

CHEMISTRY 11
Manual of Lab Activities
Spring 2010

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CHEMISTRY LABORATORY SAFETY RULES

1. Perform no unauthorized experiments.
2. No lab experiments should be performed without your instructor present.
3. All students are required to have a pair of ANSI-approved safety goggles, which must be worn in the lab when anyone is working with chemicals and/or heat.
4. Contact lens wearers: It is recommended that you avoid wearing contacts, if possible, when working with chemicals.
5. Wear appropriate clothing, avoiding bulky or loose items and open-toed shoes. Long hair should be restrained.
6. MSDS (materials safety data sheets) on all chemicals in stock are available; you have the right to request to see them. Your instructor will show how to use them.
7. Please do not eat or drink in the lab. Wash your hands frequently during lab. Also, be sure to wash hands after lab periods before touching any food (or your eyes).
8. Notify your instructor if you are pregnant; you may be asked to take additional precautions.
9. Laboratory desks and hoods should always be cleaned after use; also, be sure to wipe up spills and remove clutter as you work.
10. Wash your hands regularly during lab and when you leave lab.
11. Dispose of chemical wastes as directed by your instructor each lab period. Many chemicals may not be flushed down the sink or put in ordinary trash containers.
12. Do not taste chemicals. If checking chemicals for odors, use the wafting technique. (See instructor for method.)
13. Never heat or cool a closed system (one that is not vented to the surrounding atmosphere).
14. If caustic chemicals (e.g., strong acid and bases) are spilled on the body, wash liberally with water. If spill is over a large portion of the body, use safety shower and remove clothing quickly. For chemicals in eyes, use the eyewash and rinse continually for 15 minutes.
15. In case of any accident in the lab (spills, fires, explosions, etc.), yell "help" first to get assistance, then be sure to call the instructor. Notify your instructor no matter how small the accident seems.
16. When heating a test tube, be sure the open end is not aimed in anyone's direction.
17. Never pour water into a strong acid or base, but add the acid or base to the water.
18. It is your responsibility to act in a safe manner while in the chemistry laboratory.
19. Never use open flames in the lab when using volatile organic solvents. (Hot plates are fine.)
20. Never use force to place a stopper, cover, etc., on a glass container. When inserting glass tubing into holed stoppers, use lubricant and hand protection. In case of laceration, first priority is to curtail bleeding.

If you are not sure about any of the safety procedures or safety issues, ask your instructor immediately.

"I have read and agree to abide by these safety guidelines."

Print Name

Signature

Date

Name of Course and/or Laboratory

Course/Lab Meeting Time

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Sign this form, take the safety quiz on the reverse side and turn it in

1. When you should wear goggles?

- a. Whenever I feel like it.
- b. When the instructor is looking.
- c. Whenever anyone in lab is working with chemicals.
- d. Whenever I am working with chemicals.

2. Eating and drinking in lab is

- a. OK anytime.
- b. OK when I am not working with chemicals.
- c. OK when no one is working with chemicals.
- d. Is never OK

3. An MSDS

- a. Provides safety information on chemicals.
- b. Stands for Material Safety Data sheet.
- c. Can be looked at anytime by the student.
- d. All of the above.

4. I should wash

- *a. the lab tables before and after lab.
- b. the lab tables after lab.
- c. the lab tables before lab.
- d. I don't need to wash the lab tables ever.

5. I should dispose of chemicals

- a. By flushing them down the sink.
- b. By putting them in the paper trash.
- c. By following the lab instructors directions which may include using in waste jars.
- d. Drinking them.

6. If an accident occurs, you should first

- a. Run.
- b. Yell help or other shout for assistance.
- c. Cuss
- d. Faint

You must get 100% to remain in the lab so be careful!

7. If you spill a chemical on you body, you should

- a. Immediately wash it off with lots of water and yell help.
- b. Neutralize the chemical with other chemicals and yell help.
- c. Cry and yell help.
- d. Cuss and yell help.

8. I should wash my hands

- a. Before eating.
- b. After going to the bathroom.
- c. After lab is over.
- d. Frequently during lab and after lab is over.

9. If I have any injury not matter how minor, I should

- a. notify the lab instructor
- b. call my lawyer
- c. take care of it myself.
- d. go see the college nurse

10. If I am not sure how to proceed safely, I should

- a. go ahead anyway.
- b. ask my lab partner.
- c. ask my lab instructor.
- d. leave lab for safety reasons.

