“More Experiments with Pretty Good Chemicals”

GENERAL CHEMISTRY 1B

LABORATORY ACTIVITIES

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CCChemTeach.com (http://ccchemteach.com) for community college chemistry
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SAFETY IN THE CHEMISTRY LABORATORY

The experiments presented in this manual have been designed with your safety in mind. Nevertheless, whenever you work in a chemistry laboratory, potential hazards exist. However, a knowledge of the most common sources of hazard, as well as the safety precautions routinely observed in the laboratory, will help to avoid any serious accidents.

Safety Equipment

The safety equipment listed below are found in our chemistry laboratories. You should know the location of each piece of safety equipment and how to use it.

1. Eye Protection and Safety Glasses/Goggles

“Safety glasses are impact resistant lenses that protect the eyes from blows or other injury” (http://medical-dictionary.thefreedictionary.com/safety+glasses). Since our eyes and eyesight are precious, wearing safety glasses or safety goggles will protect your eyes from lab hazards when you are in the laboratory, even if you are not doing any experimental work. Hazards include chemicals splashing out of containers, glassware shattering upon heating, and test tubes flying out of a centrifuge. Safety glasses/goggles should be worn at all times when you are in the laboratory.

Note: (i) you are responsible for bringing your own pair of safety goggles/glasses to lab. (ii) In 1998, the American Chemical Society (ACS) made the following recommendation regarding contact lenses (http://pubs.acs.org/cen/safety/19980601.html): “contact lenses can be worn in most work environments provided the same approved eye protection is worn as required of other workers in the area.”

See also http://www.snopes.com/horrors/techno/cornea.asp

2. Eyewash Fountain

An eyewash fountain is a water fountain with two faucets directed at one another. When the eyewash fountain is turned on with your head an appropriate distance from the fountain, the two faucets flush water into both eyes. It is unlikely that chemicals will get in your eyes if you are wearing safety glasses.

However, if chemicals should get in your eyes, go immediately to the eyewash fountain and flush them for 15 minutes to wash the chemicals out of the eyes. Always report such an accident to your instructor, who may wish to have you see a doctor.

3. Fire Extinguisher

Fire extinguishers are classified and chosen based on the type of fire (http://www.fire-extinguisher101.com/). For example, a water extinguisher is suitable for Class A fires that involve ordinary combustible materials, such as paper and wood, but not for Class B fires that involve flammable or combustible liquids, like gasoline. Carbon dioxide and dry chemical extinguishers containing sodium bicarbonate or potassium bicarbonate are used for Class B and C fires (Class C fires involve electrical equipment). Class D fire extinguishers contain a dry powder such as sodium chloride or graphite (http://en.wikipedia.org/wiki/Fire_extinguisher#Class_D) and are used on combustible metals.

4. Fire Blanket

A fire blanket is a sheet of fire retardant material that is used to extinguish small fires. A fire blanket can be used to wrap a victim who has caught fire. Use the Stop, Drop, and Roll technique to smother the fire.

5. Safety Shower

A safety shower is an emergency shower that is designed to deluge continuously at 30-60 gallons per minute for at least 15 minutes (http://www.answers.com/topic/shower). The safety shower is found next to the eyewash fountain. If a large quantity of a hazardous chemical has spilled on a person, use the safety shower to flush large quantities of water on the victim. Usually, clothing needs to be removed for the water to reach the victim’s skin. Stay under the shower for at least 15 minutes to wash off the chemical

6. Fume Hood
A fume hood is a laboratory bench having a fan that will carry fumes out of the laboratory into the open air above the building. The fume hood is used to perform experiments that produce toxic fumes. Your laboratory instructor will direct you to carry out experiments in the fume hood. However, if you are doing an experiment that is producing an obnoxious or choking odor in the open lab, do not wait for your instructor and take your work under the hood. If you know you have a sensitivity to a chemical that is being used in lab that day, inform your instructor so you can work in the hood.

7. First-Aid Kit
A first-aid kit is located either in the lab or the prep room. This kit contains bandages, burn spray, antiseptic spray, cold spray, and other items. Always report any injury to your instructor that requires the first-aid kit, since follow-up measures may be needed.

Miscellaneous Hazards
The chemistry laboratory is a safe place to work as long as you and your co-workers are aware of the various hazards in the laboratory and follow lab safety rules and regulations.

1. People
Our chemistry lab has a capacity of 27 students. With so many people in the lab, it is easy to bump into another person or trip over a chair while moving about the lab. Focus on what you are doing but be aware of your surroundings and what other people are doing. You may be practicing lab safety but if another person standing next to you is not handling a chemical properly, you may be inadvertently involved in an accident.

2. Broken Glass
In the chemistry lab, we will use glassware, such as beakers to prepare hot or cold water baths, graduated cylinders to measure substances, and flasks to carry out chemical reactions. For many experiments, you will have to assemble several pieces of equipment and monitor your experiment from start to finish. Accidents occur when something tips over and glassware breaks. Use a broom and dust pan or wet paper towels to clean up the broken glass. Dispose of the broken glass in the broken glass container.

3. Fires
In the chemistry lab, we will use Bunsen burners, flammable liquids, and perform chemical reactions that generate heat. If something or someone catches on fire, act immediately and use either water, a fire extinguisher, the safety shower, or fire blanket to extinguish the fire. The method you use depends on the type of fire. See the section above on Fire Extinguisher.

4. Chemical Spills: Acids, Bases, and Other Caustic Chemicals
If you spill a small amount of chemical on a small area of your body, like your finger, flush the exposed area for 15 minutes with tap water from a sink. If a burning sensation accompanies the spill, flush the exposed area with water and report it immediately to your instructor. Some chemical burns begin with only a minor burning sensation, but develop into a more serious injury if not treated promptly. Your instructor will be able to recommend further action or send you to a doctor if the burn seems serious. If you spill a large amount of chemical over a large area of your body, use the safety shower. See the section above on Safety Shower.

5. Diluting Concentrated Acids
When preparing a dilute acid solution from a concentrated acid solution, always add the acid to water (“when you’re doing what you oughter, add the acid to the water”). If water is added to concentrated acid, the solution will become hot and acid may spatter on you.
6. Spattering from Test Tubes
Spattering may occur when heating liquids in a test tube. To minimize the danger of spattering, heat the test tube near the liquid surface, and agitate the contents to and fro. Never point a test tube being heated toward you or another person. Be aware of your surroundings and what other people are doing.

7. Flame-Drying Glassware
The glass beakers and flasks are designed to withstand the heat of your Bunsen burner. However, certain pieces of glassware, such as graduated cylinders, burets, volumetric flasks, and pipets, should never be heated with a burner, as they are likely to shatter. Hot glass looks the same as cold glass so be careful touching or approaching glass that someone else is using.

8. Inserting Glass Tubing in Stoppers
The Chemistry Stock Room has an assortment of glass tubing in stoppers that you can use. However, if you need to insert glass tubing into a rubber or cork stopper, make sure the hole is the proper size for the glass tubing and use glycerol (glycerin) or soap as a lubricant. Hold the glass near the end being inserted, and twist the glass into the hole. Never force a piece of glass tubing into a hole. The glass may snap, and the jagged edges on the broken glass can cause a serious cut.

9. Detecting Odors
If your lab instructor directs you to smell a chemical, do not place your nose directly over a container and inhale deeply. Hold the container away from your nose and use your hand to waft the odors gently toward your nose. Partially fill your lungs with air before inhaling the odors to avoid over-inhalation of the fumes. See the Material Safety Data Sheet (MSDS) of the substance for more information.

10. Tasting
Never taste chemicals prepared in a chemistry laboratory unless specifically directed to do so by your instructor. Many chemicals are toxic or hazardous to our health. Your equipment have been cleaned but still may have trace amounts of toxic or hazardous chemicals. See the Material Safety Data Sheet (MSDS) of the substance for more information.

11. Horseplay
The laboratory is no place for horseplay, since there is always the danger of breaking or spilling something. While a relaxed atmosphere is the most conducive for productive lab work, fooling around in the laboratory is an invitation for a serious accident.

General Laboratory Procedures and Conduct

The following chemistry laboratory safety procedures apply to everyone (instructors, students, and staff) using the chemistry laboratory. Disregard of these procedures will result in disciplinary action.
1. Protective goggles or safety glasses with side shields must be worn at all times in the lab.

2. Learn the locations and the use and operation of the fire extinguishers, safety shower, eyewash fountain, fire blankets, fume hoods, and first aid kit. Learn the location of the fire alarm.

3. Learn the primary, secondary, and handicapped escape routes from the laboratory in case of fire, earthquake, or other disaster. A map of the escape route from the lab is posted next to the hall door.

4. Learn the use and operation of laboratory equipment and instruments. A diagram of laboratory equipment is shown below.

5. Read chemical labels carefully. Be sure you are using the chemical required. Put the cap or lid back on the bottle. Clean up any spills.

6. Never return unused chemical to the stock bottle to avoid contamination.
7. Dispose of chemicals in the appropriate waste container. Never discard solid residues or paper into the sinks.

8. Never perform unauthorized experiments.

9. Eating, drinking, and smoking in the laboratory are forbidden. Do not bring food or drink into the laboratory. You may eat or drink in the hallway outside of the lab.

10. Never taste a chemical.

11. If instructed to smell a chemical, do so by gently wafting the vapors toward your nose.

12. When diluting, ALWAYS add acid to the water.

13. Never point a test tube that is being heated toward you or others.


15. Long pants are recommended. Footwear should cover the feet completely. No open-toe shoes. Long hair and loose clothing should be secured.

16. At the end of each lab period or when you have finished an experiment, wipe and clean your lab bench area and the balance room, clean and dry equipment; account for and put away the equipment in your locker, and lock your locker. Return all community equipment, e.g., ring stands and hot plates, to their proper places. Dispose of chemicals in the proper waste container.

**Accidents**
1. Clean up all spills or breakages immediately. Dispose of broken glass in the broken glass container. If a mercury thermometer breaks, do not touch the mercury. Notify lab staff immediately.

2. In case of contact with a chemical, wash the affected area immediately and thoroughly with water. Notify lab staff.

3. In case of an injury, no matter how minor, notify lab staff.

**Lab Policies**
1. Safety glasses or goggles are required in lab. Prescription glasses are an adequate substitute for safety glasses/goggles. For students who wear contact lenses, you will need to wear safety glasses/goggles over your contact lenses. Try to be aware of your safety as well as the safety of others in lab.
2. FAILURE TO CHECK-IN YOUR LOCKER, whether you drop the course or complete it, results in a $25 LAB FEE plus a charge for any broken or missing equipment.
3. All labs must be performed to pass this course.
4. Late lab assignments will be penalized 5% per calendar day.
5. The chemistry lab has 12 computers.
   a. You cannot store your lab data and results on the hard drive of a computer you are using. Please bring a floppy disk or flash/thumb drive to store lab files.
   b. Each computer is connected to a network printer. You will need to supply your own printer paper. You and your lab partner are asked to donate one ream of paper to lab (you and your lab partner can share the cost of paper) for your use and other student’s to use.
   c. These computers are connected to the internet so you can look up scientific information. Please do not download images, files, or software onto these computers.
Laboratory Safety Agreement

I have carefully read the instructions on good laboratory safety practices and procedures. I understand the importance of good safety practices for my own welfare and of all people in the laboratory and I, therefore, pledge to follow the safety regulations of the college.

Date: ____________________    Signature: ____________________

Drawer Number: ____________    Print Name: ________________
The ability to work with people is an important skill that many employers value. A good group or team is able to share their diverse experiences, knowledge, abilities, and opinions to work effectively and efficiently to accomplish goals that one person may not be able to do as well or as quickly. Group or team work means members work together in a non-competitive, collaborative atmosphere. Skills include listening to others, being assertive with your input but not dominating the whole group, and taking responsibility for your role on the team and making sure other members are doing their role. It helps to focus on the “big picture”, i.e., the overall goal of the group, rather than getting caught up in individual issues.

For most of the labs, you will work with a lab partner in a group of two. For other labs, you will work in a group of four. Working in a larger group requires teamwork and communication. Each group member will be assigned one of the following roles so that duties are shared equally:

**Group Leader**: responsible for supervising the group and makes sure each member contributes equally to the team.

**Communicator**: responsible for communicating with the instructor and for completing all materials to be submitted by the team that reflects the thinking of all team members.

**Record Keeper**: responsible for keeping records of all materials discussed and is for informing absent team members of work missed and progress made.

**Counselor**: responsible for making sure all members of the team agree on planning, execution, and presentation of work.

Roles should be rotated with each different lab so each member of the group has the opportunity to perform a different function.
Critical and Analytical Thinking in Science: Data Analysis

Learning how to analyze data to get meaningful results and draw conclusions from those results is one very important objective in this course. One way to learn and develop critical thinking skills is to try to prove or disprove a hypothesis. According to the American Heritage Dictionary, New College Edition, Houghton Mifflin, 1979, a hypothesis is “an assertion subject to verification or proof, as:

a. a proposition stated as a basis for argument or reasoning.
b. a premise from which a conclusion is drawn.
c. a conjecture that accounts, within a theory or ideational framework, for a set of facts and that can be used as a basis for further investigation.”

In each lab, you will be presented with one or more problems to solve. For each problem, there is a hypothesis for you to prove or disprove. In general, you will prove or disprove a hypothesis by doing the following:

1. collect data from an experiment that is relevant to the hypothesis;
2. analyze the data, usually by calculations using chemical principles, to obtain results;
3. interpret your experimental results to determine the validity of the hypothesis.

Using a hypothesis makes you focus on the important experimental variables and helps you understand how chemical principles are used in the experiment. To help you analyze data, you can represent your data and results in the form of tables or graphs or both to help you see trends and patterns in the data and results. We will also be collecting data by computer to enable you more time to analyze data in lab. Also, you will learn to find information from various information sources and use that information to help you prove or disprove hypotheses and draw conclusions.

Representing Data and Results in Tables and Spreadsheets and with Graphs

In each lab, you will be making observations - some qualitative and some quantitative. For the quantitative observations, you will attach numbers to your observations. You will analyze your numerical data by doing calculations to get results. Then, you will interpret your results, i.e., figure out what these numbers mean, to draw conclusions about your experiment. You will record all of this information in your lab notebook.

Tables

You will want to organize your data and results in a way that makes it easy for you to read and interpret as well as another person who looks at your lab notebook. Preparing a table is an excellent way to organize your data and results. In general, a table will show data in columns (as opposed to rows). For example, see Table 1 on p. 12 of this Lab Manual. Note that this table consists of two columns: Measuring Devices and Uncertainty. Each device is clearly listed with its corresponding uncertainty.

When you prepare a table, do the following:

1. Give the table a specific number and title, e.g., “Table 1. Uncertainties of Various Measuring Devices” is a good title whereas “Devices” is not specific and doesn’t state what information is shown in the table.
2. Label each column (or row) with the appropriate units. Use the appropriate number of significant figures.
3. Make your table legible.

Spreadsheets

In the lab, we will prepare tables using Microsoft Excel spreadsheets. You can program Excel to do calculations that help you analyze data. For example, you can multiply the number in one cell (A1) by a number in another cell (B1) to give you the product in a third cell (C1). Of course, you can also do this on your calculator. However, the advantage of Excel becomes evident when you have a big set of data. If you have the same calculation to do 10 times, you can have Excel to do it for you very quickly.

<table>
<thead>
<tr>
<th>Cell</th>
<th>A</th>
<th>B</th>
<th>C</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>3</td>
<td>5</td>
<td>15</td>
</tr>
</tbody>
</table>

After you prepare your Excel spreadsheet/table, print it out and staple it into your notebook.
Graphs

Often, you will want to show the data you have organized in a table in a graph. A graph is another excellent way to represent data. A graph can show trends in your data or results that you may not see from a table. The data that is represented in a graph can give you a result. For example, in Lab 1, you will graph mass of water on the y-axis and volume of water on the x-axis. The slope of this line gives you the result you are looking for. The slope includes data from all of your data points. Another way to do this is to individually calculate the mass/volume and then take the average. However, the graph gives you this information quickly and also gives you information about the error in your experiment.

For every graph that you do, do the following:

1. Give your graph a specific number and title, e.g., “Graph 1. Graphical determination of ______ from mass and volume.”
2. Label each axis with the appropriate quantity and show the units.
3. Label the divisions on your graph paper so that an individual, e.g., your instructor, who attempts to interpret the graph, may easily interpret each data point. Furthermore, the divisions should be laid out so that the graph covers most of the graph paper.
4. Usually, you will not want to just show data points. You may draw a smooth curve through your data points (do not connect the data point) or fit a straight line (with a linear regression analysis) through your data points.

You can graph your data using the Graphical Analysis software. You enter your data in a table and this software automatically graphs the data for you. You can do a linear regression using this software or fit your data to other math functions. Then, you can print out your graph and staple it into your notebook.

Computer-Assisted Data Acquisition

You will acquire data and take measurements using a computer for several experiments using Vernier Science and Technology hardware and software. You will take temperature, conductivity, pH, and pressure readings using the appropriate sensors and probes. To make these measurements, you will need the following equipment:

1. Computer with Vernier LoggerPro software
2. Vernier LabPro hardware interface
3. AC power adapter (output: 6 V DC at 600 mA) for the LabPro.
4. USB cable that connects LabPro interface to the computer
5. Appropriate probe that connects to LabPro interface

Your lab instructor will help you connect the probe to the computer.

Open the Vernier folder (which should be on the computer desktop), and double click on the LoggerPro icon. The interface and software automatically senses the probe that is connected to the computer.

You will also use the Ocean Optics Chem 2000 UV-VIS spectrophotometer to measure absorption and emission spectra (the amount of light absorbed or emitted by a sample at each wavelength). Each spectrophotometer is connected to a computer. To use the Chem 2000, you will need:

1. UV-VIS (deuterium tungsten) light source and integrated cuvette holder.
2. Spectrophotometer detector.
3. Fiber optic cable that connects light source and holder to the spectrophotometer detector.
4. Serial cable that connects the light source and holder to the spectrophotometer detector.
5. USB cable that connects the spectrophotometer detector to the computer.
6. AC power adapter (output: 12 V DC at 1.25 A) for the UV-VIS light source.
7. A cuvette for your sample.

Your lab instructor will help you connect the Chem 2000 to the computer.

Double-click on the OOIC icon on the desktop (if the OOIC icon is not on the desktop, it will be in the Ocean Optics folder) to open the Ocean Optics software to collect spectra.
Chemistry References
There are many reference sources in science and chemistry. Here are a few common references that are available in the chemistry lab:

1. CRC Handbook of Chemistry and Physics. This reference compiles a lot of science information in one book.
2. Merck Index.
3. Chemical vendor catalogs.

Internet Access in the Chemistry Lab
There are 12 computers in each chemistry lab. These computers are for chemistry students to use. To log onto the Hartnell College network, use the username shown on the top of monitor. The password is: merril.

You will use several chemistry websites to look up information. Here are some websites of interest in chemistry:

2. Los Alamos National Laboratories Periodic Table of the Elements (http://periodic.lanl.gov/default.htm). Click on an element for more information.
4. Material Safety Data Sheets (http://hazard.com/msds/). Type in a chemical name and get the MSDS of that substance.
5. Chemistry Data Tables (http://www.wiredchemist.com/chemistry/data/chem_data.html). This website includes thermodynamic properties of elements, compounds, and ions; solubilities of ionic compounds in water.
6. National Institutes of Standards and Technology Chemistry WebBook (http://webbook.nist.gov/chemistry/). This site provides thermochemical, thermophysical, and ion energetics data compiled by NIST under the Standard Reference Data Program.
Scientific Measurement and Significant Figures

KEY POINTS:
1. Observations are quantified by using scientific instruments or equipment to make measurements, e.g., mass and volume.
2. Every measurement has uncertainty associated with it.
3. The uncertainty of a measurement is reflected by the number of significant figures. The last significant digit is the uncertain digit.
4. Calculated results must reflect the uncertainty in the measurement (data collected).

Quantitative observations involve scientific measurement. For example, a ruler and a scale are used to make quantitative measurements of substances. However, each measurement has uncertainty associated with it. The amount of uncertainty in a measurement is reflected by the number of significant figures that is reported in the measurement. When you measure an object, you want to determine the digits in the measurement that you are certain about plus one additional digit which you are allowed to guess at. This last rightmost digit in a number is the digit that is uncertain. The number of significant figures in a number tells us something about the accuracy of the measurement.

