CHEMISTRY 60

LAB

MANUAL

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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 in 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, simply 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.
If an acid is spilled on the floor or lab bench, use the baking soda solution to neutralize the acid. Then, clean up the spill.

If a base is spilled on the floor or lab bench, use the boric acid solution to neutralize the base. Then, clean up the spill.

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.

14. Never pipet by mouth. Use a pipet filler bulb when using a pipet.

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.

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: _____

Laboratory Equipment

wash bottle

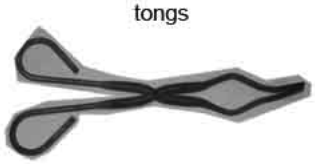
beaker

Erlenmeyer flask

graduated cylinder

spatula

tongs



crucible and cover

wire gauze

test tube clamp and test tube

test tuble rack

pinch clamp

Bunsen burner

Lab 1. Food Labels: Scientific Measurement and Significant Figures

KEY POINTS:

1. Observations are quantified by using scientific instruments or equipment, e.g., balances and graduated cylinders, 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).

Introduction

Look! Up in the sky, it's a bird, it's a plane, ... Whatever you are looking at, you wonder, how big or small is it? How fast or slow is it? How long or short is it? Observations such as these are important in science and chemistry because they allow us to describe matter. Our ability to observe is important. It allows us to discover and create, either by intent or by accident, new things.

Observations can be either qualitative or quantitative. An object that is described as long and heavy and red in color are qualitative observations. Qualitative observations are subject to interpretation. However, your interpretation of an observation may be different than another person's interpretation of the same observation. For example, the long and heavy and red object that you see may be short and light and pink to another person. To make observations less subjective or more objective, we quantify our observations with numbers. An object that is described as 10.3 meters long and has a mass of 55.2 kilograms, and absorbs 645 nanometer light (red light) are quantitative observations. Note that units are attached to these numbers. Units give numbers meaning and context. If I ask you how long this object is and you respond "10.3", I would wonder whether the object is 10.3 inches long, 10.3 feet long, 10.3 meters long or 10.3 miles long.

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 that you are allowed to guess. 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 footlong 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.

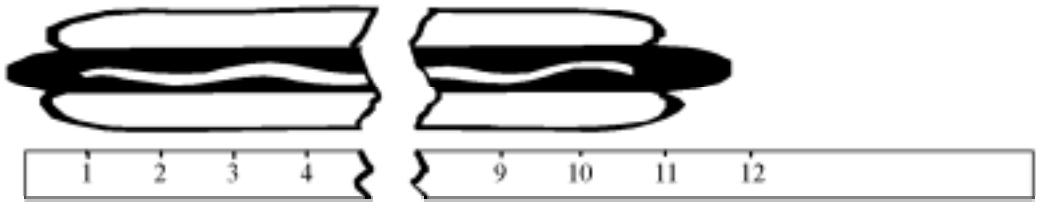


Figure 1. Measuring the length of a sandwich with a ruler and determining the number of significant figures.

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. This ruler is more accurate than ±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. The 3 is the uncertain digit.

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. The last 0 is the uncertain digit.

b. All zeros to the left of the first non-zero digit are not significant. E.g., 0.0020030 has 5 significant figures. The last 0 is the uncertain digit.

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. The 3 is the uncertain digit.

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. The 3 is the uncertain digit.

5. Significant Figures in Calculations

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.,

	22.2 cm	one decimal place - least accurate measurement
+	11.67 cm	two decimal places
	33.87 cm	answer needs to be rounded to one decimal place = 33.9 cm.

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.,

	14.0	three significant figures
x	6.000	four significant figures
	84.0	answer has three significant figures

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.,

22.2 ÷ 14 + 6.000 = ?

22.2÷14 = 1.5|857143 the digits in the intermediate answer to the left of the line between the 5 and 8 are the significant figures

1.5|857143 + 6.000 = 7.5|857143 rounded to 7.6

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. 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:

$$\text{Absolute error} = \text{experimental value} - \text{true value} \qquad (1).$$

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

$$\%\ \text{error} = \frac{\text{absolute error}}{\text{"true" value}} \times 100 \qquad (2).$$

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:

$$\% \text{ difference} = \frac{\text{high - low}}{\text{average}} \times 100 \quad (3).$$

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

Measuring Device	Uncertainty
12 cm ruler	± 0.05 cm
triple beam balance	± 0.05 g
analytical balance	± 0.0001 g
10 ml graduated cylinder	± 0.05 ml
100 ml graduated cylinder	± 0.5 ml
50 ml buret	± 0.02 ml
25 ml volumetric flask	± 0.02 ml
25 ml transfer pipet	± 0.02 ml

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).

References

1. R.A. Day and A. L. Underwood, "Quantitative Analysis", 5th ed., Chapters 1 and 2, Prentice Hall, 1986.

Materials

Nutrition Label from 2 foods

Procedure

1. For each food label:

a. record the name of the food and the nutrition information.

b. For each number you see on the label, determine the quantity (type of measurement) and the units being reported, the number of significant figures, and identify the units as either English units or metric units.

c. Convert each numbers reported in English units to the appropriate metric unit. Show the conversion factor(s) that you used in your calculation.

2. For each label, answer the following questions. Use significant figures appropriately in your calculations.

a. From the mass of fat listed on the label, calculate the calories from fat.

b. From the mass of carbohydrates listed on the label, calculate the calories from carbohydrates.

c. From the mass of protein listed on the label, calculate the calories from protein.

d. Did the company report the number of total calories correctly on the label using significant figures? Give reasons.

e. Based on your calculation in part a, did the company report the number of fat calories correctly on the label using significant figures? Give reasons.

f. Calculate the % fat in one serving based on your answer from part a and part d. Then, calculate the % fat in one serving based on the amount of fat listed on the label. Do your answers match? If not, explain why these answers do not match.

Lab 1 Report Form Name: _____

Fill out Table 1 on the next page. Show your calculations. Use significant figures appropriately.

Table 1. Food Label Calculations

	Food 1 Name:	Units (English/ Metric)	Number of Significant Figures	Food 2 Name:	Units (English/ Metric)	Number of Significant Figures
Mass of fat						
Mass of carbohydrates						
Mass of protein						
Calculated calories from fat						
Calculated calories from carbohydrates						
Calculated calories from protein						
Calculated total calories						
Total calories from label						
Calories from fat from label						
% fat according to your calculations						
% fat according to label						

Lab 2. Properties of Substances: Density

Prelab Questions: 1. a. Give the name and atomic symbol of a metal element. Name one physical property and one chemical property of this element.
b. Give the name and atomic symbol of a non-metal element. Name one physical property and one chemical property of this element.
2. a. Define mass, volume, and density.
b. Describe how you would experimentally measure the density of a liquid.
c. Describe how you would experimentally measure the density of a solid.

KEY POINTS:

1. Properties, such as density, are used to identify substances and distinguish between substances.
2. Density is the ratio of mass to volume. Measure mass and volume to calculate density.
3. Use significant figures in measurements and calculations.

Introduction

How do you identify one person from another? Each person has unique characteristics or properties, such as name, height, weight, hair color, etc. We do the same thing with substances. Substances are characterized by properties which are used to identify substances.

In the first part of this lab, you will observe the properties of various elements and identify a property that the metal elements have in common and a property that the non-metal elements have in common. Next, you will mass and volume observations of common objects and use these observations to describe and distinguish different substances from each other. However, the mass and volume of two different substances could be the same. For example, a baseball and tennis ball has approximately the same volume. One baseball and several tennis balls have the same mass. So, mass and volume are not enough. You will relate the mass and volume of a substance to its density:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad (1)$$

and use this property to distinguish between different substances.

Materials

Part 1. Copper, tin, lead, aluminum, silver, nickel, carbon (graphite), silicon, oxygen, nitrogen, iodine
Part 2. baseball, tennis ball, water, balance, graduated cylinder

Procedure

Part 1. Properties of elements
1. Various elements are exhibited in the lab.

a. For each element, describe two properties. Identify each property as a physical or chemical property. Fill in your observations in Table 1 in the Report Form.
b. Sort the elements into metals and non-metals.
c. What property do the metal elements have in common? What property do the non-metal elements have in common?

2. Your instructor will assign you to look up the boiling point, melting point, and density of one metal element and one non-metal element in the CRC Handbook of Chemistry and Physics. You may also do an internet search or use Wikipedia. Record this information on the table on the chalk board. Do you see a trend in the melting point, boiling point, or density of metals vs. non-metals?

Part 2. Density measurements.
1. Qualitative density observations.
a. Take one baseball in one hand and one tennis ball in your other hand. Which ball has the greater mass? Is the mass of the baseball much less than, slightly less than, about the same, slightly greater than, or much greater than the tennis ball?
b. Compare the volume of the two balls. Is the volume of the baseball much less than, slightly less than, about the same, slightly greater than, or much greater than the tennis ball?
c. Which ball is more dense? Explain your answer by using the density equation (1) in the Introduction.

2. Measure the density of a solid using the displacement method. Your instructor will assign you to measure the density of a:
(i) solid metal element,
(ii) solid non-metal element, and
(iii) unknown solid
For each solid,
a. Measure the mass. The solid should be clean and dry.
b. Fill a graduated cylinder approximately half-full with water. Record the volume of the water.
c. Carefully place the solid into the graduated cylinder containing the water. You may want to tilt the graduated cylinder and slide the solid down to the bottom of the cylinder. The volume of water displaced by the solid object equals the volume of the solid object.
d. Record the volume of the water and solid.
e. Calculate the density of the solid.
f. Determine the identity of the unknown solid.

3. Measure the density of water.
a. Measure the mass of a clean and dry 50 ml graduated cylinder. Record your data in Table 3 in the Report Form.
b. Add 10 ml of water to the graduated cylinder. Record the volume using the appropriate number of significant figures.
c. Measure the mass of the graduated cylinder and water.

d. Add 10 more ml of water to the graduated cylinder. Record the volume using the appropriate number of significant figures.
e. Measure the mass of the graduated cylinder and water.
f. Repeat Steps d and e three more times.
g. Using the Graphical Analysis software, graph mass of water (on y axis) vs. volume of water (on x axis). Go to the "Analyze" pull down menu and choose Linear Fit. A straight line that best fits your data points will appear. Record the slope. What does the slope represent?

Lab 2 Report Form Name: _____

Part 1. Properties of Elements

Table 1. Common Household Substances and Their Identification

Element	Atomic Symbol	Metal or Non-metal?	Property 1 Physical or chemical?	Property 2 Physical or chemical?	Boiling point	Melting Point	Density

1. List the metal elements:

2. List the non-metal elements:

3. What property do the metal elements have in common?

4. What property do the non-metal elements have in common?

Part 2. Density Measurements

1. Explain your reasons for determining the density of a baseball to a tennis ball.

2. Report your data and results in Table 2. Show your calculations below. (The % error formula is shown Lab 1, Equation (2).)

Table 2. Density Data by Displacement

Substance	Mass of sub-stance, g	Volume of H_2O, ml	Volume of H_2O and substance, ml	Volume of substance, ml	Experi-mental density of substance, g/ml	True density of substance, g/ml	% error

Use significant figures appropriately in this table.

Calculations:

Based on the density of the unknown solid, identify this solid. Give reasons for your choice.

3. Show your data and results for the density of water in Table 3.

Table 3. Mass and Volume Measurements to Determine the Density of Water

Volume of H_2O, ml	Mass of H_2O, g	Density of H_2O = mass/volume, g/ml
10 ml		
20 ml		
30 ml		
40 ml		
50 ml		

Include your graph of mass of water vs. volume of water. Explain what the slope of this line means.

Question

1. What property can you use to distinguish between metal elements and non-metal elements?

Lab 3. Properties of Compounds and Identification of Ions in Tap Water and Bone Meal

Prelab Questions: 1. What types of elements comprise an ionic compound? What types of elements comprise a molecular compound?
2. Name two properties of ionic compounds and molecular compounds. Identify each property as a physical property or chemical property.

KEY POINTS:

1. The chemical formula of a compound shows the atomic symbol of each element in the compound. The subscripts represents the ratio of each element in the compound.
2. Ionic compounds contain a metal cation and non-metal anion. An ionic compound that dissolves in water dissociates into ions and forms an electrolyte solution that conducts electricity.
3. The chemical formula of an ionic compound is determined by identifying the smallest whole number ratio of the cation charge to anion charge so the net charge is zero.
4. For ionic compounds, name the metal first followed by the non-metal. Use the "-ide" suffix for the non-metal.
5. Covalent compounds contain non-metals. A covalent compound that dissolves in water forms a non-electrolyte solution that does not conduct electricity.

Introduction

Our body is composed of 65% oxygen, 18% carbon, 10% hydrogen, 3% nitrogen, 1.5% calcium, 1.2% phosphorus, and small amounts of potassium, sulfur, chlorine, sodium, magnesium, iron, cobalt, copper, zinc, iodine, selenium, and fluorine (http://en.wikipedia.org/wiki/Abundance_of_the_chemical_elements#Human_body_elemental_abundance). These elements are not found in our body by themselves but are combined with other elements as compounds. For example, the carbon in our body is found in carbohydrates, proteins, fats, and nucleic acids.

Compounds are classified as ionic compounds or covalent (molecular) compounds. Ionic compounds are distinguished from covalent compounds by their properties. For example, ionic compounds have high melting points and form ions when dissolved in water, whereas covalent compounds have low melting points and do not form ions when dissolved in water. So, salt has a melting point of 801°C and forms sodium ions and chloride ions when it dissolves in water. Sugar has a melting point of 190°C and stays as the sugar molecule it dissolves in water.