11. The most important way to be safe is

- a. skim over procedures quickly because I will get out fast and that is the safest.
- b. read and follow procedures carefully and exactly.
- c. use my own judgement.
- d. rely on my lab partner.

You must get 100% to remain in the lab so be careful!

Lab 1

Measurement, Metric System, and Archimedes Principle

Scientists use a variety of laboratory apparatus and equipment. In this activity you will be introduced to various pieces of glassware and instruments that will be commonly used in this course. When using measuring devices that have calibration lines, **one always estimates one decimal between the demarcation lines**, always!! If your instrument has a digital read-out, then be sure to read all of the digits. Watch your instructor demonstrate the various techniques.

1. Fill a 500 mL flask to the top with water. Use the 100 mL graduated cylinder to measure the amount of water the 500 mL flask contains. Repeat with a 250 mL beaker. Record your numbers on the data sheet.
2. Using a teaspoon and a 10 ml graduated cylinder, determine and record the volume of a teaspoon.
3. Obtain a one of the metal disks that has a labeled mass. Weigh it using the triple beam balanced and then on the electronic balance. Use at least two different electronic balances. Record and compare the results for each balance.
4. Measure the length and width of the table in the center of the room. Use a short ruler, a commercial meter stick and a meter stick without any calibration lines which I made at home. Record and compare your measurements. Which do you think is the most accurate?
5. From your measures in part 4 calculate the area of the center table based on the measurements from each device. You should have three values for the area. Record them. In doing calculations, significant figures refer to the appropriate number of digits that you can use legitimately from you measurements. You just cannot copy down all of the numbers of your calculator. Consult with your instructor about the use of significant figures in your answers for the area.
6. Density is an intrinsic property of each and every pure substance. As a result it often can be used to help identify a material. Density is mass/volume. Density can be compared by floating or sinking. Stuff which is less dense floats on more dense material. Many people incorrectly use the term heavier to describe sinking.

Measuring mass is straightforward using a balance. Measuring volume, however, can be somewhat more challenging. Simplifying Archimedes' discovery, when an object is placed in water it displaces a volume of water equal to its own volume. Obtain one of the small objects for the density determination. Mass it. Now put some water in a graduated cylinder. Measure the reading on the graduated cylinder. Now add your object and measure the water level. By difference, you have the volume of the object. Calculate its density, being careful of significant figures and record it.

Look up and record the accepted value of the density of your object in one of the references provide. How did you do??

7. On the data sheet write down your height in centimeters (cm). 1 inch = 2.54 centimeters. Write down your weight in kilograms (kg). 2.2 pounds = 1 kilogram.

1. Volume of 500 mL flask _____ mL (record to the nearest 0.1)
Volume of 250 mL beaker _____ mL (record to the nearest 0.1)
2. Volume of teaspoon _____ mL (record to the nearest 0.01)
3. Mass of disc stamped on the disc _____ grams.
Mass of disc weighed on triple beam balance _____ grams (record to nearest 0.01g)
Mass of disc weighed on 1st electronic balance _____ grams (record to nearest 0.001g)
Mass of disc weighed on 2nd electronic balance _____ grams (record to nearest 0.001g)

Which balance do you think is more accurate? Why?

4. Center table measured with short ruler
L = _____ cm (record to the nearest 0.1 cm) = _____ m (divide by cm by 100)
W = _____ cm (record to the nearest 0.1 cm) = _____ m (divide by cm by 100)

Center table measured with commercial meter stick

- L = _____ cm (record to the nearest 0.1 cm) = _____ m (divide by cm by 100)
W = _____ cm (record to the nearest 0.1 cm) = _____ m (divide by cm by 100)

Center table measured with homemade meter stick

- L = _____ cm (record to the nearest 0.1 m) = _____ m (divide by m by 100)
W = _____ cm (record to the nearest 0.1 m) = _____ m (divide by m by 100)

5. Area using short ruler measurements _____ m²

Area using commercial meter stick measurements _____ m²

Area using homemade meter stick measurements _____ m²

Which of these of these distance measuring devices do you think is most accurate? Why?

6. Describe object. What type of metal do you think it is?

Mass of object _____ g (record to nearest 0.001 g)

Volume of water in graduated cylinder before adding object _____ mL (nearest 0.1 ml)

Volume of water in graduated cylinder after adding object _____ mL (nearest 0.1 ml)

Volume of object _____ mL (nearest 0.1 ml)

Density of object (divide mass by volume) _____ g/mL

Literature value of the metal used in your object _____ g/mL. A cubic centimeter (cm³ or cc) is the same as a mL.

