugs are synthesized in batches and are frequently made at different physical locations. Pharmaceutical companies assign a unique number to each batch of a drug that is produced. This number is referred to as a lot number and allows manufacturers to track any problems that may arise with a pharmaceutical product. The Food and Drug Administration (FDA) requires food and drug manufacturers to report products that pose a potential risk to the public. In November 2011, Novartis OTC issued a voluntary recall on certain lots of select sizes of Excedrin, NoDoz, Bufferin, and Gas-X. This recall was described as precautionary due to potential manufacturing errors that resulted in products containing "stray tablets, capsules, or caplets from other products."

Most commercially available pain relievers are mixtures that contain a number of different substances. To isolate the components of the tablets, you will need to take advantage of the differences in chemical and physical properties of each component. Physical properties can be observed without changing the identity or composition of a substance. Examples include solubility, boiling point, color, and density. Chemical properties, on the other hand, describe the way a substance may change, or react, to form other substances. For example, acidic or basic substances can be converted to water-soluble salts. This will allow separation from water-insoluble substances in a mixture.

The Quick Ache Relief label lists sucrose as the binder and acetaminophen as the active ingredient. In order to address Akin Omore's request, your team must separate the components of a sample of Quick Ache Relief.

The label on Quick Ache Relief lists the following percentages: 15% binder, 85% acetaminophen. For experimental purposes, ranges of 12–18% binder and 82–88% acetaminophen are sufficient to consider the label reasonably accurate. Based on your experimental results, calculate the percent by mass of each component found in your sample of Quick Ache Relief. Determine which research team (Team A or Team B) at the company obtained results consistent with your experimental findings.

Upon completion of the experiment, you will compose a report to Akin Omore at the pharmaceutical company. The report should clearly discuss your experimental results in the context of Akin Omore's original request. Your report needs to detail your methods, results, conclusions, and any sources of experimental error.

PRELAB GUIDING QUESTIONS/SIMULATIONS

Step 1: To facilitate safety and assist in designing an efficient procedure, your group should develop a table of chemicals that lists the physical properties of the chemicals involved in the experiment. For each compound, include the compound's name, molecular formula, molecular mass, and potential hazards. The websites *http://www.msdsonline.com/msds-search/* and *http://www.msds.com/* provide references to MSDS information needed to design the table. Your teacher will engage you in a whole-group discussion around the following guiding questions:

- What does each component have in common?
- What are their differences?
- What safety practices must be used when working with these substances?

Step 2: As part of the experimental design, your group will need to include a liquid-liquid extraction. This video clip *http://www.wonderhowto.com/how-to/video/how-to-do-a-liquid-liquid-extraction-in-the-chemistry-lab-259811/view/* shows students the basics of using a separatory funnel and how to perform liquid-liquid extractions. Be prepared to share your response to the following question with the whole class.

- Would you be able to separate the components using two liquids in the same container?

Step 3: This video clip *http://www.wonderhowto.com/how-to/video/how-to-dry-an-organic-solution-in-the-chemistry-lab-259802/view/* shows students the basics of using an anhydrous reagent to remove water from an organic solution. View the video clip and be prepared to share your response to the following question with the whole class.

- Is there a way to remove water from a liquid solvent?

Step 4: Answer the following questions and be prepared to share your answers in a whole-group discussion:

a. Explain the difference(s) between intramolecular and intermolecular forces.

b. Provide two examples of evidence of a chemical reaction.

c. Describe the physical property that will help you determine which layer in the separatory funnel is the aqueous layer (containing sodium hydrogen carbonate) and which is the organic layer (containing ethyl acetate).

d. A student was given a 5.00 gram sample of a mixture containing three substances, A, B, and C. Using the physical and chemical properties of each component, the student recovered 2.51 grams of substance A, 1.36 grams of substance B, and 0.98 grams of substance C. Calculate the percent composition of A, B, and C. Calculate the percent recovery.

PREPARATION

Materials

Quick Ache Relief, ~3 g	Ethyl acetate, ($CH_3COOC_2H_5$) 50 mL	10% Sodium hydrogen carbonate ($NaHCO_3$) 50mL	6 *M* hydrochloric acid (HCl) 15 mL	Anhydrous magnesium sulfate ($MgSO_4$) ~5 g
100 mL glass beakers	125 mL Erlenmeyer flasks	pH paper	Separatory funnels	Buchner funnel
Glass funnel	Side arm flasks	Vacuum filtration apparatus	Filter paper, circular to fit Buchner funnel and glass funnel	Graduated cylinders
Digital balances	Glass stirring rod	Watch glass	Hot plate	Drying oven, optional

Safety and Disposal

The acid and base solutions require careful handling as they can cause skin burns and eye damage. Ethyl acetate is flammable. You should wear safety goggles at all times and use extreme caution when heating. Chemical waste should be disposed of properly as per teacher's instructions. Liquid waste should be placed in an appropriate container.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

You should practice gravimetric analysis and separation techniques before beginning the guided inquiry of the lab. Liquid-liquid separation techniques can be practiced using an organic solvent such as hexane and water in order to visualize different layers of clear liquids. Your teacher will provide further instruction if required.

INVESTIGATION

Procedure

You should work in pairs to investigate the properties of each of the components and design an experimental procedure to separate the components of Quick Ache Relief. Using the reagents provided and information from the relevant MSDS, experimentally determine the physical and chemical characteristics of sucrose, acetaminophen, and aspirin (acetylsalicylic acid), such as solubility and reactivity.

Using the information gained in the prelab, design an experiment to separate each component of Quick Ache Relief. Outline the stages involved in the experiment. For each stage, relate what's happening chemically to what is happening experimentally. Your method should be complete enough to allow you to isolate each component in your sample of Quick Ache Relief. Include proper chemical names, sequencing of steps, times, amounts, required precision, concentrations of solutions, etc. Flowcharts are recommended, or this section can be presented in two columns.

Experimental Steps: The left column should identify the step-by-step actions of the experiment.	What's Happening: Describe what is happening chemically in each step of the experiment.

Data Collection and Computation

1. Calculate the percent by mass of each component by dividing the sum of the masses of all recovered components by the mass of the Quick Ache Relief that you started with multiplied by 100.
2. Calculate the percent recovery of Quick Ache Relief based on the total mass of the components recovered multiplied by 100.

Argumentation and Documentation

Upon completion of the experiment, you and your partner should generate individual reports to Akin Omore at the pharmaceutical company that:

1. Clearly state the goal(s) of the experiment;
2. Discuss the outcome of the experiment and whether the goal(s) were achieved;

3. Provide experimental evidence that support your conclusions;
4. Explain any potential sources of experimental error;
5. Suggest methods for additional analysis; and
6. Discuss within your team what is necessary to include in a report to the company and create a rubric to be used in grading reports.

POSTLAB ASSESSMENT

You should generate a series of statements describing the particulate level changes that occurred during each stage of separation. A summary paragraph should outline why each stage was necessary.

SUPPLEMENTAL RESOURCE

Reference

Lehman, John W. *Operational Organic Chemistry: A Problem-Solving Approach to the Laboratory Course*. Upper Saddle River, NJ: Pearson-Prentice Hall, 2009.

INVESTIGATION 10

How Long Will That Marble Statue Last?

CENTRAL CHALLENGE

Your task is to investigate how the speed of the chemical reaction between solid calcium carbonate and a solution of hydrochloric acid is affected by changing variables relating to the two reactants. You will implement the investigation by first constructing and then testing a hypothesis.

CONTEXT FOR THIS INVESTIGATION

Many historic buildings and monuments are made from limestone or marble. Limestone and marble are minerals that contain large amounts of calcium carbonate, $CaCO_3$. Since the industrial revolution, air pollutants, chiefly in the form of oxides of sulfur and nitrogen, have been absorbed into the atmosphere leading to the production of rainwater that has become significantly more acidic. This *acid rain* will react with the limestone, eroding it and causing much disfigurement and damage.

PRELAB GUIDING QUESTIONS/SIMULATIONS

The prelab guiding questions are designed to be completed as you think about the central challenge and the context of the investigation. They will help ensure that you are familiar with the vocabulary, and will help you prepare to create your own procedure and to perform the necessary calculations.

1. List the factors that you think may affect the speed at which calcium carbonate will react with acid. (The hydrochloric acid will be available to you in varying concentrations, and the calcium carbonate will be available in "chunks" of varying size. You will have access to a heat source.)
2. If a chemical reaction produces a gas, suggest a way of monitoring the production of that gas as the reaction proceeds.
3. Discuss how increasing the surface area of a solid might influence the rate of a chemical reaction.

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

The rate of a chemical reaction is explained by collision theory. Collision theory can be condensed into three large ideas that apply to the rate of all chemical reactions in which two or more molecules, atoms, or ions come together.

1. When substances are brought together in chemical reactions, the particles that make up those substances collide with one another. The collisions that occur between the particles are the first criteria for a reaction to take place. In potential reactions in which two or more molecules, atoms, or ions come together without those collisions, no actual reaction is possible.
2. The collisions that take place have to occur with a certain minimum energy. That minimum energy is called the activation energy. If the particles collide with insufficient energy — i.e., with an energy less than the activation energy — no reaction occurs and the particles simply bounce off one another without producing any products. Collisions that do not result in a reaction are called "unsuccessful."
3. The collisions must occur with the correct orientation — i.e., the particles must come together in a certain specific, physical way in which the atoms "line up" with one another, allowing a reaction to occur.

For a collision to be "successful"— i.e., for a collision to result in a chemical reaction in which reactants turn into products — the particles of the reactants must fulfill each of the three conditions of collision theory.

PREPARATION

Materials

Specific materials will depend upon the final experimental procedure determined by each student group; the likely materials for each experiment run by each group are listed below.

Marble or limestone chips (limestone/calcium carbonate), (approx. 2 large, 4 medium, or 8 small chips = 1.20–1.40 g)	Stoppers	Digital balance, with at least ±0.001g precision (as appropriate)
50.0 mL of hydrochloric acid (HCl) of varying concentration, for example , 6.0 *M*, 5.0 *M*, 4.0 *M*, 3.0 *M*, 2.0 *M*	Syringes	Glass delivery tubes and rubber tubing

Water (preferably deionized or distilled, but this isn't crucial) for dilution of a concentrated stock solution of HCl (alternatively prediluted HCl)	As an alternative to gas syringes, equipment suitable for the collection of gas over water (i.e., a eudiometer or gas jar)	Stopwatches, or students can use an online app for timing purposes
125 mL Erlenmeyer flasks	Labquest and probeware (gas pressure probes) as an alternative way of collecting data	

Safety and Disposal

Safety: 6.0 *M* hydrochloric acid is hazardous and needs to be handled with care. Goggles and protective aprons are required. You should not generate gases in closed containers since there is a risk of explosion as pressure builds. When diluting acids, acid should be added to water, not water to acid.

Disposal: Any excess solid calcium carbonate should be removed from solution and disposed of in the trash. All solutions can be diluted and washed down the sink with plenty of water.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

Your teacher will guide you through the process of identifying variables that may affect the rate of reaction of hydrochloric acid with calcium carbonate, which is meant to simulate the effects of acid rain on marble statues. Upon identification of variables, you will need to write at least two research questions pertaining to this reaction using the following format:

To what extent does A affect B?

(where A is the independent variable and B is the dependant variable)

Procedure

Your teacher may ask you to mix varying amounts of hydrochloric acid solutions and marble chips so you can get a sense of the relative speed of the reactions, and perhaps get some ideas about what you might be able to measure and observe. It might be useful to have the reaction vessel placed on an electronic balance as you make these initial observations.

INVESTIGATION

Once students develop one to two research questions (from the Practice section) to gather data for, students should design experiments for each of the questions to collect data that will allow them to draw conclusions about the effect of those variables on the rate of this reaction. The chemical reaction that will be studied is the reaction of hydrochloric acid with calcium carbonate, summarized by the equation below.

$$2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Procedure

1. Design an experimental procedure and data-collection strategy to monitor some aspect of the reaction (according to the two research questions you developed in the Practice section) in order to find a rate of reaction.
2. Formulate a hypothesis for how the independent variable will affect the dependant variable in each of your research questions.
3. Show your teacher your proposed procedure before attempting your experiment. Once your initial experimental design has been carried out and data collected, consider improvements to your procedure and perform the investigation once more, but show your procedure with changes to your teacher before trying it.

Data Collection and Computation

1. List the measurements that you took during the experiments.
2. Consider how the data or measurements should be recorded and presented to illustrate your findings and then place such data representation in a central location in the lab for the purpose of engaging in whole-group discussion about the variables that affect the rate of reaction between hydrochloric acid and calcium carbonate.
3. Use the data you have collected to comment on your original hypothesis.

Argumentation and Documentation

During whole-group discussion, using the data the students displayed in a central location of the lab, listen to each lab group present their findings and comment on whether or not their data supported or refuted their hypotheses.

Be prepared to defend the changes you made to your experimental design to ensure consistency in the data collection.

Be prepared to defend why you chose a particular treatment of the data to illustrate your findings to the class and to support or refute your hypotheses.

Your teacher will give you additional information about acid rain and marble statues. Be prepared to generalize the findings of your experiment to estimate how long it would take marble statues to disintegrate under specific conditions.

POSTLAB ASSESSMENT

Consider the following question.

An experiment was carried out in order to investigate the rate of reaction between magnesium and dilute hydrochloric acid. 0.07 g of magnesium ribbon was reacted with excess dilute acid. The volume of gas produced every 5.00 seconds was recorded.

Seconds	Volume of Gas Collected (mL)
0	0
5	18
10	34
15	47
20	57
25	63
30	67
35	69
40	70
45	70

a. Plot a graph of these results.

b. When is the reaction fastest? How can you use the graph to determine when the reaction is the fastest?

c. How long does it take for the 0.07 g of magnesium to react completely with the dilute hydrochloric acid?

d. Sketch another curve on to your graph that might have been obtained if 0.07 g of magnesium powder had been used instead of magnesium ribbon.

e. Suggest two other factors that would alter the rate of this reaction.

f. Write a chemical reaction for this process.

SUPPLEMENTAL RESOURCES

Links

Blauch, David N. “The Bromate-Bromide Reaction.” Davidson College. Accessed July 30, 2012.
http://www.chm.davidson.edu/vce/kinetics/BromateBromideReaction.html

“Investigating the Rate of a Reaction.” Creative Chemistry. Accessed July 30, 2012.
http://www.creative-chemistry.org.uk/gcse/documents/Sc1/rates.pdf

“Reactions and Rates.” University of Colorado at Boulder, PhET Interactive Simulations. Accessed July 30, 2012.
http://phet.colorado.edu/en/simulation/reactions-and-rates

References

Copper, Christine L., and Edward Koubeck. "An Experiment to Demonstrate How a Catalyst Affects the Rate of a Reaction." *Journal of Chemical Education* 76, no. 12 (1999): 1714-15. Accessed July 30, 2012. *http://pubs.acs.org/doi/abs/10.1021/ed076p1714?prevSearch=Rates%2Bof%2BReaction&searchHistoryKey=*

Eliason, Robert, and Terence McMahon. "Temperature Effect on Reaction Rates." *Journal of Chemical Education* 58, no. 4 (1981). Accessed July 30, 2012. *http://pubs.acs.org/doi/abs/10.1021/ed058p354.1?prevSearch=Rates%2Bof%2BReaction&searchHistoryKey=*

Evenson, Andy. "Putting Reaction Rates and Collision Theory in the Hands of Your Students." *Journal of Chemical Education* 79, no. 7 (2002): 822–23. Accessed July 30, 2012. *http://pubs.acs.org/doi/abs/10.1021/ed079p822?prevSearch=Rates%2Bof%2BReaction&searchHistoryKey=*

Mickey, Charles D. "Chemical Kinetics: Reaction Rates." *Journal of Chemical Education* 57, no. 9 (1980): 659. Accessed July 30, 2012. *http://pubs.acs.org/doi/abs/10.1021/ed057p659?prevSearch=Rates%2Bof%2BReaction&searchHistoryKey=*

INVESTIGATION 11

What Is the Rate Law of the Fading of Crystal Violet Using Beer's Law?

CENTRAL CHALLENGE

The purpose of this laboratory activity is to determine the rate law for the reaction of crystal violet (CV) and sodium hydroxide (NaOH).

In Part 1 of the investigation, you will prepare dilutions of a stock CV solution to generate a Beer's law calibration curve for CV using similar techniques found in Investigation 2.

In Part 2 of the investigation, you will perform a reaction of CV with NaOH while monitoring in real time the concentration of CV remaining.

This laboratory investigation will illustrate a variety of science concepts because determining the rate law for the reaction of CV with NaOH requires you to use graphical analysis and a simplifying approximation that leads to a pseudo-rate law while also integrating your prior chemistry knowledge involving spectroscopy, Beer's law, solution dilution, calibration curves, and chemical kinetics.