A chemical formula tells us the elements in a compound and describes the number of each element in the compound. An ionic compound contains a metal and non-metal and its chemical formula consists of the simplest whole number ratio of ions such that the sum of the positive charges and negative charges equal zero. For example, sodium chloride (common table salt) is made up of one sodium ion, Na^+, and one chloride ion, Cl^-. The +1 charge on one sodium ion is counterbalanced by the –1 charge on one chloride ion so a 1:1 ratio of these ions results in a neutral charge (+1 + (–1) = 0). The formula of sodium chloride, then, is NaCl:

Na^{+} + Cl^{-} form NaCl.

However, magnesium chloride, which is used as a road de-icer and coagulant to prepare tofu, is made up of one magnesium ion, Mg^{2+}, and two chloride ions, Cl^{-}. Since the magnesium ion has a +2 charge and the chloride ion only has a –1 charge, two chloride ions are needed to counterbalance the +2 charge on the magnesium ion so a 1:2 ratio of Mg^{2+} to Cl^{-} gives an electrically neutral formula (+2 + 2(-1) = 0):

Mg^{++} + 2 Cl^{-} form $MgCl_2$

Note the chemical formula shows the two is written as a subscript after the chloride to indicate that two chlorides ions are needed.

Aluminum oxide, which is the mineral corundum and is used to make aluminum metal and as an abrasive for sandpaper, is made up of two aluminum ion, Al^{3+}, and three oxygen ions, O^{2-} (the oxygen ion is called the oxide ion). Since the aluminum ion has a +3 charge and the oxide ion has a –2 charge, two aluminum ions, which have a total charge of +6, are counterbalanced by three oxide ions, which have a total charge of –6 so a 2:3 ratio of Al^{3+} to O^{2-} gives an electrically neutral formula (2(+3) + 3(-2) = 0):

2 Al^{3+} + 3 O^{2-} form Al_2O_3

Again, note the subscripts, which are written to the right of each element's symbol, tells us the ratio of Al to O in this compound.

For ionic compounds, the chemical formula shows the positive ion (cation), which is a metal, first followed by the negative ion (anion), which is a non-metal. When naming ionic compounds, name the metal first followed by the non-metal and use the "-ide" suffix for the non-metal.

The compounds shown above showed ions formed from atoms (monoatomic ions). Ions can form from molecules. These ions are called polyatomic ions. There are many common ionic compounds that contain polyatomic ions. Follow the same procedure as above to determine the chemical formula based on the ratio of ions, be they monoatomic ions or polyatomic ions.

In Part 1 of this lab, you will identify electrolyte solutions and non-electrolyte solutions by testing conductivity. Electrolyte solutions contain ions that conduct current. Non-electrolyte solutions do not contain ions. Based on the conductivity tests, you will be able to identify the compound contained in each solution as an ionic compound or molecular compound.

In Part 2, you will perform various chemical tests to identify some ions present in tap water and bone meal. Tap water contains dissolved minerals that are ionic salts. Ions are also present in the human body. You will test tap water, bone meal, and an unknown solution for chloride, sulfate, calcium, and iron (III) ions.

Finally, you will practice chemical formula writing and naming ionic compounds with the "Ionic Formula Card Game". This is a game originated by Karen Timberlake, author of the textbook used in the lecture portion of this course.

Materials

Part 1. distilled water, solid sodium chloride, sodium chloride solution, solid sugar, sugar solution, dilute ammonia water, dilute sodium hydroxide, dilute potassium nitrate, ethyl alcohol, glycerol
Part 2. bone meal, 6 M nitric acid, funnel, 0.1 M solutions of NaCl, $AgNO_3$, Na_2SO_4, $BaCl_2$, $Ca(NO_3)_2$, $(NH_4)_2C_2O_4$, $Fe(NO_3)_3$, and KSCN
Part 3. Sets of cards for "Ionic Formula Game", Ion Chart

Procedure

Part 1. Conductivity tests with ionic and molecular compounds
Your lab instructor will demonstrate the use and operation of the conductivity apparatus. Test each solution by dipping the electrodes of the conductivity apparatus into the solution. Take care not to touch the bottom of the container with the electrodes. List each substance and record the brightness of the lamp observed for each test in Table 1 on the Report Form.

Test the following substances:

sodium chloride, solid
distilled water
sodium chloride solution
sugar, solid
sugar solution
dilute ammonia water
dilute sodium hydroxide
dilute potassium nitrate
ethyl alcohol
glycerol

Part 2. Ions in Tap Water and Bone Meal
You will test the following solutions for calcium ion, chloride ion, sulfate ion, and iron (III) ion:

1. bone extract
2. An unknown solution.
3. Deionized (DI) water. DI water does not contain calcium, chloride, sulfate , or iron ions.
4. Tap water.

Record the results of each test in Table 2 on the Report Form.

A. Preparation of bone sample.
It is easier to test a liquid than a solid so you will extract ions present in bone into solution.
1. Place about 1 g of bone meal in a 150 ml beaker.
2. Add 15 ml of DI water and 15 ml dilute nitric acid.
3. Set up a ring stand in the fume hood. Place a wire gauze and the beaker on the ring stand.
4. Warm the solution for about 5 min. Do not boil.
5. Allow the solution to cool until you can handle the beaker safely.

6. Filter the solution. Divide the filtrate (liquid solution) between you and your partner and label it “bone solution”.

B. Testing of solutions.
1. Obtain an unknown containing one of the ions. Record your unknown number.

2. The test for chloride is the addition of silver nitrate ($AgNO_3$) solution. Place about 1 or 2 ml of each of the following solutions in separate clean test tubes:
a. sodium chloride (the known chloride solution). Since you know this solution contains chloride ion, you will want to compare your observation of this solution to the other solutions to determine whether chloride is present.
b. deionized water
c. tap water
d. bone solution
e. unknown solution
Add a few drops of silver nitrate solution to each test tube. Record your observations. Dispose of the solutions in the heavy metals waste container.

3. The test for sulfate is addition of barium chloride ($BaCl_2$) solution. Place about 1 or 2 ml of each of the following solutions in separate clean test tubes:
a. sodium sulfate (the known sulfate solution)
b. deionized water
c. tap water
d. bone solution
e. unknown solution
Add a few drops of barium chloride solution to each test tube. Record your observations. Dispose of the solutions in the heavy metals waste container.

4. The test for calcium ion is addition of ammonium oxalate ($(NH_4)_2C_2O_4$) solution. Place about 1 or 2 ml of each of the following solutions in separate clean test tubes:
a. calcium nitrate (the known calcium solution)
b. deionized water
c. tap water
d. bone solution
e. unknown solution
Add a few drops of ammonium oxalate solution to each test tube. Record your observations. Dispose of the solutions in the heavy metals waste container.

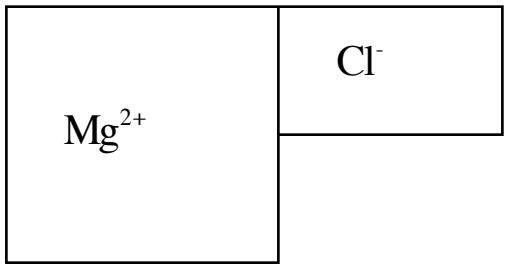
5. The test for iron (III) is addition of potassium thiocyanate (KSCN) solution. Place about 1 or 2 ml of each of the following solutions in separate clean test tubes:
a. iron (III) nitrate (the known iron (III) solution)
b. deionized water
c. tap water

d. bone solution
e. unknown solution
Add a few drops of potassium thiocyanate solution to each test tube. Record your observations. Dispose of the solutions in the heavy metals waste container.

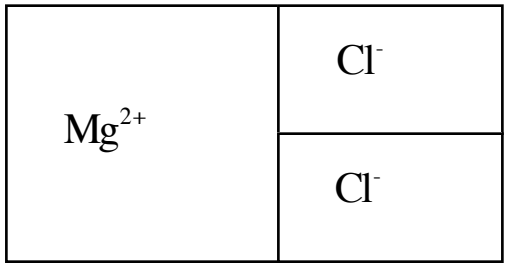
Part 3. Using the Ion Cards.
A. The cards are used to represent ions. The height of the card is proportional to the ionic charge. For example, Mg^{+2} is twice as high as Cl^{-1}.

B. The cards are placed together to form formulas using the following rules:
1. Place the positive ion on the left and the negative ion on the right.
2. Use one, two, or three cards of an ion to complete a rectangle or square.

C. Example: Determine the formula of the compound composed of magnesium ions and chloride ions.
First, place one card of each ion on the table.



Next look to see if the square or rectangle is complete. In this case it is not. You need another chloride ion to complete the figure.



In some cases you will need to add a third card of either ion to complete the figure.

D. Using the ion cards, determine the formulas for the combinations of ions listed below. As you finish each one, record the total positive charges, the total negative charges, and the electrically neutral formula on the Report Sheet using the number of cards of each ion as the subscripts in the formula.

Combinations:

1. copper (I) ion and sulfate ion
2. sodium ion and arsenate ion
3. ammonium ion and arsenate ion
4. ammonium ion and carbonate ion
5. copper (II) ion and nitrate ion
6. aluminum ion and oxide ion
7. calcium ion and phosphate ion
8. silver ion and sulfide ion
9. lead (IV) ion and chromate ion

Ion Cards:

Na^+	Na^+	Na^+	Cu^+	Cu^+	NH_4^+
NH_4^+	NH_4^+	Ag^+	Ag^+	NO_3^-	NO_3^-
Cu^{2+}	SO_4^{2-}	CO_3^{2-}	O^{2-}	O^{2-}	O^{2-}
S^{2-}	CrO_4^{2-}	CrO_4^{2-}	Ca^{2+}	Ca^{2+}	Ca^{2+}

Al^{3+}	Al^{3+}	PO_4^{3-}	PO_4^{3-}	AsO_4^{3-}	
Pb^{4+}					

Lab 3 Report Form Name: _____

Part 1. Conductivity Tests

Table 1. Conductivity Data

Substance	Chemical Formula (include state symbol)	Observation	Compound Type
sodium chloride, solid			
distilled water			
sodium chloride solution			
sugar, solid			
sugar solution			
dilute ammonia water			
dilute sodium hydroxide			
dilute potassium nitrate			
ethyl alcohol			
glycerol			

Part 2. Ions in Tap Water and Bone Meal

Table 2. Data for Ions Tested in Tap Water and Bone Meal

Ion Tested	Test	Known	Deionized Water	Tap Water	Bone Solution	Unknown
Cl^-	add $AgNO_3$					
SO_4^{2-}	add $BaCl_2$					
Ca^{2+}	add $(NH_4)_2C_2O_4$					
Fe^{3+}	add KSCN					

Conclusions:
List each ion present and its relative amount (slight, moderate, or abundant).

Tap water Bone Extract Unknown #_____

Part 3. Writing Ionic Formulas
I. Using Ion Cards

COMBINATIONS:	Total Positive Charge	Total Negative Charge	Write Electrically Neutral Formulas (count the number of each ion used for the subscripts)
copper (I) ion and sulfate ion			
sodium ion and arsenate ion			
ammonium ion and arsenate ion			
ammonium ion and carbonate ion			
copper (II) ion and nitrate ion			
aluminum ion and oxide ion			
calcium ion and phosphate ion			
silver ion and sulfide ion			
lead (IV) ion and chromate ion			

Questions

A. Write formulas for the following binary (2 element) compounds:
Check for charge neutrality after you have written the formula.

1. potassium bromide_____

2. calcium fluoride _____

3. magnesium nitride_____

4. lithium sulfide_____

5. barium oxide_____

6. aluminum sulfide_____

B. Write formulas for the following binary compounds with metals of variable charge:

7. Copper (I) iodide_____

8. Iron (II) sulfide_____

9. Iron (III) sulfide_____

10. Copper (II) iodide_____

11. Cobalt (III) chloride_____

12. lead (IV) oxide _____

C. Write formulas for the following compounds containing a polyatomic ion:

13. sodium nitrate_____

14. sodium sulfate_____

15. calcium nitrate_____

16. potassium phosphate_____

17. ammonium nitrate_____

18. ammonium sulfate_____

19. potassium chromate_____

20. ammonium chromate_____

D. Write formulas for various types:

21. Iron (III) hydroxide_____

22. copper (I) sulfate_____

23. sodium hydrogen carbonate_____

24. zinc carbonate_____

25. magnesium cyanide_____

26. lead (IV) sulfate_____

E. Name the following compounds given the chemical formula. Hint: the compounds with a * need a Roman numeral in middle of name.

27. CaF_2 _____

28. * $FePO_4$ _____

29. * $Sn(NO_3)_4$ _____

30. Na_2CO_3 _____

31. * $Pb(OH)_2$ _____

32. $ZnSO_4$ _____

33. * CuF _____

34. * CuF_2 _____

Lab 4. Chemical Reactions and Quantities: Gas Generation and Identification

Prelab Questions: 1. Find a pure substance in your kitchen, e.g., salt, sugar, baking soda.
a. Give the name, chemical formula, and molar mass of this substance.
b. On the food label, look up the mass of one serving of this substance. Calculate the moles in one serving.
2. In this lab, you will make hydrogen gas by reacting zinc metal with hydrochloric acid to make zinc chloride and hydrogen.
a. Write a balanced chemical equation that represents this reaction.
b. If 2 moles of zinc reacts, how many moles of hydrochloric acid reacts? How many moles of hydrogen are produced?
c. If 0.25 moles of zinc reacts, how many moles of hydrochloric acid reacts? How many moles of hydrogen are produced?

KEY POINTS:

1. The coefficients in a balanced chemical equation represent moles, not mass.
2. Given the mass of one reactant, you can predict the mass of the other substances in a chemical reaction by calculating moles.
3. You will generate a gas in a chemical reaction, collect it, and test it to determine its identity.
4. Each gas has unique properties that allow you to identify it.