For example, you eating a foot long sandwich and want to know if it is really a foot long. You use a ruler to measure the length of the sandwich as shown in Figure 1.

Figure 1. Measuring the length of a sandwich with a ruler and determining the number of significant figures.

From Figure 1, note that the sandwich is between 11 and 12 inches long. You know the number “11” with certainty since this number is explicitly marked on the ruler. In scientific measurement, you are allowed to guess at one additional digit. So you report the length of the sandwich as 11.5 inches long. This number has a total of three significant figures. The digit “5” is the uncertain digit that you are allowed to guess and is the last significant digit. By reporting the tenths digit, you are implying that the ruler is accurate to ±0.1 inches or ±0.2 inches depending on your ability to “eyeball” between the 11 and 12 marks on the ruler.

If you report the length of the sandwich as 11 inches, you did not guess your one allowed uncertain digit and have not reported enough significant figures. By reporting only two significant figures, you are saying the accuracy of the ruler is ±1 inch. If you report the length of the sandwich as 11.56 inches, you have guessed at two digits and have reported too many significant figures. By reporting four significant figures, you are saying the accuracy of the ruler is ±0.01 inches and this ruler is not this accurate.

Science and chemistry use computers and hand-held calculators extensively. These instruments display many digits in numbers so it is easy to include too many significant figures in your answer. The following rules will help you determine the number of significant figures and how to round numbers:
1. Zeros that are in between non-zero digits are considered significant. E.g., 2.003 has 4 significant figures.
2. For numbers that have a decimal point,
   a. All zeros to the right of the last non-zero digit are significant. E.g., 2.0030 has 5 significant figures.
   b. All zeros to the left of the first non-zero digit are not significant. E.g., 0.020030 has 5 significant figures.
3. For numbers that do not have a decimal point, all zeros to the right of the last non-zero digit are not significant. E.g., 20030 has 4 significant figures.
4. When converting numbers between the expanded (regular) notation and scientific notation, keep the same number of significant figures in each notation. E.g., \(20030 = 2.003 \times 10^4\). Each number has 4 significant figures.

5. For mathematical operations, a calculated result is no better than the experimental data from which it came. Calculated results will have to be rounded to reflect the significant figures in the quantitative measurements.

a. Rounding numbers:
   - if the discarded digit is greater than 5, increase the last retained digit by one. E.g., 15.7 (3 significant figures) rounds to 16 (2 significant figures).
   - if the discarded digit is less than 5, leave the last retained digit unchanged. E.g., 15.4 (3 significant figures) rounds to 15 (2 significant figures).
   - if the discarded digit is equal to 5, increase the last retained digit by one if this digit is an odd number or leave it unchanged if it is an even number. E.g., 15.5 (3 significant figures) rounds to 16 (2 significant figures) or 14.5 (3 significant figures) rounds to 14 (2 significant figures).

b. Addition and subtraction. The number of decimal places in the numbers that are being added or subtracted determines the number of significant figures in the answer. The answer will have the same number of decimal places as the number with the fewest decimal places that is being added or subtracted. The number with the fewest decimal places reflects the least accurate measurement. E.g.,
   \[
   \begin{align*}
   22.2 \text{ cm} & \quad \text{one decimal place - least accurate measurement} \\
   + 11.67 \text{ cm} & \quad \text{two decimal places} \\
   33.87 \text{ cm} & \quad \text{answer needs to be rounded to one decimal place = 33.9 cm.}
   \end{align*}
   \]

c. Multiplication and division. The product or quotient will have the same number of significant figures as the factor with the fewest number of significant figures. E.g.,
   \[
   \begin{align*}
   14.0 & \quad \text{three significant figures} \\
   \times 6.000 & \quad \text{four significant figures} \\
   84.0 & \quad \text{answer has three significant figures}
   \end{align*}
   \]

d. Combining operations in a series of calculations. To avoid rounding errors, carry through all the digits in intermediate calculation steps and then round your final answer. Use the addition/subtraction and multiplication/division rules to determine the number of significant figures. E.g.,
   \[
   \begin{align*}
   22.2 / 14 + 6.000 & = ? \\
   22.2/14 = 1.5|857143 & \quad \text{the digits in the intermediate answer to the left of the line} \\
   & \quad \text{between the 5 and 8 are the significant figures} \\
   1.5|857143 + 6.000 = 7.5|857143 & \quad \text{rounded to 7.6}
   \end{align*}
   \]

   To summarize, every measurement has uncertainty associated with it. The uncertainty is reflected in the last significant digit (the uncertain digit or the digit with which you are allowed to guess). By looking at the measuring device you are using, you can determine the digits you know with certainty with the next digit being the uncertain digit. The sum of the digits you know with certainty plus the uncertain digit gives you the number of significant figures. The uncertain digit tells you the sensitivity of the measuring device.

Random Error and Systematic Error

Each measurement has uncertainty associated with it, i.e., each measurement has error. Error refers to the numerical difference between a measured value and the true value. There are two types of errors: random error and systematic error.

You take 10 g of sand and weigh it 10 times. If you use a coarse mass measuring device, such as a triple beam balance that measures mass to the nearest 1 g, each measurement should give you the same result each time. In other words, your 10 measurements are reproducible. However, if you used a more sensitive balance, such as an analytical balance, each mass measurement will of slightly different in the last digit. These random fluctuations in the measured quantity are called random error. Random error is caused by unpredictable and imperceptible factors that are beyond the control of the experimenter, i.e., you.

Errors that are due to definite causes are called systematic errors. A systematic error is, in general, reproducible and always higher than the true value or always lower than the true value. In many
cases, a systematic error can be predicted or identified by a person who thoroughly understands all the aspects of the measurement. Examples of sources of systematic errors include a corroded weight, parallax reading of a buret, a poorly calibrated buret, an impurity in a reagent, an appreciably solubility of a precipitate, a side reaction in a titration, and heating a sample at too high a temperature.

During the course of Chem 1B lab, you will often compare your experimental result to a true value. Random errors are always present but you want to reduce or eliminate systematic errors in your experimental measurements.

Accuracy and Precision

Since each measurement has uncertainty associated with it, we will determine how “good” our measurements and experiments are. Error in measurement is reflected in accuracy and precision.

Recall the last time you played darts. A throw that is very close to the bull’s eye is accurate. A set of throws that is spread all over the board is not precise. Accuracy refers to the closeness of an experimental value to its “true” value. Precision refers to the closeness of a set of data to each other. Quantitatively, accuracy is represented by absolute error and percent error. Absolute error is the difference between the experimental value and the true value:

\[ \text{Absolute error} = \text{experimental value} - \text{true value} \]  

The percent error is the absolute error relative to the true value:

\[ \% \text{ error} = \frac{\text{absolute error}}{\text{"true" value}} \times 100 \]  

In science, we want our observations to be reproducible, i.e., we want to get the same result each time to tell us that what we are seeing is what we want to see. Precision can be quantified by calculating the % difference:

\[ \% \text{ difference} = \frac{\text{high} - \text{low}}{\text{average}} \times 100 \]  

There are other ways to measure precision of a set of results: average deviation and % average deviation, standard deviation and % standard deviation.

Table 1 lists the uncertainties of various measuring devices. The uncertainties are expressed in the significant figures that the device is capable of measuring.

Table 1. Uncertainties of Various Measuring Devices

<table>
<thead>
<tr>
<th>Measuring Device</th>
<th>Uncertainty</th>
</tr>
</thead>
<tbody>
<tr>
<td>12 cm ruler</td>
<td>± 0.05 cm</td>
</tr>
<tr>
<td>Triple beam balance</td>
<td>± 0.05 g</td>
</tr>
<tr>
<td>Analytical balance</td>
<td>± 0.0001 g</td>
</tr>
<tr>
<td>10 ml graduated cylinder</td>
<td>± 0.05 ml</td>
</tr>
<tr>
<td>100 ml graduated cylinder</td>
<td>± 0.5 ml</td>
</tr>
<tr>
<td>50 ml buret</td>
<td>± 0.02 ml</td>
</tr>
<tr>
<td>25 ml volumetric flask</td>
<td>± 0.02 ml</td>
</tr>
<tr>
<td>25 ml transfer pipet</td>
<td>± 0.02 ml</td>
</tr>
</tbody>
</table>

With your knowledge of scientific measurement, the next time someone asks you how much you weigh, respond qualitatively (“a little” or “a lot”) or quantitatively (“50” and remember those units unless you have ulterior motives).

Reference:
Lab 1. It Stinks! Organic Chemistry: Smells and Odors from Esters

Why do some chemicals smell? What makes chemicals smell different?

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Your instructor will assign your group to one of the reactions in Table 1 below. Put one drop of the acid and one drop of the alcohol on opposite sides of a piece of filter paper. Waft the vapors toward your nose. Describe the odor of the alcohol and acid. Do NOT put your nose directly over the filter paper and inhale deeply.
2. Read the Procedure for Part A. List the data you will collect for this experiment. What result do you want to determine?

Objective: (i) identify functional groups in organic compounds
(ii) predict the structure of an ester from the structures of an alcohol and acid
(iii) match smell to an ester structure
(iv) determine the effect of temperature and catalyst on a reaction
(v) identify bond types using IR spectroscopy
(vi) synthesize an ester

Introduction
From popcorn in a movie theater to a skunk in the woods, our nose is a very good detector of smells and odors. Smells and odors, which play a big role in taste and flavor, are chemicals. Chemists have identified substances that are responsible for odors. Figure 1 shows one way of classifying odors and an example of the substance with that odor.

Figure 1. Classification of Odors and Examples of Odors (taken from Sienko and Plane, “Chemistry”, 1976, p. 685)

Chemists have come up with three theories to odor (see also Science News, 4/10/99):
1. molecular interaction. A reaction occurs between an odor molecule and a receptor.
2. vibrational interaction. IR frequency bond vibrations stimulate a receptor.
3. radiant energy interaction. Photons of energy, perhaps in the IR, stimulate a receptor somewhat as visible light stimulates the eye.

Two scientists won the 2004 Nobel Prize in Medicine for their work in one of these theories. Esters are responsible for a variety of different smells, most of which are pleasant. In general, the smells of flowers and fruits are due to esters. In this lab activity, you will react a carboxylic acid (RCOOH) with an alcohol (R’OH) to produce an ester (RCOOR’):

\[
\text{RCOOH} + R’\text{OH} \rightarrow \text{H}_2\text{SO}_4 \rightarrow \text{RCOOR’} + \text{H}_2\text{O} \quad (1)
\]

\[
\begin{array}{c}
\text{O} \\
\text{R} \quad \text{O} \quad \text{O} \quad \text{H} \\
\downarrow \\
\text{R}’ \quad \text{O} \quad \text{H} \\
\text{O} \\
\text{R’} \quad \text{O} \\
\downarrow \\
\text{R’} \quad \text{O} \\
\downarrow \\
\text{H}_2\text{O} \\
\downarrow \\
\text{H}_2\text{O}
\end{array}
\]
where R and R’ are alkyl groups.

Sulfuric acid is used as a catalyst in this reaction. You will identify the smell of each ester as banana, fingernail polish remover, wintergreen, rum, fruity. Then, you will characterize each reactant and product using IR spectroscopy. Finally, you will investigate the shape of the smell receptor site using the Virtual Molecular Modeling Kit (chemagic.com). By the end of this experiment, you will experience the sweet smell of success.

**Materials**

Acids: acetic acid, salicylic acid, propanoic acid, benzoic acid  
Alcohols: iso-amyl alcohol, ethanol, methanol  
H₂SO₄ (conc.)  
Dropper test tubes stirring rods  
filter paper  
Bruker FT-IR  
macroscale distillation equipment  
ChemDoodle MarvinSketch Virtual Molecular Modeling kit (chemagic.com) newspapers

**Caution:** many organic compounds are flammable and toxic.  
H₂SO₄ (conc.) is corrosive.

**Part A: Ester synthesis and smell identification**

Your group will be assigned one of the reactions in Table 1. You’ve already smelled the odor of the alcohol and acid in your reaction. You will react the alcohol with the acid to produce an ester. Each reaction produces one of the following odors:

- banana, fingernail polish remover, wintergreen, rum, fruity.

**Procedure**

1. a. Prepare a hot water (70°C) bath.  
b. In this step, you will identify the odor that corresponds to the ester of your assigned reaction.  
   (i) Add the indicated number of drops of the alcohol and acid as shown in Table 1 into a small test tube. Swirl gently to mix. Then, add 2 drops of concentrated sulfuric acid to the test tube. Place the test tube in the hot water bath.  
   (ii) After two minutes, use a stirring rod to transfer a drop of your reaction mixture to a clean piece of filter paper. Waft the vapors toward your nose and describe the odor (banana, fingernail polish remover, wintergreen, rum, or fruity) of the product in Table 1. If the mixture solidifies, waft the vapors from the solid material on the end of the stirring rod.  
   If you are not sure about the odor, let the reaction run for one more minute and carefully smell the reaction mixture again.  
   (iii) Draw a conclusion from this experiment.

c. In this step, you will test the effect of temperature on your assigned reaction. (What observation tells you how much product is produced in this reaction?)  
   (i) Add the indicated number of drops of the alcohol and acid as shown in Table 1 into two small test tubes. Label these test tubes. Swirl gently to mix.  
   (ii) Add 2 drops of concentrated sulfuric acid to each test tube.  
   (iii) Place Test Tube 1 in the hot water bath. Leave Test Tube 2 at room temperature.  
   (iv) After one minute, use a stirring rod to transfer a drop of each reaction mixture to a clean piece of filter paper. Waft the vapors toward your nose. If the mixture solidifies, waft the vapors from the solid material on the end of the stirring rod. Is there a difference in smell between Test Tubes 1 and 2? Describe the odor of each test tube in Table 1. Return Test Tube 1 to the water bath.  
   (v) Wait another minute and carefully smell again. Is there a difference in smell compared to (iii)?  
   (vi) Draw a conclusion from this experiment. (As the temperature increases, what happens to the reaction rate?)

d. In this step, you will test the effect of the sulfuric acid catalyst on your assigned reaction. (What observation tells you how much product is produced in this reaction?)
(i) Label two test tubes.
(ii) What should you do next? Once you and your group figure out what do, tell your instructor what you will do. If your instructor approves, do your experiment.
(iii) Describe your observations in Table 1.
(iv) Draw a conclusion from this experiment. (As the amount of sulfuric acid catalyst increases, what happens to the reaction rate?)

e. **Share** your data with the rest of the class. Write the odor (banana, fingernail polish remover, wintergreen, rum, fruity) you identified for your assigned reaction on the table on the chalkboard.

**Waste Disposal:** ethanol, methanol, acetic acid – in sink.
Solids – in solid waste container.
Esters – in non-halogenated waste.

2. a. For your reaction, calculate the moles of acid and alcohol used. State any assumptions that you make. Which reactant is the limiting reactant? Then, calculate the theoretical yield of the ester that is produced.
b. For another reaction, calculate the moles of acid and alcohol used. State any assumptions that you make. Which reactant is the limiting reactant? Then, calculate the theoretical yield of the ester that is produced. Check your calculations with a group who did this reaction.

**Table 1. Reaction of Alcohols and Acids to Produce Esters**

<table>
<thead>
<tr>
<th></th>
<th>Reaction 1</th>
<th>Reaction 2</th>
<th>Reaction 3</th>
<th>Reaction 4</th>
<th>Reaction 5</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alcohol</td>
<td>iso-amyl alcohol</td>
<td>20 drops</td>
<td>ethanol</td>
<td>20 drops</td>
<td>methanol</td>
</tr>
<tr>
<td>Acid</td>
<td>acetic acid</td>
<td>10 drops</td>
<td>acetic acid</td>
<td>10 drops</td>
<td>salicylic acid</td>
</tr>
<tr>
<td>Ester Structure</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Odor</td>
<td></td>
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<tr>
<td>T =</td>
<td></td>
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<td>70 °C</td>
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<td>T =</td>
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<td>___ °C</td>
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<td></td>
<td></td>
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<tr>
<td>drops H₂SO₄</td>
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<td></td>
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<tr>
<td>drops H₂SO₄</td>
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<td></td>
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<tr>
<td>Moles of alcohol</td>
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<td>------------------</td>
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<td></td>
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<tr>
<td>Moles of acid</td>
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<td></td>
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<tr>
<td>Limiting Reactant</td>
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<tr>
<td>Theoretical Yield</td>
<td></td>
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</tbody>
</table>

**Questions**

1. a. Draw the structures of the reactants and products in your assigned reaction. Circle the functional group(s) in each compound and write the name of the functional group next to your circle. Box the bond that breaks in each reactant and the bond that forms in each product.

   b. In one sentence, describe the effect of temperature on the ester synthesis reaction. In other words, compare the odor of the ester reaction at each temperature. Draw a conclusion from your observations.

   c. In one sentence, describe the effect of sulfuric acid on the ester synthesis reaction. Draw a conclusion from your observations.

   d. Show your theoretical yield calculation.

2. Draw the structures of the reactants and products of the 4 other reactions. Circle the functional group(s) in each compound and write the name of the functional group next to your circle. Box the bond that breaks in each reactant and the bond that forms in each product.

3. a. You accidentally mix propanoic acid with iso-amyl alcohol with H₂SO₄. Draw the structure of the product of this reaction. What do you think the product will smell like the odor of Reaction 1 or Reaction 4?

   b. Salicylic acid has three functional groups. Salicylic acid reacts with acetic acid with a small amount of sulfuric acid. Draw the structure of the product of this reaction.

   c. The compound below smells like cherry. Draw the structures of the reactants that makes the cherry compound.

![Cherry Compound Structure]

**Part B.** Introduction to organic lab techniques: reflux, distillation, extraction, IR spectroscopy.

Scientists spend a lot of time, energy, and money making new materials or figuring out better ways to make current materials. Making materials involves synthesis, purification, and characterization. One important and common characterization technique is infrared (IR) spectroscopy. When a molecule absorbs IR radiation, the molecule will vibrate (stretch or bend), like a Jack-in-the-Box or bobble head doll. IR spectroscopy is used to identify bond types, e.g., C-H bond and C-O bond. A C-H bond requires a different amount of energy to stretch than a C-O bond so an IR spectrum can tell us what bond types are in a molecule.

**Prelab**

Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.

1. Look at each piece of equipment in the organic lab kit. Try to identify the function of each piece of equipment.
2. a. Describe how a spectrophotometer works. Draw a schematic diagram of a typical spectrophotometer.

b. Is infrared (IR) radiation at a higher energy or lower energy that visible light? Look up the range of wavelengths in nm in the IR region of the electromagnetic spectrum. Convert the lowest IR wavelength to energy in J, frequency in Hz, and wavenumbers in cm\(^{-1}\).

c. Name one source of IR light. Briefly describe how light is produced.

**Procedure**

1. Using the organic lab kit, identify the equipment that is used to do a reflux. Assemble a reflux. Draw a diagram of a reflux apparatus. Specify the function of each piece of equipment.

2. Using the organic lab kit, identify the equipment that is used to do a distillation. Assemble a distillation. Draw a diagram of a distillation apparatus. Label the distillation flask, condenser, receiving flask, thermometer, and various adapters. Specify the function of each piece of equipment.

3. Using the organic lab kit, identify the equipment that is used to do a liquid-liquid extraction. Assemble an extraction. Specify the function of each piece of equipment.

4. Your instructor will demonstrate the use and operation of the infrared (IR) spectrometer. This includes sample preparation and running an IR spectrum.

a. Measure the IR spectrum of each alcohol, solid salicylic acid, and solid benzoic acid. (Note: Do NOT use liquid acids or bases on the IR. Liquid acids and bases will damage the sample disk on the IR.)

b. Use the “Peak Pick” feature in the IR software, determine the wavenumber of the major peaks in each spectrum. Print out your spectrum.

   Share your data in Table 2 on the chalkboard. List the wave number (in cm\(^{-1}\)) and the intensity (strong, medium, or weak) of each band (peak) for each compound. Ignore peaks below 1000 cm\(^{-1}\).

**Table 2. IR Data for Various Organic Solvents**

<table>
<thead>
<tr>
<th>Organic Solvent</th>
<th>Bond Type(s) (e.g., C-H, C-O)</th>
<th>IR Peaks, cm(^{-1})</th>
<th>Peak Intensity (strong, medium, weak)</th>
</tr>
</thead>
<tbody>
<tr>
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</tr>
</tbody>
</table>

   c. Based on the data on the board, match the band (peak) in each IR spectrum to the bond type in each molecule.