CONTEXT FOR THIS INVESTIGATION

If you're making something, you might think making it to last would always be a good thing. But what if you're making a pesticide with known detrimental impacts on human health? Then you may only want it to stay intact for a few days after it has been applied to crops before it decomposes into what often are less harmful products. If its molecules stay intact for too long, the pesticide can persist in the environment and build up in drinking water.

In the year 2000, over 20 million kilograms of the pesticide 1,3-dichloropropene (1,3-D) were applied to crops in the United States. Scientists investigated the rate of decomposition of 1,3-D in acidic, basic, and neutral solutions as well as in soil. For each case they generated plots of the amount of intact 1,3-D persisting versus time and found that the reaction could be characterized as pseudo first-order. Knowing the order of the reaction allowed them to determine the half-life of intact 1,3-D. In acidic media, they found that the half-life for the decomposition of 1,3-D was about eight days but in the presence of excess NaOH the half-life was reduced to about four days. Experimentally determined data like this is vital to the ability of society to use chemicals wisely in improving food production, while not endangering the end consumers or the people who work with the chemicals during the growing process.

The Beer's law lab earlier in the manual (Investigation 2) employed the use of a colorimeter (or spectrophotometer) to obtain a calibration curve that was used to convert raw absorption data from a colorimeter (or spectrophotometer) to molar concentration of a chemical in solution. In the current investigation, you will first use a colorimeter (or spectrophotometer) to generate a calibration curve for a chemical (CV) and then use the colorimeter (or spectrophotometer) to follow the change in the concentration of CV as it reacts with NaOH. By recording these changes through time and analyzing them graphically, you will be able to obtain the rate law of the reaction, which may be used to predict the behavior of the system under *different* experimental conditions without doing the actual experiments.

PRELAB GUIDING QUESTIONS/SIMULATIONS

These questions are designed to guide, but not dictate, your development of the appropriate experimental protocol. By answering them, you will develop a procedure to accomplish the goal of determining the rate law for the reaction between crystal violet and sodium hydroxide. Questions 1 and 2 relate to the Beer's law calibration curve from Part 1 of the procedure while the remaining questions relate to Part 2, the reaction of CV with NaOH. Though some questions rely on information found in the following Explanation to Strengthen Student Understanding section and in your chemistry textbook, you will also need to employ your critical reasoning skills to address many of the questions. A simulation (Matsumoto 2012) may be useful in addressing Questions 1, 4, and 8.

1. Answer the following questions about the selection of a wavelength for your experiment.
   a. Based on the absorption spectrum of 25 μM crystal violet in Figure 1 and taking into account the considerations that follow, what wavelength should you use for the Beer's law calibration curve and subsequent reaction of CV with NaOH? Please explain your answer.

      Note that if you are using a colorimeter, you only have a few possible wavelengths to choose from, but if you are using a spectrophotometer, then there are many wavelengths available. Two main factors to consider, working in opposition to each other, are sensitivity and range. These instruments are typically not sensitive enough to reliably measure absorbance values much above 1.0 absorbance units. Yet the absorbance at the chosen wavelength should be high enough that it can vary over a wide range of values during the reaction.

   b. Simulate the instrument readings you will get in Part 1 of the experiment by doing the following: Trace Figure 1 onto your own paper. Draw a vertical line at the wavelength you have chosen, intersecting the absorbance curve at that wavelength. Where your vertical line intersects the absorbance curve is the absorbance value your instrument should read for the stock 25 μM CV solution. Keeping in mind Beer's law from Equation 1, and being mindful that the wavelength and path length are fixed, draw X's on your vertical line where you expect the absorbance values will be for the diluted solutions you prepare in Question 2. Use appropriate ratios of concentrations to determine where on the vertical line to make your marks.

2. A calibration curve requires the preparation of a set of known concentrations of CV, which are usually prepared by diluting a stock solution whose concentration is known. Describe how to prepare 10. mL of a 5, 10, 15, and 20 μ*M* CV solution using a 25 μ*M* CV stock solution.
3. During the reaction of CV with NaOH, do you expect the colorimeter's (or spectrophotometer's) absorbance reading to change? How do you expect it to change if such a change is anticipated (i.e., increase, decrease or no change) as the reaction proceeds? Explain your reasoning.
4. Answer the following questions for a reaction of CV with NaOH in these two scenarios: a solution with a 1:1 NaOH:CV mole ratio and a solution similar to what you will be using with a 1000:1 NaOH:CV mole ratio.
   a. Using your prior knowledge of reaction stoichiometry, what is the final percentage of each reactant remaining if each reaction went to completion? Show work and reasoning to justify your answer.
   b. Based on this result, describe how one gets Equations 3 and 4 from Equation 2 above.
5. Using the kinetics chapter in your textbook and websites like "Chemical Kinetics — Integrated Rate Laws" *http://www.chm.davidson.edu/vce/kinetics/IntegratedRateLaws.html*, describe the graphical analysis that can be done to determine the order (considering only 0th, 1st, or 2nd order) and the value of the pseudo-rate constant, k*, of a chemical reaction from concentration data collected through time.
6. Based on your answer to questions 3–5, design an experiment for the reaction of CV with NaOH and describe the subsequent data analysis to accomplish the central challenge, the determination of the value of (i) w, the order with respect to CV and (ii) k*, the pseudo-rate constant found in the rate law in Equation 3. For simplicity, use 10. mL for the combined volume of CV and NaOH because it is a bit more than enough to fill cuvettes appropriately.
7. As an optional extension, design an experiment to determine from Equations 2 and 4 (i) z, the order with respect to NaOH and (ii) k, the rate constant.
8. Answer the following questions after examining Figure 3 to address the issue of when to stop collecting data.
   a. For early parts of the three different reactions in Figure 3, all three curves seem relatively linear with different slopes. But as the reactions progress through time, at roughly what concentration level would you say some graphs start to look nonlinear?
   b. Given that you don't yet know the order of the reaction of CV with NaOH, how might Figure 3 help you to decide when to stop collecting data? Hint: think in terms of % completion instead of concentration.

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

As you learned in Investigation 2 earlier in the manual, Beer's law is represented by Equation 1, provided here.

$$\text{Absorbance} = \text{absorptivity constant} \times \text{path length} \times \text{concentration}$$

Equation 1

For a fixed concentration of solute and a fixed path length (e.g., fixed cuvette width), the amount of light absorbed by a solution varies directly with the absorptivity constant of the solute. Figure 1 below shows the visible light absorbance spectrum of CV for a fixed, 25 µM, concentration of CV and a fixed, 1.0 cm, path length. Because concentration and path length are both kept constant, Figure 1 reveals how the absorptivity constant for CV varies with the wavelength of light passing through the solution. Figure 1 was generated by a spectrophotometer. A colorimeter is an instrument that, like a spectrophotometer, measures how much light is absorbed when passed through a sample but does so for only a few predetermined wavelengths of light set by the manufacturer.

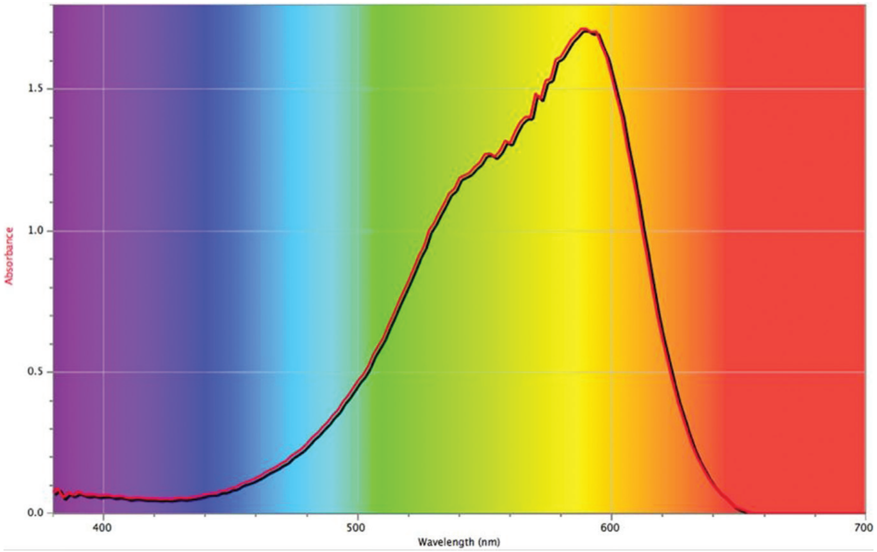

Figure 1. The visible spectrum of a 25 µM CV solution

If we still keep the path length fixed, but now choose only one particular wavelength of light to pass through the solution, thereby fixing the absorptivity constant, then we can observe how the absorbance of light at that wavelength changes as we change the concentration of CV. Under these conditions, Beer's law describes a straight-line relationship for a graph of absorbance versus solute concentration whose slope is simply the product of the molar absorptivity constant and path length.

In the reaction of CV and sodium hydroxide (see Figure 2), the dye's color will fade as it reacts with sodium hydroxide. A colorimeter (or spectrophotometer) will be used to follow the disappearance through time of CV by measuring the absorbance of a solution of CV during its reaction with NaOH. The raw absorbance measurements from the colorimeter (or spectrophotometer) can be transformed to molar concentration of CV via the use of a Beer's law calibration curve.

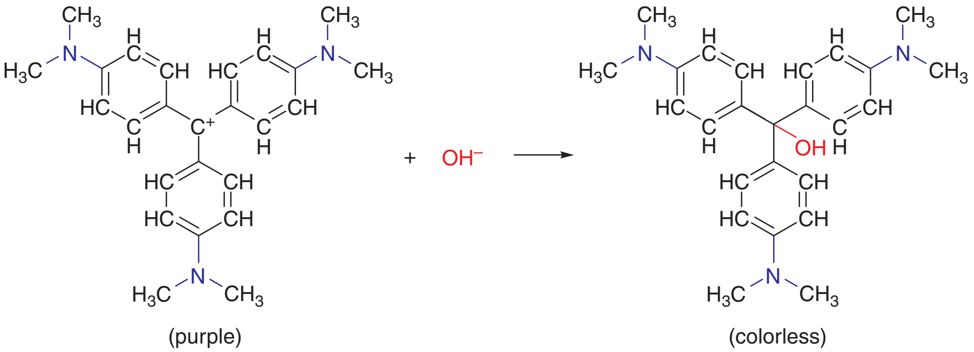

Figure 2. Chemical structures in the reaction in this laboratory activity

The net ionic equation for the reaction can be written as

$$CV^{+}\,(aq) + OH^{-}(aq) \rightarrow CVOH\,(aq)$$

The rate law of the reaction is

$$\text{rate} = k\,[CV^{+}]^{w}\,[OH^{-}]^{z}$$

Equation 2

where k is the rate constant while w and z are the order of the reaction in relation to CV^{+} and OH^{-}, respectively.

Under certain experimental conditions (see prelab Question 4), the rate law in Equation 2 simplifies to the following equation

$$\text{rate} = k^{\star}\,[CV^{+}]^{w}$$

Equation 3

where

$$k^{*} = k\,[OH^{-}]^{z}$$

Equation 4

and k* is the pseudo-rate constant. Equation 3 is referred to as the pseudo-rate law, since it is an approximation of Equation 2, the actual rate law, and significantly simplifies the analysis.

A differential rate law describes the rate of a chemical reaction as a function of the concentration of the reactants, while an integrated rate law describes the concentration of a reactant as a function of time; both types of rate laws are related to each other by the use of calculus. Equation 3 is a differential rate law, by which a graphical analysis of the corresponding integrated rate law can be used to determine the value of the parameters in Equation 3, using least-squares linear regression analysis. The degree or extent of linear fit may be evaluated using the coefficient of determination (or square of the correlation coefficient), i.e., it may be used to identify the graph that has a linear relationship.

Figure 3 shows concentration data plotted versus time for three *different* hypothetical chemical reactions. From plots like these and knowledge of integrated rate laws found in your text, one can determine the exponents in the rate law equation.

	A		B	C
1	initial [reactant] =		10	
2	rate constant =		0.04	
3				
4	Rx order =	0	1	2
5	time	[reactant]	[reactant]	[reactant]
6	0	10	10	10
7	5	9.8	8.187308	3.333333
8	10	9.6	6.7032	2
9	15	9.4	5.488116	1.428571
10	20	9.2	4.49329	1.111111
11	25	9	3.678794	0.909091
12	30	8.8	3.011942	0.769231
13	35	8.6	2.46597	0.666667
14	40	8.4	2.018965	0.588235
15	45	8.2	1.652989	0.526316
16	50	8	1.353353	0.47619
17	55	7.8	1.108032	0.434783
18	60	7.6	0.90718	0.4
19	65	7.4	0.742736	0.37037
20	70	7.2	0.608101	0.344828
21	75	7	0.497871	0.322581
22	80	6.8	0.407622	0.30303
23	85	6.6	0.333733	0.285714
24	90	6.4	0.273237	0.27027
25	95	6.2	0.223708	0.25641
26	100	6	0.183156	0.243902

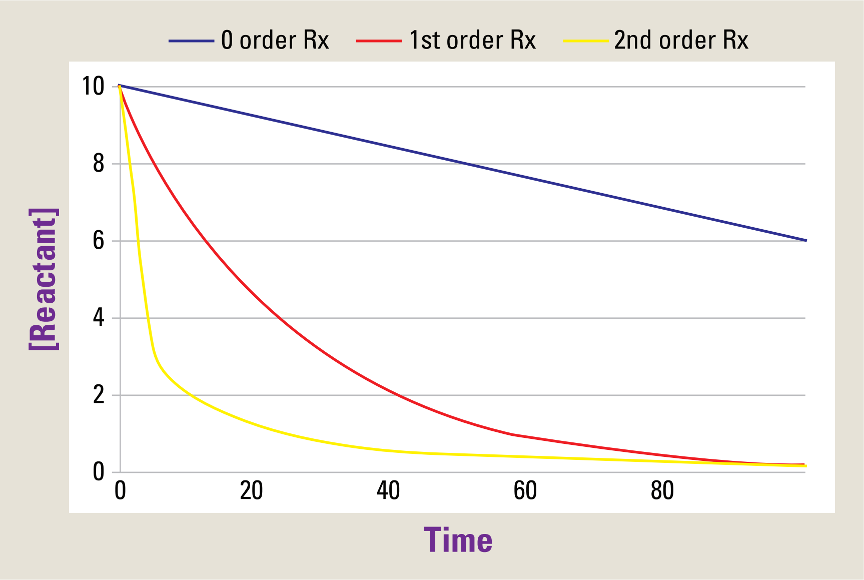

Figure 3. Concentration data plotted versus time for three *different* hypothetical chemical reactions — 0th order (blue line), 1st order (red line), and 2nd order (yellow line)

All reactions have the same numerical value for their initial reactant concentration and the rate constant.

PREPARATION

Materials

25.0 μ*M* crystal violet (CV) solution	Colorimeter or spectrophotometer
0.200 *M* sodium hydroxide (NaOH) solution	Data display device (probeware or computer software)
Distilled water	5–10 mL graduated pipettes (use separate pipettes for CV, NaOH, and water)
Cuvettes	Pipette bulb or pipette pump
	Stopwatch (if a computer won't be used)

Safety and Disposal

Crystal violet is a dye that stains skin and clothing. A sodium hydroxide solution is a corrosive skin irritant. You must wear an apron and splash-proof goggles during the lab and use disposable gloves in handling the aqueous solutions of crystal violet and sodium hydroxide. If any of the solutions spilled on your skin, wash with copious amounts of water. Glassware stained by CV may be cleaned with 70% isopropyl alcohol (rubbing alcohol). Solutions may be rinsed down the drain after the appropriate treatment — i.e., mix any remaining solutions of crystal violet and sodium hydroxide, add any additional sodium hydroxide if needed to bleach the crystal violet, neutralize the sodium hydroxide with an acid, then discard the entire solution.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

To convert a colorimeter's (or spectrophotometer's) raw absorbance values to the molar concentration of a chemical, you must determine how absorbance varies with concentration for a set of solutions of known concentration. A previous experiment in this lab manual (Investigation 2) had you do this using a Beer's law calibration curve.

Procedure: The collection of experimental data to generate a calibration curve

Use your answer to prelab Question 1 to select a wavelength that you will use in Part 1 to generate a Beer's law calibration curve relating raw absorbance values to CV concentration and in Part 2 to track absorbance during the reaction of CV with NaOH.

Use your answer to prelab Question 2 to guide your preparation of solutions that can be used to correlate absorbance with CV concentration by generating a Beer's law calibration curve. Make sure to use an appropriate blank to zero out your colorimeter (or spectrophotometer) before taking absorbance measurements.

INVESTIGATION

Here you will address the central challenge, to determine the rate law for the reaction of CV with NaOH.