Introduction

Hydrogen is considered the fuel of the 21st century. Carbon dioxide is a greenhouse gas. Oxygen is essential for life. Nitrogen dioxide is a reddish-brown gas which is a component of Southern California smog. Where do these gases come from? How is each gas made? What are the properties of each gas? How can I identify each gas?

In this lab, you will make hydrogen by reacting a metal , zinc in this case, with an acid:

zinc + hydrochloric acid ----> zinc chloride + hydrogen.

You will prepare carbon dioxide by reacting an acid with a carbonate containing base:

hydrochloric acid + calcium carbonate ---> calcium chloride + water + carbon dioxide.

You will generate oxygen by the decomposition of hydrogen peroxide. This reaction is very slow so you will use a manganese dioxide catalyst to speed up the reaction:

hydrogen peroxide -----MnO_2 catalyst-------> water + oxygen.

Nitrogen dioxide is produced by reacting copper with concentrated nitric acid:

copper + nitric acid ---> copper(II) nitrate + water + nitrogen dioxide.

For each reaction, you will measure the mass of each reactant and predict the mass of each gas produced. You will be able to make this prediction by:
1. writing a balanced chemical equation to determine the coefficients. The coefficients represent the moles of each reactant and product.
2. Converting mass of reactant to moles of reactant using molar mass as our conversion factor. The mole allow us to count the number of atoms or molecules or ions of a substance by weighing it.
3. Converting moles of reactant to moles of product by using the coefficients in balanced chemical equation which tells us the mole ratio of reactants to products.
4. Converting moles of product to mass of product.

Materials
Gas generating/collecting apparatus: ring stand, clamp, large test tube, delivery tube, trough wooden splints, small test tube, wide mouth bottles, glass squares, test tube clamp, straws
Hydrogen generation: dilute hydrochloric acid, mossy zinc
Carbon dioxide generation: marble chips (calcium carbonate), limewater (saturated calcium hydroxide solution)
Oxygen generation: manganese dioxide, 10% hydrogen peroxide
Nitrogen dioxide demonstration: concentrated nitric acid, copper wire or strip.

Procedure
1. Set up a gas generating/collecting apparatus as shown in Fig. 1. Clamp a large test tube to a ring stand. Attach a delivery tube (glass tubing inserted through a rubber stopper connected to rubber tubing) to the test tube. Place the end of the rubber tubing in the water in the trough. You will perform a chemical reaction in the large test tube to produce a gas. The gas passes through the delivery tube where it will be collected in another test tube or glass bottle by displacement of water.

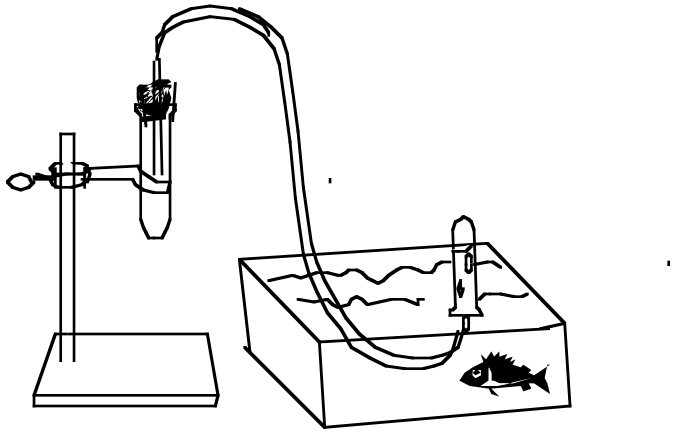


Fig. 1. Gas generating/collecting apparatus.

2. Preparation and Study of Hydrogen
a. Measure and record the mass of several pieces of mossy zinc. Place the zinc in the large test tube you have set up as your gas generator.
b. Add 3 ml of dilute 6 M hydrochloric acid to the large test tube containing the zinc. Place the stopper with the delivery tube onto the large test tube. You should see fizzing occur in the test tube, which is the hydrogen gas being produced, and gas bubbling into the water in the trough at the end of the delivery tube.
c. Quickly fill a small test tube with water. Hold your thumb over the opening of the test tube and invert it upside down. Do not let the water escape. Place the upside down test tube over the end of the delivery tube so the gas bubbles into the test tube. Hold the test tube in place until the hydrogen has displaced all of the water in the tube, then lean the tube mouth down in the corner of the trough. Then, collect a second test tube of gas.
d. Test for hydrogen gas. Light a splint. Remove the first test tube from the water. Hold the tube horizontally and quickly bring the lighted splint to the mouth of the test tube.
If air is present in a test tube with hydrogen, the hydrogen will burn explosively and you should hear a loud bark or pop.
If no air is present in a test tube with hydrogen, the hydrogen will burn quietly at the mouth of the tube.

Answer the Observation Questions (OQ) on the Report Form.
OQ 1. Did the hydrogen burn explosively or quietly?
OQ 2. Did water vapor condense on the inside walls of the tube? Did it fog up?

e. Repeat part d for the second test tube filled with hydrogen.
OQ 3. How did the second tube of hydrogen burn?
OQ 4. When you insert the flaming wood splint into the test tube, what happens to the splint?
OQ 5. What is the color of hydrogen gas?
OQ 6. Did you notice any distinct odor?

f. Feel the bottom of gas generator tube.
OQ 7.What do you notice?

g. Disconnect the gas generator tube, add a little water , and slowly pour the liquid into a beaker so the leftover zinc stays in the tube. Put the zinc in the "waste zinc" container.

2. Preparation and Study of Carbon Dioxide
a. Prepare your gas generator tube for reuse. Fill three wide mouth bottles with water. Hold a glass plate over the mouth of each bottle, turn it upside down, and place it in the water in the trough. Remove the glass plate.
b. Measure and record the mass of about a dozen marble chips (calcium carbonate). Place these marble chips in the gas generator test tube. Add 5 ml of dilute (6M) hydrochloric acid to the tube.

c. When the gas bubbling starts (it may take a few minutes because of debris on the surface of the chips), connect the test tube to the delivery tube and fill the three bottles with carbon dioxide. If the bubbling stops before a bottle is filled with gas, add more acid.
d. When a bottle has filled with gas, remove the bottle from the trough. Quickly turn the bottle right side up and place the glass plate over the mouth.
OQ 8. Observe and record the color of the gas.

e. On the first bottle of gas, test for carbon dioxide by pouring 10 ml of clear limewater (saturated calcium hydroxide solution) into the bottle of gas, replacing the glass plate, and shaking the bottle for a minute.
OQ 9. Observe and record changes in the limewater. Carbon dioxide will always cause this change in limewater. The other gases will not.

f. Test your breath for carbon dioxide by putting a few ml of limewater in a test tube and blowing (exhaling) your breath into it through a straw. Do not inhale!
OQ 10. What do you observe?

g. Light a wooden splint and insert it into the second bottle of carbon dioxide.
OQ 11. Observe and record what happens.

h. Wash out a bottle containing air and add 10 ml of limewater to it. Take your 3rd bottle of carbon dioxide and "pour" the gas into the bottle with the limewater. Place the glass plate over the mouth of the limewater bottle and shake for a minute.
Note: If carbon dioxide is denser than air, it will go downward; if it is lighter than air, it will go upward.

OQ 12. Does the limewater give a positive test for carbon dioxide or not?
OQ 13. Did the carbon dioxide have any odor?

i. Clean out your generator tube as you did before by adding water. Pour the liquid into a beaker and return the marble chips to the "used marble chips" container.

3. Preparation and Study of Oxygen
a. Set up the gas generator tube and gas collecting apparatus. The tube should be dry. Fill one bottle with water and invert it in the water in the trough.
b. Weigh out 0.1 g of manganese dioxide on a piece of weighing paper and transfer it into the gas generator tube. Measure 5 ml of 10% hydrogen peroxide solution. Pour the hydrogen peroxide onto the manganese dioxide and quickly insert the stopper. (This reaction is very fast.) Collect the gas in the bottle.
c. Light a wooden splint and blow out the flame so the splint only glows. Immediately thrust the glowing splint into the bottle containing oxygen. While the splint is burning, withdraw it from the bottle and then reinsert it deeper into the bottle where there is still oxygen.

OQ 14. What happens to the glowing splint when placed in the oxygen?
OQ 15. Compare the burning splint in oxygen and the burning splint in air.
OQ 16. Observe and record the color of oxygen.
OQ 17. Did the oxygen have any odor?

4. Preparation and Study of Nitrogen Dioxide (Demonstration}
Since nitrogen dioxide is a poisonous gas with a choking odor, the Instructor will place a demonstration setup in the hood. It is prepared from copper and concentrated nitric acid. The blue solution you see forming in the bottle is copper nitrate.
Record the mass of copper metal.

OQ 18. Describe the gas.

Lab 4 Report Form Name: _____

Complete and balance the chemical equations that represent the gas producing reactions. Under each reactant and product, write the molar mass of that substance. Record the mass of each reactant that you measured under the molar mass. Calculate the moles of each reactant and product. Calculate the mass of each gas produced.

	Zn +	___HCl --->	$ZnCl_2$ +	H_2
molar mass	_____	_____	_____	_____
mass	_____	_____	_____	_____
moles	_____	_____	_____	_____

___H_2O_2 ----------> 2 H_2O + _____

molar mass	_____	_____	_____
mass	_____	_____	_____
moles	_____	_____	_____

	___HCl +	$CaCO_3$ ------>	$CaCl_2$ +	H_2O +	_____
molar mass	_____	_____	_____	_____	_____
mass	_____	_____	_____	_____	_____
moles	_____	_____	_____	_____	_____

Observation Questions:
HYDROGEN
1. How did the Hydrogen burn?_____

2. Did water vapor condense?_____

3. How the second tube of hydrogen burn?_____

4. What happened to the splint?_____

5. Color of hydrogen_____

6. Any odor?_____

7. Feeling of tube? _____

CARBON DIOXIDE
8. Color of carbon dioxide gas._____

9. Changes in the limewater_____

10. Observation of limewater after blowing breath into it._____

11. Observation of flaming splint in carbon dioxide_____

12. Results of limewater test after pouring gas_____

13. Odor of carbon dioxide_____

OXYGEN
14. Results of glowing splint test_____

15. Comparison of burning of splint in air and in the oxygen_____

16. Color of oxygen_____

17. Odor of oxygen_____

NITROGEN DIOXIDE
18. Description of nitrogen dioxide gas_____

Conclusions:
Summarize the results of each experiment. Include one property of each gas that distinguishes it from the other gases.

Questions

1. Was each gas more dense or less dense than water? Give an observation from this experiment that supports your answer. Hint: how did you collect each gas?

2. Is carbon dioxide more dense or less dense than air? What evidence from your experiments supports your answer?

3. Since helium is less dense than air, when a balloon is filled with helium, it will rise in the air. Would a balloon filled with carbon dioxide rise or sink? Give reasons. What observation in your experiment shows that carbon dioxide is more dense or less dense than air?

4. Which gas would be helpful in extinguishing a fire? What observation from your experiment supports your answer?

5. a. Given two gases, what property, e.g., color, odor, burning, support of burning, and limewater test, could you use to distinguish between the gases?
Note: The first one has been partially done to start you out.

Gases	Property	Describe for each gas
(i) Carbon dioxide and nitrogen dioxide	Color	carbon dioxide =_____
		nitrogen dioxide =_____
(ii) Hydrogen and carbon dioxide		
(iii) Oxygen and hydrogen		

b. You are given a sample of a gas and asked to identify it. This gas could be hydrogen, carbon dioxide, oxygen, or nitrogen dioxide. You make the following observations and tests:
(iv) The gas is colorless. Which gas can you eliminate?

(v) When a sample of the gas is bubbled through limewater, nothing happens. Of the 3 remaining possible gases, which one is now eliminated?

(vi) A glowing splint test on the gas bursts into flame when thrust into a bottle of the gas. What is the gas? Which one is eliminated and why?

Lab 5. Hot Packs and Cold Packs

Prelab Question: 1. How does sweating cool your body? What phase change is occurring? Is this phase change exothermic or endothermic? What gains heat? What loses heat?

KEY POINTS:

1. Every chemical reaction is exothermic or endothermic.
2. A hot pack uses an exothermic reaction to raise the temperature of water.
3. A cold pack uses an endothermic reaction to lower the temperature of water.

Introduction

You are walking with your friend who steps on a rock and twists her ankle. Being a future health professional, you put a cold pack on her ankle to reduce swelling from the injury. But, you wonder, "How does this cold pack work?"

In every chemical reaction, heat is given off (exothermic reaction) or taken in (endothermic reaction). Heat is the energy transferred between two objects due to a difference in temperature. When heat is given off in a chemical reaction, the heat is transferred to the environment surrounding the reaction. For example, if an exothermic chemical reaction occurred in water, the chemical reaction loses heat while the water gains heat. As a result, the temperature of the water rises. If an endothermic reaction occurred in water, the chemical reaction gains heat while the water loses heat. As a result, the temperature of the water drops.

We can apply this heat concept to a cold pack. In a cold pack, ammonium nitrate (NH_4NO_3) dissolves in water. This NH_4NO_3 dissolution reaction gains heat (6.3 kcal/mole) while the water loses heat. The temperature of the water drops and makes the pack cold. The amount of heat gained by the dissolution reaction equals the amount of heat lost by the water:

heat gained by the NH_4NO_3 dissolution reaction = heat lost by the water (1)

The heat gained by the NH_4NO_3 dissolution reaction is calculated by

q = (6.3 kcal/mole) x (moles of NH_4NO_3 dissolved in water) (2).