**Questions**

1. Based on the IR results of the class, compile a table (Table 3) that correlates the bond type with the IR bands, i.e., list the wave number and intensity of the bands in an IR spectrum of an acid, alcohol, and ester.

2. Based on IR, can you definitively distinguish one solvent from the others that we looked at? Give reasons.

**Part C. Synthesis and characterization of an ester.**

Each bench will be assigned one of the reactions in Table 1. For your reaction, you will synthesize 10 g of ester, separate, purify, and characterize your product.
Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. For your assigned reaction, calculate the mass or volume of the acid and alcohol to make 10 g of ester. Include in your calculations the volume of sulfuric acid to use. Calculate the theoretical yield of product.
2. Read the Procedure for Part C. List the data you will collect for this experiment. List the results that you will determine or calculate.

Procedure
1. Synthesis of the ester. Transfer the appropriate amounts of reactants for your assigned reaction and catalyst into a 50 ml round-bottomed flask. Reflux the reaction mixture for one hour. Let the reaction mixture cool.

2. Purification, isolation, and characterization of the ester.
   a. Pour the reaction mixture into a beaker containing 25 ml of cold water. You should see two phases (layers) form. Any remaining acid and alcohol should dissolve in the water; the ester should be immiscible. Use a separatory funnel to separate the organic and aqueous layer (your instructor will demonstrate the use of a separatory funnel).
   b. Purify the organic layer by washing three times with 10 ml of distilled water. Use a separatory funnel. This step removes any traces of acid or alcohol.
   c. Measure the actual yield of your ester. Calculate % yield.
   d. Measure the IR spectrum of your ester.

3. Draw a conclusion from this experiment. (Summarize your results. Did you make the desired product?)


Questions
1. a. Show your theoretical yield calculation for the ester.
   b. Discuss your % yield of product. Include numbers.

2. Interpret your IR spectrum. Discuss whether you made the desired product. Include numbers.

Part D. Investigate the interaction of the ester at the smell receptor site. We looked at “Olfactophore Models for Structure-Odour Correlations in Fragrance Research” in lecture. These researchers identified two types of features: hydrogen bond acceptor sites and hydrophobic sites.

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. a. You put a hat on your head. Does the hat have to be a certain size and shape to fit your head? Does your head have to enter the hat a certain way to fit into the hat?
   b. You put your hand into a glove. Does the glove have to be a certain size and shape to fit your hand? Does your hand have to enter the glove a certain way to fit into the hat?
   c. How are these two examples related to how an ester interacts with a receptor site?
   d. Think of another analogy.
2. Read the Procedure for Part D. List the data you will collect for this experiment. Identify the result you want from this experiment.

Students: bring old newspapers to lab

Procedure
1. a. Construct a model of your ester from Part A using a molecular model kit.
   b. Identify the hydrogen bond acceptor sites and hydrophobic sites. How many of each are there in your ester?

2. a. Use the Virtual Molecular Modeling kit (http://chemagic.com/JSmolVMK2.htm) to draw your assigned ester from Part A.
b. Click on the “Draw” button and draw the structure of the ester.
c. Then, click on the “2D to 3D” button to see a 3D model of this compound. Use your mouse to rotate the molecule. Is the ester flat?
d. Click on the “WireFrame” button to see a different representation of this compound. Do you like this better than the ball and stick?
e. Click on the “SpaceFill” button to see a different representation of this compound. Do you like this better than the ball and stick or wire frame?
f. Click on the “van der Waals” button to see the molecule’s surface and surface area (see upper left corner of screen). This gives you an idea of what this molecule looks like and how it has to fit into a protein receptor site so you can smell the smell.
g. Click on the “Length” button and follow the instructions to measure the size of the molecule and relevant distances, e.g., distance from one hydrogen bond acceptor site to another, distance from one hydrogen bond acceptor site to hydrophobic interaction.

2. Use newspaper to make a model of the ester receptor site.
a. How do you think the ester fits into the receptor site? Does the ester enter the receptor site from one end or on its side? Where does the receptor hydrogen bond to the ester? Where does the receptor have a hydrophobic interaction with the ester?
b. Fold, shape, etc. the newspaper in a way that allows the ester (your molecular model) to enter the receptor site (newspaper). Show how your model of the receptor fits the ester (your molecular model) and ester features (hydrogen bond acceptor sites and hydrophobic sites).
c. Sketch the receptor site for your ester. Show the dimensions of the receptor site. Describe how the ester enters the receptor. Show the hydrogen bond acceptor sites and hydrophobic features in the receptor site that bind to the ester.

Waste Disposal: a mind is a terrible thing to waste.

Questions
1. Draw the Lewis structure of your assigned ester. Describe the shape of the ester. Sketch the shape of the ester receptor site. On your sketch, label the relevant distances that describe the shape. Describe the hydrogen bond site(s) and hydrophobic parts.

2. Describe the dimensions of the receptor site. Describe the hydrogen bond acceptor site(s) and hydrophobic sites where the ester fits the receptor.

Reference:
Lab 2. Drug Abuse. Solubility and Colligative Properties of Solutions: Coffee, Soda, and Ice Cream

_How do I make a stronger cup of coffee? How do I make ice cream?_

**Prelab**
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Using the coffee you brought to class, make a very small cup of coffee. Then, design an experiment to answer the following questions:
   a. Does the temperature of the water affect the amount of chemical(s) extracted from the coffee beans to the water? How can you tell?
   b. Does the amount of time the water is in contact with the coffee beans affect the amount of chemical(s) extracted? How can you tell?
   c. How will you quantify the amount of chemicals extracted from the coffee beans into the water?
2. a. Look up the amount of caffeine in coffee. Cite the reference where you found this information.
   b. Draw the structure of caffeine. What type of compound is caffeine? Is caffeine polar?
   c. Name two solvents in which caffeine is soluble.
   d. What chemical forces are involved when caffeine dissolves in each solvent? Draw the structure of each solvent next to the structure of caffeine and show the chemical forces between caffeine and solvent.
3. a. Read the Procedure for Part A. List the data you will collect for this experiment. What instrument will you use to collect each piece of data?
   b. Identify the result you want to determine from this experiment.

**Objectives:**
(i) relate solute amount to a colligative property of a solution
(ii) determine the solubility of a gas with temperature and pressure

**Introduction**
Your morning ritual: boil water, pour ground coffee into a filter, pour the hot water through the coffee and filter, fresh coffee. What happened? The hot water extracts the coffee flavor, caffeine, etc. from the solid coffee beans into the water in a solid-liquid extraction. What conditions make a strong cup of coffee? You will look at a colligative property of solutions (boiling point elevation) to determine if you have how strong your coffee is.

Caffeine is the most consumed psychoactive substance in the world. (Maybe you should reduce your caffeine intake.) You will look at solubility of caffeine in different solvents to decaffeinate coffee. Then, you’ll use chromatography to determine whether you’ve decaffeinated your coffee.

Your afternoon pick-me-up: your favorite soda. You like the whoosh sound when you open the car, the sugar, caffeine, the tart taste, and the bubbles. Where do the bubbles come from? How do I keep my soda from going flat? You will look at what makes soda go flat.

Your special dessert: ice cream. How do you make home-made ice cream? You can use the freezing point depression colligative property of solutions to lower the freezing point of ice to make ice cream.

**Materials**
Part A: Coffee (Students: bring some ground coffee to lab)
filter paper and funnel
Vernier temperature probe (range is -10°C to 110°C), Vernier LabPro, and computer
hot plate or sand bath
High Performance Liquid Chromatograph (HPLC)
ethyl acetate (5 ml per group) separatory funnel
activated charcoal

_Caution:_ hot plates get hot. Ethyl acetate is flammable.

Part B: three bottles or cans of carbonated water or soda (Students: bring in 3 small bottles or cans of soda)
Balloon 500 ml or 1 l graduated cylinder
Part C: Freezing Point Depression and Ice Cream

**We will supply:**
large ziplock baggie (gallon size)
small ziplock baggie (quart size)
crushed ice
coarse rock salt

**You need to bring:**
Ingredients to make ice cream, e.g.,
half-half milk/cream mixture
sugar
vanilla
flavoring, e.g., chocolate or strawberry syrup

English/Metric Conversions:  1 pint = 2 cups, 4 cups = 1 quart, 4 quarts = 1 gallon
1 cup = 8 ounces = 240 ml

Part A. Making coffee. Decaffeinating coffee.

**Procedure**
1. Make a strong cup of coffee by solid-liquid extraction. Determine the effect of temperature and time.
   You designed an experiment in the Prelab Activity to investigate the effect of temperature and time on the concentration (strength) of coffee.
   You and your lab class will combine data so we can identify the temperature and time that makes the strongest cup of coffee. Your instructor will assign other experimental variables (mass of coffee and ______) for the class to use.
   From Prelab Question 1c, you can quantify the amount of chemical(s) extracted by:
      i. b.p. elevation
      ii. separate the water from the rest of the chemicals and then measure the mass of solid residue.

a. Do your experiment(s). Record your data and results in Table 1.
   Which colligative property did you use to quantify the amount of chemical(s) extracted?
   b. Draw a conclusion from the class data. What temperature and time makes the strongest cup of coffee?

<table>
<thead>
<tr>
<th>Group</th>
<th>Brewing Temperature, °C</th>
<th>Brewing Time, min</th>
<th>Boiling Point of Pure Solvent, °C</th>
<th>Boiling Point of Coffee, °C</th>
<th>ΔTb, °C</th>
<th>Molality, m</th>
<th>Mass of residue, g</th>
</tr>
</thead>
<tbody>
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</tbody>
</table>

2. You will work with another group to decaffeinate coffee. One group will use an organic solvent extraction to decaffeinate coffee; the other group will use water extraction and activated charcoal (see [http://antoine.frostburg.edu/chem/senese/101/consumer/faq/decaffeinating-coffee.shtml](http://antoine.frostburg.edu/chem/senese/101/consumer/faq/decaffeinating-coffee.shtml)). You want to remove the caffeine only, not the other chemicals, such as the flavor chemicals, from coffee.

(i) Organic solvent extraction method:
a. Identify an organic solvent that removes only caffeine from coffee.
b. Once you and your lab group have identified a solvent, design an experiment to decaffeinate coffee.
c. Once you remove the caffeine, we will use HPLC to analyze your decaffeinated coffee sample. What solvent would you use to analyze caffeine by HPLC? Give reasons. What do you want to look for in the HPLC data? Record your data and results in Table 2.

(ii) Water extraction and activated charcoal method:
a. Using the class results from Part A, Step 1, make a cup of coffee.
b. Pass the extract through activated charcoal. What equipment will you use to do this step? (After using
the activated charcoal, do **NOT** dispose of the activated charcoal – place in designated container so we
can reuse it.)
c. Analyze the filtrate by HPLC. What do you want to look for in the HPLC data? Record your data and
results in Table 2.

(iii) Draw a conclusion from the class data and results. Which decaffeination method worked better?

Table 2. HPLC Coffee Decaffeination.

<table>
<thead>
<tr>
<th>Group</th>
<th>Caffeine Peak Area before decaffeination, mAu</th>
<th>Decaffeination Method</th>
<th>Solvent or Mass of Activated Charcoal</th>
<th>Caffeine Extraction Time, min</th>
<th>Caffeine Peak Area after decaffeination, mAu</th>
<th>% caffeine removed</th>
</tr>
</thead>
<tbody>
<tr>
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</tr>
</tbody>
</table>


**Questions**

1. Describe your experimental procedures for Step 1 and Step 2.

   a. Show your Table 1.
   b. Discuss the effect of temperature and time on the coffee extraction based on:
      i. b.p. elevation. What does molality tell you?
      ii. mass of residue. Compare the mass of residue to the molality.
   c. What temperature and time made the strongest cup of coffee? Give reasons.
   d. When a liquid boils, the intermolecular forces between liquid molecules are broken and these
      molecules escape into the gas phase. The b.p. of a solution is higher than the b.p. of the pure solvent.
      Does adding a solute make it easier or harder for solvent molecules to escape to the gas phase? Draw a
      picture that shows how a solute makes the b.p. higher.

3. Decaffeinating coffee.
   a. Show your Table 2.
   b. Discuss the effect of the organic solvent on the decaffeination process. Include HPLC numbers. Draw
      the structure of caffeine next to the structure of your organic extraction solvent. Identify the chemical
      force(s) between the caffeine and solvent.
   c. Discuss the effect of the activated charcoal on the decaffeination process. Describe how activated
      charcoal interacts with caffeine to decaffeinate coffee.
   d. Compare the class data. Which decaffeination method worked better? Give reasons.

**Part B. Solubility of a gas in a liquid. Carbonated beverages.**

**Prelab**

Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.

1. Soda and soda water are carbonated. How is soda or water carbonated? Consider the following reactions:
\[
\begin{align*}
\text{H}_2\text{CO}_3 (aq) & \rightarrow \text{H}_2\text{O} (l) + \text{CO}_2 (g) \\
\text{H}_2\text{CO}_3 (aq) & \rightleftharpoons \text{H}^+ (aq) + \text{HCO}_3^- (aq) \\
\text{CO}_2 (g) & \rightleftharpoons \text{CO}_2 (aq)
\end{align*}
\]

You have two bottles or cans of soda. One bottle/can will be cold. For the other bottle/can, do the following and record your observations:

(i) shake the bottle for 30 seconds.
(ii) Put the bottle in a hot water bath.

How can you determine the pressure of gas in the each unopened bottle?

2. a. Read the Procedure for Part A. List the data you will collect for this experiment.
b. Identify the result that you will calculate from a formula or equation in this experiment.

**Procedure**

1. You will experimentally determine the pressure of gas in soda three (3) ways:
   i. put a balloon or rubber glove over the soda and secure the balloon or rubber glove tightly over the lip. A twist tie should work well. Carefully open the soda. The escaping gas should inflate the balloon or rubber glove. When the balloon or rubber glove has inflated as much as you think it will inflate, carefully squeeze the balloon or rubber glove. What happens? Carefully take the balloon or rubber glove off the soda and tie it. Try not to let any gas escape. Measure the volume of the balloon or rubber glove. Calculate the moles and mass of gas in the balloon or glove. Based on this information, calculate the pressure of gas in the soda before you opened it. Record your data in Table 3.
   ii. If you have a can of soda, open the soda and quickly pour the contents into a 1 liter graduated cylinder. Measure the volume of liquid and the volume of foam. What does the volume of foam tell you? Using the volume of foam, calculate the moles and mass of gas. Calculate the pressure of gas in the soda. Record your data in Table 3.
   iii. Calculate the pressure of CO\(_2\) in your soda based on Chang, "General Chemistry: The Essential Concepts," 7th ed, Problem 13.111. “A student carried out the following procedure to measure the pressure of CO\(_2\) in a soft drink bottle. First, she weighed the bottle (______)g). Next, she carefully removed the cap to let the CO\(_2\) gas escape. She then reweighed the bottle with the cap (____ g). Finally, she measured the volume of the soft drink (____ ml). Given that Henry's law constant for CO\(_2\) in water at 25°C is 3.4x10\(^{-2}\) mol/L atm, calculate the pressure of CO\(_2\) in the original bottle. Why is this pressure only an estimate of the true value?” Measure the appropriate quantities for your soda. Record your data in Table 3.

<table>
<thead>
<tr>
<th>Method i</th>
<th>Method i</th>
<th>Method ii</th>
<th>Method ii</th>
<th>Method iii</th>
<th>Method iii</th>
</tr>
</thead>
<tbody>
<tr>
<td>Soda brand/can or bottle</td>
<td>Soda brand/can or bottle</td>
<td>Soda brand/can or bottle</td>
<td>Soda brand/can or bottle</td>
<td>Soda brand/can or bottle</td>
<td>Soda brand/can or bottle</td>
</tr>
<tr>
<td>Volume of balloon, ml</td>
<td>Volume of liquid, ml</td>
<td>Mass of soda, g</td>
<td>Mass of soda after CO(_2) escaped, g</td>
<td>Mass of CO(_2) in soda, g</td>
<td>Mass of CO(_2) in soda, g</td>
</tr>
</tbody>
</table>
Moles of gas in soda before opening | Moles of gas in soda before opening | Moles of CO₂
---|---|---
Volume of gas in soda (space between liquid and top of can/bottle), ml | Volume of gas in soda (space between liquid and top of can/bottle), ml | Volume of soda, ml
Pressure in soda, atm | Pressure in soda, atm | Pressure in soda, atm

2. a. Compare the pressures you measured in Step 1. Are these pressures the same? Calculate an average pressure and % difference. Which method do you think has the highest experimental error? b. Draw a conclusion from this experiment.

Waste Disposal: carbonated beverages – in sink.

Questions
1. a. Show your Table 3.
b. Discuss the pressures you calculated in Step 1. Include numbers.

2. a. Look up the solubility of CO₂ in water at room temperature. Cite the reference where you found this information.
b. There are at least two ways to keep soda from going flat. Draw a picture that describes each way to keep soda from going flat based on your results.

Part C. Freezing point depression. Make ice cream.

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
Ice cream is made by crystallizing a milk/cream/sugar/flavor mixture. Before you make ice cream, do the following:
(i) What is the solvent in ice cream?
(ii) At what temperature will this milk mixture crystallize to make ice cream? Estimate the temperature you think this mixture will crystallize. Then, calculate the molal concentration of solute(s) in milk. What is the solute in milk that lowers its freezing point? Give reasons.
(iii) Then, calculate the amount of rock salt and ice required to achieve this temperature. Your group can use a maximum of 50 g of rock salt and 2 cups of ice.

Ice cream is made by crystallizing a milk/cream/sugar/flavor mixture. In ice cream, milk is the “solvent”; however, milk itself is a mixture that consists mostly of water.

Procedure
1. a. Into the small ziplock baggie, place the milk, sugar, vanilla, and flavoring of your choice. Carefully seal the baggie completely to ensure that all the ingredients are safely locked inside (unless you enjoy salty ice cream!).
b. Mix the ingredients into a homogeneous mixture together by kneading gently.

2. Fill the larger ziplock baggie with the mass of ice and rock salt that you calculated. (You want the bag to be about 1/2 full of ice and rock salt.) Mix the ice/salt into a homogeneous mixture by kneading gently. Measure the temperature of your rock salt/ice mixture. Is the temperature equal to the temperature you calculated before you started the lab?
3. Carefully place the small baggie with your sealed milk mixture into the large baggie with ice/salt and seal the large baggie.

4. Shake and knead on a table or other hard surface for about 10-15 minutes or until the milk mixture hardens into ice cream.

(Or use two large ziplock baggies with the small milk mixture in between – a milk mixture sandwich!)

5. When your ice cream is done,
   a. Check the temperature of the rock salt/ice mixture. Is the temperature equal to the temperature you calculated before you started the lab?
   b. Carefully remove the small ziplock baggie from the larger baggie and gently rinse the ziplock seal on the small ziplock baggie with cold ice water. This rinses away the excess salt from your ice cream baggie.
   c. Test taste your ice cream. Give your ice cream a letter grade. Describe the criteria you used for this grade. If you don’t think any contamination has occurred, offer some to your lab technician.
   d. Draw a conclusion from this experiment.

**Waste Disposal:** ice/salt mixture - in sink.
Rinse out the large baggie with water and return in to your instructor for reuse.
Dispose of your small baggie in the appropriate receptacle after you finish eating your ice cream.

**Questions**

1. Show your calculation of the mass of NaCl to lower the freezing point of water.

2. a. Report the temperature of your salt-ice mixture.
   b. Discuss the effect of temperature of your salt-ice mixture on your ice cream making process. Include numbers.

3. Describe the texture of the ice cream you made.
Lab 3. Turn Blue: The Kinetics of the Iodine Clock Reaction

What makes a reaction go fast or slow?

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Green food coloring is a mixture of blue and yellow food colorings. Bleach is an oxidizing agent.
Add 1 drop of green food coloring to 50 ml of water. Mix.
Add bleach to green food coloring solution. Record your observations.
How will you change this reaction to make the color change occur faster? Try it out to see if it works.
Explain your observations using reaction rate principles.

2. a. Read the Procedure for Part A. List the data you will collect for this experiment. What instrument will you use to collect each piece of data?
b. Identify the result you want to determine from this experiment. What equation or formula will you use to calculate this result?