Procedure

The collection of experimental data of absorption as a function of time for the reaction of CV with NaOH

The prelab questions guide you to consider different factors involved in designing an experiment to address the central challenge, the determination of the rate law given in Equation 2, rate = k $[CV^+]^w$ $[OH^-]^z$, for the reaction between CV and NaOH.

Prelab Question 4 addresses the issue of how to ensure that Equation 3 (rate = k* $[CV^+]^w$) is a valid approximation to Equation 2.

The analysis described in prelab Question 6 requires that you know the concentration of CV throughout the course of the reaction. The concentration of CV can be obtained from raw absorbance data by applying the Beer's law calibration curve formula you obtained during the Practice.

Prelab Question 6 asks you to design an experiment to determine the value of w and k* found in Equations 3 and 4. Both w and k* can be determined by making appropriate plots of your data from the reaction of CV with NaOH and checking for linear relationships (Blauch 2010; Zumdahl & Zumdahl 2003). Use your answer to prelab Question 8 to decide during the experiment when to stop collecting absorbance data to get the clearest distinction between the 0th, 1st, and 2nd order reactions during your postlab graphical analysis.

Optional prelab Question 7 asks you to describe how one could determine the value of z and k found in Equation 2. Check with your instructor to see a) if you are doing this part of the investigation and b) if yes, whether or not you have a method that could address that goal.

Data Collection and Computation

The successful completion of this laboratory investigation involves the planning and implementation of appropriate data-collection strategies. This requires a clear understanding of the central challenge of this investigation, how it would be accomplished, and the appropriate calculations (e.g., dilution of CV) to be done *prior* to making the various experimental measurements.

The successful completion of this laboratory investigation involves the appropriate use of mathematics to analyze your experimental data to obtain the rate law of a chemical reaction. The prelab questions are designed to guide you in your analysis of the data. You could use common data-analysis software to conduct the appropriate graphical analysis and any data transformations prior to your graphical analysis.

Argumentation and Documentation

You will use guidance from your teacher as well as the sections above and discussions within your group to justify the selection of your experimental protocol and the subsequent data analysis to determine the rate law of the chemical reaction in this investigation. Justification of your experimental design and data-analysis method should appear in your final lab report.

POSTLAB ASSESSMENT

Prepare and submit a lab report according to your instructor's guidelines, which will include asking you to provide appropriate graphs and data analysis to support your answer to the central challenge. You should clearly support with evidence the values you obtained for w and k*. If you performed the optional extension to determine z and k, make sure you follow the same guidelines for the write-up of that experiment as well.

SUPPLEMENTAL RESOURCES

Links

Blauch, David N. "Chemical Kinetics — Integrated Rate Laws." Davidson College Chemistry Resources. Accessed January 6, 2012.
http://www.chm.davidson.edu/vce/kinetics/IntegratedRateLaws.html

"Inquiry and the National Science Education Standards: A Guide for Teaching and Learning." The National Academies Press. Accessed August 3, 2012.
https://www.nap.edu/catalog/9596/inquiry-and-the-national-science-education-standards-a-guide-for

Matsumoto, P. "CV Prelab." Mr. Matsumoto's Website. Accessed August 3, 2012.
http://chem-is-try.us/class/ap/labs/kinetics.htm
(The mathematical model is stored on that web page as a cdf file in the resource section at the bottom of the web page and requires the free Wolfram cdf player plugin
https://www.wolfram.com/products/player/)

Mingxin, Guo, Sharon K. Papiernik, Wei Zheng, and Scott R. Yates. "Effects of Environmental Factors on 1,3-Dichloropropene Hydrolysis in Water and Soil." *Journal of Environmental Quality* 33 (2004): 612–618. Accessed February 26, 2012.
https://pubag.nal.usda.gov/download/9114/PDF

"R-Squared or Coefficient of Determination." Khan Academy. Accessed January 6, 2012.
http://www.khanacademy.org/video/r-squared-or-coefficient-of-determination?playlist=Statistics

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Corsaro, Gerald. “A Colorimetric Chemical Kinetics Experiment.” *Journal of Chemical Education* 41, no. 1 (1964): 48–50.

Glantz, Stanton A. 1997. “Ch. 8. How to Test for Trends” in *Primer of Biostatistics*, 4th ed. McGraw-Hill. New York.

Matsumoto, Paul. “The Importance and Efficacy of Using Statistics in the High School Chemistry Laboratory.” *Journal of Chemical Education* 83, no. 11 (2006): 1649–1651.

Mingxin, Guo, Sharon K. Papiernik, Wei Zheng, and Scott R. Yates, “Effects of Environmental Factors on 1,3-Dichloropropene Hydrolysis in Water and Soil.” *Journal of Environmental Quality* 33 (2004): 612–618.

Thompson, Jonathan E., and Jason Ting. “A Simple, Inexpensive Water-Jacketed Cuvette for the Spectronic 20.” *Journal of Chemical Education* 81, no. 9 (2004): 1341–1343.

Zumdahl, Steven S. and Zumdahl, Susan A. 2003. “Ch. 12. Chemical Kinetics” in *Chemistry*, 6th ed. Houghton Mifflin. Boston, MA.

INVESTIGATION 12

The Hand Warmer Design Challenge: Where Does the Heat Come From?

CENTRAL CHALLENGE

The ideal hand warmer increases in temperature by 20°C (but no more) as quickly as possible, has a volume of about 50 mL, costs as little as possible to make, and uses chemicals that are as safe and environmentally friendly as possible. You will carry out an experiment to determine which substances, in what amounts, to use in order to make a hand warmer that meets these criteria.

CONTEXT FOR THIS INVESTIGATION

Have your fingers ever been so cold they felt numb? Wouldn't it be great if you could generate heat to warm your hands up anytime you want to? That's exactly what a "hand warmer" does. Hand warmers are small packets that people put inside gloves or mittens on cold days to keep their fingers warm. They are very popular with people who work outside in winter or do winter sports. One type of hand warmer contains water in one section of the packet and a soluble substance in another section. When the packet is squeezed the water and the soluble substance are mixed, the solid dissolves, and the packet becomes warm. In this experiment you will learn how a hand warmer works and use chemistry to design an effective, safe, environmentally benign, and inexpensive hand warmer.

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

Breaking bonds and particulate attractions absorbs energy from the surroundings, while forming new bonds and particulate attractions releases energy to the surroundings. When an ionic solid dissolves in water, ionic bonds between cations and anions in the ionic solid and hydrogen bonds between water molecules are broken, and new attractions between water molecules and anions and water molecules and cations are formed. The amount of energy required to break these bonds and form new ones depends on the chemical properties of the particular anions and cations. Therefore, when some ionic solids dissolve, more energy is required to break the cation-anion bonds than is released in forming the new water-ion attractions, and the overall process absorbs energy in the form of heat. When other ionic compounds dissolve, the converse is true, and the bond making releases more energy than the bond breaking absorbs, and therefore the process overall releases heat. When heat is absorbed, the enthalpy change, *q*, is endothermic, and the enthalpy change is positive. When heat is released, the change is exothermic, and the value of *q* is negative. The entropy change of solution formation is generally positive, regardless of whether it is endothermic or exothermic, because solutions are much more disordered than are the pure solute and solvent from which they are made. This positive entropy change is thermodynamically favorable.

PREPARATION

Materials

Thermometer	Sodium chloride (NaCl)
Balance	Anhydrous calcium chloride ($CaCl_2$)
100 mL graduated cylinder	Anhydrous sodium carbonate *Na_2CO_3)
Scoops	2–8 oz polystyrene cups with wooden cover or lid
Dixie cups or weighing boats	Anhydrous sodium acetate ($NaC_2H_3O_2$)
Magnetic stirrer with stir bars or stirring rod	Lithium chloride (LiCl)
Magnesium sulfate ($MgSO_4$)	Ammonium nitrate (NH_4NO_3)

Safety and Disposal

The solids and resulting solutions in this investigation are potential eye and skin irritants. Calcium chloride can cause skin burns. Ammonium nitrate is a powerful oxidizer that must be kept away from ignition sources and is quite toxic on ingestion. You must wear gloves and splash-proof safety goggles and must use caution throughout the experiment and adhere to all safety guidelines. If solutions are spilled on skin, wash with copious amounts of water. Follow your teacher’s directions for the disposal of solutions and any unused chemicals.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

In this experiment, you will collect data that will allow you to calculate the change of enthalpy of dissolution (also called the “heat of solution,” with symbol ΔH_{soln}, and units of kJ/mol solute) occurring in aqueous solution. The data necessary to calculate the heat of solution can be obtained using a device called a calorimeter.

A calorimeter is a container used to determine the enthalpy change that occurs during a process. Calorimetry is an important technique in chemistry, and chemists often work with devices called bomb calorimeters. For home or classroom experiments, however, a coffee cup calorimeter is sufficient to make rough measurements. This exercise will give you practice assembling and using a calorimeter so that you can use one to help you determine which solid is best to use in a hand warmer. It will also allow you to calibrate your calorimeter with a process that supplies a known amount of heat. This calibration process allows you to determine the amount of heat the calorimeter itself absorbs as the temperature of the materials inside it change, a value known as a calorimeter constant.

Procedure

Part 1 (calorimetry practice)

Assemble your calorimeter on a magnetic stirrer as shown in Figure 1. Measure out exactly 100.0 mL water in a graduated cylinder and pour into your calorimeter. Measure and record the temperature of the water. Add a stir bar and turn on the stirrer so the water is stirring without splashing. Measure 5.00 g magnesium sulfate anhydrate solid in a plastic cup. While monitoring the temperature of the water, quickly add all of the magnesium sulfate to the calorimeter. Record the highest temperature reached. Dilute the resulting solution with water and dispose of it per your teacher's instructions. Repeat the experiment. For each trial find the temperature change of the water per gram magnesium sulfate, and average this value for the two trials. Check with your instructor to make sure that your average temperature change is within 10 percent of the expected value. If it is not, repeat the experiment, measuring and monitoring the temperature carefully and ensuring that all of the solid is dissolved, and recheck your results with your instructor.

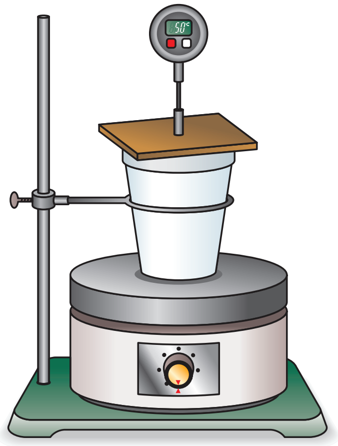

Figure 1. Calorimeter assembly

Part 2 (calorimeter calibration procedure)

Place a 100.0 mL sample of water in a clean, dry 150 mL beaker. Heat with occasional stirring to approximately 50°C. Remove the beaker from the hot plate and place on the lab bench. Meanwhile, place exactly 100.0 mL of cool water (approximately 20°C) in the clean, dry calorimeter.

Measure the temperature of the hot water and the cold water and record, then immediately pour the entire hot water sample into the calorimeter and quickly put on the cover. Wait 15 seconds then take a temperature reading. Repeat this determination twice.

INVESTIGATION

In today's experiment you will design and execute an experimental procedure to determine which of three ionic compounds is most suitable for use in a hand warmer. Review the criteria for an ideal hand warmer from the central challenge. For each solid you need to consider safety, cost, and environmental impact as well as the amount of heat released or absorbed.

Procedure

1. Safety and environmental impact: Obtain the Material Safety Data Sheets (MSDS) for your three solids from your teacher. Review each one, making notes about safety concerns, necessary precautions, and disposal.
2. Cost: Rank the solids you are given from least to most expensive.

Substance	2012 Cost per 500 g ($)
NaCl	3.95
$CaCl_2$	6.55
$NaC_2H_3O_2$	12.90
Na_2CO_3	6.15
LiCl	32.75
NH_4NO_3	9.05

3. Heat of solution: Work with your group to design a procedure to compare the solids in terms of the heat released or absorbed when they dissolve and include what materials and equipment you will use. You must include the safety precautions you will take.

Important Procedure Tips:

1. Be sure to keep detailed records of the amounts of substances used and the starting and ending temperature as you will need it later to determine the amount of solid to use in your hand warmer.
2. You will receive a maximum of 10 g of each solid for this part.

Data Collection and Computation

Calorimeter constant determination: According to the law of conservation of energy, energy cannot be created or destroyed, only changed from one form to another or transferred from one system to another. The temperature change observed when water or any substance changes temperature can be a result of a transfer of energy from the substance to the surroundings (in which case the temperature of the substance decreases) or the surroundings to the substance (in which case the temperature of the substance increases). When hot and cold water are mixed the hot water transfers some of its thermal energy to the cold water. The law of conservation of energy dictates that the amount of thermal energy lost (or the enthalpy change) by the hot water, q_{hot}, is equal to the enthalpy change of the cold water, q_{cold}, but opposite in sign, so $q_{hot} = -q_{cold}$. The enthalpy change for any substance is directly related to the mass of substance, m; the specific heat capacity (a substance-specific constant), c; and the temperature change, ΔT. The relationship is expressed mathematically in the equation $q = mc\Delta T$. The specific heat capacity of water is 4.184 J/°C°•g.

1a. Calculate the enthalpy change of the cold water using the equation $q_{cold} = m_{cold}c\Delta T_{cold}$. Assume that the density of water is exactly 1 g/mL. Is this an endothermic or exothermic process? Explain.

1b. Calculate the enthalpy change of the hot water using the equation $q_{hot} = m_{hot}c\Delta T_{hot}$. Assume that the density of water is exactly 1 g/mL. Is this an endothermic or exothermic process? Explain.

1c. These amounts are not equal because the calorimeter (the coffee cups) absorbs some of the thermal energy transferred by the hot water. Thus under the real conditions observed in the laboratory the law of conservation of energy equation becomes $q_{hot} = -(q_{cold} + q_{cal})$, where q_{cal} is the enthalpy change of the calorimeter. Use this equation to calculate the enthalpy change of the calorimeter.

1d. The calorimeter constant, C, is the heat absorbed by the calorimeter per degree of temperature change, $C = q_{cal}/\Delta T_{cal}$. Assuming the starting temperature of the calorimeter is the same as that of the cold water, calculate the calorimeter constant in units of joules per degree Celsius.

2. The solid and water, considered together, have a certain amount of internal energy as a function of the bonds that exist in the solid and in the water. The solution that is produced as a result of the dissolving has a different amount of internal energy than the solid and water did because the arrangement of particles and the bonds and attractions between the particles in the solution are different bonds and particulate attractions than the arrangement of particles and the bonds and attractions between the particles in the solid and water. The difference in energy, q_{soln}, is the reason for the difference in the thermal energy of the two systems (solid and pure water versus solution), with symbol q_{rxn}. Just as with the hot and cold water in the calorimeter constant determination, q_{soln} and q_{rxn} are equal in magnitude and opposite in sign, $q_{rxn} = -q_{soln}$. And just as in that case of the cold and hot water mixing, the calorimeter will also experience an enthalpy change during the solution formation process. To account for this enthalpy change the relationship is adjusted to $q_{soln} = -(q_{rxn} + C\Delta T)$ where C is the calorimeter constant determined above.

This difference in thermal energy of the system before and after solution formation, q_{soln}, can be calculated using the relationship $q_{rxn} = mc\Delta T$, where m is the total mass of the solution and c is the specific heat capacity of the solution and ΔT is the temperature change of the solution. It is important to note that we will assume that the heat capacity of the solutions is the same as pure water but in reality the solutions do not have exactly the same heat capacity, and this assumption affects the accuracy of this determination.

Using this information, calculate q_{soln} and q_{rxn} for all three solids you tested for your hand warmer.

3. By convention, scientists report enthalpy changes for dissolution (and many other processes) in units of kilojoules per mole of solute dissolved. Using your values of q_{soln}, calculate the enthalpy in units of kilojoules per mole. This quantity has the symbol ΔH_{soln}. Calculate ΔH_{soln} for each of the three solids you tested.
4. Based on the cost information provided, and your experimental work and calculations, select which chemical you believe will make the most cost-effective hand warmer. The hand warmer you are designing needs to increase in temperature by 20°C. Calculate the amount of the compound you selected that would be required for a hand warmer that meets this requirement.

Argumentation and Documentation

Write a paragraph in which you describe all of the factors you considered and you explain your rationale for choosing one chemical and not each of the other chemicals studied in this experiment. Your paragraph should start with a claim sentence that clearly states your choice and the amount of substance to use. The claim should be followed by evidence from your experiment and cost and safety analysis. The paragraph should conclude with reasoning explaining how your evidence supports your claim.

POSTLAB ASSESSMENT

1. Are the dissolving processes you carried out endothermic or exothermic or neither? Explain your thinking.
2. Dissolving ionic compounds involves the separation of the solid ionic compound into cations and anions in water. This process can be represented by an equation showing the solid as a reactant and the aqueous ions as products. The heat of reaction ΔH_{soln} is written after the products, typically in units of kJ/mol.