The heat lost by the water is calculated by

$$q = m\ s\ \Delta T \quad (3)$$

where q = heat
m = mass of object
s = specific heat = the amount of heat required to raise one g of a substance 1^oC,
and ΔT = the change in temperature = $T_f - T_i$.

Note that the amount of heat transferred between two objects depends on the mass, specific heat, and temperature of the object.

In this lab, you will make a cold pack. You will dissolve 10 g of ammonium nitrate in 50 ml of water and record the initial temperature and final temperature of the water. You will calculate the heat gained by the NH_4NO_3 dissolution reaction using Eq. (2) and the heat lost by the water using Eq. (3) and compare these two heats. Then, you will make a second cold pack with 20 g of ammonium nitrate dissolved in 50 ml of water. When the mass of salt is doubled, what happens to the temperature?

Then, you will make a hot pack using calcium chloride ($CaCl_2$). This $CaCl_2$ dissolution reaction loses heat (18 kcal/mole) while the water gains heat.

Not only can you help your friend who twisted her ankle feel better, you can take her mind off her pain by explaining how a hot pack works. (Note: she may need an aspirin afterwards for two reasons.)

Materials

ammonium nitrate
calcium chloride
thermometers or Vernier stainless steel temperature probe, LabPro, and computer
Styrofoam calorimeter cups and cover

Procedure

1. You want to make a 0°C cold pack using 50 ml of water.

a. Obtain a Styrofoam cup and cover. Add 50 ml of water to the Styrofoam cup. Measure and record the temperature of the water. This temperature is the initial temperature of the water, T_i.

b. Add 10 g of ammonium nitrate to the water in the Styrofoam cup. Record the temperature every 15 seconds until the temperature levels off. This temperature is the final temperature of the water, T_f. How close is the final temperature to the desired temperature?

c. Repeat the experiment using 20 g of ammonium nitrate and the same amount of water.

2. You want to make a 40°C hot pack using 50 ml of water.

a. Repeat Step 1a.

b. Add 3 g of calcium chloride to the water in the Styrofoam cup. Record the temperature every 15 seconds until the temperature levels off. This temperature is the final temperature of the water, T_f. How close is the final temperature to the desired temperature?

c. Repeat the experiment using 6 g of calcium chloride and the same amount of water.

Lab 5 Report Form Name: _____

1. Record your data in Table 1. Show your heat calculations below.

Table 1. Hot and Cold Pack Data.

	Mass of salt, g	Initial temperature of water, ^{o}C	Final temperature of water, ^{o}C	ΔT, ^{o}C	Heat gained/lost by water , J	Heat lost/ gained by salt, J
Cold Pack: Run 1						
Cold Pack: Run 2						
Hot Pack: Run 1						
Hot Pack: Run 2						

Questions

1. Did the heat gained/lost by water equal the heat lost/gained by the salt in the cold pack? If not, what is one source of error that could be responsible for this difference?

2. When the mass of salt was doubled in Run 2 for each pack, did your calculated heat lost/gained by water double? Did ΔT of the water double? If not, what is one source of error that could be responsible for this difference?

3. In a hot/cold pack, why is a plastic bag used instead of styrofoam?

Lab 6. The Breathing Process and Gas Laws

Prelab Questions: 1. You have a big balloon and small balloon. Which balloon has more air in it?
2. You put the small balloon in an oven and heat up the balloon. What happens to the balloon? What gas law is applied in this case?
3. You fill a 20 ml syringe with 10 ml of air and then plug the end of the syringe. You push the syringe piston to lower the volume of air in the syringe. What happens to the pressure of air in the syringe? What gas law is applied in this case?

KEY POINTS:

1. Gases are described by pressure, volume, temperature, and moles.
2. Changing one of the quantities in (1) will change the other quantities, e.g., changing the volume of gas changes the pressure.

Introduction

Inhale. Exhale. Breathing is automatic. But how do we breathe? The breathing process can be explained by applying gas laws, specifically the pressure-volume relationship. In this lab, you will simulate the breathing process with a syringe serving as your lungs and the syringe piston as your diaphragm. Using a fixed amount of gas, you will vary the volume and measure the resulting pressure of gas.

The temperature of the air in your car tires rises after you drive a few miles. Does anything happen to the air pressure inside the tires? Or, you may have been warned not to put an aerosol can of hair spray in a microwave oven. Why shouldn't you do this? In the next part of this lab, you will study the pressure-temperature relationship. Using a fixed amount of gas, you will vary the temperature and measure the resulting pressure of gas.

Materials

syringe, Vernier pressure sensor, LabPro, computer

Procedure

Part 1.Pressure-volume relationship of a gas. Using a fixed amount and temperature of gas, you will vary the volume and measure the resulting pressure of gas.
1. a. Connect the pressure sensor the Vernier LabPro (green box) and the LabPro to the computer using a USB cable. Turn on the computer and open the LoggerPro program. The LoggerPro program should automatically detect the pressure sensor. You should see a box on the lower left corner of the screen showing a pressure reading in atm. If the pressure is not in atm, click on the green icon on the top left of the screen (or go to the "Experiment" pull down menu and click on "Set Up Sensors"). A dialog box will appear. Click on the box that shows the pressure sensor. Choose atm for units. Then, click on "Close".
b. To set up your experiment, click on the data collection button (this button has a graph on it) immediately left of the "Collect" button on the upper right side. In the dialog box that appears, select or type in the following options:
Mode: Events with Entry

Column Name: Volume
Short Name: V
Units: ml
Then click “Done”.
c. Move the piston of the 20 ml syringe to the 10.0 ml mark. Attach the 20 ml syringe to the white stem at the end of the pressure sensor box with a gentle half turn.

2. Mimic the breathing process with the syringe as your lungs and the piston as the diaphragm.
a. When you inhale, what happens to your diaphragm? Do this with the syringe and piston. What happens to the volume of your lungs/syringe? How does this change in volume affect the pressure? Does this pressure change correspond to what happens inside your lungs? Explain how air gets into your lungs based on this pressure change.
b. When you exhale, what happens to your diaphragm? Do this with the syringe and piston. What happens to the volume of your lungs/syringe? How does this change in volume affect the pressure? Does this pressure change correspond to what happens inside your lungs? Explain how air leaves your lungs based on this pressure change.

3. a. Click the “Collect” button to start collecting data. Move the piston in the syringe to the 5.0 ml mark. Hold the piston in this position until the pressure reading stabilizes. When the pressure reading is stable, click “Keep”. Type the volume reading in the edit box. Press the Enter key to keep this data pair. If you want to redo a point, you can redo a point by pressing the ESC key (after clicking Keep, but before entering a value).
b. Repeat for volumes 7.5 ml, 10.0 ml, 12.5 ml, 15.0 ml, 17.5 ml, and 20.0 ml.
c. Click the “Stop” button when you have finished collecting data.

4. Look at the graph of pressure vs. volume. Is there a direct or inverse relationship between pressure and volume? To determine the math relationship, go to the “Analyze” pull-down menu,
a. click on the Curve Fit button
b. choose Variable Power (y = Ax^n) from the list at the lower left. Enter the value of n in the Degree/Exponent edit box that represents the relationship shown in the graph, e.g., type 1 if direct relationship, -1 if inverse. Click “Try Fit”.
c. A best-fit curve will be displayed on the graph. If you made the correct choice, the curve should match up well with the data points. If the curve does not match, try a different exponent and click Try Fit again.
d. Once you have determined the relationship, print this graph.

5. Using the equation $P_1V_1 = P_2V_2$, solve for P_2. Use $P_1 = 1$ atm and $V_1 = 10$ ml. For V_2, use the volume reading for each data point. Calculate P_2. Compare your calculated P to the experimental P.

<u>Part 2.</u> Pressure-temperature relationship. Using a fixed amount and volume of gas, you will vary the temperature and measure the resulting pressure of gas.

1. a. Attach the pressure sensor and temperature probe to the Vernier LoggerPro, turn on the computer, and open the LoggerPro program. Set the pressure units to atm. Set the temperature units to ^{o}K. Click on the Data Collection button (next to the Collect button). In the dialog box that appears, select or type in the following options:
Mode: Events with Entry
Column Name: Temperature
Short Name: T
Units: K
Then click "Done".
b. Obtain a rubber stopper assembly with a piece of plastic tubing connected to one of its two valves. Attach the connector at the free end of the plastic tubing to the open stem of the pressure sensor with a clockwise turn. Leave the two-way valve on the rubber stopper open.
c. Insert the rubber stopper assembly into a 125 ml Erlenmeyer flask. Make sure the stopper fits tightly. Then, close the two-way valve above the rubber stopper.

2. a. Fill a 1 liter beaker with about 800 ml of water. Place the flask and temperature probe into the water bath. Make sure the entire flask is immersed in the water bath. Cool the water bath to freezing. Allow the gas in the flask to reach the same temperature as the water bath. When the pressure and temperature readings in the Meter window stabilize, click Keep and type the temperature reading in the edit box to save the pressure-temperature data pair.
b. Heat the water bath about 20^{o}C, let the gas in the flask reach the same temperature as the water bath and record the pressure and temperature readings. Heat the water another 20^{o}C and record the P-T data until you have four or five P-T data points. Make sure you record P-T data from 0^{o}C to 100^{o}C. Do not take your first data point at room temperature, take your second data point at 0^{o}C, and next point at 40^{o}C; you won't be able to do Step 3.

3. a. Look at your graph of pressure vs. temperature. Is there a direct or inverse relationship between pressure and temperature? Determine the math relationship between P and T (in Kelvin) the same way you did in Part 1, Step 3. Remember to use T instead of V.
b. Using the equation $P_1/T_1 = P_2/T_2$, solve for P_2. Use P_1 = 1 atm and T_1 = 298 K. For T_2, use the temperature reading for each data point. Calculate P_2. Compare your calculated P to the experimental P.

Lab 6 Report Form Name: _____

1. Show your graph of P vs. V. Based on your graph, is the pressure of a gas directly proportional or inversely proportional to volume? Give reasons.

2. For your Part 1 data, show a table of your experimental and calculated pressures and volumes. Compare the experimental pressure to the calculated pressure for each data point. Is the experimental pressure the same as the calculated pressure? If not, give one reason for this discrepancy.

3. For your Part 2 data, show a table of your experimental and calculated pressures and temperatures. Compare the experimental pressure to the calculated pressure for each data point. Is the experimental pressure the same as the calculated pressure? If not, give one reason for this discrepancy.

4. Describe the breathing process using gas laws. Why is the diaphragm important in breathing?

Questions

1. In Part 1, you looked at the pressure-volume relationship of a gas and found that gases are compressible. Can you compress a liquid by applying pressure to it? How about a solid?
2. You pour 1 cup (240 ml) of vinegar (0.85 M acetic acid) into a 1 liter bottle. You add 1 g of baking soda to a balloon and fit the balloon snugly over the lip of the bottle. You lift the balloon so that the baking soda in the balloon falls into the vinegar.
a. Write a chemical equation that represents the reaction between baking soda and vinegar. What gas is produced?
b. Describe what happens to the balloon when the baking soda and vinegar are mixed inside the bottle.

Lab 7. Solubility and Osmosis In The Kitchen

Prelab Questions: 1. Write the electron-dot structures of water, NaCl, sugar, and oil ($C_{10}H_{22}$). Determine its polarity of each substance.
2. You add 1 g of salt to 50 ml of water. Calculate the concentration of this solution in molarity and % concentration by mass.

KEY POINTS:

1. The "like dissolves like" rule allows you to predict solubility.
2. The amount of water in fruits or vegetables can be controlled by osmosis.
3. % concentration tells us the amount of solute, such as salt, in a solution.

Introduction

You are cooking a pot of rice (or beans). You add salt to the water and the salt dissolves. You add oil to vinegar to make salad dressing but you see the oil doesn't mix with the vinegar. Why does salt dissolve in water whereas oil does not? In this lab, you will look at the structure and polarity of various substances and investigate the "like dissolves like" principle.

Crispy lettuce makes that salad look very appetizing but the limp lettuce makes you skip the salad. Fruits and vegetables contain water and other substances, such as salt, to make a solution. You will prepare a series of salt solutions and make observations based on osmosis to determine the salt concentration in lettuce or celery. Then, you'll apply your knowledge of osmosis to control the flow of water into or out of lettuce or celery to figure out how to make limp lettuce or celery crisp again.

Materials

water, vinegar, salt, sugar, oil
lettuce or celery

Procedure

Part 1. Solubility and polarity.
1. In each of four test tubes, add 1 ml of water.
a. To test tube 1, add a small spatula of salt. Shake. Record your observations in Table 1 in the Report Form.
b. To test tube 2, add a small spatula of sugar. Shake. Record your observations in Table 1.
c. To test tube 3, add a few drops of vinegar. Record your observations in Table 1.
d. To test tube 4, add a few drops of oil. Record your observations in Table 1.

2. Add 1 ml of vinegar to a test tube. Add 1 ml of oil to the vinegar. Shake. Record your observations in Table 1.

Part 2. Limp or crisp lettuce or celery by osmosis
1. Prepare four 50 ml NaCl solutions of the following concentrations: 0.5%, 1.0%, 2.0%, 5.0% by mass.

2. Take a piece of lettuce or celery. Is the lettuce or celery crisp or limp? Cut the lettuce or celery into five pieces.
a. Place one piece in each of the four solutions you prepared above. Place the fifth piece into deionized water.
b. Wait for 10 minutes. Take each piece of lettuce or celery out of the solution. Is the lettuce or celery crisp or limp?
c. Take a piece of crisp lettuce or celery. What solution would you use to make it wilt? Try this solution and record your observations.
d. Using your wilted piece of lettuce or celery from part c, determine a way to make it crisp again.