Objectives:
(i) Measure the rate of a reaction
(ii) Determine the order of a reaction
(iii) Calculate the rate constant from rate
(iv) Determine the rate law of a reaction
(v) Relate rate, order, rate constant to temperature and concentration
(vi) Calculate the activation energy of a reaction
(vii) Speculate on a reaction mechanism based on the rate law

Introduction
The iodine clock reaction is a very popular reaction with which to study reaction rates. In this reaction, two colorless solutions are mixed together. After a certain period of time, iodine is produced and a dramatic blue color appears. Since you can predict the time to turn the solution from colorless to blue, this reaction can be used as a clock.
A chemical reaction occurs when reactant atoms or molecules collide with sufficient energy for bonds to break or form. Three factors affect how fast or slow a reaction occurs: concentration, temperature, and a catalyst.
(i) Increasing the concentration increases the number of collisions increases and results in a faster reaction. We will study the effect of concentration on rate to determine the rate law for this reaction.
(ii) Increasing the temperature makes atoms/molecules move faster and results in more energetic collisions between reactants. The minimum energy for a reaction to occur is called the activation energy. We will also study the effect of temperature on rate to determine the activation energy of this reaction.
(iii) We will look at the effect of a biological catalyst (enzyme) on reaction in Lab 5 with food preservatives.
The iodine clock reaction is a combination of three reactions:

\[ IO_3^- + 3 HSO_3^- \rightarrow I^- + 3 SO_4^{2-} + 3 H^+ \] (1)
\[ 5 I^- + 6 H^+ + IO_3^- \rightarrow 3 I_2 + 3 H_2O \] (2)
\[ I_2 + \text{starch} \rightarrow \text{blue solution (indicator)} \] (3).

Reaction (1) is the reaction whose rate we are measuring. When all of the HSO_3\(^-\) has reacted, the I\(^-\) reacts with the remaining IO_3\(^-\) to produce I\(_2\) as shown in Reaction (2). The I\(_2\) that is produced in Reaction (2) then reacts with starch to form the blue starch-iodine complex as shown in Reaction (3). Reactions (2) and (3) tells us by way of the blue color that Reaction (1) is complete. This allows us to measure the time for Reaction (1) to be complete and therefore allows us to determine its rate.

Materials
Green food coloring
Bleach
Solution A: 0.020 M potassium iodate, KIO_3
Solution B: sodium hydrogen sulfite, NaHSO_3/starch (0.404 g Na_2S_2O_5/liter)
Lab Practical: bleach and green food coloring

**Caution:** bleach and sodium hydrogen sulfite are oxidizing agents

**Part A.** Determine the rate law, order, and rate constant at room temperature.

**Procedure**

1. **a.** Measure 5.0 ml of Solution A into a small flask or beaker. (What volume measuring device should you use?) Measure 5.0 ml of Solution B into a clean dry test tube. Record the concentrations of each solution.

   \[ [\text{KIO}_3] = \ldots \quad [\text{NaHSO}_3] = \ldots \]

   b. Add Solution B to Solution A and record the time to the nearest second it takes for the solution to turn blue. Start recording the time the instant both solutions come in contact with each other (t=0). Stir the solutions to ensure uniform mixing. Keep stirring until the solution turns blue. Record this time in Table 1. Record the temperature of this solution.

2. **a.** You will be assigned three of the diluted following solutions of KIO$_3$. Prepare your solutions as described in Table 1. Calculate the [KIO$_3$] for Solutions 1-8. Fill in your answers in Table 1.

   **Table 1. Solution Preparation, Concentrations, and Rates. Temperature = \ldots**

<table>
<thead>
<tr>
<th>Solution</th>
<th>Vol of Solution A, ml</th>
<th>Vol of Water, ml</th>
<th>Total Volume, ml</th>
<th>[KIO$_3$], M (Sol. A)</th>
<th>[NaHSO$_3$], M (Sol. B)</th>
<th>Time, sec</th>
<th>rate</th>
<th>k</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>5.0</td>
<td>0</td>
<td>5.0</td>
<td>[KIO$_3$]</td>
<td>[NaHSO$_3$]</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>4.5</td>
<td>0.5</td>
<td>5.0</td>
<td>[KIO$_3$]</td>
<td>[NaHSO$_3$]</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>4.0</td>
<td>1.0</td>
<td>5.0</td>
<td>[KIO$_3$]</td>
<td>[NaHSO$_3$]</td>
<td></td>
<td></td>
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<tr>
<td>4</td>
<td>3.0</td>
<td>2.0</td>
<td>5.0</td>
<td>[KIO$_3$]</td>
<td>[NaHSO$_3$]</td>
<td></td>
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<tr>
<td>5</td>
<td>2.5</td>
<td>2.5</td>
<td>5.0</td>
<td>[KIO$_3$]</td>
<td>[NaHSO$_3$]</td>
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<td></td>
</tr>
<tr>
<td>6</td>
<td>2.0</td>
<td>3.0</td>
<td>5.0</td>
<td>[KIO$_3$]</td>
<td>[NaHSO$_3$]</td>
<td></td>
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<tr>
<td>7</td>
<td>1.5</td>
<td>3.5</td>
<td>5.0</td>
<td>[KIO$_3$]</td>
<td>[NaHSO$_3$]</td>
<td></td>
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<td></td>
</tr>
<tr>
<td>8</td>
<td>1.0</td>
<td>4.0</td>
<td>5.0</td>
<td>[KIO$_3$]</td>
<td>[NaHSO$_3$]</td>
<td></td>
<td></td>
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</tbody>
</table>

   b. Repeat Step 1b. Record the time it takes for your assigned solution to turn color in Table 1.
   c. Your group will share your data and results with your lab class. Record your times on the Table on the board. Calculate the average time for each solution.

**Waste Disposal:** iodine clock reaction solutions – in iodine clock reaction waste

3. From the class data for each solution, calculate the change in \([\text{KIO}_3] = \Delta[\text{KIO}_3] = [\text{KIO}_3]_{t=0} - [\text{KIO}_3]_{t=\text{blue}}\). This is the change in the concentration of KIO$_3$ while the reaction is taking place.

   a. For each solution, calculate the rate at which KIO$_3$ disappears = \(\Delta[\text{KIO}_3]/\Delta t\). \(\Delta t\) is the time it takes for the reaction to turn blue. Record these rates in Table 1.
   b. Determine the order with respect to KIO$_3$. Assume the order with respect to HSO$_3^-$ is 1.
   c. Write the rate law for the iodine clock reaction.
   d. From your rate law, calculate the rate constant, \(k\), for this reaction. Include units for \(k\). Specify \(T\). Record these rate constants in Table 1. Calculate the average rate constant.
4. a. Using the data in Table 1 and the Graphical Analysis software, plot a graph of [KIO₃] vs. time. Label this Graph 1. Draw a smooth curve through the data points.
b. Choose one point on your Graph 1. Draw a tangent to the curve. What does the slope of the tangent represent? (For those of you who have take Calculus, refer to the derivative of a curve.)

5. Draw a conclusion from this experiment. (Summarize your results.)

Questions
1. a. Show your Table 1.
b. Report the rate law and rate constant for the iodine clock reaction. Include numbers and units.
c. Discuss the effect of concentration on the rate of this reaction based on your results. Include numbers. Draw a picture that shows how an increase in concentration affects rate.
d. Why is the rate constant, k, for this reaction the same for each of the solutions in Part A?
e. Show your Graph 1. Based on your Graph 1, what happens to the reaction rate as the reaction proceeds?

2. a. You do a 9th solution in which you take 0.5 ml of Solution A, add 4.5 ml of water and then add this solution to Solution B. Calculate the time it takes for the solution to turn blue.
b. You run an iodine clock reaction with [IO₃⁻] = 0.1 M and [HSO₃⁻] = 0.1 M at the same temperature that you ran earlier. Calculate the rate.

3. Refer to reactions (1) and (2) in the Introduction section. In reaction (1), IO₃⁻ reacts with HSO₃⁻ to produce I⁻. In reaction (2), the I⁻ that is produced in reaction (1) reacts with IO₃⁻ to produce I₂.
a. In reaction (1), which reactant must be the limiting reactant? Give reasons.
b. Why doesn’t I₂ form immediately if all reactants are present in the solution? (Hint: compare the rate of each reaction.)

4. a. Refer to reaction (1) in the Introduction section. 5 ml of 0.0201 M IO₃⁻ reacts with 5 ml of 0.002 M HSO₃⁻. Calculate the moles of each reactant that reacts and the moles of each product that forms. Identify the limiting reactant.
b. In carrying out the various dilutions of Solution A, why did the total volume have to be kept constant at 5.0 ml for each dilution? (Hint: One variable needed to remain constant and would have changed if the volume had changed.)

Part B. Determine the activation energy of the reaction.
To determine the activation energy of this reaction, you will determine the rate constant at several temperatures. You will be assigned two temperatures from Table 2. Note that a range of temperatures is given. When you do your experiment, record the actual temperature to the nearest °C in Table 2.

Procedure
1. Prepare a water bath at one of your assigned temperatures. Use a large enough beaker to hold the flask, beaker, or test tube that contains Solutions A and B. Use a hot plate or Bunsen burner to warm water above room temperature or ice to cool water below room temperature.

2. a. Repeat Part A, Step 1a.
b. Immerse the flask and test tube in your water bath. Leave them in the bath until the solutions are at the desired temperature. Make sure both solutions are at the same temperature before starting the reaction.
c. Repeat Part A, Step 1b except record the time in Table 2.
d. Again, the class will share results. Calculate the average time for the reaction to turn blue for each temperature.

Waste Disposal: iodine clock reaction solutions – in iodine clock reaction waste

3. a. From the class data, calculate the rate at which KIO₃ disappears at each temperature. (Should the rate be faster or slower with increasing temperature?)
b. Calculate the rate constant, k, at each temperature. (Should k be bigger or smaller with increasing temperature?)
c. To Table 2, add new columns for rate, k, ln k, 1/T (T in K).
4. a. Plot a graph of ln k vs. 1/T. Label this Graph 2. Do a linear fit of your data points.
b. Calculate the slope of the line. From the slope, calculate the activation energy.
c. Determine the y-intercept of the line. What does the y-intercept represent?

<table>
<thead>
<tr>
<th>Temperature, °C</th>
<th>Time, sec</th>
<th>Average Time, sec</th>
<th>Rate</th>
<th>k</th>
<th>ln k</th>
<th>1/T</th>
</tr>
</thead>
<tbody>
<tr>
<td>(0-10)</td>
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<td>(10-20)</td>
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<td>(20-30)</td>
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<td>(30-40)</td>
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</tbody>
</table>

5. Draw a conclusion from this experiment. (Summarize your results.)

Questions
1. a. Show your Table 2 and Graph 2.
b. Discuss the effect of temperature on the rate of this reaction based on your results. Include numbers. Draw a picture that shows how temperature affects collisions between molecules and the rate of a reaction.
c. Discuss the effect of temperature of the rate constant in this reaction based on your results. How are temperature and k related based on your results? How are rate and k related based on your results? Include numbers.

2. Based on your $E_a$, calculate $k$ at 75°C. Then, calculate the rate and the time to turn the solution blue.

3. Speculate on the mechanism of this reaction based on your rate law. Identify the rate determining step in your mechanism.

Lab Practical
As a class, design an experiment in which you make a solution turn blue in 0.5 min, 1 min, 1.5 min, and so on. Each group is responsible for a specific time. Every group will perform their experiment in front of the class at the same time to see how well the timing works. Use your knowledge of the rate law and the rate constant at different temperatures for the iodine clock reaction. Part of your grade on this lab will be based on getting the specified time (practice) based on your calculations (theory). We will see whether practice matches up with theory!
Lab 4. It’s About Control: Shifting the Direction of a Reaction Using LeChatelier’s Principle

How can I control a reaction? How can optimize the yield of a reaction?

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Explain how each picture is related to equilibrium.

http://sc6.blogspot.com/2010_02_01_archive.html

http://ingrimayne.com/econ/Keynes/DifferentViews.html

2. a. Read the Procedure for Part A. List the data you will collect for this experiment.
b. Identify the result you want to determine from this experiment.

Objectives:
(i) Apply equilibrium principles (LeChatelier’s principle) to make a reaction go in the forward and reverse direction.
(ii) Identify the reaction conditions to produce the highest yield of product.

Introduction
Scientists know how to make a chemical reaction go faster (or slower) by adjusting the amount of reactants, temperature, and pressure. You looked at the effect of these factors in Lab 3. These three factors are also used to control the direction (forward or reverse) of a reversible, or equilibrium reaction. In other words, adding or removing reactant or product, changing the temperature or pressure changes the direction of a reaction. This concept is known as LeChatelier’s principle: adding a stress (reactant, product, T, or P) to a reaction shifts the reaction in the direction that relieves the stress. At equilibrium, the ratio of the concentrations of products to reactants (the equilibrium constant) remains fixed at a specific T and P.

In this experiment, your lab group will investigate an equilibrium reaction. You will use LeChatelier’s principle and apply a stress to make this reaction go in one direction and then apply a different stress to make the reaction go in the opposite direction. Your group will teach this reaction to your lab class.
Materials

Part A
Reaction A. 0.1 M K$_2$CrO$_4$, 6 M HCl, 6 M NaOH
Reaction B. 0.1 M Al(NO$_3$)$_3$, 6 M NaOH, 6 M HCl
Reaction C. 0.1 M Co(NO$_3$)$_2$, solid NaCl, AgNO$_3$, ice bath, hot water bath, conductivity probe
Reaction D. phenolphthalein, 1 M HCl, 1 M NaOH
Reaction E. CaCl$_2$ (s), 1 M NaOH, 1 M HCl
Reaction F. 250 ml filter flask, NaHCO$_3$ (s), 1 M HCl, 1 M NaOH, phenolphthalein. **Students:** bring in an unopened colorless carbonated beverage in a clear, colorless bottle (don't bring a can)

Part B
Salicylic acid Methanol H$_2$SO$_4$ (conc.)

Caution: Chromium, cobalt, and silver are heavy metals. Acids and bases are corrosive.

Procedure

Part A. Your instructor will assign a reaction to each group. For your assigned reaction, do the following:
1. answer the following questions:
   a. What is the state of matter of each reactant and product? Is this reaction exothermic or endothermic?
   b. What is the name of each reactant and product? What is the structure of each reactant and product?
   Identify properties of each reactant and product.
   c. Write a net ionic equation for the reaction if applicable. Identify the spectator ions.
   d. What observation (color; formation of solid, liquid, or gas; temperature; smell) tells you reactants are present? What observation tells you products are present?
   e. What materials do you need to do this reaction?
   f. What safety precautions do you need to take in handling and using these materials?
   g. What “stress” will you place on the reaction to shift the reaction in one direction or other?
   h. Is the equilibrium constant for your reaction greater than 1 or less than 1? Give reasons. How would you experimentally measure $K_{eq}$?

2. Do an experiment. Starting from an equilibrium state, make the reaction shift in one direction. Then, make it shift in the opposite direction.
   a. Describe an experimental procedure.
   b. Do the experiment to see if it works the way you want it to work.

3. Teach the class your reaction in a 5-10 minute presentation. (This is your Report for this part.)
   a. Prepare a lesson or teaching plan.
      (i) Identify what you want your students to learn.
      (ii) What activity will you use to teach this reaction? E.g., lecture, demonstration, movie.
      (iii) How will you evaluate what your students learned? E.g., quiz, survey, other.
   b. Teach your lesson. You may want to practice your lesson first.
   c. Evaluate what your students learned. Give your students an assignment that should take 5 minutes to complete. Collect this assignment and grade it using a letter grade (A, B, C, D, or F).
   d. Summarize the results of your assignment (number of students with each grade). Give this summary to your instructor. Part of your grade for Lab 4, Part A will be based on what your students learned from your presentation.

The Reactions!

Reaction A.

$$2 \text{HCl (aq)} + 2 \text{K}_2\text{CrO}_4 (aq, \text{yellow}) \rightleftharpoons \text{K}_2\text{Cr}_2\text{O}_7 (aq, \text{orange}) + 2 \text{KCl (aq)} + \text{H}_2\text{O (l)}.$$  

1. Place 10 drops of 2 K$_2$CrO$_4$ (aq) in a test tube. Note the color.
2. Add HCl dropwise until you see a color change.
3. Then, make the solution go back to its original color.
4. Make this reaction shift in one direction or the other with heat.

How can you use this reaction to determine pH?

Reaction B.

$$\text{Al(NO}_3\text{)}_3 (aq) + 3 \text{NaOH (aq)} \rightleftharpoons \text{Al(OH)}_3 (s) + 3 \text{NaNO}_3 (aq)$$  

(B1)
\[
\text{Al(OH)}_3 (s) + \text{NaOH (aq)} \rightleftharpoons \text{NaAl(OH)}_4 (aq)
\]  

(B2)

1. Place 10 drops of \(\text{Al(NO}_3\text{)}_3\) (aq) in a test tube.
2. Add NaOH until you see \(\text{Al(OH)}_3\) (s) form (Reaction B1).
3. Then, add more NaOH until you see \(\text{NaAl(OH)}_4\) (aq) form (Reaction B2).
4. Then, make the reaction shift in the other direction.

How can you use this reaction to dissolve \(\text{Al(OH)}_3\)? Will your method work to dissolve any insoluble hydroxide compound?

Reaction C.
\[
\text{CoCl}_4^{2-} (aq, \text{blue}) + 6 \text{H}_2\text{O (l)} \rightleftharpoons \text{Co(H}_2\text{O)}_6^{2+} (aq, \text{pink}) + 4 \text{Cl}^- + \text{heat.}
\]

Co\((\text{NO}_3\text{)})_2\) (aq) is the source of \(\text{Co(H}_2\text{O)}_6^{2+}\) (aq).

What is a source of Cl\(^-\)?
1. Place 10 drops of Co\((\text{NO}_3\text{)})_2\) (aq) in a test tube. Note the color.
2. Add or remove a reactant or product to make the reaction shift to \(\text{CoCl}_4^{2-}\) (aq).
3. Then, make the solution go back to its original color.

Which solution, the pink one or the blue one, contains more ions? Design an experiment using the Conductivity probe that shows this reaction shift in one direction or the other.

Drierite is used as a dessicant (drying agent). It contains a small amount of \(\text{Co}^{2+}\) as \(\text{CoCl}_2\). How is \(\text{Co}^{2+}\) used in Drierite as a color indicator to tell you if it has absorbed water?

Reaction D.
\[
\text{phenolphthalein (colorless)} + \text{NaOH (aq)} \rightleftharpoons \text{phenolphthalein (pink)}
\]

Look up the structure of phenolphthalein. What change in structure occurs when phenolphthalein turns from colorless to pink? Draw each structure.

1. Place 10 drops of water in a test tube. Add one drop of phenolphthalein. Note the color.
2. Add or remove reactant or product to make the reaction shift in one direction.
3. Then, make the reaction shift in the other direction.

How can you use this reaction in invisible ink?

Reaction E.
\[
\text{CaCl}_2 (aq) + 2 \text{NaOH (aq)} \rightleftharpoons \text{Ca(OH)}_2 (s) + 2 \text{NaCl (aq)}
\]

1. Dissolve a small amount of Ca\(\text{Cl}_2\) (s) in water in a test tube.
2. Add NaOH dropwise. What do you see?
3. Then, make the reaction shift in the other direction.

If you have hard water (water that contains high amounts of \(\text{Ca}^{2+}\) and \(\text{Mg}^{2+}\)), will lime stains form on your kitchen cups, glasses, and sink by using soap? By drinking orange juice?

How do you use this reaction to get rid of lime stains from your kitchen cups, glasses, and sink?