Example: sodium hydroxide dissolves exothermically, releasing 44.2 kilojoules per mole dissolved. This process is represented as $NaOH(s) \rightarrow Na^+ (aq) + OH^- (aq)$, $\Delta H_{soln} = -44.2 kJ/mol$.

Write an equation to represent the dissolving process for each salt you studied. Include your calculated heat of reaction as in the example.

3. Changes in matter are generally classified as physical or chemical based on whether new substances are formed through the process. Does dissolving represent a physical change, a chemical change, or an intermediate change? Explain your reasoning, including evidence from the animation you viewed.

4. Share your calculated values of ΔH_{soln} with your classmates and obtain their values.
   a. Determine the class average value and standard deviation for each solid.
   b. Find the published value of ΔH°_{soln} for each solid and determine the percent error in the class average value.
5. What possible sources of error could affect the accuracy of your calculated value of the amount of solid in your hand warmer? List at least two and what effect they would have on the temperature change.

SUPPLEMENTAL RESOURCES

Links

"Calorimetry." Sparknotes SAT Chemistry Common Experiments. Accessed August 3, 2012. *http://www.sparknotes.com/testprep/books/sat2/chemistry/chapter11section4.rhtml*

"Heats of Reaction." University of Massachusetts Boston. Accessed July 31, 2012. *http://alpha.chem.umb.edu/chemistry/genchem/103/files/103lab/7CoffeeCupCalorimeterRevised.pdf*

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Brouwer, Henry. "Small-Scale Thermochemistry Experiment." *Journal of Chemical Education* 68, no. 7 (1991): A178.

Marzacco, Charles J. "The Enthalpy of Decomposition of Hydrogen Peroxide: A General Chemistry Calorimetry Experiment." *Journal of Chemical Education* 76, no. 11 (1999): 1517.

Vannatta, Michael W., Michelle Richards-Babb, and Robert J. Sweeney. "Thermochemistry to the Rescue: A Novel Calorimetry Experiment for General Chemistry." *Journal of Chemical Education* 87, no. 11 (2010): 1222–1224.

INVESTIGATION 13

Can We Make the Colors of the Rainbow? An Application of Le Châtelier's Principle

CENTRAL CHALLENGE

What is Le Châtelier's principle? How and why does it work? The central challenge is to investigate this principle by testing several systems at equilibrium and then selecting specific ones to produce the colors of the rainbow based on applications of Le Châtelier's principle. An additional challenge involves selecting which reaction system to use for which color in producing the rainbow while trying to only use a given "stress" once. Several systems produce similar colors so not all of you will have the same results.

CONTEXT FOR THIS INVESTIGATION

Until now, most of the reactions studied have been assumed to go to completion. In these reactions, a color change was observed, a precipitate formed or a gas evolved with bubbling or fizzing. These reactions had been carried out under conditions that favor product formation.

In fact, though, for many chemical changes, under the correct conditions, the reaction does not go all the way to completion; rather, chemists say that an equilibrium state is reached, with some amounts of all reactants and products present in varying amounts.

One equilibrium system that is constantly being stressed by changes in reaction conditions is responsible for the transport of oxygen and carbon dioxide in your bodies. The oxygen, O_2, is in equilibrium with the oxygenated (HbO_2) and deoxygenated (Hb) forms of hemoglobin, the iron containing molecule in the blood that "carries" oxygen to your cells and carbon dioxide back to the lungs.

The equilibrium can be represented as: $Hb + O_2 \rightleftharpoons HbO_2$

In the lungs, the pressure of the oxygen is relatively high, so the reaction conditions, the "equilibrium," is said to favor the formation of HbO_2 there. The oxygen-rich blood then leaves the lungs and is carried to the cells of the body. Once the oxygenated hemoglobin reaches the cells where it is needed, the pressure of oxygen is much lower and the equilibrium no longer favors HbO_2, but rather the reverse (Hb). Thus, as a result of these new reaction conditions, the oxygen is "released" into the cell as the equilibrium is said to now favor the reactants.

At the same time, the pressure of carbon dioxide is elevated in the cell and so the CO_2 molecule binds to the Hb molecule in an equilibrium similar to the one responsible for oxygen transport. When the blood reaches the lungs again, the carbon dioxide is “released” from the hemoglobin as the pressure of CO_2 is now lower and the reactants (Hb and CO_2) are now favored.

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

In the introduction, the equilibrium system that is responsible for the transport of oxygen and carbon dioxide in your bodies was introduced. A related equilibrium involves carbon monoxide, a colorless, odorless gas that is produced by incomplete combustion and is poisonous to living organisms.

Carbon monoxide (CO) reacts with hemoglobin in an equilibrium similar to that of oxygen:

$$\text{Hb} + \text{CO} \rightleftharpoons \text{HbCO}$$

The reason that carbon monoxide is so poisonous is that the affinity between hemoglobin and carbon monoxide is approximately 200 times stronger than the affinity between hemoglobin and oxygen so carbon monoxide binds more strongly to hemoglobin than oxygen does. This means that a carbon monoxide poisoning victim will be starved of oxygen because carbon monoxide rather than oxygen is being transported in the blood.

Because the affinity between carbon monoxide and hemoglobin is so strong, low concentrations of carbon monoxide can be harmful. In the United States, the Occupational Safety and Health Administration (OHSA) limits long-term workplace exposure levels to less than 50 ppm averaged over an 8-hour period, and if a level of 100 ppm is reached employees must be evacuated from the area as it is unsafe. The symptoms include headaches and drowsiness at lower exposure levels, and unconsciousness and/or death with levels in excess of 1000 ppm for an hour.

Treatment for carbon monoxide poisoning involves the administration of oxygen at high pressures, which increases the amount of oxygenated hemoglobin (HbO_2) in the bloodstream.

Chemical equilibrium means that there are forward and reverse reactions occurring simultaneously at equal rates. This means that the amounts of both the reactants and the products remain constant. That is, the reactants are consumed to generate products at the same rate the products react to regenerate the reactants. No net change in amount occurs. To the outside observer it looks as if nothing is happening because on the macroscopic scale, no properties of the reaction system change. This is known as a dynamic equilibrium.

Equilibrium systems can be described mathematically using the concentrations of solutions or pressures of gasses in the mass action expression. The equilibrium constant is the mathematical result of these calculations. This equilibrium constant's specific value is dependent upon the temperature at which the reaction occurs.

For the generalized reaction: $aA(aq) + bB(aq) \rightleftharpoons cC(aq) + dD(aq)$

the equilibrium constant, or mass action expression, is: $K_c = \dfrac{[C]^c\ [D]^d}{[A]^a\ [B]^b}$

where K_c is the equilibrium constant, and [ ] denotes the concentration of each substance.

The only substances included in the expression are those that have a concentration — that is substances that are pure (solids or liquids) are not included in the expression. The solvent, water, is also omitted from the expression for dilute solutions as the concentration of water does not change appreciably.

Le Châtelier's Principle

Henri Le Châtelier studied and published extensively on the subject of equilibrium in solutions. His principle has been interpreted in many different chemical systems, but essentially it states that:

"If a change in concentration, temperature, pressure or volume is imposed on a chemical system at equilibrium, then the equilibrium shifts by changing the concentration or pressure to counteract the imposed change and establish a new equilibrium."

These changes in conditions are often referred to as "stresses." We say the equilibrium system has been "stressed" by the change and can predict the subsequent changes in concentration or pressure by examining the "stress." The principle is applied to predict changes in the relative amounts of reactants and products, called the equilibrium position, in the following manner.

For the generalized chemical reaction:

$$aA(aq) + bB(aq) \rightleftharpoons cC(aq) + dD(aq)$$

At constant temperature and pressure, adding substance A to the system when it is at equilibrium will cause a "stress" on the system by increasing the concentration of the reactant. An increase in the concentration of A will increase the rate of the forward reaction and the result will be an increase in the amounts of C and D, until their respective concentrations have risen to the point where the reverse reaction rate equals the forward reaction rate and equilibrium is restored. When this occurs we say that the equilibrium position "shifts to the right." Adding substance B would have the same effect for the same reason.

The "stress" of removing either substance A or B would have the reverse effect. In this case, the reduction in the concentration of the reactants would cause the rate of the forward reaction to decrease and thus the momentarily faster reverse reaction would cause an increase in the concentration of the removed substance until it has risen to the point where the reverse reaction rate equals the forward reaction rate and equilibrium has been restored. In this case Le Châtelier's principle states that the equilibrium has adjusted to the stress by "shifting to the left."

Alternately, removing C or D, the products, would cause the equilibrium position to "shift to the right" and adding C or D would "shift it to the left." These "shifts" would be due to increases or decreases in the rates of the reactions in one direction due to changing conditions.

In reactions involving gases, the pressure of the gas is important. The partial pressure of a gas is analogous to concentration so increasing the pressure of a gas is the same as increasing the concentration of the substance. Reducing the pressure is the same as reducing the concentration of the substance. Changes in volume are commonly employed in chemistry to adjust gas pressures and in applying Le Châtelier's principle special attention must be paid to the stoichiometry of the system. If the moles of gas present in the balanced equation are the same for both products and reactants there will be no change in the position of the equilibrium because of a volume change; however, if they are different, then the equilibrium position will shift. This shift will happen because there will be changes in the rate of the reaction in response to the unequal changes in the partial pressures of the gasses involved when the volume changes.

Finally, the equilibrium constant is temperature dependent. In your study of chemical kinetics you observed that reaction rates depend upon temperature and since equilibrium involves reaction rates, it naturally follows that the K_c value, and thus the position of an equilibrium, is temperature dependent. An increase in the temperature at which an endothermic reaction is performed will increase the rate of the forward reaction more than the rate of the reverse reaction and thus shift the reaction to the right. At the increased temperature, more reactant species will possess sufficient energy (the activation energy) to react and the rate for the forward reaction will exceed the rate of the reverse, shifting the equilibrium position. The value of the equilibrium constant, $K_{c,}$ is increased under the new higher temperature conditions. The opposite is true for an exothermic reaction. In this case, lowering the temperature for an exothermic reaction will cause a shift to the right because, at the lower temperature, fewer reacting species will possess the activation energy and thus the rate of the reverse reaction will be greater, causing the shift to the right.

In this lab, you will study a variety of equilibrium systems and use Le Châtelier's principle to predict shifts that you will then use to create a color display.

PREPARATION

Materials

Materials (per tray)

Tray 1	Tray 2	Tray 3	Tray 4	Tray 5	Tray 6
Bromothymol blue indicator solution in a small dropper bottle (30 mL)	0.10 M KSCN solution in a bottle (250 mL)	0.25 M $CuSO_4$ in a bottle (250 mL)	10 g of solid $CuCl_2 \bullet 2H_2O$	10 g of solid $CoCl_2 \bullet 6H_2O$ in a small vial	Cold, fresh seltzer water/soda water/club soda (small commercial bottle) on ice
0.10 M HCl in a small dropper bottle (15 mL)	0.20 M $Fe(NO_3)_3$ solution in a small dropper bottle (15 mL)	Concentrated ammonia in small dropper bottle (15 mL)	12 M HCl in a small dropper bottle (15 mL)	95% ethanol in a bottle (250 mL)	Methyl red indicator solution in a small dropper bottle (15 mL)
0.10 M NaOH in a small dropper bottle (15 mL)	10 g of solid KSCN in a small vial	1.0 M HCl in a small dropper bottle (15 mL)	1 labeled, small spatula	Distilled water in a small dropper bottle (30 mL)	Modified 60 mL or larger syringes with Luer Lock valves
0.10 M NaCl in a small dropper bottle (15 mL)	10 g of solid $FeCl_3$ in a small vial			Acetone in a small dropper bottle (30 mL)	Nails to lock syringe open
	10 g of any solid sodium phosphate (Na_2HPO_4; NaH_2PO_4; Na_3PO_4) in a small vial			10 g of solid NaCl in a small vial	
	3 labeled, small spatulas			0.10 M $AgNO_3$ in a small dropper bottle (15 mL)	

Tray 1	Tray 2	Tray 3	Tray 4	Tray 5	Tray 6
	0.10 *M* KNO_3 in a small dropper bottle (15 mL)			1 labeled, small spatula	
Equipment (per lab group)					
6–50 mL beakers	2–25 mL graduated cylinders	2 test-tube racks	Ice for cold water bath	Deionized or distilled water in wash bottles	Hot water for hot water bath
2–250 mL beakers for water baths	24–10 × 75 mm test tubes and/or 13 × 100 mm test tubes	6 glass stirring rods			

Safety and Disposal

The solutions used in this lab require careful handling and adherence to all safety guidelines, as some of them can cause skin burns and eye damage. When working with a syringe or gases under pressure, direct the syringe away from people and toward the wall or floor. You must take normal laboratory precautions, including wearing splash-proof goggles at all times. You may wish to wear gloves. If solutions are spilled on skin, wash with copious amounts of water. Once used in the experiment, all the used solutions can be safely disposed following standard procedures as directed by your instructor.

INVESTIGATION

You have been asked by the chemistry department to design a display for the department's showcase illustrating the use of Le Châtelier's principle to produce the colors of the rainbow. At a minimum, you must have red, orange, yellow, green, blue, and violet. As you are aware, the stresses that can be applied include adding or removing a reactant or product, increasing or lowering the temperature of the system, and increasing or lowering the pressure on a gas sample by adjusting the volume. Your goal is to use as many different stresses as possible in producing your display.

Procedure

Possible equilibrium systems for your use are described below. You will need to investigate some or all of them and make predictions.

Be sure to discuss your proposed stresses with your teacher before attempting them. In most cases you will prepare a “stock” solution and then divide it up into 13 × 100 mm or 10 × 75 mm test tubes to study the effects of your stresses. Be sure to use small quantities of reagents, add these reagents dropwise while stirring, and use hot plates and/or water baths to heat your samples should you wish.

Be sure to keep detailed written records, including a step-by-step procedure, a list of materials used, all data, and observations. You need to clearly indicate the reactant and product species in each system, the stressor you applied, and the resulting color.

Once you have completed all your investigations and decided which reactions and stresses you will use to prepare your display you may be asked to produce the rainbow display for your teacher.

You are to take one tray of materials relating to an equilibrium system to be examined from a central location to the lab bench and study the systems one at a time. Once you have completed the investigation of a particular system, return the tray to the central location and obtain another. Continue to exchange trays until all the systems have been studied.

Possible Equilibria

An Acid-Base Indicator Equilibrium

Acid-base indicators are large organic molecules that can gain and lose hydrogen ions to form substances that have different colors. The reaction of the indicator bromothymol blue (BTB) is well known. Investigate its behavior in water with dilute acid and base. To create the equilibrium system, add 1 mL of BTB to 25 mL of water. Investigate the shifts that can be observed by applying a stress to small amounts of this solution in the test tubes provided. You have 0.10 *M* NaOH(*aq*), 0.10 *M* HCl(*aq*), and 0.10 *M* NaCl(*aq*) for your use.

Some Complex Ion Equilibria

a. An equilibrium system can be formed in solution with iron(III) chloride and potassium thiocyanate. The ion that forms from their combination is the $FeSCN^{2+}$ ion. To create the equilibrium mixture add about 20 mL of 0.10 *M* KSCN solution to a beaker and then add 20 mL of distilled water and 5 drops of 0.20 *M* $Fe(NO_3)_3$ solution. Investigate the shifts that can be observed by applying a stress to small amounts of this solution in the test tubes provided. You have 0.10 *M* KNO_3(*aq*), solid $FeCl_3$, solid KSCN and a sodium phosphate, either Na_2HPO_4, NaH_2PO_4, or Na_3PO_4 for your use.

b. An equilibrium system can be formed in a solution with copper (II) sulfate and ammonia. To create the equilibrium mixture, take 25 mL 0.25 *M* $CuSO_4$ and add concentrated NH_3 dropwise to observe the production of a precipitate. Continue adding the ammonia until the precipitate disappears. Investigate the shifts that can be observed by applying a stress to small amounts of this solution in the test tubes provided. You have HCl(*aq*) and your original reagents.

c. An equilibrium system can be formed in a solution with copper(II) chloride in water. The copper ion bonds to six water molecules to form the hydrated complex ion, $[Cu(H_2O)_6]^{2+}(aq)$, while the Cl^- ion remains in solution. To create the equilibrium mixture add about 2 grams of copper (II) chloride to 25 mL of water. Investigate the shifts that can be observed by applying a stress to small amounts of this solution in the test tubes provided. You have concentrated HCl (caution) and water.

Hydrated Cobalt Complex Ions in Alcohol Solution Equilibrium

Caution – Ethanol is flammable. Turn off all flames.