How close is your experimental value?

7. Your height in inches _____; in centimeters _____.

Your weight in pounds _____; in kilograms _____.

8. List glassware used to measure the volume of liquids with good accuracy.

9. List glassware used to just contain liquids which have little accuracy in measuring volume.

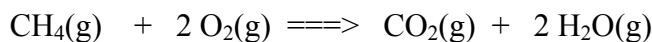
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Lab 2

Exploration of Bunsen Burner Flame

In this experiment you will determine the temperature of different portions of a Bunsen burner flame by an indirect method. Since the burner flame is too hot for an ordinary mercury thermometer, you will use a metal with a high melting point and note the color of the metal in different portions of the burner flame to give you an approximate temperature. Next you will explore the burner flame by noting what part of the flame will melt various solids. From known melting points of various solids you can determine what the approximate temperature is in the different portions of the flame.

The chemical reaction that produces the flame is the combustion of methane gas to produce carbon dioxide gas and water gas. Combustion always refers to the chemical reaction of some material with molecular oxygen, O₂.



If one reduces the amount of oxygen available, the flame changes and so does the products of the reaction. With insufficient oxygen, carbon monoxide (CO) gas and elemental carbon, C, also known as soot are produced.

Procedure A

1. Light a Bunsen burner. Close the air intake and observe the flame. Place a ceramic dish over the flame and observe what collects on the bottom of the dish. Write down your observations
2. Now open the air intake fully and observe the flame. Place a ceramic dish over the flame and observe. Write down your observation.

Procedure B

1. Obtain a 15 cm. piece of Nichrome (alloy of 60% nickel, 24% iron, 16% chrome, 0.1% carbon) wire and bend the end of it into a tiny loop. With the air intake fully open hold the loop steady in a portion of the burner flame until the color of the loop remains constant, note the color and the position of the loop on a drawing of the flame in your data. Use the same procedure in other parts of the flame until you have a complete diagram of the various loop colors in different parts of the flame. Record any observations, and put them in your data.
2. Repeat step 1 with the air intake closed. Which flame is hotter, the one with more air or less air?
3. Repeat your exploration in step 1 (air intake open) using various solids. Place a tiny sample of each available solid, one at a time, on your loop. Hold the solid in the flame until it melts. Record in your data, the identity of the solid and the position of the flame where it melted.

Results

Draw a map (picture) of the burner flame for each of the above flames indicating the approximate temperature at various points and showing the color of the wire or what solids melted at various points in the flame.

<u>Color of wire</u>	<u>Temperature °C</u>	<u>Compound</u>	<u>Melting point °C</u>
Incipient red	500	Barium nitrate	592
Dark red	700	Potassium iodide	723
Bright Red	900	sodium chloride	801
Orange	1100	sodium carbonate	851
Incipient White	1300	Brass wire	940
White	1500	Copper wire	1083

A.1. Observation off ceramic dish over Bunsen burner with air intake closed.

A.2. Observation off ceramic dish over Bunsen burner with air intake open.

B.1. Picture of flame with air intake open and with temperatures and color of wire shown.

B.2. Picture of flame with air intake closed and with temperatures and color of wire shown.

B.3. Picture of flame with air intake open and with temperatures and melting of solids shown

Lab 3

Observation and Identification of Metal Ions Using a Flame Test

In class you have seen the gas-filled discharge tubes emit light. The model to explain this phenomenon was that electrons got excited by absorbing electrical energy and jumped from a low energy orbital to a higher energy one. When the electron returned to the ground state energy level, it emitted light which has the energy equal to the difference between the energy levels.

Other types of energy can be used to excite electrons. One often used is the heat of a flame. Different metal ions of salts can be put in a flame and they will emit a spectrum of light unique for each metal. What you see is a color which is characteristic of a specific metal. Besides being very pretty, the characteristic color can be used to identify certain metal ions.

1. Place a small sample of the following salts in its own individual well.

LiCl, NaCl, KCl, CaCl₂, SrCl₂, BaCl₂, CuCl₂

2. Light a Bunsen burner and put a small amount of each salt into the flame and record your observations.
3. Obtain an unknown salt and using a flame test, identify the metal ion in it.

Fireworks: A redox lab

Chlorates (as in potassium chlorate, KClO₃) were once used in fireworks (and still are in some places) until a few years ago. This reaction involves oxidation and reduction, but the balanced equations for this reaction have not been determined.