Reaction F.
You looked at these reactions in Lab 2.
\[
\begin{align*}
\text{H}_2\text{CO}_3 (aq) & \rightleftharpoons \text{H}_2\text{O (l)} + \text{CO}_2 (g) \\
\text{H}_2\text{CO}_3 (aq) & \rightleftharpoons \text{H}^+ (aq) + \text{HCO}_3^- (aq) \\
\text{H}_2\text{O (l)} + \text{CO}_2 (g) & \rightleftharpoons \text{H}^+ (aq) + \text{HCO}_3^- (aq) \\
\text{CO}_2 (g) & \rightleftharpoons \text{CO}_2 (aq) \\
\text{HCO}_3^- (aq) & \rightleftharpoons \text{OH}^- (aq) + \text{CO}_2 (g)
\end{align*}
\]

Do this first:
1. Prepare a saturated baking soda (NaHCO\(_3\)) solution. Dissolve 32 g of NaHCO\(_3\) (s) in 200 ml of water. What observation tells you the solution is saturated?
2. Pour 150 ml of the NaHCO\(_3\) solution in a 250 ml filter flask. Add 3-4 drops of phenolphthalein. If the solution is pink, add dilute HCl until the solution is colorless. (What information does the color tell you?)
3. Stopper the flask (make sure it’s tight) and connect vacuum tubing from the side arm to the vacuum line.

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4. Turn on the vacuum. (What happens to the pressure inside the flask?) You should see a gas form (what is the gas?) and the solution will eventually turn pink. Explain what is happening based on Equation (F5).

Now try the same experiment except substitute soda for the saturated NaHCO₃ (aq).
Is a carbonated beverage acidic or basic? What happens to the pH after you open the beverage?
Make the reaction go in one direction by applying a stress.
Then, make the reaction shift in the other direction.

**Waste Disposal:** Cr, Co, and Ag solutions - in heavy metals waste.
Neutralize acids and bases – in sink.

**Part B. You made wintergreen in Lab 1. You will identify the reaction conditions to optimize the yield of wintergreen.**

**Prelab**
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Wintergreen is produced by reacting salicylic acid with methanol. H₂SO₄ is used as a catalyst.
   a. What observation tells you how much products are present?
   b. Is this reaction exothermic or endothermic? Determine ΔH for this reaction. (ΔHᵣ of salicylic acid = -590 kJ/mole, ΔHᵣ of methyl salicylate= -531 kJ/mole)

2. a. Read the Procedure for Part B. List the data you will collect for this experiment. What instrument will you use to collect each piece of data?
b. Identify the result you want to determine from this experiment. What equation or formula will you use to calculate this result?

**Procedure**
You made Wintergreen (methyl salicylate) in Lab 1 by reacting 0.7 ml of methanol with 0.1 g of salicylic acid and 2 drops of H₂SO₄. You heated the reaction mixture in a 70°C water bath for 1 or 2 minutes.

1. Make wintergreen.

2. Optimize the yield.
   a. Using equilibrium principles, identify the reaction conditions, e.g., amounts of reactants (limiting and excess), temperature, pressure, you want to adjust to increase the yield of wintergreen. Describe the conditions you changed and the reasons.
   b. Make a second batch of wintergreen using your new reaction conditions. Did you make more wintergreen? What observation tells you that you made more wintergreen?

3. Draw a conclusion from this experiment. (Summarize your results.)


**Questions**
1. For the wintergreen synthesis, what are the optimal reactions conditions to make a high yield of product? Relate these conditions to equilibrium principles. Include numbers.

2. Aspirin is synthesized by reacting salicylic acid with acetic anhydride (ΔHᵣ = -625 kJ/mole) to make aspirin (acetylsalicylic acid, ΔHᵣ = - 758 kJ/mole) and acetic acid. Discuss reaction conditions to maximize the yield of aspirin.
Lab 5. Change is Difficult: Buffers As Food Preservatives and the Fruit Browning Reaction

What makes fruit turn brown? How can I prevent fruit from turning brown?

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Bring an apple and a banana and carrot to lab for Part B.

2. a. Draw a picture of HCl in aqueous solution. Use your picture to explain why HCl can't be used as a buffer.
   b. Draw a picture of acetic acid in aqueous solution. Use your picture to explain why acetic acid can be used as a buffer.
   c. You want to make a 10 ml of a pH 5 buffer using 0.1 M acetic acid and 0.1 M sodium acetate. Calculate the volumes of 0.1 M acetic acid and 0.1 M sodium acetate to prepare this buffer. Use the Henderson-Hasselbach equation.

3. a. Read the Procedure for Part A. List the data you will collect for this experiment. What instrument will you use to collect each piece of data?
   b. Identify the result you want to determine from this experiment. What equation or formula will you use to calculate this result?

Objectives:
(i) apply equilibrium principles to acid-base reactions - use the Henderson-Hasselbach equation to calculate the amount of acid and base to use in a buffer.
(ii) Prepare buffer solutions.
(iii) Identify the conditions that make fruit turn brown.

Introduction
When you bite into an apple or banana, the inside of the apple or banana turns brown upon exposure to air. Fruit turns brown due to the oxidation of phenolics in the fruit with oxygen. This oxidation reaction is catalyzed by enzymes in the fruit. The action of this enzyme can be prevented, inhibited, or interfered by using ascorbic acid, sodium bisulfite, citric acid, NaCl, sugars, chilling, or blanching (Reference: http://www.food-info.net/uk/colour/enzymaticbrowning.htm).

In this lab, you will try to inhibit fruit browning by adjusting the pH at which the fruit browning enzyme is not active. You will do this by using buffers. Buffers are substances that resist change in pH and consist of a weak acid and its salt. Based on the pKₐ of a weak acid, you can design and prepare a pH buffer. You will prepare two buffers and test them as a food preservative to see if fruit browns slower or faster. You will identify the pH at which fruit stays fresh. You will also test salt as a preservative and compare the results of these tests to your buffers.

Materials
benzoic acid, sodium benzoate
0.1 M acetic acid, sodium acetate
0.1 M H₃PO₄, NaH₂PO₄, Na₂HPO₄, and Na₃PO₄
apples, bananas, carrots
Sodium chloride
1 M HCl
1 M NaOH
Vernier LabPro, pH probe

Caution: strong acids and bases are corrosive

Part A. Buffer preparation.

Procedure
1. a. Your group will be assigned to make two different pH buffers to use in your fruit browning tests. Determine which acid and salt you will use to make each pH buffer.
b. Calculate the volume and concentration of acid solution and salt solution to make 15 ml of buffer solution at your assigned pH. Make the concentration of the acid and salt solutions between 0.1 M and 0.2 M.

c. Check your calculations to your instructor. Make any changes as needed.

d. Prepare your buffer solution. Test the pH of your buffer solution with a pH meter. Make sure you calibrate the pH meter before your test the pH of your buffer. Record the pH.

e. If the pH of your buffer is different than the pH you predicted, dropwise add HCl or NaOH until you attain the desired pH.

**Waste Disposal:** neutralize liquid acids and bases – in sink.
Solids – in trash.

**Report**

1. Report the acid and base you used to make your assigned buffers. Show your calculation of the amounts of acid and base used for each buffer.

2. Draw the structure of the acid part of your buffer and the base part of your buffer. At your buffer’s pH, what is the charge of your buffer?

**Part B. Fruit browning tests**

**Procedure**

1. a. Take several slices of an apple and a banana. Design an experiment to test the rate at which each fruit turns brown with the two pH buffers that you just made and NaCl. Think of this experiment as a reaction rate experiment. What variables can you control to help you determine a rate law for the fruit browning reaction? How could you determine the rate law for the fruit browning reaction?

b. Show your method to your instructor. Make any modifications as needed.

c. Record the time it takes for each fruit to turn brown. (It may take a while). Record your data in Table 1. What data or result goes in the last two columns of Table 1?

d. Share your data with the rest of the class so we can identify the buffer that works best for each fruit.

**Table 1. Fruit Browning Tests with Various Buffers**

<table>
<thead>
<tr>
<th>Fruit</th>
<th>Preservative (e.g., buffer)</th>
<th>Start Time</th>
<th>End Time</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
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</tr>
</tbody>
</table>

2. While you are waiting for your fruit to turn brown, try this lab class experiment with a carrot:

(i) cut a carrot into several equal size pieces. Measure the mass of each carrot piece.

(ii) Place one carrot piece in 50 ml of water. Place a second piece in a _____ g NaCl in 50 ml of water solution. Each group will use a different mass of salt.

(iii) Wait 20 minutes. (Did the carrot turn brown?)

(iv) Remove each carrot from the liquid. Dry and measure the mass.

(v) Compare the mass of each carrot before and after. If the mass has changed, explain why. Summarize your data and results in Table 2. What goes in the last column?
Table 2. Carrot Jerky Experiment

<table>
<thead>
<tr>
<th>Carrot</th>
<th>Original Carrot Mass, g</th>
<th>Mass of NaCl in 50 ml water, g</th>
<th>% NaCl (m/V)</th>
<th>Carrot Mass after soaking, g</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td>0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

(vi) Is there a salt solution in which a carrot’s mass won’t change? If so, what is the concentration of this solution?

3. Draw a conclusion from this experiment. (Summarize your results. Which preservative worked best? What is the % NaCl in a carrot?)

**Waste Disposal:** liquid acids and bases should be neutralized – in sink. Solids – in trash.

**Questions**

1. a. Show your Tables of data and results.
   b. Discuss the effect of pH on fruit browning based on your results. Include numbers. Describe how pH works as a food preservative.
   c. Identify the buffer that works best in preventing each fruit from turning brown. Give reasons. Include numbers.
   d. Discuss the effect of salt on fruit browning based on your results. Include numbers. Describe how salt works as a food preservative. Cite the reference where you found this information.
   e. Speculate on the rate law for the fruit browning reaction. What is the order of reaction with respect to each reactant?

2. a. What substance makes the carrot’s mass increase after soaking in salt water for 20 minutes? What sort of “tonic” solution is this?
   b. What substance makes the carrot’s mass decrease after soaking in salt water for 20 minutes? What sort of “tonic” solution is this?
   c. What is the NaCl concentration that won’t change the carrot’s mass? What sort of “tonic” solution is this?
Lab 6. Who Cut the Cheese? Acid-Base Properties of Milk

How do I make cheese? Should I drink milk if I have an upset stomach? Can I make milk from cheese?

PreLAB
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.

Bring in a sample of milk to lab.
1. a. Is milk a pure substance or mixture?
   b. If milk is a mixture, what is the solvent? What is the solute?

2. Read the “Say Cheese” article in the Appendix on p. 42.
   Identify the experimental variables (data) that you will observe or measure to make cheese. Describe the piece of equipment or instrument that you will use to observe or measure each variable.

3. a. Read the Procedure for Part A. List the data you will collect for this experiment.
   b. Identify the result you want to determine from this experiment.

Objectives:
(i) apply equilibrium principles to acid-base and solubility reactions
(ii) separate the protein from milk and make cheese
(iii) determine the pI and buffer capacity of milk

Introduction
According to the milk industry, “Milk does a body good”. Milk is one of the more complete foods that are found in nature. Whole milk contains vitamins (mainly thiamine, riboflavin, pantothenic acid, and vitamins A, D, and K), minerals (calcium, potassium, sodium, phosphorus, and trace metals), proteins (which include all the essential amino acids), carbohydrates (mainly lactose), and lipids (fats). The only important elements that milk does not have are iron and Vitamin C.

The fat in non-skim milk is dispersed in milk as very small (5-10 microns in diameter) globules. Since the fat is so finely dispersed, it is digested more easily than fat from any other source. Milk is homogenized by forcing it through a small hole. This breaks up the fat globules into smaller ones about 1-2 microns in diameter. The fat in homogenized milk will not separate.

There are three kinds of proteins in milk: caseins, lactalbumins, and lactoglobulins. Casein is the most abundant protein in milk and exists in milk as the calcium salt, calcium caseinate, which forms a micelle. Lactalbumins are the second most abundant protein in milk and are denatured and coagulated by heat.

Lactose is the principal carbohydrate in milk and is the only carbohydrate that mammals synthesize. The enzyme lactase is needed to digest lactose into its two component sugars, glucose and galactose. People who lack this enzyme are lactose intolerant. Lactose is converted to sour lactic acid when milk stands at room temperature. Cultured milk products, such as buttermilk, are made by allowing milk to sour before it is processed.

In this lab, you will investigate the acid-base properties of milk. First, you will make cheese by separating the protein from milk by precipitating the protein and filtering the solid. Then, you will measure the isoelectric point of milk and see how this pH is related to cheese. Also, you will determine the buffer region(s) of milk to see if milk is a good drink to cure an upset stomach.

Finally, you will apply your knowledge of acids and bases to convert monosodium glutamate (MSG), a common flavor enhancer, to either glutamic acid or disodium glutamate.

Materials
Part A
Milk (Students: bring some milk (cow, soy, almond, or other) to lab)
dilute acetic acid, glass stirring rod with an attached rubber policeman, rennin, cheese cloth or equivalent

Part B
Milk (Students: bring some milk (the same you used in Part A) to lab)
Monosodium glutamate (MSG)
0.1 M HCl, 0.1 M NaOH, buret, Vernier pH probe, computer
Caution: acids and bases are corrosive

Part A. De-proteining milk. Make cheese.

Procedure
   Each lab partner should do Step 1 individually. After you have precipitated the casein, combine your portions together for Step 2.
   a. Place 10 ml of milk (also works with skim milk or de-fatted milk) in a 100 ml beaker. Heat the milk in a water bath to 40°C.
   b. Add dilute acetic acid dropwise to the warm milk. After every 5 drops, stir the mixture gently using a glass stirring rod with an attached rubber policeman. Using the rubber policeman, push the casein up onto the side of the beaker so that most of the liquid drains from the solid. Then, transfer the congealed casein to another small beaker in portions. If any liquid separates from the casein in the small beaker, use a Pasteur pipet to transfer the liquid back to the reaction mixture. Continue to add dropwise until you have added 1.0 ml of acetic acid or until you do not see any more casein precipitate out of the milk. Avoid adding an excess of acetic acid to the milk solution as this will cause the lactose in the milk to hydrolyze into glucose and galactose.

2. Combine your precipitated casein samples into one beaker.
   a. Measure the mass of casein.
   b. See the “Say Cheese” article (see Prelab Question 2). What did you substitute for lactic acid in Step 1? Once you figure that out, follow the article to make cheese. What step requires cheese cloth? What determines whether hard cheese or soft cheese is made?
   c. Record your cheese making method in writing.
   d. When you finish making your cheese, measure the mass of cheese. Calculate the mass of cheese per g of milk that you started with. Please don’t eat your cheese. Safety first!

3. Draw a conclusion from this experiment.


Questions
1. a. Name the three types of proteins in milk. Which protein is the principal protein found in milk? Cite the reference where you found this information.
   b. Report the mass of casein (milk protein) collected from your milk sample.
   c. Calculate the percent protein in your milk sample. Show your calculation.
   d. Compare the amount that you found in experiment to the amount of protein on the label of the milk. Calculate the % error. Show your calculation.

2. a. Prepare a table like Table 1 below that compares your method to the “Say Cheese” article method. Describe the function of lactic acid and rennin in cheese making.

<table>
<thead>
<tr>
<th>Step</th>
<th>Your Method</th>
<th>“Say Cheese” Method</th>
</tr>
</thead>
<tbody>
<tr>
<td>Coagulation and precipitation</td>
<td></td>
<td>Heat milk to $T = 29-41 , ^\circ C$</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Add starter culture of bacteria to make lactic acid to drop pH.</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Add rennin.</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Etc.</td>
</tr>
<tr>
<td>Concentration of curd</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ripening</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

b. Discuss your cheese making process. Describe how your cheese making method is different from the “Say Cheese” article cheese making method.
Part B. Determination of the isoelectric point and buffer region of milk by titration.

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. a. Is milk an acid or base? Look up the pH of milk to confirm your answer. Cite the reference where you found this information.
   b. Milk protein has many acidic protons. What is the charge of milk protein at the normal pH of milk? Is milk protein soluble in water at this pH?
   c. What is isoelectric point? Starting from the normal pH of milk, would you add acid or base to get to the isoelectric point of milk? Give reasons.

2. a. A buffer can be used as a food preservative. Describe how you would measure or determine the buffer region of milk.
   b. Is milk a good cure for an upset stomach?

3. a. Read the Procedure for Part B, Step 1. List the data you will collect for this experiment. What instrument will you use to collect each piece of data?
   b. Identify the result you want to determine from this experiment.

Procedure
1. Determine the isoelectric point and buffer region(s) in milk. Try to convert cheese to milk.
   You’ve done many titrations in all of your chemistry classes. You will titrate milk with NaOH and monitor the pH and volume of base added. Then, you will titrate milk with HCl and monitor the pH and volume of acid added.
   a. Prepare and standardize a 0.1 M NaOH solution. Do the same with a 0.1 M HCl solution.
   b. Titrate your milk sample.
      (i) Place _____ ml of milk in a flask.
      (ii) Fill a clean buret with either your standardized NaOH or HCl. (Note: Do the concentrations of the NaOH and HCl have be the same? In other words, should 1 ml of NaOH neutralize the same amount of milk as 1 ml of HCl?)
      (iii) Set up the Vernier LabPro using the pH probe. Calibrate the pH probe. Then, measure the pH of your milk sample.
      (iv) Before you start your titration, set up your experiment in the “Events with Entry” collection mode. When you are ready to start the titration, click on the “Collect” key.
      (v) During your titration, you will add a volume of base or acid from a buret (an “aliquot”) and record the pH after each aliquot is added. You will be prompted to record the volume of base or acid that corresponds to the pH. The LabPro software will automatically graph pH vs. volume of base or acid added.

NOTE: (i) for your milk-NaOH titration, continue adding NaOH to a pH of 11.
(ii) for your milk-HCl titration, record the pH at which cheese starts to form. Continue adding HCl to a pH of 3 (which is the pH of stomach acid). Then, see if you can convert the cheese back to milk. What can you try?
(iii) when you do your titration with HCl, record the volume of HCl as a negative volume, e.g., -1.00 ml instead of 1.00 ml. This will make it easier to combine the two titration curves into one curve in the next step.
   c. Combine your two titration curves (milk titration with base and milk titration with acid) into one titration curve. Once you have your titration curve, determine the end points (at least three), pI, and buffer regions. Take the first derivative and second derivative (for those of you who have taken Calculus, you know what this means) of the titration curve to help you determine the end points.
   d. If you have time, repeat your titration for reproducibility.

2. Acid-base properties of monosodium glutamate (MSG).
   Situation 1: You have low blood pressure so you figure it’s ok to consume a little more sodium. Since you love the taste of MSG, you figure maybe disodium glutamate (DSG) tastes twice as good.
   Situation 2: You have high blood pressure so you figure it’s not ok to consume too much sodium. Since you love the taste of MSG, you figure maybe using glutamic acid instead of MSG will be ok.
   We have solid MSG.
a. Describe an experimental procedure to make DSG from MSG. (Hint: glutamic acid is a ___protic acid. Look up the p___ of this acid. Draw a t_____tion curve.) How will you determine whether you made this compound?

b. Describe an experimental procedure to make glutamatic acid from MSG. (Hint: glutamatic acid is a ___protic acid. Look up the p___ of this acid. Draw a t_____tion curve.) How will you determine whether you made this compound?

3. Draw a conclusion from this experiment. 


Questions
1. a. Show your titration curve of milk. Label each endpoint. Show the charge on the milk protein at each endpoint. Show the pH at which cheese forms.
b. Report the pl of milk. Show how you determined pl on your titration curve. Compare your experimental pl to the true pl. Calculate a % error.
c. You made cheese in Part A. What does the pl of milk have to do with making cheese? In other words, what is the charge on milk protein when curds form? What makes milk protein soluble in water at the normal pH of milk but insoluble at another pH?
d. In Part A, you converted cheese to milk. Is this reaction a reversible reaction? Report your observations to support your answer.
e. State the pH range(s) for the buffer region(s) in milk. When you drink milk and the milk goes in your stomach, does milk work like a buffer? Will milk help an upset stomach? Give reasons.
f. The pH of stomach acid ranges from 1-3 and can rise to 4-5 if food is in the stomach (http://scienceline.ucsb.edu/getkey.php?key=275). What happens to milk when it goes in your stomach?

2. a. Show your titration curve of MSG. Show the structure and charge of MSG (give sign and magnitude) at each end point.
b. At what pH does DSG form from MSG? Support your answer with your titration curve.
c. At what pH does glutamatic acid form from MSG? Support your answer with your titration curve.
d. Report the pl of MSG. Show how you determined pl on your titration curve. Compare your experimental pl to the true pl. Calculate a % error.
e. Soy sauce and Worcestershire sauce contain glutamate (https://en.wikipedia.org/wiki/Glutamate_flavoring#Taste_perception). Besides taste, does glutamate preserve these sauces? (Hint: look up the pH of each sauce (http://www.foodscience.caes.uga.edu/extension/documents/FDAapproximatenHoffoodslacf-phs.pdf)).