To create the equilibrium mixture, add 2 grams of cobalt(II) chloride hexahydrate to 25 mL of 95 percent ethanol in a 100 mL beaker. There is a small amount of water dissolved in the ethanol. The equilibrium reaction is endothermic.

$$\underset{\text{pink}}{[Co(H_2O)_6]^{2+}(alc)} + 4\,Cl^-(alc) + \text{Energy} \rightleftharpoons \underset{\text{blue}}{[CoCl_4]^{2-}(alc)} + 6\,H_2O(alc)$$

Investigate the shifts that can be observed by applying a stress to small amounts of this solution in the test tubes provided. You have solid NaCl, water, acetone, and $AgNO_3(aq)$. You also have an ice and hot water bath.

Soda Water and Methyl Red Indicator Equilibrium

To create the equilibrium mixture, place 20 mL of cold soda water in a beaker and add 1 mL of methyl red indicator. Draw 10 mL of the mixture up into a large syringe with a Luer Lock valve. Invert the syringe and depress the piston to expel gas and then close the stopcock on the Luer Lock valve. Investigate the shift produced by pulling back on the piston and holding it. You can put a small nail through the hole in the piston to keep it extended (see Figure 1 below).

$$CO_2(g) \rightleftharpoons CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$$

The methyl red is red in the original soda water but once the CO_2 has moved into the gas phase the methyl red becomes yellow, indicating the shift to the left. You may have to shake the syringe to get the CO_2 out of solution. This can be the source of the yellow- or orange-colored solution.

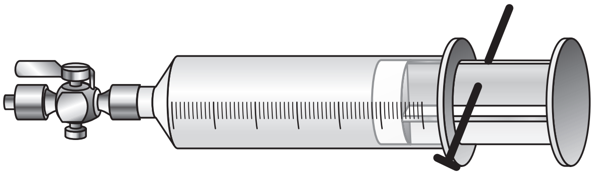

Figure 1. Large syringe with a Luer Lock valve

Data Collection and Computation

For each equilibrium system studied, create a data table to present the following components:

- Write the chemical equation for each reaction
- identify the reactants, products, and result of each change of conditions
- indicate the type of stress each change of conditions imposed, the color produced, and the reasons for your assertions
- explain how the applied stress resulted in the observed color
- use Le Châtelier's principle in your explanations; however, do not simply quote it as a substitute for a proper chemical explanation
- list the chemical species responsible for color of any kind in order of increasing concentration

As a final step, show your teacher your rainbow!

POSTLAB ASSESSMENT

1. A student obtained a test tube with a suspension of white, slightly soluble calcium hydroxide in water. This system was at equilibrium as represented by the following equation:

$$\mathrm{Ca(OH)_2}(s) \rightleftharpoons \mathrm{Ca^{2+}}(aq) + 2\ \mathrm{OH^-}(aq)$$

   a. Write the equilibrium constant expression for this reaction.
   b. What would you expect to observe if hydrochloric acid, HCl(aq), was added? Explain your answer using Le Châtelier's principle.
   c. What would you expect to observe if calcium nitrate was added? Explain your answer using Le Châtelier's principle.
   d. When the solution was placed in an ice bath and cooled it was observed that more solid calcium hydroxide was produced.
      i. Based on this observation would you expect the reaction to be exothermic or endothermic? Explain your answer using Le Châtelier's principle.
      ii. If the solution was placed in a hot water bath and heated, what would you expect to observe? Explain your answer using Le Châtelier's principle.

2. Propose a series of equations to represent the equilibria described in the prelab discussion involving the transport of oxygen from the lungs to the cells and the carbon dioxide back to the lungs. Use Hb^+ for deoxyhemoglobin, HbO_x for oxyhemoglobin, and $HbCO_2$ for the carbon dioxide-hemoglobin structure.
   a. Use your equations to explain how the oxygen transport system works using Le Châtelier's principle.
   b. Carbon monoxide is a poison that prevents the hemoglobin molecule form binding with the oxygen in the lungs. This is because of the structure and polar nature of the CO molecule, which causes it to bind strongly with the Hb^+. Using Le Châtelier's

principle, explain how this strong binding would affect the oxygen-carrying capacity of the blood.

c. Propose a method, using Le Châtelier's principle, that could effectively treat a victim of carbon monoxide poisoning.

3. The following questions concern this equilibrium system.

$$Br_2(l) + Cl_2(g) \rightleftharpoons 2\ BrCl(g)\ \Delta H = +29.4\ \text{kJ/mol}$$

How will the following factors influence the equilibrium listed above? Indicate whether the system will shift left, right, or remain unchanged, and give a short explanation for your choice. Simply citing Le Châtelier's principle is not an adequate answer; rather, explain why the system does or does not respond to the stress using your knowledge of the collision theory and chemical kinetics. In all cases, the listed change is the only change — all other variables (P, V, or T) remain constant.

a. Increasing the temperature

b. Increasing the pressure in the flask by adding Ar

c. Increasing the volume of the flask

d. Adding $Br_2(l)$

e. Removing $Cl_2(g)$

f. Adding $BrCl(g)$

g. Adding a catalyst

4. Indicating Drierite is a material used in the laboratory to remove water vapor from gasses and as a dessicating agent. When purchased it is blue in color and it changes to pink upon absorbing moisture.

a. Based upon your observations in this lab, explain how Indicating Drierite works.

b. Indicating Drierite can be regenerated and used over and over again. Propose a method of regenerating Drierite quickly, efficiently, and at low cost. Explain how your method will work, using Le Châtelier's principle.

5. Nickel forms a green complex ion, $[Ni(H_2O)_6]^{2+}$, in water. The addition of a small amount of ethylenediamine, results in a light-blue-colored solution. Ethelynediamine is a bidentate ligand, meaning it bonds to metals through two atoms of the ligand. Further addition of ethylenediamine results in a royal blue solution while the addition of even more ethylenediamine produces a violet solution. Further addition produces no additional color changes.

a. Propose a series of equilibrium reactions that would illustrate the observations.

b. Predict the effect of adding solid nickel(II) nitrate to the violet solution. Explain your answer using Le Châtelier's principle.

c. A small amount of hydrogen sulfide gas is bubbled through the violet solution resulting in a black precipitate and the solution turns royal blue. Using LeChâtelier's principle, explain this observation.

SUPPLEMENTAL RESOURCES

Links

"Chapter Fourteen — Gas Phase, Solubility, Complex Ion Equilibria: Animations." The North Carolina School of Science and Mathematics. Accessed July 31, 2012. *http://www.dlt.ncssm.edu/core/c14.htm*

Greenbowe, Tom. "Bromine Gas/Liquid Equilibrium Animation." Iowa State University. Accessed July 31, 2012. *http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/equilvpBr2V8.html*

Greenbowe, Tom. "Cobalt Chloride Equilibrium Animation." Iowa State University. Accessed July 31, 2012. *http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/CoCl2equilV8.html*

Greenbowe, Tom. "N_2/N_2O_4 Equilibrium Animation." Iowa State University. Accessed July 31, 2012. *http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/no2n2o4equilV8.html*

"Reversible Reactions." University of Colorado at Boulder, PhET Interactive Simulations. Accessed July 31, 2012. *http://phet.colorado.edu/en/simulation/reversible-reactions*

"Salts and Solubility." University of Colorado at Boulder, PhET Interactive Simulations. Accessed July 31, 2012. *http://phet.colorado.edu/en/simulation/soluble-salts*

INVESTIGATION 14

How Do the Structure and the Initial Concentration of an Acid and a Base Influence the pH of the Resultant Solution During a Titration?

CENTRAL CHALLENGE

While there are times when you only need to know if a solution is acidic, basic, or neutral, often the exact concentration is important, such as when making biodiesel fuel from vegetable oil. When vegetable oil degrades it becomes acidic. A base such as lye is added to neutralize the acid. The exact concentration of the acid must be known because if too much base is added, you will get soap instead of biodiesel fuel! A titration of the acid with a base will determine the exact concentration. Besides doing the titration you must be able to analyze the resultant titration curve.

CONTEXT FOR THIS INVESTIGATION

Many foods taste as they do due to the presence of acidic or basic content. All foods, beverages, pharmaceuticals, biofuels, water in aquariums, drain cleaners, surface cleaners, and vitamins contain acids or bases, or a mixture of acids and bases. The amount of acid, base, and the pH of solutions and solids must be maintained at an optimal level. If a solution is too acidic, some base can be added to react with some of the acid. For example, hydrochloric acid reacts with sodium hydroxide to produce sodium chloride and water.

$$\mathrm{HCl}(aq) + \mathrm{NaOH}(aq) \rightarrow \mathrm{NaCl}(aq) + \mathrm{H_2O}(l)$$

By carefully controlling the amount of base added while doing an acid-base titration, and knowing when to stop adding base by using an indicator or a pH meter, one can determine the amount of acid present in the substance. The food industry uses titrations to determine the amount of sugar, free fatty acid content, and the concentration of vitamin C or E present in products.

PRELAB GUIDING QUESTIONS/SIMULATIONS

Part I: Questions

Choose from the following questions, as indicated by your teacher, or generate your own questions to start your investigation.

1. Given 50 mL of 0.10 *M* HCl and 50 mL of 0.10 *M* acetic acid, will the amount of 0.10 *M* NaOH required to neutralize each solution be the same, more, or less?
2. Will the pH at the equivalence point of 50 mL 0.10 *M* HCl be the same, more, or less as the pH at the equivalence points for 50 mL of 0.10 *M* acetic acid?
3. What are some structural features that might help us classify an acid as a strong acid or weak acid?
4. Draw a molecular and particulate view of what is happening in the steep part of a general acid-base titration curve (such as Figure 1).
5. How does the structure of an acid affect the shape of the titration curve?
6. How can a pH titration curve be used to help classify the resultant solution at the end-point, as acidic, basic, or neutral?
7. How do the structure and the initial concentration of an acid and a base influence pH of the resultant solution during a titration?

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

Everyone has tastes and textures of food they prefer. Taste and texture are often linked to the acidity or alkalinity of a food or beverage detected by the tongue, which has sensors for different tastes. A sour, acidic lemon has a different taste compared to a drink of green tea or herbal tea, which are alkaline.

Acids and bases are unique compounds that play an important role in influencing the pH of a solution. Chemists use several definitions to help classify compounds as an acid or a base. Svante Arrhenius defined acids as compounds containing the hydrogen ion, H^+, and bases as compounds containing the hydroxide ion OH^-. Brønsted-Lowry acids are defined as proton donors in a reaction and bases are proton acceptors in a reaction. The proton referred to is an H^+ ion. A hydrogen atom has one proton and one electron and when the electron is removed to form an H^+ ion, only a proton remains. It is not possible though for a single H^+ ion to exist in water. The H^+ combines with a water molecule to form the hydronium ion, H_3O^+. The pH of an aqueous solution is a measure of the amount of hydronium ion $[H_3O^+]$ species which is also simply represented as the hydrogen ion H^+ in the solution; $pH = -\log[H^+]$.

Acids and bases can be considered weak or strong by the amount of ionization occurring in solution. Strong acids will ionize nearly 100 percent into ions while weak acids will ionize only a small percentage. For example the strong acid HBr will ionize almost completely into H^+ and Br^-, while the weak acid CH_3COOH will remain primarily CH_3COOH even though some CH_3COO^- and H^+ form. The common strong acids are HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_3$, and $HClO_4$. The strong bases also ionize completely and are the Group 1 and some of the Group 2 hydroxides: LiOH, NaOH, KOH, RbOH, CsOH, $Ba(OH)_2$, $Sr(OH)_2$, and $Ca(OH)_2$. Even though some Group 2 hydroxides are only slightly soluble, the amount dissolving ionizes completely.

In acid-base titrations, the titrant in the buret is the chemical solution added to the chemical solution in the flask or beaker called the titrate. The objective of this experiment is to determine the concentration or molarity of a solution by doing a titration. Often in titrations, the base is added to the acid. When the moles of acid (really the moles of H^+ released) present are equal to the moles of base added (or H^+ consumed), the reaction has reached the equivalent point. The calculation of the unknown molarity involves finding the moles of acid and the moles of base. If the known molarity is that of the base then the base molarity times its volume in liters times the number of hydroxides in its formula will equal the total moles of hydroxide present: $M_{base} \times V_{base} \times \#OH^-$ ions in the formula = moles OH^-. This will be equal to the total moles of acid present at equivalence. The total moles of acid is then set equal to the molarity of the acid times the volume of the acid used in liters times the number of hydrogen ions in the acid: moles acid = $M_{acid} \times V_{acid} \times \#H^+$ ions in the formula. The equivalence point can be determined by graphical means or by using an indicator.

An indicator is a solution containing an organic compound, either a weak acid or weak base, which exhibits a different color in certain pH ranges. A common indicator used in reactions of strong acids and strong bases is phenolphthalein which is colorless in acidic solutions and pink in basic solutions. Choosing the right indicator is important since the indicator color change is supposed to indicate the equivalence point. If the equivalence should be at a pH of 8 the indicator needs to change color around pH = 8. Done this way, the titration stops at the indicator change point, which is called the end point. This method does not monitor the pH throughout the titration. A titration curve cannot be made using a single acid-base indicator (universal indicator, a mixture of multiple indicators would work).

If a graph is made of how the pH changes as the titrant is added, this is called a titration curve. The graph below shows how this might look.

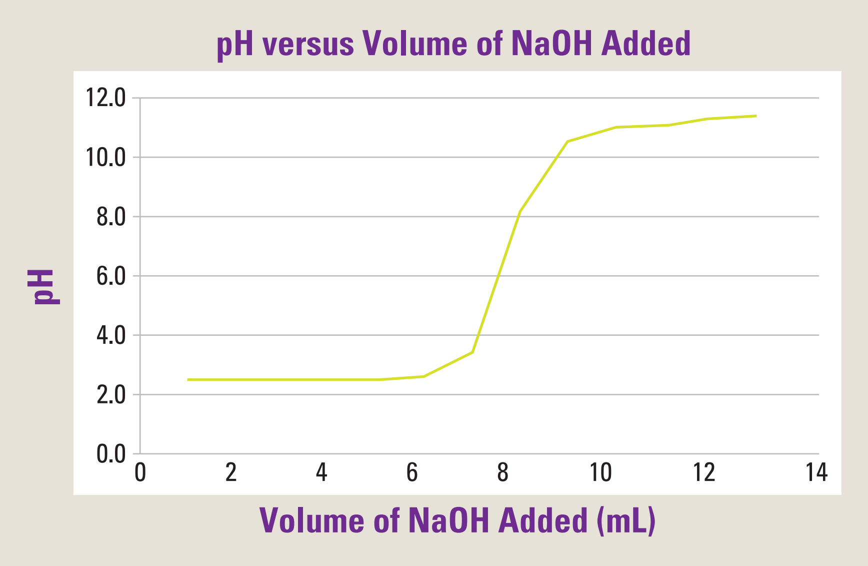

Figure 1. Acid-base titration curve

PREPARATION

Materials

Hydrochloric acid (HCl)	Ammonia (NH_3)	2–250 mL Erlenmeyer flasks or beakers	2–50 mL burets
Sulfuric acid (H_2SO_4)	Sodium hydroxide (NaOH)	Utility stand	Stirring rod
Nitric acid (HNO_3)	Calcium hydroxide ($Ca(OH)_2$)	pH meter or pH probe	2–100 mL graduated cylinders
Acetic acid (CH_3COOH) or other acids as noted by your teacher			

Safety and Disposal

Acidic and basic solutions can be dangerous. Working with acids and bases requires you to adhere to all safety guidelines including wearing gloves. You need to look up the specific MSDS of your possible acids and bases prior to doing this activity. The following URL has freely accessible MSDS for the acids and base used in this activity. *http://www.ehso.com/msds.php*

Acids and bases can cause skin damage and eye damage. Some of the acids you will be working with are extremely corrosive and hygroscopic. Acid-base reactions are exothermic. When preparing dilute acid or base solutions from concentrated acids and bases, take care because the solution process is exothermic. Do not add water to any acid. This may causes the water to sputter and some acid may splash out of the container. If you are uncertain of any process or procedure, check with your teacher.

Please wear splash-proof goggles and rubber gloves at all times when you and your fellow students are working with acids and bases. If solutions are spilled, inform your teacher immediately. If on your skin, rinse with running water for 15 minutes, and follow your lab safety procedures. At the end of the lab, the solutions should be neutralized and the pH tested so that the waste can be safely disposed of following the procedures outlined by your instructor.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

Prior knowledge in this lab includes knowing how to perform an acid-base titration. You should follow a general procedure for doing an acid-base titration. The following procedures are to develop an understanding of not only a titration but why it must be done when calculations are required.

Procedure

Acids and bases can be tested in several ways. One way is just to test to see if it is an acid or base using litmus paper or pH Hydrion Paper.

Step 1: Test an acid and a base with litmus paper. Does the litmus test provide you any quantitative data about the substance? What does the pH Hydrion Paper test indicate that the litmus paper did not?

Step 2: Measure 5.0 mL of acid and 5.0 mL of base. Pour them together. What can you tell is happening with the acid and base just by observing the reaction at this point?