Lab 7 Report Form Name: _____

Part 1. Solubility and polarity.
1. Draw the electron-dot structures of water, salt, sugar, vinegar, and oil.

2. Record your solubility observations in Table 1.
Table 1. Solubility Observations.

	polar or non-polar?	Soluble in water?	Soluble in oil?
water		x	
salt			x
sugar			x
vinegar			
oil			x

3. Does the "like dissolves like" rule hold for each mixture? Explain. Then, explain why salt dissolves in water whereas oil does not dissolve (immiscible) in vinegar.

Part 2. Osmosis

Table 2. Lettuce or Celery Crispness in Various Salt Solutions.

Solution	Mass of salt to prepare 50 ml of solution	Molarity	Osmosis observations
water			
0.5% NaCl			
1.0% NaCl			
2.0% NaCl			
5.0% NaCl			

1. Based on your observations, what is the concentration of salt in lettuce or celery? Give reasons.

2. To make limp lettuce or celery crisp, does salt pass into the vegetable or out of the vegetable? Give reasons.

Questions

1. Describe how to make pickles from cucumbers using your knowledge of osmosis.
2. Can you make raisins from grapes the same way pickles are made from cucumbers? Give reasons.

Lab 8. Properties of Acids and Bases and the Determination of Acid in Soda

Prelab Questions: 1. Bring food or beverage or another substance from home to test pH. Also, bring a colorless or light colored soda, e.g., 7-Up, to lab.
2. Name three properties of an acid. Name three properties of a base.
3. a. What is the difference between a strong acid and weak acid? Give one example of each.
b. Will a strong acid have more ions (higher conductivity) in solution than a weak acid? Give reasons.

Chemistry Concepts: acid, base, proton, pH, acid-base strength, acid-base reaction, titration

KEY POINTS:

1. Acids: proton (H^+) donor, sour taste, pH < 7, turns litmus paper red, colorless in phenolphthalein, reacts with metals and bases.
2. Bases: proton (H^+) acceptor, bitter taste, slippery feel, pH > 7, turns litmus paper blue, pink in phenolphthalein, reacts with acids.
3. pH is a measure of the H^+ concentration (pH = - log [H^+]).
4. You can use the properties of acids and bases to identify an acid or base.
5. Conductivity tells you the amount of ions in a solution.

Introduction

A lemon tastes sour. Soap is slippery. These substances and many substances found around your home are either acids or bases. An acid is a substance that donates a proton whereas a base is a substance that accepts a proton. The properties of acid are different than a base, e.g., an acid tastes sour whereas a base tastes bitter; an acid has a pH less than 7 whereas a base has a pH greater than 7; an acid turns blue litmus red whereas a base turns red litmus blue; a base feels slippery whereas an acid does not. An acid reacts with a base; a base reacts with an acid.

Do you freely spend your hard earned money (spender)? Or do you hold on and save your money (saver)? Some acids are “spenders” and easily gives up its proton; these acids are strong acids. Other acids are “savers” and won’t easily give up its proton; these acids are weak acids. Your favorite clothing store or bank are like strong bases; they will easily accept your money just like strong bases easily accept protons.

In this lab, you will test various household substances with litmus paper and for pH to determine whether the substance is an acid or base. You will test the conductivity of various acids and bases to determine its strength. Lastly, you will determine the acid content in soda by doing an acid-base titration. A titration is a quantitative analytical technique in which a known volume of a base is added to an acid until the acid is completely neutralized.

Materials

red and blue litmus paper, pH probe, Vernier LabPro, computer
conductivity apparatus, distilled water, sodium chloride solution, concentrated (glacial) acetic acid, dilute acetic acid (vinegar), dilute hydrochloric acid, dilute sulfuric acid, dilute nitric acid, dilute ammonia water, dilute sodium hydroxide, dilute potassium nitrate

standardized 0.1 M (to 3 or 4 significant figures) NaOH solution, phenolphthalein, buret

Procedure

1. Test the Acidity of Common Household Substances. You will be assigned several household substances to test. Once you make your observations, write your household substance and observations on the board to share with the rest of the class.

a. If your assigned substance is a liquid, go to step b. If your substance is a solid, place a small amount of this substance in a test tube and add about 10 ml of water. Shake the test tube to mix. If you see solid settle to the bottom of the test tube, carefully decant the liquid above the solid into a small beaker. Use this liquid for step b.

b. Test the liquid with red and blue litmus paper. Record your observations in Table 1 in the Report Form.

c. Carefully place the pH probe into the liquid and record the pH in Table 1.

d. Test the substances you brought from home with litmus paper and for pH.

2. Determine the Strength of Acids and Bases by Conductivity.

a. Test the conductivity of each solution by dipping the electrodes of the conductivity apparatus into the solution. For each substance, record the brightness of the lamp observed for each test in the report sheet for this experiment. Record your results in Table 1.

Test the following substances:

distilled water
sodium chloride solution
concentrated (glacial) acetic acid
dilute acetic acid (vinegar)
dilute hydrochloric acid
dilute sulfuric acid
dilute nitric acid
dilute ammonia water
dilute sodium hydroxide
dilute potassium nitrate

b. You will be assigned to test two of the substances on this list with litmus paper and to measure the pH. Record your results in Table 1.

c. Test the conductivity of the household substances that you brought from home. Record your results in Table 1.

3. Determination of the acid content in soda. Your instructor will demonstrate a titration to you and how to use a buret.

a. You will be given the molar concentration a NaOH solution to three or four significant figures. Record this concentration in your lab notes.

b. Obtain a buret, buret clamp, and ring stand.

c. Rinse the buret with a small volume of the standardized NaOH solution. Dispose of the NaOH. Fill the buret with the standardized NaOH solution. Record the initial volume of NaOH in the buret.

d. Measure 25 ml of your soda using a pipet and dispense the soda into a clean Erlenmeyer flask.
e. Add 2 drops of phenolphthalein to the soda.
f. Titrate the soda with the NaOH. When the soda solution turns pink and stays pink for at least 30 seconds, the end point of the titration has been reached and you can stop adding NaOH. Record the final volume of NaOH in the buret.
g. Calculate the concentration of acid in the soda.

Lab 8 Report Form Name: _____

Table 1. Acid-Base Observations of Common Household Substances

Substance	Red Litmus	Blue Litmus	pH	$[H^+]$	Conduc-tivity	Strong or weak?
(your food)						
distilled water						
sodium chloride solution						
concentrated (glacial) acetic acid						
dilute acetic acid (vinegar)						
dilute hydrochloric acid						
dilute sulfuric acid						
dilute nitric acid						
dilute ammonia water						
dilute sodium hydroxide						
dilute potassium nitrate						

2. a. Rank the acids by strength. Explain how acid strength is related to conductivity.

b. Rank the bases by strength.

3. a. What is the acid in your soda that you titrated with NaOH?

b. Write the chemical equation that represents the reaction between the acid in your soda and NaOH.

c. Calculate the concentration of acid in your soda. Show your data, e.g., initial and final volume of NaOH, and calculations.

Questions

1. Is it better to eat a strong acid or weak acid? Give reasons.
2. 25 ml of 0.110 M HCl is titrated with 0.0944 M NaOH.
a. Calculate the pH of 0.110 M HCl. Calculate the pH of 0.0944 M NaOH.
b. Write a balanced chemical equation that represents the reaction between HCl and NaOH.
c. Calculate the volume in ml of 0.0944 M NaOH that is required to react with the 25 ml of 0.110 M HCl.

Lab 9. Nuclear Radiation: Determination of the Half-Life of ^{40}K and The Effect of Shielding

Prelab Questions: 1. Write a balanced nuclear equation for the radioactive decay of the radioisotope of potassium.
2. Where is potassium found in our body?
3. Name the type of shielding required for alpha particle, beta particles, and gamma rays.
4. Bring a stick of licorice or celery stalk for this experiment.

KEY POINTS:

1. Radioactive isotopes are found in our body and in our surroundings.
2. Radioactive isotopes undergo radioactive decay and emit particle radiation, e.g., alpha particles and beta particles.
3. The amount of particle radiation emitted depends on half-life.
4. Shielding is used to protect us from radiation.

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.

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 1.6×10^4 g of carbon and 140 g of potassium.

In 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 1.28×10^9 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:

$$\text{Rate (activity)} = k\,N \qquad \text{where } N = \text{number of radioactive nuclei.}$$

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

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

Then, we will test various types of shielding to determine the relative effectiveness. Finally, we will do an exercise to help us understand the half-life of radioactive decay.

Materials

Computer, Vernier LabPro, Vernier radiation monitor
KCl
paper, lab coat, Al foil
licorice stick or celery stalk (students will bring these materials)

Procedure

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. (Connect the cable from the radiation monitor to the DIG/SONIC1 connector on the LabPro. Connect the LabPro to the computer using the USB cable.)

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

a. 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.

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: 1.5 minutes/sample

Then, click on the Done button. You are now ready to start collecting data.

c. 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.

d. Measure 1 g of KCl to 0.01 g.

e. Place your KCl sample directly beneath the radiation monitor. The sample should be approximately 1 inch from the monitor. Measure the activity of this sample for 15 min.

2. The effect of shielding on radiation.
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.

3. Adapted from K. Timberlake, "Chemistry: An Introduction to General, Organic, and Biological Chemistry", 9th ed., p. 317, "Explore Your World: Modeling Half-Lives".
a. On a piece of paper, draw a vertical axis and horizontal axis. Label the vertical axis as radioactive atoms and the horizontal axis as minutes.
b. Take a licorice stick or celery stalk. Place the licorice stick or celery stalk against the vertical axis and mark its height for zero minutes. Measure the length of the licorice stick or celery stick.
c. In the next minute, cut the licorice or celery in half. Place the shortened licorice or celery at 1 minute on the horizontal axis, mark and measure its height. Every minute cut the licorice or celery in half again and mark and measure the shorter height at the corresponding time. Keep reducing the length by half until you cannot divide the licorice or celery in half any more.
d. On your graph, connect the points you made for each minute. Then, take your measured height and time data and graph your data using the Graphical Analysis software.

Lab 9 Report Form Name: _____

1. Report the activity of K-40 unshielded and with shielding in Table 1. Use the appropriate units.

Table 1. Unshielded Radiation and Half-Life Data for the Potassium-40

		Run 1	Run 2 (from another group)	Average
A	background radiation counts/interval			
B	K-40 + background counts/interval			
C=B-A	K-40 unshielded counts/interval			
D	interval, min			
E=C/D	counts/min			
F	counts/year			
G=F/efficiency	true counts/year			
H	mass of KCl			
I=H/74.6	moles of KCl			
I	moles of K			
J	nuclei of K-40 in KCl sample			
k=G/J	rate constant, k			
L=0.693/k	half-life			

Calculate the half-life of K-40 based on your data. Show your calculations below. For the true counts/yr, account for the efficiency of the radiation monitor.

Table 2. Unshielded and Shielded K-40 Data.

	background	K-40 unshielded	K-40 shielded with paper	K-40 shielded with lab coat	K-40 shielded with Al foil
counts/ interval					
interval					
counts/min					
counts/year					

2. Which type of shielding worked best for K-40? Give reasons.

3. Show your half-life graph with the licorice stick or celery stalk. Based on your graph, what is the half life of the licorice or celery?

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Questions

1. Name one radioisotope that is used in medicine. What type of radiation is emitted by this radioisotope? Briefly describe the use of this radioisotope in medicine.

Lab 10. Organic Chemistry: Structure, Functional Groups, and Properties

Prelab Questions: 1. Draw the electron dot structure of water, NH_3, and CO_2. Use VSEPR theory to determine the geometry of each molecule. Then, determine the polarity of each molecule.

KEY POINTS:

1. Organic compounds are classified by functional groups.
2. Functional groups have different physical and chemical properties.
3. The shape of a molecule determines its polarity.
4. Polarity determines solubility.

Introduction

Natural gas, vinegar, sugar, vegetable oil, plastic. All of these substances are organic compounds. Organic compounds contain carbon; carbon has the ability to bond to other carbon atoms to form a variety of molecules with different shapes, which include chains, rings, and sheets. There are millions and millions of organic compounds. As a result, organic chemistry is the most studied field of chemistry.

Although there are many organic compounds, certain groups of atoms in organic compounds show similar physical and chemical properties. These groups of atoms within organic compounds are called functional groups. Understanding the structure and shape of functional groups allows one to predict properties, such as reactivity and function. For example, structure and geometry studies have led to an understanding of biomolecules and their functions, industrial catalysts such as zeolites and solid surfaces, and synthetic polymers. Furthermore, our understanding of shape enables us to design and create molecules with properties that we desire. We can synthesize almost any shape needed for a specific purpose. For example, scientists have designed and created cages that trap ions of a particular size, molecules that have the shape necessary to bind to only one specific type of molecule (molecules that "recognize" each other), and long-chain molecules that conduct an electric current and thus behave as molecular wires. This amazing variety and complexity of molecules that chemists can now create illustrates a very important aspect of chemistry, namely, that it is a creative science in the material sense. Chemists make new structures that never existed before. Molecular modeling programs now make it even easier to understand and become familiar with the shapes of molecules.

In this lab activity, you will apply your knowledge of bonding, such as valence electrons and the octet rule, to draw a two-dimensional electron-dot structure of various organic compounds. From your electron-dot structure, you will look at the bonds in the molecule and determine the polarity of each bond. From the bond polarity, you will determine the polarity of the molecule as a whole. From the polarity, you can predict whether this substance dissolves in another substance. Then, you will construct a three-dimensional molecular model of your electron-dot structure. Does your electron-dot structure look like your molecular model or vice versa?