3. a. Milk protein at its pl is insoluble in water. Is MSG at its pl soluble or insoluble in water?
b. Are all acids or bases soluble in water at their pl?
c. Casein and glutamic acid has a charge of 0. Why is casein insoluble in water whereas glutamic acid is soluble in water?

Reference
Appendix. “Say Cheese”
Roberta Baxter, ChemMatters, February 1995

Slap it on a burger, melt it onto macaroni, sprinkle it on a pizza, savor a gooey cheesecake. Americans say “cheese” millions of times a day. According to the American Dairy Association, Americans consumed 8.5 billion pounds of cheese in 2000.

The term cheese comes from caseus, the Latin word for cheese. Truly an ancient food, there is evidence of cheese making as long as 8000 years ago. Residues of cheese were discovered in an Egyptian pot that is 2300 years old. Where did cheese originate? Probably in many places, because is likely that people around the world discovered that milk carried in pouches made of animals' stomachs coagulated into cheese. In its simplest form, the curds found in soured milk, when drained from the liquid, are cheese.

There are hundreds of varieties of cheese, but they all are made with the same fundamental process. Milk is coagulated and the whey (liquid) is squeezed out of the curd (solid), which becomes the cheese. We usually think of cheese made from cow's milk, but in some parts of the world, goat and sheep milk are more popular. Bleu cheese and Roquefort are made from goat's milk.

The two broad groups are fresh cheese and aged cheese. Fresh cheese is made from milk that is coagulated by acid. Some examples are cottage cheese, ricotta, and mozzarella. These cheeses have high moisture content and must be eaten soon after manufacture because they spoil easily.

Aged or ripened cheese comes from milk that is fermented (soured) by lactic acid bacteria and also coagulated by enzyme action. The curd is pressed, salted, and then ripened (aged) for months. Cheddar, Parmesan, and Swiss are ripened cheeses. Very hard cheeses are aged up to a year, and the flavor usually gets stronger over time.

Worldwide, there are hundreds of varieties of cheese, but they all begin as milk.

Milk parts
Milk contains water, protein, sugar, fat, and some dissolved ions. Milk is about 87% water by weight, but cheese depends on the nonwater parts, mainly protein and fat. Practically all the water and sugar are removed, which is why it takes 100 kg of milk to make 9.5 kg of cheese.

Milk protein, or casein, is a mixture of molecules with molecular weights ranging from 75,000 to 375,000 daltons (amu). The molecules are folded together to form spheres, called micelles, which are fairly large, with diameters of 30 to 300 nm. Calcium and phosphate ions are interspersed through the micelles (see Figure 1).

The sugar that is found in milk is a variety known as lactose, which is a combination of two smaller molecules, glucose and galactose. Cow's milk contains about 4.5–5.0% lactose.

The fat in milk is found in minute droplets, each surrounded by a thin membrane of protein that helps the oily droplet dissolve in the water. If cow's milk is allowed to stand, the fat droplets float to the top and form a layer of cream.

The manufacturing of cheese causes changes in both the lactose and the casein in milk. During ripening, the fat is also changed somewhat and the protein is changed further, which accounts for the particular texture and flavor of the cheese.

Curds and whey
Let’s follow a cheese maker preparing cheddar cheese, a ripened variety developed in the town of Cheddar, England, in the 17th century. Early colonists brought the process with them to North America and cheddar has become America’s favorite cheese.

Making ripened cheese is a three-step process: coagulation and precipitation of the curd, concentration of the curd, and ripening. Coloring (usually yellow carotene) is added to the milk at the beginning. Another
color the cheese maker could add is the dye annatto (from a tropical plant), which also lends a different shade of yellow.

The curd is first precipitated by warming the milk to 29–41 °C (about 85–106 °F). Then a starter culture of lactic streptococci and lactobacilli bacteria is added. These bacteria are also found naturally in milk and are responsible for turning milk sour. The bacteria change lactose sugar into lactic acid. The two bacteria work in tandem. The action of the streptococci declines as the pH drops. The lactobacilli keep working into the more acidic range.

The acid produced by the bacteria causes the spheres of casein to unfold and begin to link into long chains that are not soluble in water. The casein precipitates into a soft gel filled with the watery whey.

Once the pH of the milk has dropped somewhat from the normal pH of 6.5, the cheese maker adds the enzyme rennin. Rennin (also called chymosin) was originally produced by soaking the fourth stomach of a calf in brine (salt water). Now the enzyme can be produced by fungi or bacteria—much cheaper sources than calf stomachs. Rennin is an acid protease, meaning it breaks down protein bonds in an acidic environment. It is a very efficient enzyme. In pure form, one gram of rennin will coagulate five kilograms of milk. The rennin used in commercial cheese processing is diluted to about 1 part in 4500.

The lactic acid has already caused the casein micelles to unfold. Now the rennin attacks the casein molecules, breaking some of the protein bonds. Under the combined action of the lactic acid and the rennin, the casein forms a precipitate—the curd.

The curd is allowed to set into a gel for an hour or two at temperatures from 21 °C to 35 °C depending on the type of cheese being manufactured. Low temperatures cause soft curds (like those in cottage cheese), while high temperatures give rubbery curds. Because cheddar is a hard cheese, its curd is formed at about 35 °C.

Now it is time to remove the whey, or water solution, that is trapped in the casein gel. The cheese maker cuts the gel into tiny chunks and allows the whey to drain out. Next, the gel is cooked and stirred to squeeze out more whey. Cheddar cheese is cooked at 37 °C for about an hour.

Cheddaring
The heat and the continued enzymatic action of rennin cause the protein to fuse into stringlike filaments, making the curd denser. The mass of curd is cut into large chunks, which are piled on top of one another. As they sit in a vat, more whey oozes out and the filaments of casein are forced closer together. This process is called cheddaring. The pressure and warmth compress the protein filaments into horizontal fibers.

An important step is salting. The salt may be added after the curd is cut into chunks or, if the cheese is pressed into shapes such as wheels or blocks, it can then be soaked in brine. In making cheddar, the salt is added directly to the cut curds. Besides adding to the taste of the cheese, the salt has two other effects: It brings even more whey out of the gel structure and slows the activity of the starter bacteria.

After salting, the cheese is pressed into large blocks and the “green,” or newly made, cheese is placed in a controlled environment to ripen. The atmosphere is like a cave, with temperature at 10 °C, to keep the microbial growth slow and steady, and humidity at 80% to keep the surface from drying out or the ripening organisms from dying.

Ripening is a very complex process in which the changes that have began in the fresh cheese continue. Generally, complex molecules are broken into simpler ones. Any remaining lactose is changed into lactic acid and carbon dioxide; fats into fatty acids; proteins into smaller peptides, single amino acid molecules, and even ammonia, NH₃. Cheddar cheese gives off carbon dioxide gas as it ripens. The cheese begins to develop the distinctive flavor and texture that we know. The flavor of the cheese depends on a balance of the components produced during the ripening stage. Too much of one of the compounds may result in unwanted bitter, rancid, or sulfuric tastes. This is why cheese making is an art as well as a science.

If the cheese maker is producing a mild-flavored cheddar, the cheese will be allowed to ripen for about three months. For a medium flavor, six months is required, while a sharp taste takes even longer.
Cheese making is an ancient art, but it also produces one of our most popular modern foods. If you are a typical cheese eater, you consume over 30 pounds each year—and savor each bite.
Lab 7. You Can’t Win, You Can’t Break Even, You Can’t Quit

Where does the energy come in food? How is burning alcohol like a car engine?

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Bring one food item to lab. Your food item should be burnable.
   a. Look up the nutrition information of your two foods on the food label or on http://www.nat.uiuc.edu/mainnat.html.
   b. How is energy stored in food?
2. a. Read the Procedure for Part A. List the data you will collect for this experiment. What instrument will you use to collect each piece of data?
   b. Identify the result you want to determine from this experiment. What equation or formula will you use to calculate this result?

Objectives:
(i) describe energy transformations in chemical reactions using thermodynamics
(ii) relate and perform calculations using thermodynamic quantities
(iii) measure the energy content (ΔH) of a food
(iv) make a reaction that does not occur (ΔG > 0) occur

Introduction
    Ice cream tastes good (we looked at the thermodynamics of ice cream in Lab 2). BBQ smells good. A candy bar gives you quick energy. We eat food for taste, smell, and for energy. Foods are fuels; we can find energy and nutrition information on food labels.

In this lab, you and your group will investigate the thermodynamics of foods and design experiments to answer the following questions:
   A. How much energy is stored in food?
   B. Is it better to burn a solid, liquid or gas fuel in an engine?
   C. What’s the best way to cook an egg?
   D. Is baking soda stable at room temperature? What temperature should I store baking soda?
   E. How can I use a rubber band to keep me warm?

Materials
Part A: food items that burn (Students: bring one burnable food to lab)

Part B: dry plastic bottle (approximately 0.5 liter)
ethanol
rubbing alcohol (70% isopropanol)
(Students: if you have hair spray, bring it to lab)
Matches

Part C: baking soda
eggs (Students: bring a few eggs to lab)
1 M HCl
0.1 M Na₂CO₃ (aq) solution
gas collection device, e.g., glass tubing in rubber stopper, balloon, rubber glove
ethanol
1 M Pb²⁺ (aq) solution
Vernier temperature probe, computer

Part D: thick rubber bands or rubber tubing

Caution: acids and bases are corrosive

Part A: Determine the energy content in a food.
You will burn food. The burning food heats up water. Use a Vernier temperature probe to measure T.
    1. How will you hold the food item?
2. How much of the food item will you burn?
3. What will you use to hold the water. What properties do you need to consider for this container?
4. How much water will you use?
5. How can you maximize the heat transferred to the water and minimize the heat lost to the environment?

Procedure
1. a. Design an experiment to determine the energy content in food. Describe your experimental procedure.
b. Do your experiment. Calculate the energy content in food in J/g.
c. See the food label. What is the true energy content in this food? You’ll have to convert the energy content in Cal per serving to J/g. Compare this true energy content to your experimental energy content. Calculate % error.
d. (i) Is your experimental energy content close to the true energy content? If not, what part of your experiment do you think led to this large % error? (ii) What part of your experiment would you change or modify to reduce the % error? (iii) Use these experimental changes to measure the energy content of the food.
e. Summarize your data and results in a table. Label this Table 1. What goes in columns 5, 7, and 8?

<table>
<thead>
<tr>
<th>Food</th>
<th>Mass of ___, g</th>
<th>Mass of _____, g</th>
<th>T, °C</th>
<th>q gained by ___, J</th>
<th>True Energy content, J/g</th>
</tr>
</thead>
</table>

2. Share your data with the class. Write your foods and data on the table on the chalkboard.

3. Compare the class results. Draw a conclusion from this experiment. (Which food item has the highest energy content?)


Questions
1. Show your Table 1.

2. a. Compare your experimental energy content to the energy information on the food label. Include numbers. Report % error.
b. Compare the burning of the foods in the class. Describe which food burned best. Include numbers.

3. a. Is the burning of food spontaneous? Give reasons.
b. Is the burning of your food item favored by enthalpy? Give reasons.
c. Is it possible for all of the heat lost by the burning food to be transferred to the water? If not, how is entropy or enthalpy involved? Explain.
d. How does food store energy? How is this energy released?

4. For your food, how much food is needed to heat one cup of water to its boiling point? Show your calculations.

Part B. Fuel testing for an engine or rocket. Sugar, a fuel, is used to make ethanol, another fuel. If your car ran out of gasoline, which fuel, sugar or ethanol, would you put in your gas tank?
Prelab
Spend 15 minutes doing the following activity. Assign a notetaker. Report to class.

1. See how ethanol (liquid) burns.
Pour 5 drops of ethanol on the benchtop. Light a match. Set the ethanol on fire. Should any flammable substance be near the ethanol? What burns better, your food from Part A or ethanol? What color was the flame?

2. A car engine is a type of heat engine. Heat engines need a fuel.
   a. When a fuel burns in a heat engine, should the reaction be:
      (i) exothermic or endothermic?
      (ii) fast or slow?
      (iii) produce work or not?
   b. See Fig. 1. What happens in the hot reservoir? How is work produced? What happens in the cold reservoir?

Fig. 1. Schematic diagram of a heat engine.

3. a. Read the Procedure for Part B. List the data you will collect for this experiment. What instrument will you use to collect each piece of data?
   b. Identify the result you want to determine from this experiment. What equation or formula will you use to calculate this result?

Procedure
1. See how ethanol (gas) burns. Do you think ethanol (gas) burns differently than ethanol (liquid) in Prelab Question 1?
   a. Measure the mass of a clean and dry 0.5 liter plastic bottle. (Make sure the bottle is dry.)
   b. Add 5 drops of ethanol to the bottle. Measure the mass of the bottle and alcohol. Calculate the mass of ethanol. Account for the concentration if the alcohol is not pure. Record your data in Table 2.
   c. You want to vaporize the alcohol in the bottle.
      (i) Cap the bottle. Shake the bottle. Invert it several times. Turn the bottle on its side. Roll the bottle so the alcohol coats the walls of the bottle. Hold the bottle in your warm hands.
      (ii) Take off the cap. Carefully smell (waft) the contents of the bottle. Does it smell like "alcohol"? If so, go to the next step. If not, what should you do?
   d. Place the bottle upright on the bench top. Should any flammable substance be near the bottle? Light a match. Carefully move the match to the opening of the bottle. Drop the match in the bottle.
      (i) What happened? What color was the flame?
      (ii) Carefully touch the bottle. How does it feel?
      (iii) Do you see any liquid condensing on the sides of the bottle? If so, what is the liquid?
      (iv) What chemical reaction occurred?
      (v) Let the bottle cool to room temperature. Measure the mass of the bottle. Calculate the mass of liquid in the bottle.
   e. Based on the mass of alcohol that was in the bottle, calculate the mass of water produced (theoretical yield). Does this mass match the mass of liquid you measured in Step 1d(v)? If not, what is the reason for this discrepancy?
   f. Calculate ΔH of reaction (see Step 1d(iv)) in kJ/mole. Then, calculate the heat in J based on the mass of alcohol that burned in your experiment.
   g. Calculate work for this reaction under standard state conditions. Use H₂O (g). w = - p ΔV = -(Δn) R T
      (i) Based on the mass of alcohol that burned, calculate theoretical moles of water, CO₂, and O₂.
      (ii) Calculate Δn = (moles of gas products) – (moles of gas reactants)
      (iii) Calculate w. (Use R = 8.31 J/mole K)
Table 2. Alcohol in a bottle data and results. (Identify state of matter for isopropanol and sugar.)

<table>
<thead>
<tr>
<th>Ethanol (gas)</th>
<th>Ethanol (liquid)</th>
<th>Isopropanol ( )</th>
<th>Sugar, C_{12}H_{22}O_{11}, ( )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Concentration, %</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of bottle</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of bottle + alcohol</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of alcohol</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Moles of alcohol</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Reaction</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Flame color</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Fast/slow?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hotter/cooler?</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of bottle after reaction</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of water</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Theoretical yield of water</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\Delta H_{\text{reaction}}$, kJ/mole</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$q$, J (from mass of alcohol burned)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Theoretical moles of CO$_2$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Theoretical moles of H$_2$O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Theoretical moles of O$_2$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\Delta n$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$w$, J</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2. Repeat Step 1 except use rubbing alcohol. Was this reaction faster or slower than the ethanol (gas) reaction? Did rubbing alcohol burn hotter or cooler than ethanol?

3. If you have time and an extra bottle or two (ask your instructor),
a. repeat except place the bottle on its side. Find a safe place where the bottle won’t hit anyone. Ignite the alcohol vapor. Measure the distance the bottle travels. OR
b. repeat except use hair spray instead of alcohol.

4. Calculate $\Delta H$, $q$, and $w$ for the ethanol (liquid) burning reaction from the Prelab assignment. Use the same mass of ethanol (liquid) as ethanol (gas) in your calculations. (Remember: look up the $\Delta H_f$ for ethanol (l) and look at the gases only when you calculate $\Delta n$.)

5. In Part A, you burned solid food. Your food likely contained sugar. Calculate $\Delta H$, $q$, and $w$ for the sugar (C$_{12}$H$_{22}$O$_{11}$) burning reaction. Use the same mass of sugar as ethanol in your calculations.
6. Draw a conclusion from this experiment. (Which fuel and state of matter burns best?)

Questions
1. Show your Table 2.

2. a. How is burning alcohol in a bottle like a heat engine? Draw a picture of the bottle with burning alcohol. In your picture, show the hot reservoir and cold reservoir. Label \( q_{\text{hot}} \), work, and \( q_{\text{cold}} \).
   b. All of the heat produced in the alcohol combustion reaction was not converted to work.
      (i) What observation supports this statement?
      (ii) What thermodynamic property (\( \Delta G \), \( \Delta H \), or \( \Delta S \)) supports this statement? In the picture of a cylinder and piston shown below, show how the gas atoms are moving inside the cylinder that supports your answer. (Hint: are the gas molecules moving in the same direction as they collide with the piston?)

   ![Cylinder and Piston Diagram]

   c. Did the theoretical yield of water match the mass of water you measured in Step 1d(v)? If not, what is the reason for this discrepancy?

3. a. Compare \( q \) for the ethanol (gas) combustion reaction to the ethanol (liquid) combustion reaction. Based on \( q \), should you use a gas or liquid to burn in an engine? Give reasons.
   b. Compare \( w \) for the ethanol (gas) combustion reaction to the ethanol (liquid) combustion reaction. Based on \( w \), should you use a gas or liquid to burn in an engine? Give reasons.
   c. Compare how fast or slow the ethanol (gas) burned to the ethanol (liquid) from the Prelab assignment. Based on rate, do you want to burn a liquid or a gas in an engine? Give reasons.
   d. Compare the flame color of the ethanol (gas) combustion reaction to the ethanol (liquid) combustion reaction from the Prelab assignment. Which reaction was hotter? Based on reaction temperature, do you want to burn a liquid or a gas in an engine? Give reasons.
   e. Based on \( q \), \( w \), rate, and temperature, do you want to burn a liquid or a gas in an engine? Give reasons.
   f. Would you use the food you burned in Part A as a fuel in your car? Give reasons.

Part C. Make a non-spontaneous reaction occur.

Prelab
Spend 15 minutes doing the following activity. Assign a notetaker. Report to class.
1. Egg shells are made of CaCO\(_3\).
   (i) Place an egg in a small beaker. Add enough water to completely cover the egg.
   (ii) Place another egg in a second small beaker. Add enough 1 M HCl (aq) to completely cover the egg.
   a. Which reaction is spontaneous? What observation tells you a reaction is occurring?
   b. Confirm your observation in Question 1a by calculating \( \Delta G \) for the following reactions:
      \[
      \text{CaCO}_3 (s) \rightarrow \text{Ca}^{2+} (aq) + \text{CO}_3^{2-} (aq) \tag{1}
      \]
      \[
      \text{CaCO}_3 (s) + 2 \text{H}^+ (aq) \rightarrow \text{Ca}^{2+} (aq) + \text{H}_2\text{O} (l) + \text{CO}_2 (g) \tag{2}
      \]
   c. Is there a temperature at which Reaction (1) occurs or does not occur? If so, calculate this temperature. (Is this a temperature we can reach in CHM 1B lab?)
   d. You’ve seen Reaction (1) before. What is the K for Reaction (1) called? Look up the numerical value of K for this reaction.
   e. Using equilibrium principles, name one way to shift Reaction (1) to make more products.
   f. The substance in eggshell is very similar to the substance in our teeth. Are acidic drinks good or bad for our teeth? Give reasons.