Step 3: Do Step 2 again, but this time add 1 drop of the indicator phenolphthalein to the acid and slowly pour the base into the acid. When you see a color change, test the pH. Pour the rest of the base into the acid and test the pH again. Compare the two trials. How was the data different between the trials?

Many experimental procedures require exact concentrations, and, to get this type of quantitative data, you need to take more exact measurements. A titration produces this type of data.

In titration there will be two solutions, an acid and a base. A solution whose molarity is known is called a titrant, and this titrant is added to another solution until the chemical reaction is complete. Pour a measured volume (such as 25 mL) of the unknown solution to be titrated into an Erlenmeyer flask. Rinse a buret with the titrant and then pour the titrant into a buret held up by a ring stand. The buret is set up over the Erlenmeyer flask so the titrant can be slowly added to the unknown solution to be titrated. Monitor the pH throughout the reaction with a pH meter or a probe. Continue the titration until the pH remains constant after a steep change in pH. Make a graph of the data (pH versus titrant added in mL). From the equivalence point on the graph, determine the amount of titrant added to reach equivalence.

INVESTIGATION

Your group will write a procedure that uses different combinations of two acids: one with a known molarity and the other with an unknown molarity, and two bases, one with a known molarity and the other with an unknown molarity. The procedure will require that your group perform acid-base titrations to collect data to draw titration curves, calculate the unknown molarities, and answer the question your group is investigating, which you chose in the prelab Guiding Questions/ Simulations section.

Procedure

Write a procedure to test the question your group has decided on using an acid-base titration. After completing your procedure, have your instructor check it and initial it before proceeding.

Two of the four samples given to the group will have their molarity identified, and two will not. After collecting all your data, make a titration graph. All major points should be identified on the curves. If a sample includes a weak acid and/or base, percent ionization needs to be calculated, K_a and/or K_b needs to be calculated, and the percent error of the calculated K_a or K_b to the accepted value should be calculated.

In-Lab Discussion Questions

a. What is happening at the particulate level during a titration of a weak acid with a strong base?

b. What is happening at the particulate level when there is a steep part of the titration curve?

c. How can the steep part of the curve be used in calculations?

d. Does the steep part tell you anything about the endpoint or the equivalence point? Explain

e. Using one of your pH curves, predict and explain what the shape of the pH curve will look like if the experiment was repeated with a lower concentration of analyte.

Data Collection and Computation

Complete the data table(s) you have prepared.

Draw titration curves for each titration performed and label appropriately.

Determine the percent error of K_a and/or K_b if appropriate.

After completing the data collection and drawing titration curves, pool data with other groups and compare the titration curves.

Argumentation and Documentation

Answer the following questions after you have completed your graphs and calculations.

1. How do the process and the titration curves drawn vary if the acids or bases are weak or strong? Justify your answer.
2. What would a titration curve look like if an indicator were used to signal when to stop the titration?
3. Commercials about antacids are on television all the time. How would you go about investigating bases like antacids that are solid?
4. How would you investigate which antacid neutralizes the most acid or is the most cost effective?
5. Does it matter whether you start with pure acid or pure base as the titrate? Why or why not?

POSTLAB ASSESSMENT

1. Explain how rinsing the buret with water instead of the titrant before starting the investigation will affect the calculated unknown molarity of the titrate.
2. Explain why there is a steep slope in a section of the titration curve and explain how it can be used in calculations.

3. What types of data need to be collected to perform molarity calculations of the unknown?
4. Does the presence of weak or strong acids and weak or strong bases make a difference to when the equivalence point occurs? Justify your answer.
5. Using one of your titration curves, explain how the ratio of the acid species to the conjugate base concentration changes as the titration proceeds and draw particulate representations to show these changes at the beginning, half equivalence point, equivalence point, and beyond the equivalence point.
6. Explain how to determine the K_a of an acid and the K_b of a base from a titration curve.
7. Your car's battery blows up, spraying sulfuric acid all over the engine's hoses and yourself. Explain how you might neutralize the acid using available household chemicals.
8. Include possible equations to help explain why taking an antacid is recommended when a person has heartburn from consuming too many acidic foods or has acid reflux.
9. Challenge question: Amino acids are essential to carbon-based sentient life forms. Isoleucine is an α-amino acid with the chemical formula $HO_2CCH(NH_2)CH(CH_3)CH_2CH_3$. Since carbon-based life forms cannot synthesize isoleucine, this amino acid must be obtained through eating various foods. The IUPAC name for isoleucine is 2-amino-3-methylpentanoic acid. Draw the structure for isoleucine. Given the following acid-base titration curve for the titration of isoleucine, determine the pK_a values for isoleucine.

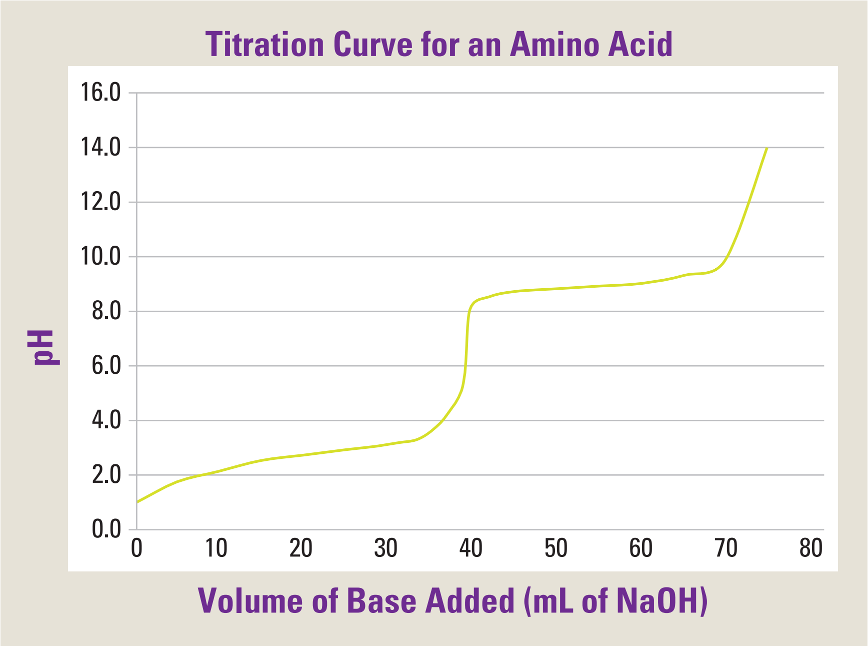

Figure 2. Amino acid titration curve

SUPPLEMENTAL RESOURCES

Links

“Acid-Base Interactions.” Oklahoma State University Chemistry Department. Accessed July 31, 2012. *http://genchem1.chem.okstate.edu/ccli/CCLIDefault.html*

“Acid-Base Solutions.” University of Colorado at Boulder, PhET Interactive Simulations. Accessed July 31, 2012. *http://phet.colorado.edu/en/simulation/acid-base-solutions*

“Acid-Base Titrations.” About Chemistry. Accessed July 31, 2012. *http://chemistry.about.com/od/chemistryquickreview/a/titrationcalc.htm*

“Titration.” 101 Science. Accessed July 31, 2012. *http://www.101science.com/Chemistry.htm#TITRATION*

References

Barnum, Dennis. W. “Predicting Acid-Base Titration Curves without Calculations.” *Journal of Chemical Education* 76, no. 7 (1999): 938.

Glaister, Paul. “A Unified Titration Formula.” *Journal of Chemical Education* 76, no. 1 (1999): 132.

Wildman, Randall. J., and Coleman, William. F. “Acid-Base Equilibria in Aqueous Solutions.” *Journal of Chemical Education* 79, no. 12 (2002): 1486.

INVESTIGATION 15

To What Extent Do Common Household Products Have Buffering Activity?

CENTRAL CHALLENGE

Many of the substances we come in contact with on a daily basis are weak acids or bases, which can act as buffers. In many instances, the buffering activity of these substances is what makes them useful to us. The objective of this experiment is for you to experimentally determine whether or not a household substance has buffering activity and to identify the potential buffer system within the household substance.

CONTEXT FOR THIS INVESTIGATION

Many of the chemicals we come in contact with on a regular basis, particularly foods, contain compounds that act as buffers. Some common buffers in foods include citrate, carbonate, phosphate, and tartrate. The dissolved carbon dioxide in carbonated beverages forms carbonate, and most sodas also contain citric acid and citrate. Some examples include lemon-lime sodas (such as Sprite, 7-Up, and Sierra Mist), ginger ale, and root beer. Colas, such as Coca-Cola and Pepsi-Cola, contain phosphoric acid. (In fact, Pepsi-Cola contains both phosphoric and citric acids, whereas Coca-Cola does not contain citric acid.)

PRELAB GUIDING QUESTIONS

1. Give a definition of a buffer.
2. If you titrate acetic acid (CH_3COOH) with sodium hydroxide (NaOH), the resulting products are the acetate ion (CH_3COO^-), the sodium ion, and water (see Figure 1). At a certain point, the reaction mixture contains acetic acid (the weak acid) and the acetate ion (its conjugate base) in solution, producing a buffer effect. If you were to continue adding NaOH to this reaction, what would eventually happen to this buffer effect? Explain.

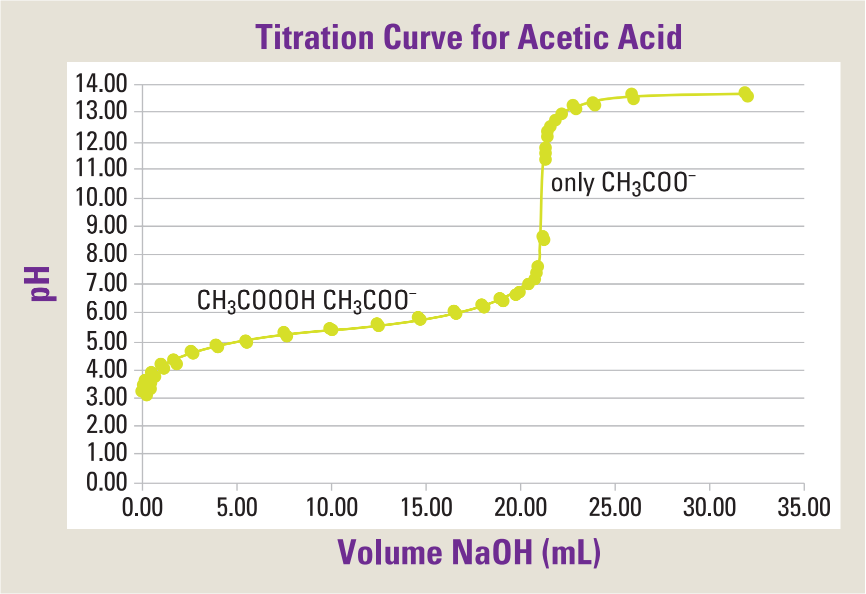

Figure 1. Titration of acetic acid with sodium hydroxide

3. Which titration curve (see Figure 2) represents the titration of a strong acid with a strong base? Which curve represents the titration of a weak acid with a strong base? Do either of these curves display a buffer effect? Explain.

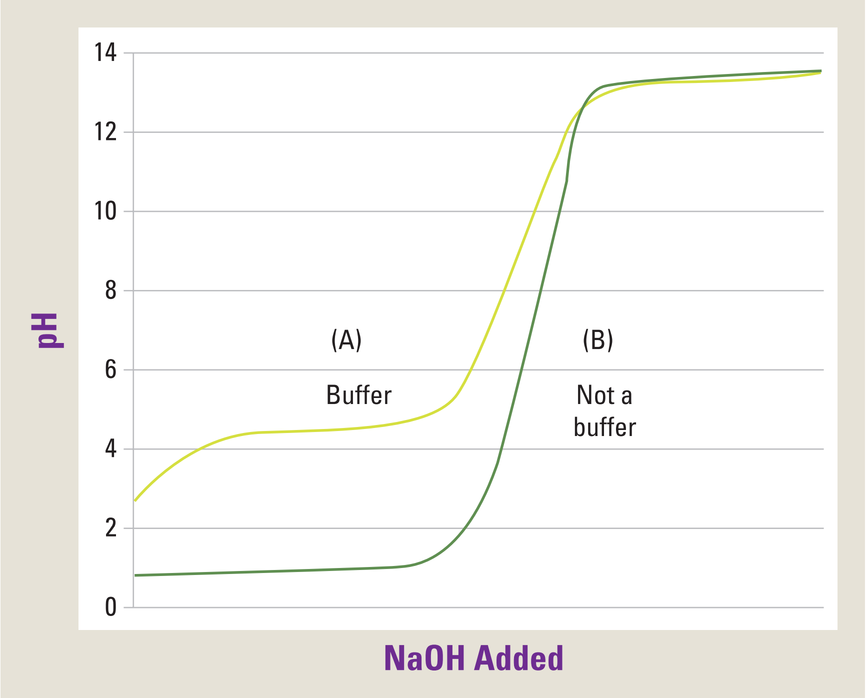

Figure 2. Titration curves

4. What is meant by the term "equivalence point?" What are the approximate equivalence points for curves A and B in Figure 2?

5. At what point along a titration curve is $pK_a = pH$ for a weak acid?
6. Identify on the titration curve (Figure 2, curve A) the place where you could estimate the pK_a of the acid. What is your best estimate?
7. What is the difference, in terms of the buffering region and the equivalence point, between acidic and basic buffers?
8. Write the neutralization reaction (in net ionic form) for a titration of acetic acid (CH_3COOH) with sodium hydroxide (NaOH). If you had 1 mole of buffer that was exactly 50 percent acetic acid and 50 percent acetate, how much NaOH would it take to neutralize all of the buffer?
9. Describe how the titration curve for a polyprotic acid would look different from a monoprotic acid.
10. Do all polyprotic acids act as buffers in aqueous solution? What determines the buffer activity of a polyprotic acid? Explain.
11. Why do titration curves of strong acids and strong bases "flatten out" around pH 1 and around pH 13?
12. When can a salt act as a buffer? Give an example.

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

A buffer is a solution of a weak acid-base pair that prevents the pH of a solution from changing drastically until the buffer has been neutralized. The weak acid has a proton that dissociates in aqueous solution, forming the conjugate base. The buffer acid's ability to give up H^+ ions and the conjugate base's ability to accept them is what enables the conjugate acid-base pair to maintain a roughly constant concentration of H^+ ions (i.e., a constant pH).

At a pH equal to the pK_a of the acid, the acid and its conjugate base are present in equal amounts. As the pH shifts away from the pK_a of the acid, the ratio of the acid to its conjugate base changes. Thus the pH of a buffer solution can be specified by adjusting the specific concentrations of the acid and conjugate base.

In this investigation, you will investigate the behavior of one or more unknown solutions, which may or may not have buffer activity. You will titrate the solution to determine whether it acts as a buffer. If it does, you will attempt to identify the potential buffer system, based on the table of pK_a values for common acids given in the Data Collection and Computation section of this document.

PREPARATION

Materials

0.1 *M* hydrochloric acid (HCl),	Buret, 50 mL	Mouthwash (Listerine contains benzoate/benzoic acid and methyl salicylate)
0.1 *M* sodium hydroxide (NaOH),	Buret funnel	Diet Coke (contains phosphate/phosphoric acid)
0.01 *M* citric acid ($C_6H_8O_7$), (optional)	Wash bottle with distilled water	Diet noncola soft drink, such as Sprite, ginger ale, etc. (contains citrate/citric acid)
pH meter and probe (or computer interface with pH probe)	Magnetic stirrer with stir bar (optional)	Tonic water (contains quinine)
Beakers, 150 or 250 mL	pH paper (broad and narrow range) (optional)	Noncarbonated diet soft drink, such as lemonade or iced tea
Graduated cylinders, 25 and 100 mL	Dishwasher detergent (Cascade contains sodium carbonate)	Powdered diet drink mix, such as lemonade or sugar-free punch
Ring stand	Citrus-based cleaning product (e.g., Lemi Shine contains citric acid)	
Double buret clamp	Isopropyl alcohol, 91%	

Safety and Disposal

All of the test materials are household products, which pose minimal risk. The titrant solutions of 0.1 *M* hydrochloric acid and 0.1 *M* sodium hydroxide are irritants. Although these solutions are dilute, extra care should be taken to avoid contact with skin or eyes. Appropriate protective clothing, including chemical safety goggles, should be worn, and in case of contact with skin or eyes, instructions in the MSDS for the applicable chemical should be followed.

Waste solutions can be combined and, if the pH is within the range permitted by applicable laws (usually between 4 and 10, but check with your teacher), flushed down the drain with a copious amount of water.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

If this experiment is your first experience with titration, your teacher may provide the following opportunities for you to practice performing a titration:

- Use of a buret, including rinsing of the buret with titrant before filling it
- pH measurement with a pH meter, indicator solution, and/or pH paper
- Graphing titration data and using the graph to find the equivalence point

For a simple titration practice, your teacher may give you an unknown concentration of hydrochloric acid solution and a 0.1 *M* sodium hydroxide solution and then require you to titrate the unknown concentration of the hydrochloric acid solution with the 0.1 *M* sodium hydroxide solution.