You will also look at some chemical properties of alcohols and carboxylic acids. First, you'll test the acid-base properties of alcohols and acids. Then, you will react an alcohol with an

acid in the presence of a catalyst to produce an ester. Many esters have distinctive odors. You'll be able to identify the odor of the esters you make in lab.

Materials
Molecular model kits
ChemDraw and Chem3D software
Ethanol, glacial acetic acid, methanol, salicylic acid, concentrated sulfuric acid
Litmus paper

Procedure
1. a. For each molecule or ion listed in the table below, write the chemical name of the substance beneath its chemical formula, determine the number of valence electrons in each atom, the total number of valence electrons, and draw the electron-dot structure. Circle the functional group(s) in your structure and identify each functional group by name.
Example: H_2O has a total of 8 valence electrons (2 H = 2 x 1 valence electron = 2 valence electrons, O = 6 valence electrons). The electron-dot structure of water is shown in Table 1.

b. List the number and type of each bond in the substance. Determine the polarity of each bond.
Example: H_2O has two O-H bonds, each O-H bond is polar.

c. Construct the molecular model of each molecule or ion to help you visualize the shape of the molecule. Describe the shape of the molecule as best you can.
Example: H_2O is a bent molecule.

d. From the polarity of each bond and the shape, determine the polarity of the molecule.
Example: H_2O has two polar O-H bonds and is bent, so it is a polar molecule.

2. Isomers are two different substances that have the same chemical formula but a different arrangement of atoms in their structure, i.e., each isomer is bonded together differently. $C_2H_2Cl_2$ has three isomers. Draw the electron-dot structure of each isomer. What property would distinguish between each isomer?

3. Acid-base properties of alcohols and carboxylic acids. Preparation of esters.
a. Ethyl acetate
(i) Place 1 ml of ethanol and 1 ml of glacial (concentrated) acetic acid in separate test tubes. Test each solution with litmus paper.
(ii) Add the acetic acid to the ethanol. Add 10 drops of concentrated sulfuric acid. Carefully note the odor by wafting the odor toward your nose. Place the test tube in hot water bath. Heat until the mixture begins to boil. Remove the test tube. Note the odor again. What does this odor smell like?
b. Methyl salicylate

Place about 0.1 g of salicylic acid in a dry test tub. Add 1 ml of methanol. Shake until the acid dissolves. Cautiously add dropwise 1 ml (about 20 drops) of concentrated sulfuric acid. Warm gently in a boiling water bath. Carefully note the odor by wafting the odor toward your nose.

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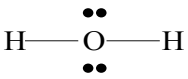
Table 1. Structure Information of Various Substances

Substance	Total # of valence electrons	Electron-dot structure/ functional group(s)	Bond type/ bond polarity	Geometry	Polarity
H_2O water	8		2 O-H bonds polar	bent	polar
CH_4					
C_2H_6					
C_3H_8					
C_2H_4					
$C_2H_2Cl_2$					
CH_3OH					

CH_3CH_2OH					
CH_3CHO					
CH_3COOH					
C_2H_2					
C_6H_6					
$C_6H_4(COOH)(OH)$					

1. Describe the results of the litmus tests on ethanol and acetic acid. Which substance is an acid? Base? Neutral? Based on these results, predict whether salicylic acid and methanol is an acid, base, or neutral.

2. a. Write a chemical equation that describes the reaction between ethanol and acetic acid to make ethyl acetate. Draw the electron-dot structure of ethyl acetate. Circle the ester functional group in your structure. Identify the bond that breaks in ethanol and in acetic acid and the bond that forms in ethyl acetate.

b. Write a chemical equation that describes the reaction between methanol and salicylic acid to make methyl salicylate. Draw the electron-dot structure of methyl salicylate. Circle the ester functional group in your structure. Identify the bond that breaks in methanol and in salicylic acid and the bond that forms in methyl salicylate.

Questions

1. a. Which molecules in Table 1 are unsaturated?
b. CH_4 looks like a square molecule on paper. What is the shape of this molecule based on the model that you made?
c. C_3H_8 is a straight molecule on paper (look at the carbon chain). But this molecule really is not straight. Which atom(s) give this molecule its kinks? Give reasons based on your answer in part b.

2. An object that is polar has a north and south pole, like the earth, or a positively charged end and a negatively charged end, like a magnet. Consider the polarity of a bond in a molecule and the polarity of a molecule as a whole.
a. What does bond polarity mean?
b. How does electronegativity determine the type of bond, polar or non-polar, in a molecule?
c. H_2O and CO_2 are similar in that each molecule has one central atom surrounded by two atoms. However, H_2O is polar whereas CO_2 is non-polar. Describe how the shape of a molecule affects polarity.
d. Discuss one trend in shape and polarity that you have observed.
e. Describe how you can apply your knowledge of polarity of a molecule to dissolve one substance in another.

3. "Like dissolves like" means that a polar substance dissolves in a polar substance or a non-polar substance dissolves in a non-polar substance. Name two molecules on the above list that dissolves in water.

Lab 11. Aspirin Synthesis

Prelab Questions: 1. Annual consumption of aspirin in the U.S. is 700 tablets per person. Calculate the mass of aspirin consumed each year in this country in kg/yr. One aspirin tablet has a mass of 325 mg and the population of this country is 260 million people.
2. You want to make aspirin. Solid salicylic acid reacts with liquid acetic anhydride in the presence of phosphoric acid to produce aspirin. You use 4.57 g of salicylic acid and isolate 3.83 g of aspirin.
a. Calculate the theoretical yield and the percentage yield of aspirin.
b. What mass of acetic anhydride reacts with 4.57 g of salicylic acid? Show your calculations.
c. Acetic anhydride is a liquid. Why is it important to know the density of acetic anhydride?
d. What is the function of the phosphoric acid in this experiment?

KEY POINTS:

1. Aspirin is synthesized by reacting an acid with an alcohol to produce an ester.
2. Organic synthesis involves synthesis of a crude product, purification of the crude product, and characterization of the pure product to determine its identity.

Introduction

You've had a long and fruitful day but all that thinking has given you a headache and all that exercise during your afterwork workout has left your muscles sore. Just like long and fruitful go together, pain and inflammation seem to go together as well. You need pain relief. You need an aspirin.

Aspirin is a common, over-the-counter non-steroidal anti-inflammatory drug (NSAID). Its medicinal properties include its effects as an analgesic (relieves pain), an anti-inflammatory agent (reduces inflammation), and an antipyretic (reduces fever). Although the use of aspirin (chemical name: acetylsalicylic acid, $C_9H_8O_4$) goes back to the late 1800's, it was known from before the time of Hippocrates in 400 B.C. that fevers could be lowered by chewing the bark of willow trees. In 1827, the active ingredient in willow bark was found to be an aromatic compound (a compound that contains benzene) called salicin. This compound, when reacted with water, produces salicyl alcohol; oxidation of the salicyl alcohol produces salicylic acid. Salicylic acid was found to be more effective than salicin for reducing fevers and also had analgesic and anti-inflammatory properties. Unfortunately, a side effect of salicylic acid, being an acid, is that it is too corrosive to the walls of the stomach for daily use. Acetylsalicylic acid is used now because it is less corrosive to the stomach.

In spite of its benefits, aspirin is more dangerous than commonly believed. It can cause stomach bleeding and allergic reactions in long-term users. Aspirin given to children recovering from the flu may cause a potentially fatal condition called Reye's syndrome. As a result of these problems, numerous other NSAIDs have been developed in the last two decades, such as Ibuprofen and Naproxen. Ibuprofen (Advil, Motrin, Nuprin) has about the same potency as aspirin but is much less prone to cause an upset stomach. Naproxen (Alleve) also has about the same potency as aspirin but remains active in the body six times longer than aspirin. (See J. McMurry, "Organic Chemistry", 4th ed., Brooks-Cole, 1996, p. 611-612.)

In this lab activity, you will react salicylic acid, $C_7H_6O_3$, with acetic anhydride, $C_4H_6O_3$, to make acetylsalicylic acid or aspirin. Phosphoric acid, H_3PO_4, will be used as a catalyst to speed up the reaction:

$C_7H_6O_3$	+	$C_4H_6O_3$	--- H_3PO_4--->	$C_9H_8O_4$	+	$HC_2H_3O_2$	(1).
salicylic acid		acetic anhydride		acetylsalicylic acid aspirin		acetic acid	

When you do this reaction, you will end up with a liquid solution that contains aspirin. The aspirin is separated from the reaction mixture by adding water, which precipitates the aspirin. After filtering, you'll have a sample of crude, impure aspirin. To purify the crude aspirin, you will dissolve it in a heated mixture of ethyl alcohol and water. When the hot solution is cooled, the purified aspirin will precipitate out of solution as white crystals. This purification method is called recrystallization and is a common method to purify solids. Then, you'll collect and dry the pure aspirin and weigh it to determine the percentage yield in your experiment.

If your have a headache after all the thinking you've done in this lab activity, please don't take your aspirin sample. You'll be testing the purity of your aspirin sample in the next lab activity. Console yourself by realizing that chemistry is truly all around us every day.

Materials

salicylic acid, acetic anhydride, ethyl alcohol (95%), phosphoric acid
125 ml Erlenmeyer flask, 600 ml beaker, Buchner funnel, filter flask, and vacuum hose
ice

Procedure

Preparation and purification of aspirin.

1. Preparation of crude aspirin.

a. Measure about 3 g of salicylic acid in dry 125 ml Erlenmeyer flask to the nearest 0.01 g. Record the mass of the salicylic acid in Table 1 in the Results and Calculation section.

b. Add 6 ml of acetic anhydride, followed by about 5-8 drops of phosphoric acid.

c. Swirl the flask to mix the reagents, and place the flask in a beaker of warm water (about 75^{o}C) for 15 - 20 minutes. At the end of this period, the reaction will be complete. (What observation indicates that the reaction is complete? Note any changes in appearance of the reagents during the reaction.)

d. Slowly add about 1 ml of water to the reaction mixture to destroy any excess acetic anhydride. Then add 15 ml of water to the flask and cool the mixture in an ice bath. (What is the reason for adding the additional 15 ml of water? Think about the solubility of aspirin in water.)

e. When no further crystal formation is evident, filter the crystals using vacuum filtration. (Your instructor will demonstrate the use of the Buchner funnel and filter flask.) Wash the crystals with ice-cold water. Note how the crude aspirin looks.

2. Purification of the crude aspirin.

a. Dissolve your crude aspirin in 10 ml of ethyl alcohol. (Is aspirin soluble in ethyl alcohol?) Then, add 25 ml of warm distilled water (about 70°C) to the aspirin-alcohol mixture.
b. Allow the solution to cool to room temperature. If time permits, slow cooling will enhance the purity of the crystals formed. However, fast cooling will provide a reasonably pure product. To achieve a maximum recovery of product, chill the solution below room temperature using an ice bath.
c. Assemble a Buchner funnel and filter flask. Weigh a piece of filter paper and insert it into the Buchner funnel. Filter the crystals by vacuum filtration, and allow them to air dry on the filter paper overnight. Note how the pure aspirin looks.

Mass of filter paper _____ g

d. Weigh a dry beaker. Transfer the dried crystals to the beaker and reweigh. Determine the mass of aspirin isolated and calculate the percentage yield.

Mass of filter paper and crystals _____ g

Usually, you would characterize your pure product next, i.e., run a test, such as a melting point measurement, to identify whether you made the compound you wanted. We will skip the characterization step in this lab.

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1. a. Write a chemical equation that represents the reaction of salicylic acid to aspirin.
b. Draw the electron-dot structure of each reactant and product. For each compound, circle each functional group and write the name of the functional group next to your circle.
c. Calculate the molar mass of each substance and record the molar mass in Table 1. Look up the density of acetic anhydride and record it below. Cite the reference where you found this information.

2. a. Record the following information in Table 1. Show your calculations.

Table 1. Aspirin Synthesis Data

Substance	salicylic acid	acetic anhydride	aspirin (theoretical)	aspirin (actual)	acetic acid (theoretical)
Chemical Formula					
Molar Mass, g/mole					
Mass, g					
Moles					

b. Calculate the percent yield of aspirin. Show your calculations.

c. From the theoretical yields of products, do a calculation to confirm that the conservation of mass law is obeyed.

Questions

1. Refer to Step 2 of the Procedure. See the electron-dot structures of aspirin, water, and ethanol.
a. Is aspirin polar or non-polar? Is water polar or non-polar? Is ethanol polar or non-polar? What is the solubility of aspirin in water? What is the solubility of aspirin in ethanol? Give reasons.
b. What is the purpose of the crystallization of aspirin from ethyl alcohol/water mixture?

2. If your percent yield of aspirin is less than 100% (and it probably will be), suggest what you could have done differently in the procedure to improve the yield.

3. One of your fellow students claims that using 10 g of salicylic acid instead of 5 g and 10 ml of acetic anhydride should double the theoretical yield of aspirin. Calculate the theoretical yield of aspirin using 10 g of salicylic acid and 10 ml of acetic anhydride. Is this student correct in saying that the theoretical yield will double?

Lab 12. Properties and Reactions of Carbohydrates

Prelab Questions: 1. Glucose, fructose, sucrose, and lactose are common carbohydrates.
a. Draw the electron-dot structure of each carbohydrate. Circle the functional groups in each structure and write the name of each functional group next to the functional group.
b. Name one food or plant source of each carbohydrate.
c. Identify each carbohydrate as a monosaccharide, disaccharide, or polysaccharide.
d. Name the bond that joins monosaccharides together or breaks polysaccharides into smaller saccharides. Name two substances that are used to break this bond.
e. Which carbohydrate(s) is/are reducing sugars?
2. What is the difference between glucose and starch?