2. You observed that Reaction (1) did not occur but Reaction (2) did. You can use thermodynamics to predict how to get CaCO\(_3\) to dissolve by combining the non-spontaneous Reaction (1) with the energy from a spontaneous reaction to get a new reaction which has a \( \Delta G \) less than 0. In other words, add
   \[
   \text{CaCO}_3 (s) \rightarrow \text{Ca}^{2+} (aq) + \text{CO}_3^{2-} (aq) \tag{1}
   \]
2 $\text{H}^+$ (aq) + $\text{CO}_3^{2-}$ (aq) $\rightarrow$ $\text{H}_2\text{CO}_3$ (aq)  \hspace{2cm} (A)

$\text{H}_2\text{CO}_3$ (aq) $\rightarrow$ $\text{H}_2\text{O}$ (l) + $\text{CO}_2$ (g)  \hspace{2cm} (B)

a. You’ve seen the reverse of Reaction (A) before. What type of reaction is the reverse of Reaction (A)?
b. Add Reactions (1), (A), and (B) together. What equation do you get?
c. Calculate $\Delta G$ for Reactions (A) and (B). Which reaction(s) is/are spontaneous?
d. Add $\Delta G$ for Reactions (1), (A), and (B) together. Do you get $\Delta G$ for Reaction (2)?
e. Where does the energy come from to make Reaction (1) occur?

Note: Place the egg in water in 0.1 M HCl (aq). Leave your eggs in the 0.1 M HCl (aq) until the next lab period. You’ll do something with these eggs in Part C, Step 5.

**Procedure**

You may have baking soda at home. Is it stable to store? What is the shelf life? At what temperature should baking soda be stored?

Baking soda is used as a leavening agent – it reacts with an acid to form $\text{CO}_2$ gas. But when heated without an acid, baking soda decomposes to form $\text{CO}_2$ gas:

$$2 \text{NaHCO}_3 (s) \rightarrow \text{Na}_2\text{CO}_3 (s) + \text{H}_2\text{O}$ (g) + $\text{CO}_2$ (g)

1. a. Does baking soda decompose at room temperature? In other words, is this reaction spontaneous at 25°C? Calculate $\Delta G$ for this reaction to confirm your answer. If this reaction does not occur, at what temperature will this reaction occur? Show your calculations. Record your values in Table 3.
b. Is this reaction favored by enthalpy or entropy or both? State $\Delta H$ and $\Delta S$ to confirm your answer.
c. Using equilibrium principles, name one way to shift this reaction to make more products.

2. a. Measure the mass of a clean and dry test tube.
b. (i) Measure approximately 0.5 g (make sure you know this mass) of baking soda and add it to your clean and dry test tube. Record your data in Table 3.
   (ii) Calculate the theoretical yield of Na$_2$CO$_3$ (s).
c. Clamp the test tube to a ring stand. Attach a rubber stopper with glass tube to the test tube.
d. Figure out a way to collect the gas product of this reaction in a test tube.
e. Decompose the baking soda. Make sure no flammable substances are around.
   (i) Use a Bunsen burner to heat the baking soda in the test tube. Move the flame back and forth along the length of the test tube. Collect the gas that is produced.
   (ii) After about a minute, stop heating the test tube. What observation tells you water is produced in this reaction?
   (iii) Continue heating the baking soda. What observation tells you to stop heating the test tube?

3. a. Test the gas you collected. (Hint: think of a fire extinguisher.)
b. After the test tube has cooled, measure the mass of the remaining solid in the test tube. Does this mass match the theoretical yield you calculated in Step 2b(ii)?

4. Share your data and results with another group.

**Table 3. Baking soda decomposition data and results. What goes in Row 5 and Row 9?**

<table>
<thead>
<tr>
<th>Run 1 (your group)</th>
<th>Run 2 (another group)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\Delta G_{\text{reaction}}$</td>
<td></td>
</tr>
<tr>
<td>$\Delta H_{\text{reaction}}$</td>
<td></td>
</tr>
<tr>
<td>$T$ at which reaction occurs/does not occur, °C</td>
<td></td>
</tr>
<tr>
<td>Mass of baking soda, g</td>
<td></td>
</tr>
</tbody>
</table>
5. Prelab Continued. What is the fastest way to cook an egg (denaturing proteins)? Work with another group so your combined group will have four eggs. Your combined group will cook eggs the following ways:
   i. in hot water
   ii. in steam
   iii. in ethanol
   iv. in a heavy metal (we’ll use Pb\(^{2+}\) (aq))

An egg white is about 90% water and 10% protein and has a pH of 9 to 9.5. An uncooked egg white is not white; it’s translucent. A cooked egg white is white. What’s happening?

From “On Food and Cooking”, Harold McGee, 2004, p. 84: “The protein chain is folded up into a compact wad whose shape is maintained by bonds between neighboring folds of the chain. In the chemical environment of the egg white, most of the protein molecules accumulate a negative electrical charge and repel each other ... so the proteins in a raw egg mostly remain compact and separate from one another as they float in water. When we heat the egg, all its molecules move faster and faster, collide with each other harder and harder, and eventually begin to break the bonds that hold the long protein chains in their compact, folded shape. The proteins unfold, tangle with each other, and bond to each other into a kind of three-dimensional network. There’s still much more water than protein, but the water is now divided up among countless little pockets in the continuous protein network, so it can’t flow together any more. The liquid egg thus becomes a moist solid. And because the large protein molecules have clustered together densely enough to deflect light rays, the initially transparent egg albumen becomes opaque.”

a. In the first sentence of the last paragraph, what bond type holds the protein chain together? To break these bonds, is energy released or absorbed?

b. Check your eggs that you left in HCl. The eggs should be “naked” eggs.
   (i) Carefully, separate the “naked” egg from the liquid. If the egg is not naked, remove the egg from the liquid. Carefully crack the egg shell and place the raw egg in a beaker.
   (ii) The liquid contains HCl. The HCl reacts with _____. What ions should be in this liquid? Determine a way to identify this ion? (Hint: What substance reacts with this ion?)

c. (i) Design an experiment to cook an egg in hot water, steam, ethanol, and Pb\(^{2+}\) (aq)) solution. What observation tells you the egg is cooked? When do you stop your experiment?
   (ii) For the cooking conditions that you think will cook an egg the fastest, carefully insert a temperature probe in the uncooked egg white. You want to measure the temperature at which the egg white proteins denature.
   (iii) Record your data and results in Table 4. Set up Table 4 on your own.
   (iv) Do your experiments.

6. Draw a conclusion from this experiment. (Were you able to make these non-spontaneous reactions occur? How? Which method worked best?)


Questions
Baking soda decomposition:
1. Show your Table 3.

2. a. Compare the theoretical yield of Na\(_2\)CO\(_3\) to the experimental mass of Na\(_2\)CO\(_3\). Does this show you were able to decompose baking soda? Give reasons.
   b. What observation told you water is produced in this reaction?
   c. What observation told you to the reaction was complete?
   d. Briefly describe your gas test observations. What gas was in the test tube? Give reasons.
e. How could you determine the molar mass of the solid decomposition product by freezing point depression or boiling point elevation?

b. Does baking soda have a long shelf life or short shelf life? What quantity or equation supports your answer?
c. You are baking soda bread. You mix flour, baking soda, salt, and milk (instead of buttermilk). What does the buttermilk do? Will your bread still rise if you bake your bread at 375°F? Give reasons.
d. Baking soda is used in some types of fire extinguishers. How does baking soda extinguish a fire?

Egg cooking:
4. a. Show your Table 4.
b. Which conditions cooked an egg the fastest? Include numbers.

5. a. Which cooking conditions did you think would cook an egg the fastest? At what temperature did the egg white cook?
b. Water boils at 100°C. This means the ____ break at this temperature. Based on your answer to Question 4a, what bond type is involved when an egg cooks? Give reasons.
c. In Part B, you compared the burning of liquid fuel to gaseous fuel. Compare the cooking of an egg in hot water to steam. Which egg cooked faster? Explain why an egg cooked faster under these conditions based on the heat equation \( q = ms\Delta T \). (Specific heat of steam = 2.0 J/g °C)
d. Describe how ethanol cooks an egg based on chemical forces. Is entropy increasing or decreasing? Give reasons.
e. Describe how a heavy metal like lead cooks an egg based on chemical forces.

Practice Problems:
6. Glucose (C\(_6\)H\(_{12}\)O\(_6\)) is the carbohydrate we use for energy:
   \[
   C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O + \text{energy (4 Cal/g)}.
   \]
   Glucose undergoes a series of reactions in our body. In the first step of glycolysis, glucose reacts with phosphate to make glucose-6-phosphate:
   \[
   \text{Glucose} + P_i \rightarrow \text{Glucose-6-phosphate} \quad \Delta G = +3.3 \text{ kcal/mole}
   \]  
   ATP is used to supply the energy for biological reactions to occur:
   \[
   \text{ATP} + H_2O \rightarrow \text{ADP} + P_i + H^+ \quad \Delta G = -7.3 \text{ kcal/mole}
   \]
   Show how Reaction (2) combines with Reaction (1) to make the first step of glycolysis occur. Show the overall reaction and calculate \( \Delta G \) for the overall reaction.

7. a. Mg(OH)\(_2\) is a base and antacid (in Milk of Magnesia) but is not soluble in water. Look up the equilibrium constant for the Mg(OH)\(_2\) dissolution reaction or calculate \( \Delta G \) to support this statement.
b. Is there a temperature at which Mg(OH)\(_2\) is soluble? If so, calculate this temperature.
c. Based on equilibrium principles, name one way to shift this reaction toward products.
d. The energy from a spontaneous reaction can be used by a non-spontaneous reaction to make it occur. Reaction (1) is not spontaneous. Reaction (2) is spontaneous. Add Reaction (1) to Reaction (2) to get Reaction (3):
   \[
   \text{Mg(OH)}_2(s) \rightarrow \text{Mg}^{2+}(aq) + 2 \text{ OH}^-(aq) \quad \Delta G = +3.3 \text{ kcal/mole}
   \]
   \[
   \text{ATP} + H_2O \rightarrow \text{ADP} + P_i + H^+ \quad \Delta G = -7.3 \text{ kcal/mole}
   \]
   What spontaneous reaction can you use to make Mg(OH)\(_2\) dissolve in water? Calculate \( \Delta G \) for this spontaneous reaction. Then, add \( \Delta G \) for Reaction (1) to \( \Delta G \) for Reaction (2). Did you get \( \Delta G \) for Reaction (3)?

Part D. Thermodynamics of a Rubber Band

Procedure

Stretching a rubber band can be represented by the following chemical equation:
Rubber band (unstretched) → Rubber band (stretched).

1. a. Hold a thick rubber band to your upper lip and quickly stretch it. How does your lip feel?
   b. Determine whether the following quantities are greater than 0 or less than 0 for stretching a rubber band. Give reasons.
      (i) q
      (ii) w
      (iii) ΔH
      (iv) ΔG
      (v) ΔS
      (vi) ΔE

2. Draw a conclusion from this experiment. (Summarize your results.)

Questions
1. Show your answers to Procedure Step 1b.

2. Your lab partner just did her first bungee jump off a bridge! But, the winch to haul her up to the bridge isn’t working and she’s stuck in mid-air hanging from the bungee cord (assume the bungee cord is a big thick rubber band). Based on your knowledge of the thermodynamics of rubber bands, you figure that cooling the bungee cord should lengthen the bungee cord and allow your lab partner to reach the ground. Explain why cooling the bungee cord makes the cord longer.

3. How can you use a rubber band to produce useful work? Give one example and provide details.
Lab 8. Energizer Bunny: Batteries and Electroplating

How does a battery work? Can I gold plate my watch?

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Use of a voltmeter or the Vernier voltage probe.
   a. Measure the voltage of a battery. Record the battery brand and type, e.g., AA, C, D. Some of the batteries may be “dead”. Are these batteries really “dead”?
   b. Switch the leads from the battery to the voltmeter. What happens to the voltage?
   c. Using two batteries for which you measured the voltage, connect the two batteries together. Draw a picture of how you connected the batteries together. Make sure you show how the (+) and (-) terminals are connected. Are you connecting the batteries in series or in parallel? Record the voltage.
   d. Based on the measured voltage of each battery, calculate the series or parallel voltage of your connected batteries.
   e. Compare your experimental voltage to the calculated voltage. Are these voltages the same? If not, what could cause the discrepancy?

2. a. Read the Procedure for Part A. List the data you will collect for this experiment. What instrument will you use to collect each piece of data?
   b. Identify the result you want to determine from this experiment. What equation or formula will you use to calculate this result?

Objectives:
(i) identify and balance oxidation-reduction reactions
(ii) apply redox principles and relevant calculations to galvanic cells - make a battery. Label the parts. Determine the voltage produced.
(iii) apply redox principles and relevant calculations to electrolytic cells - use an electrolytic cell to plate a metal onto an object and to split water.

Introduction
From batteries to chrome plating to galvanized nails to the manufacture of aluminum, sodium, magnesium, and chlorine gas to many reactions in the body, electrochemistry plays a significant role in our lives. All electrochemical reactions are electron transfer or oxidation-reduction (redox) reactions.

Electrochemical reactions occur in an electrochemical cell. All electrochemical cells utilize an electrode at which a substance is reduced (the cathode) and an electrode at which a substance is oxidized (the anode). Electrochemical cells also include an electrolyte solution with which to carry current. The electrolyte solution may be a liquid, as it is for a lead-acid battery, but could also be a solid, as in a lithium-iodine battery which is used in heart pacemakers.

Redox reactions that occur spontaneously can be used, in theory, to produce electrical energy. These electrochemical cells are called voltaic cells. The lead-acid battery in your car is a voltaic cell. For those redox reactions that do not occur spontaneously, electrical energy can be supplied to that reaction to make the reaction occur. These cells are called electrolytic cells. So by applying your knowledge of thermodynamics and redox reactions, you can use electrochemistry in this lab activity to make batteries, plate one metal onto another, split water, and convert light to electricity.

Materials
Part A: Cu, Fe nails, Ni, Al, Zn and Pb rods or strips
1M sulfate or nitrate solutions of Cu^{2+}, Ni^{2+}, Fe^{2+}, Al^{3+}, Zn^{2+}, and Pb^{2+}
NOTE: use a MAXIMUM OF 25 ml of each metal ion solution for this experiment. You can use the same solution for different parts of this experiment.
Porous ceramic cups
voltmeters
DC power supply
Vernier LabPro, voltage-current probes
Lemon or potato (Students: bring one to lab)

Part B: Electrolysis of water: H_2SO_4 or NaOH
Part C: PV cell: Pennies (pre-1982)
8 M HNO₃
NaCl

Caution: heavy metals are toxic. HNO₃ is corrosive.
Waste Disposal: if possible, re-use the metal ion solutions in the next part of the lab. Clean solid electrodes and return these electrodes and batteries to community bench for other students to use.

Combine your group with another group (6 students maximum) in this experiment.

Part A. Make batteries (voltaic cells) from scratch.

Procedure

NOTE: use a MAXIMUM OF 25 ml of each metal ion solution for this experiment. You can use the same solution for different parts of this experiment.

1. Construct a Zn/Cu cell. Record your data in Table 1.
   a. (i) Clean the Zn and Cu electrodes with steel wool to remove ____. Measure the mass of each electrode.
   (ii) Use a 50 ml beaker to hold one electrode and solution; use the porous ceramic cup to hold the other electrode and solution. Place the ceramic cup in the beaker. The ceramic cup allows ions to pass through (ionic current) and serves as a ____.
   (iii) Measure the cell voltage with the voltmeter for 5 minutes. Which electrode is the (+) terminal? Is the (+) terminal the more active metal or less active metal? Is the (+) terminal the anode or cathode?
   (iv) Disconnect the wires from each electrode and dry the electrodes. Measure the mass of each electrode.
   b. Draw a diagram of your cell. Label the anode and cathode, charge on each electrode, electrolyte solutions, salt bridge, (+) terminal, and (-) terminal. Indicate the direction of electron flow in the wire.
   c. Next to your diagram, show the half reactions that occur in each half cell. Calculate the cell voltage using standard electrode potentials and compare this voltage to the actual voltage.
   d. For the electrode at which oxidation occurs, what happened to the mass? Did you expect this mass to change? What about the electrode at which reduction occurs?

<table>
<thead>
<tr>
<th>Battery</th>
<th>Anode 1/2 reaction</th>
<th>Cathode 1/2 reaction</th>
<th>Overall Reaction</th>
<th>Calculated $E_{cell}$</th>
<th>Actual $E_{cell}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zn/Cu</td>
<td>Mass before: ___</td>
<td>Mass before: ___</td>
<td>Mass after 5</td>
<td>Mass after 5</td>
<td>Start:</td>
</tr>
<tr>
<td></td>
<td>Mass after 5</td>
<td></td>
<td>minutes: ___</td>
<td>minutes: ___</td>
<td>After 5 minutes:</td>
</tr>
</tbody>
</table>

2. Make at least two different batteries with different combinations of metals. Work with the other groups on your bench.
   a. (i) Using standard reduction potentials, which metal is the best reducing agent? Which metal is the worst reducing agent?
   (ii) Which combination of metals will produce a battery with the highest voltage? Draw a cell diagram. Write the half reactions and overall reaction.
   (iii) Which combination of metals will produce a battery with the lowest voltage? Draw a cell diagram. Write the half reactions and overall reaction.
   b. Assemble the metal and metal ion solution to make each battery. Measure the voltage.
   NOTE: use a MAXIMUM OF 25 ml of each metal ion solution for this experiment. You can use the same solution for different parts of this experiment.
   c. Compare the measured voltage to the calculated voltage for each battery. If the voltages are not the same, account for any discrepancy. Calculate % error.
   d. Record your data in Table 1.
e. Which materials produced the highest voltage battery?

3. Make a lemon battery or potato battery.
   a. Record the materials you used to make your battery. Draw a cell diagram of your battery. Record the voltage. Record your data in Table 1.
   b. Use a different combination of materials to make a second lemon or potato battery. Draw a cell diagram of your battery. Record the voltage. Record your data in Table 1.

   c. Was the voltage from part a the same as part b? How does the choice of electrodes affect voltage?

4. Using your data from Steps 1 and 2, construct a 2 V battery. Record your data in Table 1.

   Note: If you cannot construct a battery of your desired voltage with one cell, try connecting two cells together. Using standard reduction potentials, identify the metals and metal ion solutions that you need to make two batteries that when connected together produce 2 V. Draw a cell diagram.

5. Draw a conclusion from this experiment. (Summarize your results. Relate activity (reducing strength) to battery voltage.)

   Waste Disposal: if possible, re-use the metal ion solutions in the next part of the lab. Clean solid electrodes and return these electrodes and batteries to community bench for other students to use.

Questions
1. Show your Table 1.

2. For your Zn/Cu cell,
   a. draw a picture that shows the flow of electrons through the electrochemical cell (from the electrode through the wire to the other electrode through the solution).
   b. (i) Compare the experimental cell voltage to the calculated cell voltage. Include numbers. Were these voltages the same? If not, give reasons for the difference in voltage.
   (ii) Is the reaction that occurs in this cell spontaneous? Calculate %G to support your answer.
   c. What experimental observation told you the electrode at which reduction occurs? Give numbers.
   d. What happens to the concentration of each electrolyte solution as the battery is used to produce energy? Give reasons.
   e. Using the Nerst equation, calculate the cell voltage when the reactant concentration has dropped to half its original concentration.
   f. Using the Nerst equation and the starting voltage and voltage after 5 minutes, calculate the % discharge of this cell.

3. a. For Step 2, compare your experimental cell voltage of the cell you made to the calculated cell voltage based on standard reduction potentials. Include numbers. Discuss any differences in the experimental and calculated cell voltages.
   b. From the Materials list, which metals should make the highest voltage battery? Support your answer by showing the half reactions, standard reduction potentials, and E_{cell}.

4. a. Draw a diagram of your lemon or potato battery.
   b. What was the salt bridge?
   c. What voltage was produced? Was this voltage close to the voltage based on the two metals you used?

Part B. Using an electrolytic cell for electroplating and electrolysis. (Start Part C while you are doing Part B.)

Procedure
1. You want to plate zinc onto a penny.
   a. (i) Design an experiment to plate a Cu penny with Zn. What is the anode? Cathode? Electrolyte?
   (ii) Show your diagram of your cell and your calculation of the applied voltage you will use to your instructor before you do your experiment. Make any changes as needed.
   (iii) Measure the mass of the penny before you plate it with Zn and after it is plated with Zn.
   (iv) Record your data in Table 2. What goes in the last two columns?
NOTE: use a MAXIMUM OF 25 ml of each metal ion solution for this experiment. You can use the same solution for different parts of this experiment.

b. Using a battery from Part A as your power supply, construct your electrolytic cell and run the electroplating reaction. What observation tells you Zn is plating the Cu? Is this a good time to measure the mass of the penny?
   If you don’t observe anything happening, what should you do?
   c. Substitute the DC power supply for the battery.
   Initially, set the voltage to 0 on the power supply. Gradually increase the voltage until you see something happening. Record your observations.
   Increase the voltage on the power supply until you see bubbles evolving from one or both of the electrodes. Again, record your observations.