After you have performed your practice titrations, you should graph the results (pH versus amount of base added) and identify the equivalence point on the graph.

INVESTIGATION

Procedure

For the experiment, your group will be assigned 1–2 household substances. Your group will need to determine whether or not your assigned substance has buffering activity by titrating and plotting the pH versus amount of titrant added. If a substance shows buffering activity, determine the pK_a of the substance from your titration curve and use Table 1 to predict what the buffering ingredient may be.

Suggestion: Remember that you do not know how much titrant will be needed (if any) to neutralize the buffer. It may help to do a “quick and dirty” titration with different dilutions of your test substance to find out approximately how much titrant will be needed. After you have made your estimate, you can follow up with a careful titration to get accurate data.

Data Collection and Computation

If your titration curve shows buffering activity, estimate the pK_a from your graph, and decide which of the following common buffer systems (if any) in Table 1 may be responsible for the buffering effects.

Table 1. pK_a Values for Common Acids

Acid	pK_{a1}	pK_{a2}	pK_{a3}
$C_8H_8O_3$ (methyl salycilate)	9.8		
NH_4^+	9.2	38	
H_2CO_3	6.4	10.2	
$C_6H_8O_7$ (citric acid)	3.1	4.8	5.4
$C_{20}H_{24}N_2O_2$ (quinine)	5.1	9.8	
$C_2H_4O_2$ (acetic acid)	4.8		

Acid	pK_{a1}	pK_{a2}	pK_{a3}
$C_4H_6O_6$ (tartaric acid)	3.2	4.8	
$C_7H_6O_2$ (benzoic acid)	4.2		
H_3PO_4	2.2	7.2	12.6
H_2SO_4	−12	2.0	

Argumentation and Documentation

Your group will be required to report to the class your findings, including your titration curves, calculated pK_a values, neutralization reactions between the buffer and titrant, and your reasoning as to why you believe your substance does or does not show buffering activity.

POSTLAB ASSESSMENT

1. Look at the structure for citric acid (see Figure 3). Various soda drinks will have citric acid and sodium citrate as two of their ingredients.

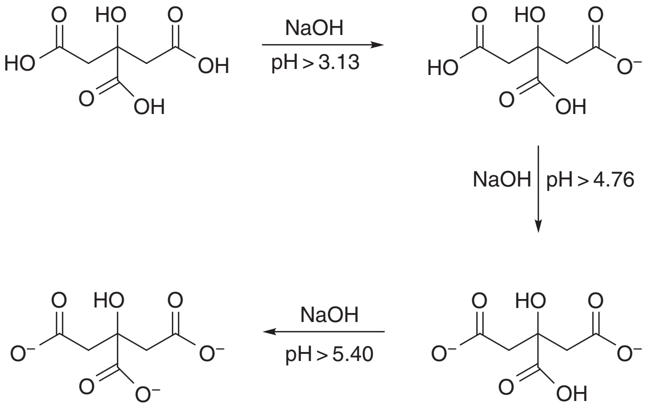

Figure 3. Citric acid has three ionizable protons.

   a. Write a net ionic equation for citric acid and sodium citrate in solution.
   b. Offer a possible function for these two chemicals in the soda when an acid or base is added to the soda.

2. Phosphoric acid and citric acid are both triprotic. The titration curve for phosphoric acid shows three distinct buffering regions (see Figure 4):

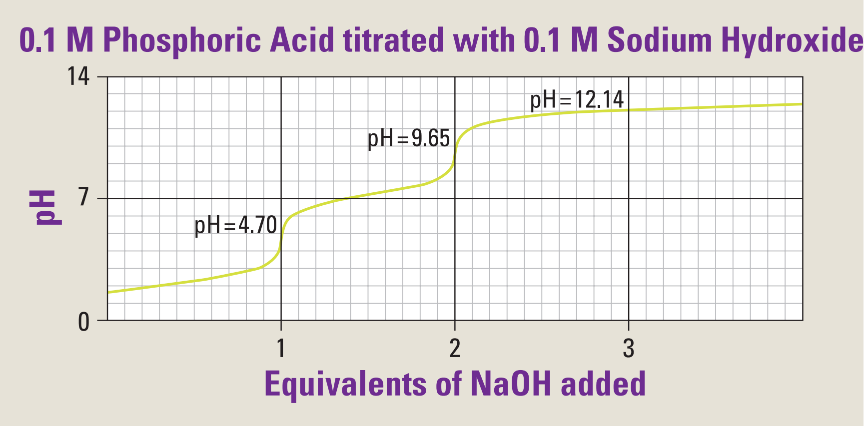

Figure 4. Titration curve for phosphoric acid

whereas the titration curve for citric acid shows only one (see Figure 5):

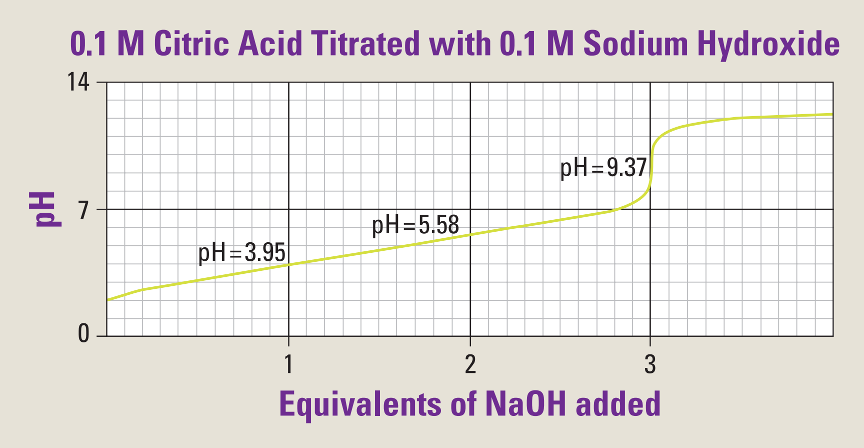

Figure 5. Titration curve for citric acid

Explain why this difference is observed. Write the neutralization reaction in net ionic form for phosphate buffer and sodium hydroxide at pH 7.5.

3. Explain the role of a buffer in an aquarium.

SUPPLEMENTAL RESOURCES

Links

Abraham, Michael, John Gelder, and Tom Greenbowe. "Molecular Level Laboratory Experiments (MoLE)." Accessed July 27, 2012. *http://genchem1.chem.okstate.edu/CCLI/CCLIDefault.html*

"Chem Guide — Buffer Solutions." Chem Guide. Accessed July 27, 2012. *http://www.chemguide.co.uk/physical/acidbaseeqia/buffers.html*

“Chem Guide — pH (Titration) Curves.” Chem Guide. Accessed July 27, 2012. *http://www.chemguide.co.uk/physical/acidbaseeqia/phcurves.html*

“Titration of Polyprotic Substances and Mixtures.” ChemBuddy. Accessed July 27, 2012. *http://www.titrations.info/acid-base-titration-polyprotics-and-mixtures*

References

Abraham, M.R. “What Can Be Learned from Laboratory Activities? Revisiting 32 Years of Research.” *Journal of Chemical Education* 88, no. 8 (2011): 1020–1025.

Eick, Charles, Lee Meadows, and Rebecca Balkcom. “Breaking into Inquiry: Scaffolding Supports Beginning Efforts to Implement Inquiry in the Classroom.” *Science Teacher* 72, no. 7 (2005): 49.

Singer, Susan R., Margaret L. Hilton, and Heidi A. Schweingruber. *America’s Lab Report: Investigations in High School Science*. Washington, DC: The National Academies Press, 2006.

INVESTIGATION 16

The Preparation and Testing of an Effective Buffer: How Do Components Influence a Buffer's pH and Capacity?

CENTRAL CHALLENGE

You and your classmates are being asked to prepare and test a series of buffers to be used in an important biochemical project. Each group will be given an assignment card for one of the buffers. Once you've performed your experimental design to achieve the assignment on the card, you will share and compare your results with those of your classmates. You will then argue, with evidence, whether you've completed your assignment or need to rely on the product of another group. Scientists often share data and information with other researchers. The need to interact with others is not an indication of failure provided your final argument is sound.

CONTEXT FOR THIS INVESTIGATION

Many important biochemical reactions occur only over a small range in pH. Living organisms dependent upon these reactions rely on chemical systems called buffers to maintain a relatively constant pH when acids or bases are added to their environment. The company contracting your chemistry class aims to produce a variety of bacteria designed to destroy harmful living agents. The bacteria must be grown in a medium with a pH similar to that of the environment they will be functioning in. This medium must be able to maintain a pH within plus or minus one unit of the target pH for the bacteria it will support when a strong acid or base is added. Your job will be to produce a buffer for such a medium.

As buffers must neutralize both acids and bases, they must contain both a base and an acid. The question is, how do we prevent the acid and base in the buffer from simply neutralizing one another, thus rendering the buffer useless? We must also consider how much acid and base our buffer should be able to neutralize. Buffers are most effective when they've been produced so they may neutralize a reasonable amount of either acid or base. The quantity of acid or base that may be added to a buffer while maintaining a relatively constant pH is a function of the buffer's capacity.

PRELAB GUIDING QUESTIONS/SIMULATIONS

1. How does a buffer solution resist a change in pH?
2. Why would HCl and NaOH be a poor choice for an acid-base pair to make a buffer?
3. Go to the animation at:
   http://introchem.chem.okstate.edu/DCICLA/pHbuffer20.html

Create a "buffer solution" using equal volumes (100 mL each) of 0.10 *M* nitric acid and 0.10 *M* sodium nitrate. Insert the probes and record the pH in the table below. Remove the probes. Then go to Part II of the simulation and add first 0.001 moles of HCl and then the same amount of NaOH. Record the pH in each case. Remove the probes, return to Part I, and try the next combinations. Complete the first three lines of the table below.

Remake the "buffers" using the same volumes of 1.0 *M* components (increased molarity by 10X). Add the same 0.001 moles of HCl and NaOH to the "stronger" buffers. Complete the middle of the table.

Finally, repeat the process adding 0.011 moles (additional 10X more) of strong acid and base. Complete the last three lines of the table.

Acid-Base Pair	pH of "buffer"	pH with 0.001 mol HCl added	pH with 0.001 mol NaOH added
HNO_3 and $NaNO_3$			
$HC_2H_3O_2$ and $NaC_2H_3O_2$			
NH_4Cl and NH_3			
Increase molarity 10X			
HNO_3 and $NaNO_3$			
$HC_2H_3O_2$ and $NaC_2H_3O_2$			
NH_4Cl and NH_3			
		pH with 0.011 mol HCl added	**pH with 0.011 mol NaOH added**
HNO_3 and $NaNO_3$			
$HC_2H_3O_2$ and $NaC_2H_3O_2$			
NH_4Cl and NH_3			

a. Write a general chemical equation to represent the equilibrium that exists in an aqueous system of the weak acid, HA as it ionizes in water. Assume the weak acid to be $HC_2H_3O_2$ (0.10 *M*), with a K_a value of 1.8×10^{-5}. What is the pH? What does the addition of $NaC_2H_3O_2$ do to the equilibrium you just represented? Use these chemical principles to explain why the first pH recorded from the simulation is so much larger than your calculated value for pure 0.10 *M* acetic acid.

   b. How do the pK_a values for acetic acid and ammonium ion ($K_a = 5.6 \times 10^{-10}$) compare to the pH values for the first two buffers in the simulation? Explain this phenomenon.

   c. Why don't the pH values change when the component concentrations are increased?

4. Write a general equation to show how a buffer containing an acid HA and the salt of its conjugate base, NaA, would respond to the addition of each of the following.

   a. The strong acid, HCl

   b. The strong base, NaOH

5. Why are the pH changes so noticeable with the last two additions of strong acid and base in the simulation?

6. Which of the two concentration combinations would be most *effective* for an antibiological agent's buffer?

EXPLANATION TO STRENGTHEN STUDENT UNDERSTANDING

A buffer is a solution designed to resist the changes in pH that occur when small amounts of acid or base are added. Many important chemical reactions occur only over a small pH range. If it were not for the presence of buffer systems in the blood of human beings, the intake of a small glass of orange juice could lead to a dangerous condition called acidosis that would lead to seizures, loss of consciousness, coma, and eventually death. Clearly buffers are very important.

You were reminded in the Context for this Investigation section that because buffers must neutralize both acids and bases, they must contain a base and an acid. In that section you were asked how we prevent the acid and base in the buffer from simply neutralizing one another, thus making the buffer useless. You may have recalled or determined during the Prelab Guiding Questions section that the answer to this problem is to use an acid and base that form a conjugate pair. A conjugate acid-base pair is a set of two species that differ from one another by one easily removable hydrogen ion (a proton). Should the members of a conjugate pair neutralize one another, they will simply reform the same species, hence there is no neutralization.

It is critical that both members of the conjugate pair are weak as the conjugate of a strong species will be so weak it will be essentially neutral and could not effectively neutralize any added acid or base. A good example of a conjugate pair for a buffer would be hydrofluoric acid (HF) and fluoride ion (F^-), as HF can neutralize added base:

$$HF(aq) + OH^-(aq) \rightarrow F^-(aq) + H_2O(l)$$

and F^- can neutralize added acid:

$$F^-(aq) + H_3O^+(aq) \rightarrow HF(aq) + H_2O(l)$$

It is also important to remember that we cannot simply obtain an independent anion (or cation) from our stockroom shelf. Rather we require a soluble salt as the source of the required ion. In the case of our sample HF/F$^-$ buffer, we might combine hydrofluoric acid with some sodium fluoride salt. In this experiment, you will combine different pairs of chemicals in an effort to create a buffer of a particular pH.

Since buffers contain weak acids, it is possible to calculate the pH as you would for any weak acid, using the K_a expression, with one significant difference: There is some common ion (actually the conjugate base of the weak acid) in the solution. Consequently, use the general weak acid equation:

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq) \text{ produces the following expression}$$

$$K_a = \frac{[H_3O^+]\,[A^-]}{[HA]}$$

As there is already conjugate base, A$^-$, present in the solution, when equilibrium concentrations are inserted into the expression, the $[H_3O^+]$ will not be equal to the $[A^-]$.

Buffers are most effective when they are equally prepared to neutralize acids or bases. For this to be true, the [HA] should approximately equal the $[A^-]$. Examining the expression for the general buffer equation above, we see that such a circumstance would result in the K_a of the conjugate acid in the buffer being equal to the $[H_3O^+]$. The negative logarithmic form of this equation indicates that the pK_a is equal to the pH of the buffer. In summary, when $[HA] = [A^-]$:

$$K_a = [H_3O^+] \text{ and } pK_a = pH$$

The effectiveness of a buffer — that is, its ability to neutralize added base or acid — is called the buffer's capacity.

PREPARATION

Materials

0.10 *M* acetic acid (CH_3COOH) $K_a = 1.8 \times 10^{-5}$	Sodium acetate ($NaCH_3COO$)	0.20 *M* sodium hydroxide (NaOH)	2–150 mL beakers	Stirring rod or magnetic mixer (use care around pH electrodes)
0.10 *M* ammonia (NH_3)	Ammonium chloride (NH_4Cl) K_a of NH_4^+ = 5.7×10^{-10}	0.20 *M* hydrochloric acid (HCl)	2–250 mL beakers	pH meter (or pH probe with an interface) or pH Hydrion Paper

0.10 *M* sodium dihydrogen phosphate (NaH_2PO_4) K_a of $H_2PO_4^-$ = 6.3×10^{-8}	Sodium hydrogen phosphate (Na_2HPO_4) (also called sodium biphosphate)	Milligram (or centigram) balance	2–50 mL burets	
0.10 *M* citric acid ($H_3C_6H_5O_7$) K_a = 7.1×10^{-4}	Sodium dihydrogen citrate ($NaH_2C_6H_5O_7$)	Scoopula or spatula	2–100 mL volumetric flasks	
0.10 *M* sodium monohydrogen citrate ($Na_2HC_6H_5O_7$) (also called sodium bicitrate) $K_a = 4.1 \times 10^{-7}$	Sodium citrate ($Na_3C_6H_5O_7$)	100 mL graduated cylinder	2 buret clamps	
Sample pH 7 buffer	Sodium chloride (NaCl)	10 mL graduated cylinder	Utility stand	

Safety and Disposal

Be sure to check the appropriate Material Safety Data Sheets (MSDS) for all chemicals before beginning the laboratory.

Splash-proof safety goggles should be worn as when handling any chemical in the lab. Aprons or lab coats should also be worn if available.

Neutralize acids and bases with the corresponding bases and acids before flushing down the sink. Neutralized solutions should be preceded and followed by large quantities of water during flushing. These disposal procedures should be followed unless your local state/jurisdiction directs you otherwise.