KEY POINTS:

1. Carbohydrates contain alcohol, ether, aldehyde, and ketone functional groups.
2. Carbohydrates are classified as simple carbohydrates or complex carbohydrates.
3. Complex carbohydrates are chains of monosaccharides connected together by a glycosidic bond.
4. Complex carbohydrates are broken down to simple carbohydrates by an enzyme that breaks the glycosidic bond.
5. A reducing sugar contains an aldehyde group (for a sugar chain) or hemiacetal group (for a sugar ring).

Introduction

Carbohydrates, such as sugar and starch, are found in many foods. As you know from looking at food labels earlier this semester, one gram of carbohydrate stores 4 Cal of energy. Carbohydrates are a fast source of energy; it takes the body about 0.5 to 2 hours to metabolize carbohydrates to use as fuel.

Carbohydrates are biomolecules based on organic chemistry. The alcohol, ether, aldehyde, and ketone functional groups are found in carbohydrates. Carbohydrates are classified as either simple carbohydrates or complex carbohydrates. Simple carbohydrates are monosaccharides and disaccharides; complex carbohydrates are polysaccharides. Monosaccharides exist as either chain or cyclic (ring) structures; disaccharides and polysaccharides are cyclic structures. A glycosidic bond joins monosaccharide rings together to form a disaccharide or longer chained polysaccharide.

In this lab, you will look at the reducing and nonreducing properties of sugars. Some carbohydrates are called reducing sugars because they reduce an oxidizing agent , such as Benedict's reagent or Fehling's reagent. Reducing sugars contain an aldehyde group in a monosaccharide chain or a hemiacetal in a cyclic sugar as shown in Fig. 1.

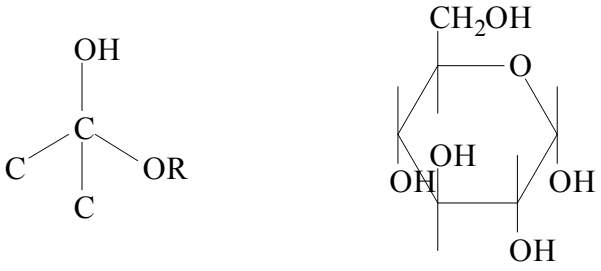

Fig. 1. A hemiacetal has –OH and –OR groups bonded to the same carbon as shown on the left. Glucose, shown on the right, is a hemiacetal .

Many polysaccharides have a hemiacetal group but do not give a positive test with Benedict's or Fehling's reagent . There is only one hemiacetal group for every few thousand monosaccharide units; this low concentration of reducing groups does not give a positive test.

Next, you will hydrolyze disaccharides and polysaccharides to monosaccharides. Disaccharides and polysaccharides are hydrolyzed into monosaccharides by using an acid catalyst or enzyme. Iodine, which turns blue-purple in the presence of starch, will be used as an indicator to indicate the hydrolysis of a polysaccharide. These chemical properties of carbohydrates are important tests to identify the functional groups contained in these compounds.

Reference

1. F.A. Bettelheim and J.M. Landesberg, "Laboratory Experiments for General, Organic, and Biochemistry", 4th ed., Harcourt, 2001, p. 395-404.

Materials

2% solutions of: glucose, sucrose, lactose, fructose, starch
milk, grape juice, NutraSweet (Sweet 'N Low), flour, unsalted crackers
Fehling's reagent (Solution A: CuSO4 (aq); Solution B: sodium potassium tartrate + NaOH)
3 M H_2SO_4, 3 M NaOH, 0.01 M iodine in KI
Amylase or saliva
Spot plates

Procedure

1. Identifying reducing and nonreducing carbohydrates using Fehling's solution.
a. Place about 2 ml of Fehling's solution into five test tubes. Label the test tubes 1-5.
b. Add 10 drops of glucose, sucrose, lactose, fructose, starch solutions to Test Tubes 1, 2, 3, 4, and 5, respectively.
c. Place the test tubes in a boiling water bath for 5 min.
d. A red precipitate of Cu_2O is a positive test for a reducing sugar. Identify the carbohydrates that are reducing sugars.
e. Use Fehling's solution to test the reducing properties of milk, grape juice, NutraSweet, and flour.

f. Record your observations in Table 1.

2. Hydrolysis of sucrose.
a. Place 3 ml of 2% sucrose solution in two test tubes.
b. To the 1st test tube, add 3 ml of water and 3 drops of 3 M H_2SO_4. To the 2nd test tube, add 3 ml of water and 3 drops of 3 M NaOH.
c. Heat both test tubes in a boiling water bath for 5 min. Then, cool each solution to room temperature.
d. To the 1st test tube, add 3 M NaOH until the solution is slightly basic. Test the solution with litmus paper.
e. Test a few drops of each solution with Fehling's solution like you did in Step 1. Record your observations in Table 2.
f. Repeat Steps 2a through 2e with lactose.
g. Repeat Steps 2a through 2e with milk.
h. Repeat Steps 2a through 2e with NutraSweet.

3. Hydrolysis of starch using an enzyme and acid catalyst.
a. Place 5 ml of 2% starch solution in two test tubes.
b. To the 1st test tube, add 1 ml of your own saliva or amylase. Use a 10 ml graduated cylinder to collect and measure your saliva. To the 2nd test tube, add 1 ml of 3 M H_2SO_4.
c. Heat the test tube that contains H_2SO_4 in a boiling water bath. Keep the other test tube that contains saliva or amylase at room temperature. (Why?)
d. After 5 min, transfer 3 drops of each solution to separate wells of a spot plate. Use a clean medicine dropper for each solution. Add 2 drops of the iodine solution to each sample. If the solution turns blue or purple, starch is present and the hydrolysis reaction is not complete. Continue heating the solutions in the test tubes. In 5 min intervals, transfer a few drops of each solution to the spot plate and add 2 drops of iodine. Continue taking samples until each hydrolysis reaction is complete. Determine the time needed for the hydrolysis reaction to be complete. Record your observations in Table 3.
e. Once the hydrolysis reaction for each test tube is complete, test a few drops of each solution with Fehling's solution like you did in Step 1.

4. Hydrolysis of flour using an enzyme and acid catalyst.
a. Prepare a 2% flour solution. Take a test tube and pour 10 ml of water into the test tube. Use a marking pen, e.g., a Sharpie, to mark the 10 ml level on the test tube. Pour out the water. Measure 0.2 g of flour into this test tube. Stopper and shake this mixture. Add sufficient water to make 10 ml of solution. Divide this solution into two test tubes.
b. Repeat Steps 3b through 3e.

5. Hydrolysis of an unsalted cracker using an enzyme and acid catalyst.
a. Prepare a 2% cracker solution. Take a test tube and pour 10 ml of water into the test tube. Use a marking pen, e.g., a Sharpie, to mark the 10 ml level on the test tube. Pour out the water. Crush or grind up a cracker. Measure 0.2 g of this cracker into this test tube. Add sufficient

water to make 10 ml of solution. Stopper and shake this mixture. Divide this solution into two test tubes.
b. Repeat Steps 3b through 3e.

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Table 1. Reducing and Nonreducing Properties of Sugars

Test Tube	Solution	Observations/Fehling's Test (positive or negative?)
1	2% glucose	
2	2% sucrose	
3	2% lactose	
4	2% fructose	
5	2% starch	
6	Milk	
7	Grape juice	
8	NutraSweet	
9	Flour	
10	cracker	

List the reducing sugars. Draw the electron-dot structure of each reducing sugar. Circle the aldehyde or hemiacetal functional group that undergoes oxidation with Fehling's reagent.

Table 2. Hydrolysis of Sugars Under Acidic and Basic Conditions

Sample	Conditions	Fehling's Test (positive or negative?)
Sucrose test tube 1	H_2SO_4	
Sucrose test tube 2	NaOH	
Lactose test tube 1	H_2SO_4	
Lactose test tube 2	NaOH	
Milk test tube 1	H_2SO_4	
Milk test tube 2	NaOH	
NutraSweet test tube 1	H_2SO_4	
NutraSweet test tube 2	NaOH	

Under what conditions is each sugar hydrolyzed?

For sucrose and lactose, draw the electron-dot structures of the products of the hydrolysis reaction.

Is milk hydrolyzed? What observation tells you that milk is hydrolyzed or not hydrolyzed?

Is NutraSweet hydrolyzed? What observation tells you that NutraSweet is hydrolyzed or not hydrolyzed?

Table 3. Hydrolysis of Carbohydrates Using an Enzyme and Acid

	Starch		Flour		Cracker	
Heating Time, min	Saliva enzyme Iodine test	Acid Iodine test	Saliva enzyme Iodine test	Acid Iodine test	Saliva enzyme Iodine test	Acid Iodine test
5						
10						
15						
20						
25						
Fehling's test results						
Time required to complete reaction						

Under what reaction conditions did each hydrolysis reaction go faster?

Question

1. You chew on an unsalted cracker. It does not taste sweet. Let's say you chewed on the cracker for five minutes. What do you think you would observe based on your data and observations in Step 5 of this lab?

Lab 13. Extracting Fat From Food and Making Cheese

Prelab Questions: 1. a. What is the solubility of a fat in water?
b. What is the difference between an oil and fat?
2. Give one example of how to denature a protein. Describe what happens to the protein when it is denatured.

KEY POINTS:

1. Fats are not soluble (immiscible) in water.
2. Fats are classified as saturated or unsaturated.
3. Proteins are denatured by changing pH, heat, or organic solvents.

Introduction

Potato chips washed down with milk. Except for all the fat in potato chips, this would be a delicious and healthy snack. In this lab, you will test the solubility of fats in various solvents and use the results of your solubility tests to extract fat from potato chips and milk. Once you extract fat, you will test for unsaturated fat by adding bromine to the fat. A C=C double bond an unsaturated fat reacts rapidly with a red bromine solution in cyclohexane to produce a colorless solution. In the presence of saturated hydrocarbons, which have C-C single bonds, the red bromine color does not change.

The protein in milk is casein. Casein is soluble in milk at pH 6.6 but precipitates out of milk as the pH is lowered to about pH 4 by adding acid. This milk solid (curds) is cheese. In this process, the milk protein is denatured. You will look at several ways to denature milk and then try to "re-nature" milk back to its original form by doing the opposite of what you did to denature the milk. You may never look at potato chips and milk the same way again!

Materials

Vegetable (soybean) oil, corn oil, lard or margarine or butter
Water, ethanol, hexane, vinegar, 1 M NaOH, liquid soap, 1% Br_2 in cyclohexane, Biuret reagent
Acetate sheets or spot plates, Buchner funnel and flask
Potato chips, whole or 2% milk

Procedure

1. Test the solubility or miscibility of various oils and fats in several solvents.
a. Place three drops of each oil or a tip of a spatula of fat into four separate wells in a spot plate or four separate squares in an acetate sheet.
b. Add three drops of water to each oil or fat. Stir. Observe and record whether the oil dissolves or the fat is miscible in water.
c. Add three drops of soap to the mixture. Observe and record whether the oil dissolves or the fat is miscible in soap.
d. Test the solubility or miscibility of each oil or fat in ethanol, hexane, vinegar, and 1 M NaOH by repeating Steps 1a, 1b, and 1c except replacing water by the appropriate solvent.

e. Test the solubility of the ethanol and hexane in water. Place three drops of water in separate wells in a spot plate or separate squares in an acetate sheet. Add three drops of ethanol to the first well and hexane to the second well. Observe and record whether the solvent is miscible in water.

2. Extract fat from a potato chip. Based on your results from Step 1, what solvent would you use to extract the fat from potato chips? Try out your chosen solvent:
a. Break a potato chip into a size and shape that fits into a test tube. Add 2 ml (40 drops) of your chosen solvent. Stopper the test tube and shake for one minute. Unstopper the test tube and place the test tube in a warm water bath for 15 minutes. Add more solvent as needed to keep the total volume constant.
b. Repeat Step 2a except use water instead of your chosen solvent.
c. Let each test tube mixture cool. Decant off the liquid from the potato chip. Test the liquid for the presence of unsaturated fat by carefully adding dropwise 1% Br_2 in cyclohexane. Count the number of drops needed for the solution to turn colorless. Do not add more than 10 drops of Br_2 solution.

3. Extract fat from milk. Based on your results from Step 1, what solvent would you use to extract the fat from milk? You want to use a solvent in which oil or fat is soluble or miscible but is immiscible in water. Try out your chosen solvent:
a. Place 2 ml of milk in a large test tube. Add 2 ml of your solvent of choice to the milk. Which substance is the top layer? Why? Stopper the test tube and invert gently several times. Do <u>not</u> shake since this will cause milk bubbles to form.
b. Carefully decant off the top layer. Save the layer that contains the extracted fat.
c. To the solvent/fat layer, test for the presence of unsaturated fat by adding 1% Br_2 in cyclohexane like you did in Step 2c. Record your observations.

4. Denaturation of milk. Making cheese.
a. Add 5 ml of milk to three separate test tubes.
b. To the first test tube, add vinegar dropwise to the milk. Count the number of drops of vinegar that you add until you see curds (solid) start to form. Once you start to see curds, add 10 more drops of vinegar. Record the total number of drops of vinegar added.
c. Collect the milk solid by vacuum filtration using a Buchner funnel. Test for protein by adding a small amount of the milk solid (<u>not</u> all of it) you collected into a clean test tube. Add 2 ml of Biuret reagent. If the solution turns purple or purple-blue, protein is present. If no color develops, place the test tube in a hot water bath for 2 minutes. Record your observations.
d. Add another small amount of milk solid to another clean test tube. Add the same number of drops of 1 M NaOH as the number of drops of vinegar you added in Step 4b. Does the milk solid dissolve? Do you think the milk is “re-natured”?
e. Put the second test tube of milk in a hot water bath. Heat the milk until you see a change in the appearance of milk. Continue to heat the milk for an additional two minutes. Then, let the milk cool back to room temperature. Does the milk return back to its original state when it cools?

f. To the third test tube of milk, add ethanol dropwise. Add the same number of drops of ethanol as the drops of vinegar you used in Step 4b. If curds form, collect the solid and test the solid for protein with Biuret reagent like you did in Step 4b.