Put on goggles. Follow safety precautions. For each experiment describe what you see. Identify what compound is being oxidized, which one is reduced, and what causes the color. Reactants include KClO₃ and C₁₂H₂₂O₁₁. Products include CO and K₂O, CO₂, H₂O among others.

1. With goggles on, carefully mix the following in a small, clean, dry beaker:
2.5g of sugar + 2.5 g KClO₃

Pour the mixture in a test tube and clamp the test tube to a ring stand pointed to the back of the fume hood.

Carefully add 1 drop of 18M sulfuric acid (H₂SO₄). Stand back with goggles on.

Now each group gets to do one of the following. It is your choice.

2. Repeat the above procedure using 2.5g of sugar + 2.5 g KClO₃ and sprinkle some powdered Mg.
For the following you may include powdered Mg as you wish.
3. Repeat the above procedure using 2.25g of sugar + 2.25 g KClO₃ + 1.5 g powdered Fe
4. Repeat the above procedure using 2.25g of sugar + 2.25 g KClO₃ + 1.5 g SrNO₃
5. Repeat the above procedure using 2.5g of sugar + 2.5 g KClO₃ + 1 g CuCl₂
6. Repeat using 2.5g of sugar + 2.5 g KClO₃ + 0.5 g SrNO₃ + 0.5 g CuCl₂ and sprinkle some powdered Mg,

Chem 11 Lab 3 Flame Test Data Name _____

Color of flame with

LiCl

NaCl

KCl

CaCl₂

SrCl₂

BaCl₂

CuCl₂

Unknown 1

Unknown 2

Unknown 3

Unknown 1 is the compound _____

Unknown 2 is the compound _____

Unknown 3 is the compound _____

Describe your fireworks

Lab 4

How Much Does a Penny Weigh?

Every scientific study begins with a question. Science is observing, experimenting, and analyzing in order to answer or at least try to answer the question. Today's initial question is how much does a penny weigh?

Procedure for physical properties

1. Work in groups of 2. Each group should get a container full of pennies.
2. Devise a step-by-step experimental procedure and write it down. Then perform your procedure and write down your answer on your paper and on the front board.
3. Next write a step-by-step procedure to determine the density of your penny using the formula $\text{density} = \text{mass}/\text{volume}$. Perform the task and write your answers on the paper and board. Let's discuss our results.

Chemical properties of your pennies

Now let us look at chemical properties of pennies. This means chemical reactions will occur.

1. Obtain some different pennies. Write down the dates. Put a large nick on the pennies using a file.
2. Put some hydrochloric acid (HCl) (**caution strong acid!!! Will burn your skin!!**) in a 250 mL beaker and add your pennies. Observe any reaction and write down what you see.
3. After your observations are complete, pour off the hydrochloric acid into the sink and rinse the pennies with lots of water. Which dates reacted?? Which did not? Take those pennies that did not react and put them in a 250 mL beaker. **UNDER THE HOOD**, add some concentrated nitric acid (**CAUTION!!**) and observe.
4. Place some HCl in a two different test tubes. To one tube add some solid zinc (Zn) and to the other add solid copper (Cu). Put some Cu in a test tube and under the hood add concentrated nitric acid. Write down your observations. Dispose of the Cu/nitric acid mixture in the large beaker in the hood.
5. Now compare these observations with your observations in step 3. Any correlation? Explain on paper. Check out the density of zinc and the density of copper from a reference book. Compare these densities with the densities of your pennies. Any correlation?

Fun Chemistry in Action

Your instructor has done the following: Placed about 5 grams of Zn in your 250 mL beaker and to it add 70 mL of sodium hydroxide (NaOH). In the hood put the beaker on a hot plate and heat the solution to nearly boiling.

You now should **Be careful!! Sodium hydroxide (lye) is extremely caustic. It hydrolyzes (eats) the proteins in your skin.**

3. Clean some pennies using the steel wool and put them in the beaker using tongs or forceps.
4. In about 5 minutes, the pennies should be coated with a silvery coating. When completely coated, remove the pennies using tongs, rinse well and blot (not rub) dry using a paper towel.
5. Remove the beaker from the hot plate, let it cool, and pour the contents into the beaker on the center bench.
6. Place your silver penny on the hot plate. Remove it (tongs) when the color changes. Rinse and dry.

1. Procedure for determined the mass of a penny.

Data from your procedure.

2. Procedure for determining the density of a penny.

Date from your procedure.