PRACTICE WITH INSTRUMENTATION AND PROCEDURE

Procedure

1. Rinse a buret with 10 mL or so of 0.20 *M* HCl. Fill the buret and clamp it to the utility stand. Repeat this process with a second buret and the 0.20 *M* NaOH.
2. Set up and calibrate your pH meter according to your teacher's instructions.
3. Place 50.0 mL of water into each of two 150 mL beakers. Lower the pH meter's electrode or probe into one of the beakers. Determine the pH. Record this value neatly in a prepared table in your lab notebook.
4. Run in as close as possible to 1.00 mL of HCl solution, stirring the solution in the beaker thoroughly. Do not use the electrode to stir. Record the new pH value.

5. Repeat Steps 3 and 4 with the second beaker adding 1.00 mL of the NaOH solution.
6. Finally, safely discard the solutions and fill the two beakers with 50.0 mL each of the buffer. Repeat Steps 3–5 using the buffer solutions instead of the water.

Give an explanation for what you've observed.

INVESTIGATION

Your teacher will provide you with a card that indicates your mission. Consider the list of materials available and hypothesize how best to fulfill your mission using the materials available. The list includes a number of chemicals that are not required to complete your task.

Procedure

The company contracting your class has requested that each group prepare a 100 mL sample of their buffer and test it in the hope that it falls within +/- 0.5 of the target pH stated on the "Mission Card" and that a 50 mL sample can maintain a relatively constant pH (within one pH unit of the initial value) with the addition of up to 20 mL of 0.20 *M* HCl or NaOH.

1. Neatly record your hypothesis in an "If we combine … then our buffer will …" format. The procedure, including the masses and volumes of all materials used, should be neatly recorded in your lab notebook. Use a data table for all quantitative values. Be sure to show all supporting calculations in an adjacent column or section of your lab report. Have your teacher check your hypothesis before you begin to create your buffer and test it.
2. Once your teacher has given you permission (you may have to modify your hypothesis before you are allowed to proceed), go ahead and test your hypothesis. First check your target pH by placing 50.0 mL of your buffer into a 250 mL beaker and lowering the pH meter's electrode or probe into the sample. Record the initial pH in the first line of a table with columns headed "Volume of Acid Added" and "pH."
3. Each lab group must follow the same procedure for capacity testing. Measure a second 50.0 mL portion of your buffer. Why might your total volume of buffer slightly exceed 100 mL? All measurements involve some degree of uncertainly. As the company is allowing you to be within plus or minus 0.5 units of your target pH, do you think a slight increase in volume is significant in this case?
4. Begin your capacity testing by using the buret from the Practice with Instrumentation and Procedure section to add about 5.00 mL (record the actual ongoing volume precisely) increments of 0.20 *M* HCl to your buffer. Use your table to record the pH after each addition (followed by thorough stirring). Once about 10.00 mL of acid has been added, change to approximately 2.00 mL increments until you reach a total volume of 16.00 mL of acid added. Now add the acid in approximately 1.00 mL increments until you've reached a total volume of 30.00 mL of acid added. You may now increase the volume of each increment to about 2.00 mL until you reach a total of 40.00 mL added. The last couple of increments may be about 5.00 mL each.

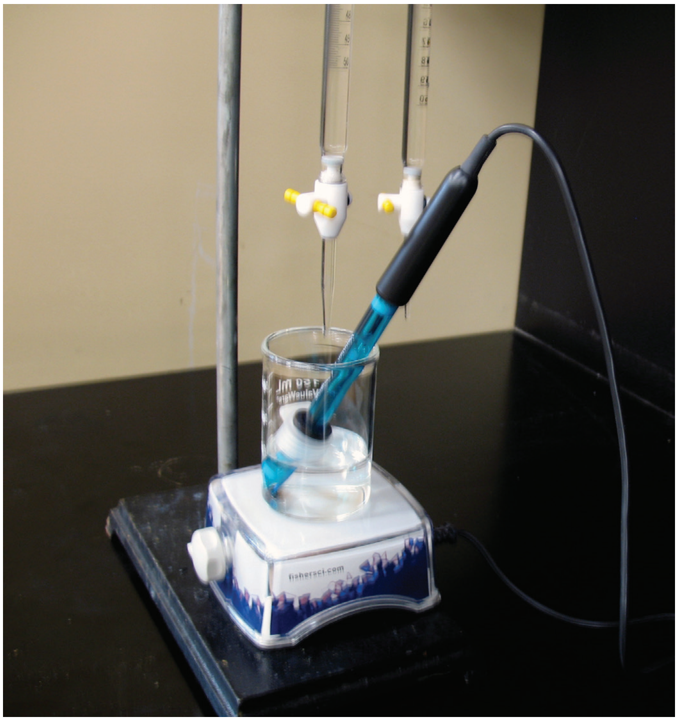

Figure 1. pH probe, hot plate, and buret setup

5. Repeat your capacity testing for base addition by following the same procedure with the addition of 0.20 *M* NaOH to your second 50.0 mL sample of buffer. The volumes added should follow the same pattern as before, but base is being added. These precise volumes and pH values may be recorded in two new columns added to the data table you've already prepared.
6. Two graphs will be plotted, one of "pH versus volume of acid added to a buffer" and one of "pH versus volume of base added to a buffer."
7. All student data and graphs will be displayed. You will share and compare your results with those of your classmates. You will then argue with evidence whether you've completed your assignment or need to rely on the product of another group. Be sure to give credit where it is due and cite the source of any ideas or information you've borrowed.

Data Collection and Computation

1. Complete neatly labeled graphs for the data tables you've prepared.
2. Calculate the anticipated pH for the buffer you were assigned. Include neat calculations in your report.
3. Determine the volume of acid and base your buffer was able to neutralize before allowing a pH change of more than one unit. Use the volume with the HCl/NaOH molarities to determine the number of moles of acid and base your buffer was able to neutralize.

4. Consider 50.0 mL of the buffer you prepared and calculate the ratio of (moles of added acid neutralized) to (moles of base component in the buffer). Do the same calculation for (moles of added base neutralized) to (moles of acid component in the buffer). State both ratios in the form of x:1. Report the value for x to two significant figures.

Argumentation and Documentation

Examine classmates' graphs. Was each buffer more effective for buffering added acid or base?

POSTLAB ASSESSMENT

1. Suppose, during preparation, an additional 10 mL of distilled water was added to your buffer by mistake.
   a. What effect would this have on your buffer's pH? Explain.
   b. Would this affect your buffer's capacity? Explain.
2. Given a solution of hydrocyanic acid (HCN), what additional reagent or reagents are needed to prepare a buffer from the hydrocyanic acid solution?
   a. Explain how this buffer solution resists a change in pH when moderate amounts of strong acid are added. Use a chemical equation in your explanation.
   b. Explain how this buffer solution resists a change in pH when moderate amounts of strong base are added. Use a chemical equation in your explanation.
3. A buffer solution contains 0.20 moles of methanoic acid, HCOOH, and 0.30 moles of sodium methanoate, NaCOOH, in 1.00 L of the buffer. The acid ionization constant, K_a, of methanoic acid is 1.8×10^{-4}.
   a. Calculate the pH of this solution.
   b. Compare the capacity of this buffer to neutralize added acid to its capacity to neutralize added base. Explain your answer completely.
   c. If 0.10 moles of HCl gas solution were bubbled through a liter of the buffer, what would happen to the pH? How would this addition affect the buffer's capacity to neutralize added acid and base in the future? Answer the question fully, including equations and calculations where necessary.

SUPPLEMENTAL RESOURCES

Links

Abraham, Michael, John Gelder, and Tom Greenbowe. "Buffered Solution." Molecular Level Laboratory Experiments (MoLE). Accessed August 16, 2012. *http://introchem.chem.okstate.edu/DCICLA/pHbuffer20.html*

“Chem Guide — Buffer Solutions.” Chem Guide. Accessed August 15, 2012. *http://www.chemguide.co.uk/physical/acidbaseeqia/buffers.html*

“Environmental Science Projects: Learn About the Environment.” Science Made Simple. Accessed August 16, 2012. *http://www.sciencemadesimple.com/environmental_projects.html*

References

Abraham, Michael. “What Can Be Learned from Laboratory Activities? Revisiting 32 Years of Research.” *Journal of Chemical Education* 88, no. 8 (2011): 1020–1025.

Meadows et al. “Breaking into Inquiry.” *The Science Teacher* 72, no. 7 (2005): 49–53.

Singer, Susan R., Margaret L. Hilton, and Heidi A. Schweingruber. “America’s Lab Report: Investigations in High School Science.” National Research Council, 2005.

APPENDIX A

DO NOT DETACH FROM BOOK.

PERIODIC TABLE OF THE ELEMENTS

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.97	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.95	43 **Tc**	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57–71 *	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po**	85 **At**	86 **Rn**
87 **Fr**	88 **Ra**	89–103 †	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112 **Cn**	113 **Nh**	114 **Fl**	115 **Mc**	116 **Lv**	117 **Ts**	118 **Og**

*Lanthanoids	57 **La** 138.91	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm**	62 **Sm** 150.36	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
†Actinoids	89 **Ac**	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

APPENDIX B

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL = liter(s), milliliter(s)
g = gram(s)
nm = nanometer(s)
atm = atmosphere(s)
mm Hg = millimeters of mercury
J, kJ = joule(s), kilojoule(s)
V = volt(s)
mol = mole(s)

ATOMIC STRUCTURE

$$E = h\nu$$
$$c = \lambda\nu$$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$$K_c = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b}, \text{ where } a\,\mathrm{A} + b\,\mathrm{B} \rightleftarrows c\,\mathrm{C} + d\,\mathrm{D}$$

$$K_p = \frac{(P_\mathrm{C})^c(P_\mathrm{D})^d}{(P_\mathrm{A})^a(P_\mathrm{B})^b}$$

$$K_a = \frac{[\mathrm{H}^+][\mathrm{A}^-]}{[\mathrm{HA}]}$$

$$K_b = \frac{[\mathrm{OH}^-][\mathrm{HB}^+]}{[\mathrm{B}]}$$

$$K_w = [\mathrm{H}^+][\mathrm{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\mathrm{C}$$
$$= K_a \times K_b$$

$$\mathrm{pH} = -\log[\mathrm{H}^+], \ \mathrm{pOH} = -\log[\mathrm{OH}^-]$$

$$14 = \mathrm{pH} + \mathrm{pOH}$$

$$\mathrm{pH} = \mathrm{p}K_a + \log\frac{[\mathrm{A}^-]}{[\mathrm{HA}]}$$

$$\mathrm{p}K_a = -\log K_a, \ \mathrm{p}K_b = -\log K_b$$

<u>Equilibrium Constants</u>

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$$[\mathrm{A}]_t - [\mathrm{A}]_0 = -kt$$

$$\ln[\mathrm{A}]_t - \ln[\mathrm{A}]_0 = -kt$$

$$\frac{1}{[\mathrm{A}]_t} - \frac{1}{[\mathrm{A}]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_{\text{A}}, \text{ where } X_{\text{A}} = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_{\text{A}} + P_{\text{B}} + P_{\text{C}} + \ldots$$

$$n = \frac{m}{\boldsymbol{M}}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE_{\text{molecule}} = \frac{1}{2}mv^2$$

$$\text{Molarity, } M = \text{moles of solute per liter of solution}$$

$$A = \varepsilon bc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
ε = molar absorptivity
b = path length
c = concentration

Gas constant, R = 8.314 J $\text{mol}^{-1}\,\text{K}^{-1}$
= 0.08206 L atm $\text{mol}^{-1}\,\text{K}^{-1}$
= 62.36 L torr $\text{mol}^{-1}\,\text{K}^{-1}$

1 atm = 760 mm Hg = 760 torr

STP = 273.15 K and 1.0 atm

Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\begin{aligned} \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ \\ &= -RT \ln K \\ &= -nFE^\circ \end{aligned}$$

$$I = \frac{q}{t}$$

$$E_{cell} = E^\circ_{cell} - \frac{RT}{nF}\ln Q$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)
Q = reaction quotient

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

APPENDIX C

Directions for Using Various Spectrophotometers

Making a Measurement with the SPECTRONIC 20D+

1. Turn on the instrument and allow it to warm up and stabilize for 15 minutes. Select %T as the data mode. Check with your instructor to determine whether your SPEC 20 can use square cuvettes. Obtain a square plastic cuvette or glass test-tube cuvette, as appropriate, to use in your experiments. If using a test-tube cuvette, use a pen to place a 1 cm vertical line near the top. Set the wavelength you wish to measure. If the instrument has a filter (your teacher will point this out if it applies), make sure that it is set to the correct wavelength range.
2. Turn the on/off/0%T knob to set the %T reading to zero.
3. Add liquid to the cuvette until there is ~3 cm of liquid in the bottom (4 cm for test tubes). If plastic transfer are available, use one. The exact liquid level in the cuvette is not important if it is approximately 3 cm. Do not waste solution or risk spills by overfilling the cuvette.
4. Place the cuvette in the sample compartment of the SPEC 20. If using a plastic cuvette, the clear sides should be on the right and left and you will need to mount it in a plastic caddy then slide the caddy into the instrument. If using a test-tube cuvette, place it so that the mark faces to the right.
5. After the warm-up period, follow Steps 3 and 4 using water. Zero the instrument by turning the 100%T knob until the %T readout shows 100.0. Set the data mode to Absorbance after you have zeroed it. NOTE: If you change the measurement wavelength you will need to reset 0%T and 100%T.
6. For each subsequent measurement, empty and rinse your cuvette, shaking out as much of the rinse water as possible. When obtaining samples, never return excess solution to the stock bottle. Pour all waste or excess into a waste beaker or down the sink. Follow Steps 3 and 4 to measure each sample. Record readings in a table in your laboratory notebook

Making a Measurement with the SPECTRONIC 200

1. Turn on the instrument and allow it to complete its startup sequence. Let the instrument warm up and stabilize for 15 minutes. Set up the experiment you want to perform in the spectrophotometer software — your teacher will show you how. Obtain a square plastic cuvette or glass test-tube cuvette to use in your experiments. If using a test-tube cuvette, use a pen to place a mark near the top.
2. Add liquid to the cuvette until there is ~3 cm of liquid in the bottom (4 cm for test tubes). If plastic transfer are available, use one. The exact liquid level in the cuvette is not important if it is approximately 3 cm. Do not waste solution or risk spills by overfilling the cuvette.

3. Place the cuvette in the sample stage of the SPEC 200. If using a plastic cuvette, the clear sides should be on the right and left. If using a test-tube cuvette, place it so that the mark faces to the right.
4. After the warm-up period, follow Steps 2 and 3 using water. Zero the instrument by pressing the autozero button.
5. For each subsequent measurement, empty and rinse your cuvette, shaking out as much of the rinse water as possible. When obtaining samples, never return excess solution to the stock bottle. Pour all waste or excess into a waste beaker or down the sink. Follow Steps 2 and 3 using your sample.

If making measurements in scan mode, you will need to press the Enter button to make a measurement. If measuring in live display mode, the measurement will update automatically every few seconds.

Making a Measurement with the SPECTRONIC 20 with an Analog Meter

Your teacher will show you how to read the scale in both transmittance and absorbance units.

1. Turn on the instrument and allow it to warm up and stabilize for 15 minutes. Obtain a glass test-tube cuvette to use in your experiments. Use the same test tube for all measurements. If your tube doesn't already have one, place a 1 cm vertical line near the top with a Permanent. Set the wavelength you wish to measure. If the instrument has a filter (your teacher will point this out if it does), make sure that it is set to the correct wavelength range.
2. With nothing in the sample compartment, turn the on/off/0%T knob to set the reading to zero%T.
3. Add liquid to your test tube until there is ~4 cm of liquid in the bottom. If plastic transfer are available, use one. The exact liquid level in the cuvette is not important if it is approximately 4 cm. Do not waste solution or risk spills by overfilling the tube.
4. Place the tube in the sample compartment of the SPEC 20. Orient the tube so that the mark faces to the right. There should be a line on the lip of the sample compartment — line your mark up with this line.
5. Follow Steps 3 and 4 using water. Zero the instrument by turning the 100%T knob until the readout shows 100.0%T. **NOTE: If you change the measurement wavelength you will need to reset 0%T and 100%T**.
6. For each subsequent measurement, empty and rinse your cuvette, shaking out as much of the rinse water as possible. When obtaining samples, never return excess solution to the stock bottle. Pour all waste or excess into a waste beaker or down the sink. Follow Steps 3 and 4 to measure each sample. Record readings in a table in your laboratory notebook

Since the SPEC 20 cannot scan, you will need to measure at individual wavelengths and then make a plot. Your teacher will assign a number of wavelengths to your group to measure, and you will combine your data with data from the rest of the class to allow everyone to plot the complete spectrum of the sample. When connecting the dots to make the scan curve, do not use straight lines. Try to draw a single, smooth, curving line that passes through all the points.

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