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Table 1. Solubility or Miscibility of Oil/Fat in Various Solvents

Fat or Oil	Soluble in Water?	Soluble in Ethanol?	Soluble in Hexane?	Soluble in Vinegar?	Soluble in 1 M NaOH?
Vegetable oil					
Corn oil					
Lard					
Margarine or butter					
Ethanol		XX		XX	XX
Hexane			XX	XX	XX

1. Based on your solubility tests, determine the polarity of each type of oil or fat.

2. What solvent did you use to extract fat from a potato chip? Give reasons for your choice. Were you able to extract fat from the potato chip? If so, what type of fat, saturated or unsaturated fat, did you find in potato chips?

3. Explain why you chose your solvent to extract fat from milk. Were you able to extract fat from the milk? If so, what type of fat, saturated or unsaturated fat, did you find in milk?

Table 2. Denaturation of Milk by Acid, Heat, and Organic Solvent

	Number of Drops Added	Curds Form?	Protein Test (positive or negative)
Vinegar			
Heat	XXX		
Ethanol			

4. a. Which method worked best to denature milk?

b. Speculate on the specific chemical forces in milk protein (casein) that were broken during denaturation.

c. Did adding 1 M NaOH to the milk solid from Step 4c "re-nature" the milk? What did you observe to help answer this question?

d. Did cooling the milk solid from Step 4c "re-nature" the milk? What did you observe to help answer this question?

Questions

1. Is denaturing proteins a reversible process? Give reasons.
2. a. What happens to the milk protein once it gets in your stomach? In other words, what does stomach acid do to milk protein?
b. What does stomach acid do to fat?
3. Look up how cheese is made. Cite the reference where you found this information. How is Step 4b in the Procedure similar to the real cheese making process?

Lab 14. Peanut Brittle: What's in the Food We Eat?

Prelab Questions: 1. The ingredients to make peanut brittle are peanuts, sugar, salt, baking soda, vanilla, corn syrup, and margarine. Match each ingredient to the chemical name listed in the Materials section.
2. For as many of the ingredients as you can, identify the substance as either organic or inorganic. For the organic compounds, name the organic functional group(s) in that substance.

KEY POINTS:

1. Chemical reactions occur when food is cooked.
2. A cooking recipe is a stoichiometry problem.
3. Peanut brittle tastes good.

Introduction

Cooking involves chemical reactions. During the cooking process, the structure and composition of foods change (see The Science of Cooking web site: http://www.exploratorium.org/cooking). Hopefully, the food will taste and smell good (a lot of chemistry is involved here) and be good to eat (taste, smell, digestion and metabolism). A cooking recipe lists the amounts of each ingredient, the order in which the ingredients are mixed, the heating or cooling temperature, and the cooking time -- just another chemical reaction and stoichiometry problem!

Materials

600 ml beaker, porcelain spatula, 250 ml beaker, watch glass, thermometer, graduated cylinder sucrose, 3 M glucose solution, solidified mixed esters, sodium chloride, sodium bicarbonate, protein pellets, 4-hydroxy-3-methoxy-benzaldehyde

Procedure

Do **NOT** use any of the equipment in your locker. Clean pots and beakers for food will be supplied. Cleanliness is **essential** in this experiment. So begin by **scrubbing** (!) all of the equipment you plan to use as well as cleaning your lab bench. (Remember: cleanliness is next to godliness.)

1. Weigh out 113 g of sucrose and 93 g of 3 M glucose solution into the 600 ml beaker.

2. To this mixture, add 29 ml of water.

3. Heat the mixture carefully with a Bunsen burner. Stir constantly with the porcelain spatula until the mixture begins to boil. While this step is going on, one lab partner should be weighing out the rest of the components (see Step 4).

4. Weigh out:
a. 14 g of "solidified mixed esters" onto a watch glass.
b. 0.5 g of sodium chloride onto a paper towel.

c. 5.5 g of sodium bicarbonate onto a paper towel.
d. 83 g of "protein pellets" into a 250 ml beaker.
e. 2 ml of 4-hydroxy-3-methoxy-benzaldehyde.

5. When the mixture (in Step 3) has reached the boiling point, add the "solidified mixed esters". Continue heating until it reaches 110^{o}C, and stir frequently after that. The mixture will probably start to foam at this point. Either (I) letting the mixture alone or (II) stirring the top layers quickly will usually solve the problem, but the same thing will not work every time.

6. The mixture will begin to thicken at 125^{o}C and turn brown at 120^{o}C. At this point, stirring will become difficult. Paper towels can be used to hold the beaker while stirring.

7. When the mixture reaches 138^{o}C, add the sodium chloride and the "protein pellets".

8. When the mixture reaches 150^{o}C, turn off the burner and set the beaker down. While the mixture is being heated, you should prepare a one foot square of aluminum foil by coating lightly with some of the "solidified mixed esters."

9. a. This step requires coordination of two people. One person should hold the beaker ready to pour while the second person adds the 2 ml of 4-hydroxy-3-methoxy-benzaldehyde and the 5.5 g of sodium bicarbonate. The second person has to stir very fast until the foaming mixture starts to flow over the top of the beaker. At this point, the mixture is poured out of the beaker, scraping with the spatula and spread very gently over the aluminum foil.
b. Clean out the beakers and utensils with plenty of hot water and soap.

10. When cool, the mixture will harden and can be broken into pieces. Savor its effect upon the palate. "Buen apetito!"

Questions

1. Classify each ingredient as a carbohydrate, protein, or fat. Draw the electron-dot structure of each substance. Identify the functional group(s) in each compound.

Lab 15. DNA Extraction from Wheat Germ

Prelab Questions: 1. Name three reasons why DNA is so important to life.
2. a. DNA is a large biomolecule and is a polymer. What is the monomeric unit of DNA?
b. Each DNA monomer consists of three parts. What is the name of each part?
c. Why is base pairing important in DNA?

KEY POINTS:

1. DNA is a polymer with a nucleotide as its monomeric unit.
2. DNA is found in cells of living organisms.
3. The cell membrane or cell wall must be broken to release DNA from it.
4. DNA can be denatured by heat or organic solvents.

Introduction

From DNA fingerprinting to the cloning of sheep to DNA sequencing protein synthesis, DNA or deoxyribonucleic acid is a vital substance to all of life. DNA consists of a chain of nucleotides. Nucleotides consist of a phosphate, a sugar, and an amine base and are joined together by phosphate-ester bonds as shown Fig. 1.

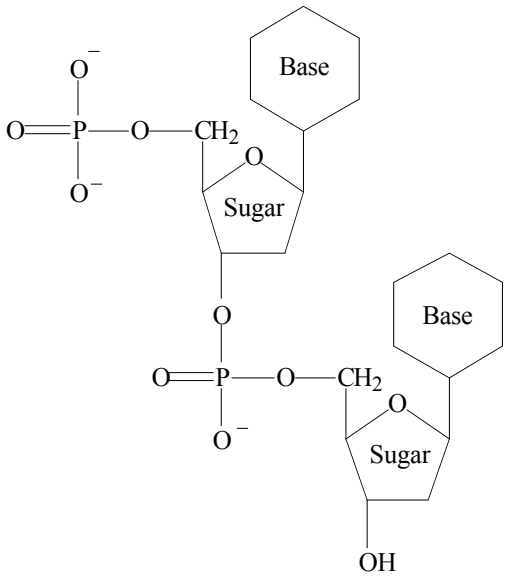

Fig. 1. A two nucleotide chain. Nucleotides consist of a base, sugar, and phosphate group.

In other words, DNA is a polymer with a nucleotide being its monomer unit. In1953, Watson and Crick proposed that DNA consists of two polymer strands coiled around each other in a double helix (and later shared the Nobel Prize for their work). The polymer strands run in opposite directions and are held together by hydrogen bonds between specific pairs of bases.

Base pairing in DNA is very important. DNA from different tissues of the same species have the same proportions of amine bases, but samples from different species can have greatly differing proportions of bases. For example, human DNA contains about 30% each of adenine

and thymine and about 20% each of guanine and cytosine. (Adenine, thymine, guanine, and cytosine are the amine bases in DNA.) However, the bacterium *Clostridium perfringens* contains about 37% each of adenine and thymine but only 13% each of guanine and cytosine.

In this lab, you will extract DNA from wheat germ. You will determine the % DNA in wheat germ and then take a sample of your DNA and denature it. Remember that this DNA comes from wheat germ and not from a pound of flesh or an ounce of blood.

Materials

0.5 g or 7 g wheat germ (DNA source)
1 ml or 5 ml Woolite or Joy, Sunlight, or Dawn detergent
3 g of Adoph's Meat Tenderizer (unseasoned)
0.5 g baking soda
14 ml (ice cold)or 70 ml (room temperature) 95% ethanol
tap water
large test tube, 100 ml graduated cylinder, 250 ml beaker, 10 ml graduated cylinder or pipet
thermometer
glass stirring rod
glass rod with hooked tip

Procedure

There are two methods to extract DNA from wheat germ. See the Genetic Science Learning Center web site (http://learn.genetics.utah.edu/archive/wheatgerm/background.html) for a description of the science behind this experiment.

Method 1.

1. Add 20 ml of water to a large test tube. Place the test tube with water in a water bath and heat the water to 45°C. It is critical to stay within this temperature range at all times. DNA denatures at higher temperatures.

2. Record your observations in each step. To the warm water,
a. Add 0.5 g of wheat germ.

mass of wheat germ = _____

Stir thoroughly for 3 minutes, keeping the temperature at 45°C.
b. Add 1 ml of Sunlight or Dawn detergent. This detergent contains sodium lauryl sulfate (SLS). SLS is the detergent in dish soap, shampoo, etc. You have seen SLS in Lab 4A, Set 2 of the compounds that you looked at in your solubility observations. The SLS will dissociate the lipids of the membranes. Stir for 5 minutes while maintaining the temperature at 45°C. Try not to create foam.

3. a. Separate the liquid from the solid wheat germ. How will you do this?
b. Tip the test tube at an angle. SLOWLY pour 14 ml of ice cold ethanol down the inside of the beaker so it forms a layer on top of the water/DNA/detergent solution. Do not mix the 2 layers.

4. DNA will begin to "appear" at the interface of the 2 layers. DNA is soluble in the aqueous solution but is insoluble in ethanol.

5. Use a glass rod to lift the lower layer up to the upper layer without mixing them together. "Stir" with a smooth constant motion for 3 minutes.

6. a. Collect the DNA precipitate from the beaker with the stirring rod.
b. Weigh the DNA that you collected.
c. Cut off a small portion of your DNA sample. Heat it up until it denatures. What did you observe that tells you that DNA is being denatured?
d. Figure out another way to denature DNA. What did you observe that tells you the DNA is being denatured?

7. If you want to keep the DNA, store it in 50% ethanol in a sealed tube or air dry it and observe the strands.

Method 2.
1. In a 250 ml beaker, heat 100 ml of tap water to 50-60^{o}C. It is critical to stay within this temperature range at all times. DNA denatures at higher temperatures.

2. Record your observations in each step. To the warm water,
a. Add 7 g of wheat germ.

mass of wheat germ = _____

Stir thoroughly for 3 minutes, keeping the temperature between 50-60^{o}C.
b. Add 5 ml Woolite. Woolite contains sodium lauryl sulfate (SLS). SLS is the detergent in dish soap, shampoo, etc. You have seen SLS in Lab 4A, Set 2 of the compounds that you looked at in your solubility observations. The SLS will dissociate the lipids of the membranes. Stir for 5 minutes while maintaining the temperature between 50-60^{o}C.
c. Dissolve 3 g unseasoned Adolph's Meat Tenderizer. The tenderizer contains papain, an enzyme from papaya fruit. The papain will digest proteins clinging to the DNA. Maintain the temperature between 50-60^{o}C for an additional 15-20 minutes. Only stir the mixture occasionally. Continuous stirring will break the DNA strands.
d. Mix 0.5 g of baking soda.

3. Tip the beaker at an angle. SLOWLY pour 70 ml of ethanol down the inside of the beaker so it forms a layer on top of the water/DNA/Woolite solution. Do not cause the 2 layers to mix.

4. DNA will begin to "appear" at the interface of the 2 layers. DNA is soluble in the aqueous solution but is insoluble in ethanol.

5. Use a glass rod to lift the lower layer up to the upper layer without mixing them together. “Stir” with a smooth constant motion for 3 minutes.

6. Collect the DNA precipitate from the beaker with the stirring rod and store it if you wish as described in Method 1, Steps 6 and 7.

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1. Describe what you observed in each step of the Procedure.

2. Calculate the % DNA in wheat germ.

3. a. What does DNA look like?

b. What does denatured DNA look like?

Questions

1. Look up the structure of sodium lauryl sulfate (SLS).

a. Draw the electron dot structure of SLS.

b. From the structure, explain why SLS dissociates lipids in membranes.

2. Adolph's Meat Tenderizer contains papain, which digests proteins clinging to the DNA. What does digestion of proteins mean?

3. Speculate on why baking soda was used in Step 2d.

4. You know that DNA is contained in cells. Speculate on how DNA was extracted from the cells of wheat germ using this procedure.

5. DNA denatures at high temperatures. When DNA denatures, what happens to the DNA?