Chemical Properties

Observations after adding HCl

Observations after adding HNO₃

Observations with Zn and HCl

Observations with Cu and HCl

Observations with Cu and HNO₃

Density of Zn _____ g/cc. Density of Cu _____ g/cc. Density of penny _____ g/cc

Are there correlations with your penny observations? Can you deduce the metals used in different pennies?

Lab 5

Distillation

Introduction

Distillation is a physical process used to separate different substances in a mixture. The distillation process takes advantage of differences in the property of volatility or boiling point. Volatility refers to how easily a substance evaporates – the higher the volatility of a substance, the faster it evaporates. A highly volatile material has a low boiling point. A substance with low volatility has a high boiling point. The quantitative measure of volatility is vapor pressure.

In any mixture, the gases above the mixture have more of the substance that evaporates easily (more volatile substance). In other words if you have 50/50 mixture of two different liquids in a mixture and analyze the gases above the mixture of liquids, you would find that the gases from the more volatile liquid would be more than 50% and the gases from the less volatile liquid would constitute less than 50%. If you would collect the gases above the mixture and condense them (turn them back to liquid), the condensed liquid would no longer be a 50/50 mixture. The condensed liquid would be “enriched” with the more volatile liquid. Enriched means the condensed liquid would contain more than 50% of the more volatile liquid and less than 50% of the less volatile liquid. This is the exact process that occurs in a distillation. In a distillation, the original mixture is heated so the liquids evaporate and the separation occurs more quickly.

For example distillation can separate pure water from a solution that contains table salt (NaCl) and water. The “normal” boiling point of is 100°C and the boiling point for NaCl is 1413°C. Water has a lower boiling point, is more volatile, and evaporates more easily than NaCl. In a distillation of salt water, you would heat the mixture until the water boils. The water is turning into a gas but the NaCl is not turning into a gas to any measurable extent. The water gas leaves the original container and is then condensed back into liquid water. Since the NaCl never turned into a gas, it stays in the original container. The condensed liquid is pure water.

This process is one method used to desalinate water. Boiling salt water is very energy intensive so distillation is an expensive way to desalinate salt water. Primarily oil rich countries which have a shortage of fresh water use it.

You can also separate a mixture of ethanol and water using distillation. Ethanol is more volatile than water, so it evaporates faster. The condensed liquid after a distillation of a mixture of ethanol and water is enriched in ethanol. The boiling point of ethanol is 78 °C and the boiling point of water is 100 °C. Because these boiling points are so close, it takes a number of distillations to achieve maximum separation. Due to the fundamental nature of ethanol and water, 100% separation is not possible by distillation. The maximum separation of ethanol and water by distillation is a 95% ethanol and 5% water solution.

Distillation of Crude Oil

Upon delivery to a refinery, crude oil which is primarily a mixture of organic compounds is distilled. The major component of crude oil is alkanes. The boiling point and volatility of alkanes is a function of the number of carbons in the alkane - the more carbons give a high boiling point and a lower volatility. The fewer the number of carbons leads to higher volatility and lower boiling point. As a result, C1 to C4 alkanes distill out first, then come C4 to C10, then C10 to C16, next are C16 to C22, followed by C22 to C30 and finally alkanes with more than 25 carbons.

To get best separation, a mixture needs to be performed a number of times. By using a technique called fractional distillation, many distillations can occur at the same time and place. In a refinery, this occurs in a fractionating tower. This is the tall tower that you see at refineries.

1. The distillation apparatus is already assembled. Some of you will be in charge of the distillation of crude oil, others will be in charge of the colored salt water, and others will be in charge of wine if that distillation is performed.
2. Each you should observe all the distillations. Draw a diagram of the distillation apparatus below.
3. Turn on the condensing water. Turn on the heat and heat the liquid. **DO NOT STAND IN FRONT OF THE CONDENSER!!**
4. Crude Oil: Monitor the temperature of the gases. When the temperature reaches 175 °C, remove the receiving beaker and replace with another beaker. When the temperature reaches 275 °C, turn off the heat but **NOT THE CONDENSING WATER**. The first beaker contains the first cut or gasoline fraction; the second beaker contains the second cut or the kerosene fraction. We will go outside and burn these.

What is the color and smell of first fraction?

What is the color and smell of second fraction?

- b. For the colored salt water: Monitor the temperature of the gases. Does it change? Distill until you have about half the original amount left. Observe the condensed liquid.

Is there color in it?

Is there salt in it? How could you test to see if the salt is still in the condensed liquid?