Table 2. Electrolytic Cell Data

<table>
<thead>
<tr>
<th>Battery</th>
<th>Anode 1/2 reaction</th>
<th>Cathode 1/2 reaction</th>
<th>Overall Reaction</th>
<th>Calculated $E_{cell}$</th>
<th>Applied $E_{cell}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zn plating of Cu penny</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2. Design an experiment to unplate the Zn from the Zn-plated Cu penny.
   Show your diagram of your cell and your calculation of the applied voltage to your instructor before you do your experiment.
   Try to remove only the Zn and not the Cu. Is it easier to oxidize Zn or Cu? What number gives you this information?
   Measure the mass of the penny after you unplate the Zn from it.
   Record your data in Table 2.
   NOTE: use a MAXIMUM OF 25 ml of each metal ion solution for this experiment. You can use the same solution for different parts of this experiment.

3. Design an experiment to split water by electrolysis.
   Show your diagram of your cell and your calculation of the voltage you will use to your instructor before you do your experiment. When you do your experiment, set the applied voltage to your calculated voltage. Did you see any bubbles form? If not, what does this mean? What should you do?
   Include in your experiment a way to collect the products and test that these products have formed using the burning splint test.
   Compare the amount of each product produced.
   Record your data in Table 2.
   NOTE: use a MAXIMUM OF 25 ml of each metal ion solution for this experiment. You can use the same solution for different parts of this experiment.

4. A lead acid battery uses Pb and PbO$_2$ electrodes in sulfuric acid and is rechargeable. PbO$_2$ is expensive so you will make a lead battery using two Pb electrodes.
   But here’s the problem: Will a battery work with two electrodes of the same material?
   Record your data in Table 2.
   a. Charge up the battery.
   Place two Pb electrodes in sulfuric acid. Connect the power supply to the Pb electrodes. Set the voltage and current on the power supply to apply to your cell. Turn on the power supply. Charge your battery for at least 5 minutes.
   b. Use your battery to run a motor (current = 100 mA, V = 1 V).
   Turn off the power supply. Connect your battery to a motor and the Vernier voltage probe. As the motor runs, record the voltage with time using the LabPro software. At what voltage does the motor stop running?
   c. Use your battery to run a light bulb (current = 100 mA, V = 1 V).
Recharge your battery again. Turn off the power supply. Connect your battery to a light bulb and the Vernier voltage probe. As the light bulb runs, record the voltage with time using the LabPro software. At what voltage does the light bulb stop running?

5. Draw a conclusion from this experiment. (Summarize your results. Relate applied voltage to _____.)

Waste Disposal: if possible, re-use the metal ion solutions in the next part of the lab. Clean solid electrodes and return these electrodes and batteries to community bench for other students to use.

Questions
1. Show your Table 2.
2. In Step 1, you plated a copper penny with Zn. In Step 2, you unplated the Zn plated penny.
   a. Show your cell diagrams for Steps 1 and 2.
   b. How could you use color to tell you zinc is plating the copper?
   c. How is the Zn plating reaction in Step 1 related to the Zn unplating reaction in Step 2?
   d. You measured the masses of the penny before plating, after plating, and after unplating. How do you know you plated the penny with Zn? Was the mass of the penny after unplating the same as the mass of the penny before you plated it? If not, explain why.
   e. In Step 1c, you set the DC power supply and your battery to the same voltage. Did you see any difference in the plating reaction between the DC power supply and your battery? Include your observations to support your answer.
   f. In Step 1d, you increased the voltage until you saw bubbles form on one or both of the electrodes. What are these bubbles? What is the source of these bubbles? Write a chemical equation that shows how these bubbles were produced.
   g. Are the bubbles you saw in this cell the same as the bubbles you saw in the electrolysis of water? At what electrode, anode or cathode, were these bubbles formed in each cell?
3. a. Show your cell diagram for the electrolysis of water. Show your calculation of the applied voltage to split water.
   b. Is the electrolysis of water a spontaneous reaction? Calculate ΔG to support your answer.
   c. Compare the calculated applied voltage to the experimental voltage you applied to split water. If you have to apply a higher voltage to split water than your calculated voltage, explain why.
   d. How did you collect the products of this reaction? Describe your observation of each test. Write a chemical equation that represents your observation as needed.
   e. What is the relative amounts of each gas produced? Does this amount match the balanced chemical equation? Give reasons.
4. a. Before you charged up your lead acid battery, what was the potential difference (voltage) between the two Pb electrodes?
   b. When you charged up your lead acid battery, what was the voltage between the two Pb electrodes?
   c. When you connected the motor or light bulb to your lead acid battery, did your battery discharge fast or slow? Was your lead acid battery producing more current or less current than the Zn/Cu battery you made in Part A?

Part C. Make a penny photovoltaic (PV) cell.
Most solar cells use semiconducting silicon (Si) because there is a lot of Si on the earth’s surface and there is a good energy match between solar energy and the Si energy bands (Si band gap = 1.1 eV). But Si solar cells require very high purity Si and high temperatures to fabricate. The penny PV cell is made of another semiconductor, Cu$_2$O, which has a band gap of 2.0 eV (corresponds to 620 nm), and is easy to make from a copper (pre-1982) penny.

Procedure
Use pre-1982 copper pennies. Pennies made after 1982 are mainly zinc covered with copper and don’t work well.
1. a. Clean several pennies with fine sand paper or steel wool until they shine. In the fume hood, immerse the pennies for 10 seconds in 8 M HNO$_3$. (Caution: HNO$_3$ is corrosive.) Then, rinse the pennies three times with deionized water.
b. Immerse a clean penny in 5 mL of a 0.1 M CuSO₄ solution in a small beaker. Place this beaker in a 45-50°C water bath for 2 hours. Then, remove the penny with tweezers, rinse it twice with deionized water, and allow it to dry in air. The penny's surface should be a reddish color from the formation of Cu₂O. The reddish Cu₂O is photoactive.

**NOTE:** use a **MAXIMUM OF 25 ml of each metal ion solution for this experiment. You can use the same solution for different parts of this experiment.**

2. a. Assemble your photocell in a small beaker. Your photocell consists of the following:
   - electrolyte = 0.1 M NaCl,
   - electrode 1 = Cu₂O penny,
   - electrode 2 = clean penny or Cu strip.

Make sure the alligator clip that is connected to the Cu₂O penny electrode is not touching the electrolyte. Connect the electrodes to a voltmeter. Record the initial voltage.

b. Shine a light bulb at the Cu₂O penny electrode. (Some phones have a flashlight.) Record the voltage.

c. Turn off the light bulb. What happens to the voltage?

d. Repeat Step b. Did you get the same voltage as in Step b?

3. Draw a conclusion from this experiment.

**Waste Disposal:** if possible, re-use the metal ion solutions in the next part of the lab. Clean solid electrodes and return these electrodes and batteries to community bench for other students to use.

**Questions**

1. You made the photoactive penny electrode by reacting Cu (s) with CuSO₄ (aq) to form Cu₂O (s):
   
   \[ \text{Cu (s) + CuSO}_4 (aq) \rightarrow \text{Cu}_2 \text{O (s)} \]

   Identify the oxidizing agent and reducing agent.

2. a. How does a photovoltaic cell work? Cite the reference where you found this information.

   b. In the picture below, the left electrode is the Cu₂O penny and the right electrode is the Cu strip. When sunlight strikes the Cu₂O penny electrode, describe what happens to the electron and how electricity is produced.

   c. What color light is needed to operate the penny PV cell?

[Image of a diagram showing a solar cell with sunlight and a band gap labeled ΔE.]
Lab 9. Want to Live Near a Nuclear Power Plant? Half-Life of $^{40}\text{K}$ and Radiation Protection

Is a banana dangerous? How is radiation measured?

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. a. Look at the Vernier Radiation Monitor. How does it measure radiation?

2. a. List the data you will collect for this experiment. What instrument will you use to collect each piece of data?
   b. Identify the result you want to determine from this experiment. What equation or formula will you use to calculate this result?

Objectives:
(i) distinguish between chemical and nuclear reactions
(ii) describe nuclear radiation sources
(i) measure the half-life of $^{40}\text{K}$
(ii) describe safety, protection - investigate the effect of shielding on beta radiation
(iii) determine the effect of distance on radiation intensity

Introduction
The word “nuclear” usually has a bad connotation. Nuclear bombs, nuclear reactors, nuclear waste, Three Mile Island, and Chernobyl are often associated with this word. As with many things, there are good things that go along with the bad things. For example, nuclear radiation is used to date ancient objects and is used in medicine to diagnose illness, detect tumors, scan images of organs, and treat diseases, such as cancer. However, when you get an X-ray, remember that your vital organs are covered with a _____ apron and the X-ray technician leaves the room.

We are exposed to nuclear radiation every day from natural and man-made sources. Background radiation includes external sources as well as internal sources. External sources of background radiation include naturally occurring radioactive isotopes in the soil, food, and water. Internal sources of background radiation include radioactive carbon-14, of which there is one radioactive C-14 isotope for every $10^{12}$ carbon atoms, and radioactive potassium-40. Each gram of potassium contains 0.00012 g of K-40 (% abundance = 0.012%). In a 70 kg human, there is $1.6 \times 10^4$ g of carbon and 140 g of potassium.

In the first part of this lab, we will look at background radiation from potassium-40. The radioisotope of this element is a beta emitter with a half-life of $1.28 \times 10^9$ years. You will use a radiation monitor to measure the activity, or the number of nuclei that disintegrate over time. Since the Geiger counter is not completely efficient, you will not be able to count every K-40 nucleus that disintegrates. Some nuclei that disintegrate will escape detection for three reasons. First, some of the radiation will be stopped by the sample before it reaches the Geiger counter. Second, not all of the radiation will enter the Geiger counter. Last, the radiation that enters the Geiger counter may not have sufficient energy to produce an electrical signal in the counter. For beta radiation, the radiation monitor has a detection efficiency of 1%. From the activity, you can calculate the rate constant, k, for the radioactive decay of K-40:

\[
\text{Rate (activity)} = k \cdot N
\]

where \(N\) = number of radioactive nuclei.

The half-life of K-40 is then calculated:

\[
t_{1/2} = \frac{0.693}{k}.
\]

Next, we will study various methods to protect oneself from radiation exposure. You will test several types of shielding to determine their relative effectiveness. Then, you will study the effect of distance from a radiation source and develop a mathematical model for the intensity of radiation as a function of distance from the source.
Materials
Computer, Vernier LabPro, Vernier radiation monitor
KCl
paper, lab coat, Al foil

Caution: K-40 is a beta emitter.

There are six radiation monitors. Your lab will divide into six groups to do this lab. Your instructor will demonstrate how to set up the computer and radiation monitor.

Part A. Measure the activity and calculate the half-life of K-40.

Procedure
1. Attach a ring on a ring stand. Attach a clamp to the ring so the clamp is pointed downward in a vertical position. Clamp the radiation monitor to the clamp so the radiation monitor is vertical with the detector pointing toward the bench. Adjust the height of the radiation monitor so the detector is 1 inch from the bench top.

2. a. (i) Turn on the Digital Radiation Monitor. Set the mode switch to “Total/Timer.” Make sure the timing period is set to at least 3 hours (see p. 6-7 on http://www.vernier.com/files/manuals/drm-btd_2.pdf)
(ii) Connect the radiation monitor to the DIG/SONIC1 port on the LabPro.
(iii) Connect the LabPro to the computer via the USB port.
b. Open the Vernier LoggerPro software.
(i) Go to the “Experiment” pull-down menu and select “Set Up Sensors” and then select “Show All Interfaces”. (Or click on the green icon that looks like the LabPro box in the upper left corner.)
(ii) In the dialog box, click on the “DIG/SONIC1” box, go down to “Choose Sensor” and then choose “Radiation”. Close the dialog box.
(iii) Click on the button that looks like a graph (the “Data Collection” button) just to the left of the green “Collect” button in the upper right corner.
(iv) In the dialog box, select the following options:
Mode: Time-based
Length: 15 minutes (you may have to change the Time box at the bottom of the dialog box from seconds to minutes)
Sampling rate (Interval): 1.5 minutes/sample
Then, click on the Done button. You are now ready to start collecting data.

3. Measure the background radiation for 15 min. Click on the “Collect” button when you are ready to start collecting data. After 15 minutes, the computer should automatically stop collecting data. If it has not, click on the “Stop” key. Then, go to the “Analyze” pull-down menu and select “Statistics”. The LoggerPro software will calculate the average counts/interval. Record this value in Table 1.

4. a. Measure 1 g of KCl to 0.01 g.
b. Place your KCl sample directly beneath the radiation monitor. The sample should be approximately 1 inch from the monitor.
c. Measure the activity of this sample for 15 min. When your experiment is completed, go to the “Analyze” pull-down menu and select “Statistics”. The LoggerPro software will calculate the average counts/interval.
d. Look at your activity vs. time graph. Do a linear fit of the data. Calculate % difference based on the highest and lowest data points and average.

5. Report the activity of K-40 in Table 1. Use the appropriate units. Calculate the half-life of K-40 based on your data. Show your calculations. For the true counts/yr, account for the efficiency (1%) of the radiation monitor.

6. a. Refer to Step 4d.
b. How would you change your Run 1 experiment to reduce the scatter in your activity data? In other words, what experimental variables would you change so the precision is better?
c. Using the changes from Step 6b, measure the half-life of K-40 again.
d. Calculate the average half-life of your two runs.
e. Quantify accuracy and precision.
7. Draw a conclusion from this experiment.

**Waste Disposal:** KCl – in used KCl container.

<table>
<thead>
<tr>
<th>Table 1. Unshielded Radiation and Half-Life Data for the Potassium-40 Decay</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Run 1</strong></td>
</tr>
<tr>
<td>background radiation counts/interval</td>
</tr>
<tr>
<td>K-40 + background counts/interval</td>
</tr>
<tr>
<td>K-40 unshielded counts/interval</td>
</tr>
<tr>
<td>interval, min</td>
</tr>
<tr>
<td>counts/min</td>
</tr>
<tr>
<td>counts/year</td>
</tr>
<tr>
<td>true counts/year</td>
</tr>
<tr>
<td>mass of KCl</td>
</tr>
<tr>
<td>moles of KCl</td>
</tr>
<tr>
<td>moles of K</td>
</tr>
<tr>
<td>nuclei of K-40 in KCl sample</td>
</tr>
<tr>
<td>rate constant, k</td>
</tr>
<tr>
<td>half-life</td>
</tr>
</tbody>
</table>

**Questions**

1. Write a balanced nuclear equation for the radioactive decay of the radioisotope of potassium. What is the reaction order of this reaction? Write a rate equation that represents this reaction.

2. a. Show Table 1.
b. Calculate the average half-life of $^{40}$K from your two runs.
c. Report your % error in the half-life. Discuss the accuracy of this experiment. Include numbers.
d. Does a long half-life mean a fast decay reaction or slow decay reaction? Give reasons.
e. Does $k$ change if you change $N$? If $k$ does not change, what changes?

3. In Step 4d, you did a linear fit of your activity vs. time graph. Do all the data points fit on the line? Based on your % difference, comment on the precision in your data.

**Part B.** The effect of shielding and distance on radiation.

**Procedure**

1. a. Place a piece of paper between the sample and the detector and measure the activity for 15 minutes.
b. Place a lab coat between the sample and the detector and measure the activity for 15 minutes.
c. Place a sheet of aluminum foil between the sample and the detector and measure the activity for 15 minutes.
d. Report your shielding data in Table 2.
Table 2. Unshielded and Shielded K-40 Data.

<table>
<thead>
<tr>
<th>counts/interval</th>
<th>background</th>
<th>K-40 unshielded</th>
<th>K-40 shielded with paper</th>
<th>K-40 shielded with lab coat</th>
<th>K-40 shielded with Al foil</th>
</tr>
</thead>
<tbody>
<tr>
<td>interval</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2. Design an experiment to measure the intensity of radiation as a function of distance from the source. Plot your data in a graph. Develop a mathematical model for the intensity of radiation as a function of distance from the source. Calculate the distance a beta particle travels in air. What does this distance depend on?

3. Draw a conclusion from this experiment.

**Waste Disposal**: KCl – in used KCl container.

**Questions**

1. a. Show your Table 2.
   b. Rank the shielding types from best to worst. Which type of shielding worked best for K-40? Give reasons.
   c. A heavy lab coat or thick rubber gloves shield (block) beta particles. Discuss whether this shielding of beta particles matches your data and results. Include numbers. Provide an explanation if there is not a match.

2. a. Show your radiation vs. distance graph. Report your mathematical model for the intensity of radiation as a function of distance from the source.
   b. Based on your radiation vs. distance graph, determine the distance a beta particle travels in air. Include numbers. What does this distance depend on?
   c. People in the health care and science field who use radiation sources are often told that the best way to minimize exposure to radiation is to stay as far away from the radiation source as possible. Discuss how your results are related to this statement.

**Part C.** Bananas are a good source of potassium. Determine the mass of potassium in a banana.

**Procedure**

Using the radiation monitor, design an experiment to determine the mass of potassium in a banana. Identify the equation or formula that you will use to determine this mass. Identify the experimental variables you want to measure based on the equation or formula. Do your experiment. Summarize your data and results in a table. Show how you calculated the mass of potassium in a banana from the data you collected. Draw a conclusion from this experiment.

**Questions**

1. a. Show your table of data and results.
   b. Look up the mass of potassium in a banana. Compare this mass to the mass of potassium you found in your experiment. Calculate % error.
Lab 10. Heavy Metal: The Color of Meat

Is red meat still raw?

Prelab
Spend 5 minutes doing the following activity. Assign a notetaker. Report to class.
1. Roses are red, just like raw meat.
   a. What compound causes this red color?
   b. What metal is in this compound? Include charge.
   c. What is the name of the organic compound that is bonded to this metal? How many atoms in this compound are bonded to this metal?
   d. The red color is due to the binding of an O\textsubscript{2} molecule to this metal. If O\textsubscript{2} is replaced by water, the color turns _____.

2. a. Read the Procedure for Part A. List the data you will collect for this experiment.
   b. Identify the result you want to determine from this experiment.

Objectives:
(i) relate bonding and structure of coordination compounds to properties, e.g., color
(ii) prepare transition metal compounds
(iii) investigate the color using UV-VIS spectroscopy

Introduction
Transition metal compounds are often colored. The specific color of a transition metal compound is due to the ligand (an atom or molecule bonded to the metal). As a result, scientists can make compounds of desired colors by using different ligands bonded to transition metals. In addition, since each ligand can produce a different color, scientists can tune the color of a substance by using different ligands.

Unlike the main group elements which use only the s and p atomic orbitals, transition metals use the s, p, and d atomic orbitals. The electrons in transition metals that occupy the d orbitals are not involved in bonding but these d orbitals will determine the shape of a transition metal compound. These transition metal compounds have shapes ranging from tetrahedral, square planar, octahedral, and trigonal bipyramid.

In Chem 1A, we learned that when a metal combines with a non-metal, an ionic compound is formed which is held together by an ionic bond. We also learned that a covalent bond forms when one unpaired electron from one atom combines with one unpaired electron from another atom. Transition metals bond to ligands by a coordinate covalent bond. In this type of bond, two electrons from the ligand are shared with the metal.

In this experiment, you start with some meat juice and you will add or substitute a ligand to change its color. Then, you will measure the absorption spectrum of each compound. Based on color and peak wavelength, you will compare the ligand field strength and develop a spectrochemical series.

Materials
Raw meat juice (beef or pork)
NaCl
O\textsubscript{2}
CO\textsubscript{2} (from exhaled breath)
1 M HCl
1 M NaOH
UV-VIS spectrometer

Caution: raw meat has bacteria.

Procedure
Effect of ligand on color
1. Design an experiment to substitute the O\textsubscript{2} in myoglobin with different ligands, e.g., H\textsubscript{2}O.
   Do your experiment. Record your observations, e.g., color.
   Measure the UV-VIS spectrum of each solution.
Waste Disposal: meat juice – in sink.

Questions
1. a. Show your Table of Data and Results.
b. Rank the compounds by wavelength.
c. Which ligand has the highest ligand field strength. Rank the ligands by field strength.