Wine:

1. Obtain a thermometer distilling flask and condenser
2. Assemble the distilling apparatus
3. Remove the thermometer and place a boiling chip and about 75ml of wine using a long-stem funnel into the distilling flask.
4. Turn on the cooling water and have the instructor check your setup, then apply the heat to the wine. After boiling begins, reduce the heat so the wine boils at a moderate rate. Note the temperature of the boiling liquid. Collect the distillate. This is the head or tête in France. This is discarded.
5. Watch the temperature closely while distilling and collect the distillate coming from the condenser while the temperature is in the range 80-85°C. The temperature should remain nearly constant in this range. This is the heart or the coeur on France. This is the good stuff.
6. When the temperature rises above 85°C, begins to increase, quickly change collecting flasks then continues distilling until the temperature is between 98 and 100°C. This is the tail or queue in France. This is discarded.
7. Record the temperature range over which you collected the alcohol enriched distillate.
8. Turn off the heat but **NOT THE CONDENSING WATER**.

What is the color of the condensed liquid?

What is the smell?

Questions:

1. Put in order from lowest boiling point to highest boiling point: butane, octane, hexane
2. Put in order from lowest boiling point to highest boiling point: ethanol, NaCl, water

Diagram of distillation apparatus.

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Lab 6

Crystal Growing

Making crystals depends on how soluble each particular substance is in water. Some substances such as sugar (sucrose) are so soluble that 100g of substance can dissolve in 50g water. Other substances are so insoluble that they can't dissolve in water at all, sand, SiO_2 is one of these substances.

In the experiment below you will be making close to saturated solutions of a variety of compounds. If the solution isn't saturated with the maximum amount of compound in water, the crystals won't form until some of the solvent (water) evaporates off. Some crystals are slow to form even with saturated solutions. Most substances are more soluble in hot water than cold water, as the solution cools the crystals will form because in hot water the solution is not saturated, but in cold water the solution becomes saturated. As a general rule, the slower the crystal formation the larger the crystal. Crystal formation is affected by impurities in tap water. **Use distilled water for making crystals.** Distilled water theoretically is pure.

Procedure:

Heat water in a small 50 mL beaker to boiling. Add the amount of weight suggested with stirring until the substance dissolves completely. Cover beaker with aluminum foil or plastic wrap and poke some small holes in it. Label substance and your name in beaker with chem. tape. Let solution stand and watch for crystals. At your work station, each person needs to make a different crystal of each of the substances listed in A-D below. If you want to make more than one different crystal, go for it. Later we will use the crystals to make beautiful things with casting resin. Observe your solutions for crystal growth over the next few weeks. Store your creations on top of reagent shelf until further notice. Observe your crystals weekly and put observations on Observation Report Sheet. Store this sheet in your lab drawer for weekly observations.

The following are different salts. Note the amount of salt and amount of distilled water needed in each case.

- A. NaClO_3 sodium chlorate 4 g. Use 30 mL of water.
- B. NaNO_3 sodium nitrate 4 g. Use 30 mL of water.
- C. $\text{K}_3\text{Fe}(\text{CN})_6$ potassium ferricyanide 4 g. Use 30 mL water.
- D. $\text{Cu}(\text{CH}_3\text{COO})_2 \cdot \text{H}_2\text{O}$ copper II acetate monohydrate 1g. Use 30 mL water.
- E. CuSO_4 copper II sulfate 2.5g. Use 40 mL water in a 100mL or 150 mL beaker. Everyone needs to make this crystal

Chem 11 Lab 6 Data Sheet for Crystal Growth Name _____

Keep this sheet in your lab drawer to record your observations over time. Record everything you see, color, shape. Using a microscope make a drawing of the shape of the crystals you made.

Name of substance _____ Formula _____
Date _____ Observations _____

Name of substance _____ Formula _____
Date _____ Observations _____

Name of substance _____ Formula _____
Date _____ Observations _____

Name of substance _____ Formula _____
Date _____ Observations _____

Name of substance _____ Formula _____
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Date _____ Observations _____

Name of substance _____ Formula _____
Date _____ Observations _____

Name of substance _____ Formula _____
Date _____ Observations _____

PAPER CHROMATOGRAPHY

Introduction

One of the most important analytical procedures in chemistry and other branches of science is chromatography. The word chromatography which comes from the Greek words “color” and “to write” was coined by the Russian botanist Mikhail Twsett. He was the first to use this procedure in 1906 to separate and identify plant pigments.

Chromatography refers to a process for separation and subsequent identification that employs a mobile phase and a stationary phase. The stationary phase does not move. The mobile phase moves. Molecules which are attracted to the chemical that moves separate from molecules which are attracted to the chemical that remains stationary. Different migration rates are caused by the relative amount of attraction of molecules to either the stationary or mobile phase. The greater the attraction a molecule has for the mobile phase over the stationary phase, the further it moves in a given time and vice versa.

Typically the attraction between the sample molecules and the stationary phase or the mobile phase is a function of the polarity of the sample molecules and the relative polarity of the stationary phase and the mobile phase. The more polar molecules are better attracted to the more polar phase. The less polar molecules are better attracted to the less polar phase. Water is a polar molecule. Other molecules which readily dissolve in water are also polar. Sometimes these molecules are called **hydrophilic** meaning water loving.

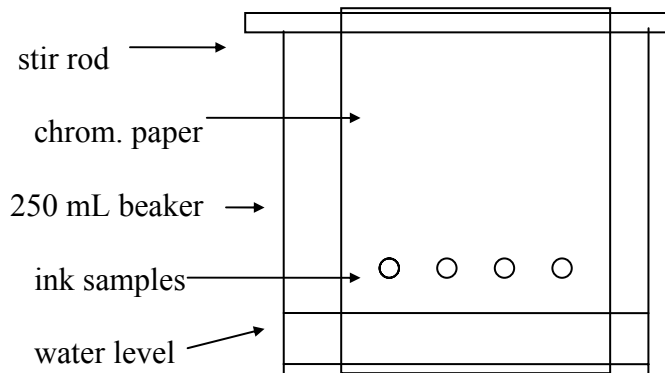
Molecules that do not readily dissolve in water such as oil or gasoline are typically nonpolar and referred to as **hydrophobic** – water fearing.

In this experiment water is the mobile phase and the paper is the stationary phase. Ink is a mixture of different pigments. Each pigment typically corresponds to one kind of molecule. Each color of ink may have one or more pigments in it which you will attempt to separate. For example green ink may consist of one pigment molecule that is green or a mixture of two pigment molecules, one yellow pigment molecule and one blue pigment molecule, that produce green when mixed.

The more a pigment molecule is attracted to the water, the further it moves up the paper. Furthermore, if two of the same colored pigments travel the same distance on the paper, this becomes evidence that they may be the same molecule. If you put molecules of which you know the structure (called standards) on the paper and they are of the same color and move the same distance as one of the pigments in the ink pens, then you probably know that the unknown is the same as your known standard. Viola, you have identified the previously unknown pigment.

Chromatography is used to separate and identify chemical compounds. These compounds can be nearly of any type, not just pigments. It can involve the solid, liquid, or gaseous states of matter.

1. Obtain some chromatography paper.
2. Set up the following according to the diagram and use the two different sets (by brand) of pens. Each set consists of a black pen, a green pen, and a blue pen. Label in pencil the brand of pen and the ink color of each original dot that you used on each sheet to be sure you know which pen corresponds to which mark. So each dot should have a label such as "papermate green" in pencil. Also do a black unknown which you will identify.



3. Let the water wick up the paper but not over the top.
4. Remove the chromatogram, let it dry, and tape or staple it to this paper.
5. Make something creative using the ink pens and chromatography.

Partners: _____

1.a. What was the mobile phase? _____

1.b. What was the stationary phase? _____

2. Label each black pen by name brand and then list the different colors of pigment that you have observed.

pen 1 name: _____ colors: _____

pen 2 name: _____ colors: _____

pen 3 name: _____ colors: _____

pen 4 name: _____ colors: _____

pen 5 name: _____ colors: _____

unknown colors: _____

3. For each black pen, which color of pigment was at the top of the chromatogram and at the bottom?

pen 1 name: _____ top color: _____ bottom color: _____

pen 2 name: _____ top color: _____ bottom color: _____

pen 3 name: _____ top color: _____ bottom color: _____

pen 4 name: _____ top color: _____ bottom color: _____

pen 5 name: _____ top color: _____ bottom color: _____

unknown: _____ top color: _____ bottom color: _____

4. For each black pen, which color of pigment was most attracted to the water and which was least attracted?

pen 1 name: _____ most attracted color: _____ least attracted color: _____

pen 2 name: _____ most attracted color: _____ least attracted color: _____

pen 3 name: _____ most attracted color: _____ least attracted color: _____

pen 4 name: _____ most attracted color: _____ least attracted color: _____

pen 5 name: _____ most attracted color: _____ least attracted color: _____

unknown: _____ most attracted color: _____ least attracted color: _____

Conclusion: My unknown black pen, number _____, is: _____

5. Label each green pen by name brand and then list the different colors of pigment that have observed.

pen 1 name: colors:

pen 2 name: colors:

pen 3 name: colors:

pen 4 name: colors:

6. For each green pen, which color of pigment was at the top of the chromatogram and at the bottom?

pen 1 name: top color: bottom color:

pen 2 name: top color: bottom color:

pen 3 name: top color: bottom color:

pen 4 name: top color: bottom color:

7. For each green pen, which color of pigment was most attracted to the water and which was least attracted?

pen 1 name: most attracted color: least attracted color:

pen 2 name: most attracted color: least attracted color:

pen 3 name: most attracted color: least attracted color:

pen 4 name: most attracted color: least attracted color:

8. Label each blue pen by name brand and then list the different colors of pigment that have observed.

pen 1 name: colors:

pen 2 name: colors:

pen 3 name: colors:

pen 4 name: colors:

9. For each blue pen, which color of pigment was at the top of the chromatogram and at the bottom?

pen 1 name: top color: bottom color:

pen 2 name: top color: bottom color:

pen 3 name: top color: bottom color:

pen 4 name: top color: bottom color:

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Lab 8

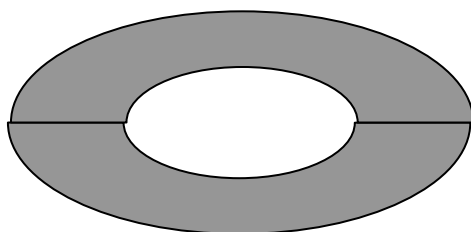
Polymers

Polymers are types of materials that indispensable to life. Polymers make up parts of our body and other living things. These are natural polymers. Examples are hair, protein, the genetic material RNA and DNA, cotton, wood, silk, wool, starch and natural latex rubber.

Chemists have been making polymers for over a century. One of the earliest partially synthetic polymer was produced in the 1800's to replace ivory in billiard balls in order to save elephants from being slaughtered. It is now called celluloid. It once was also used to make the actual film for recording movies but now it is only used to make ping pong balls. The first entirely synthetic polymer is Bakelite and it is still used for handles of pots, pans, and toasters as well as billiard balls and wall covers for electrical switches. Other examples of synthetic polymers are fabrics such as Dacron, nylon, polyester and polypropylene; virtually all plastics that we use in our daily lives; glues; caulks; and even fillings for your teeth.

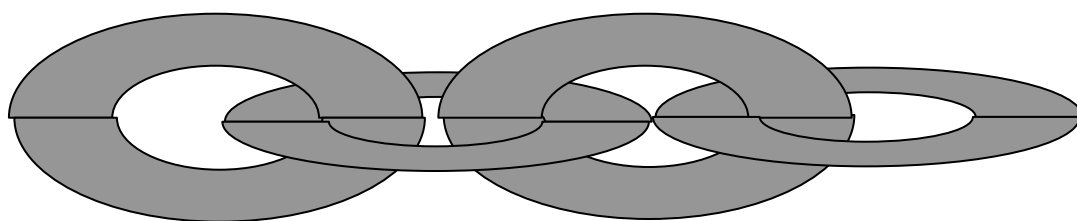
Polymers have a variety of different properties and chemists develop new ones regularly to meet the demand for new products with specific new properties. All synthetic polymers come from oil.

We can classify polymers in a variety of ways but they all share one common theme, they are chains of repeating units. Poly means many and mer means part. The repeating unit is called a monomer.



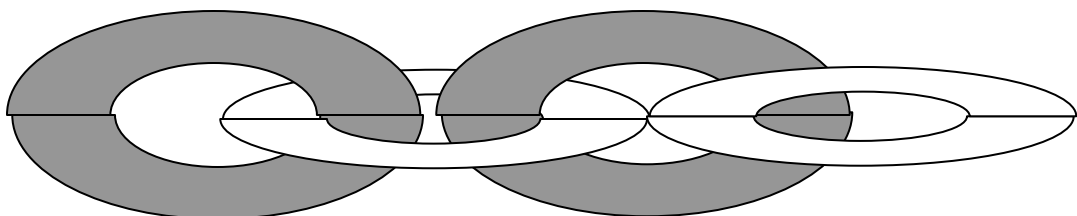
monomer

If there is only one repeating unit, the polymer is called a homopolymer.



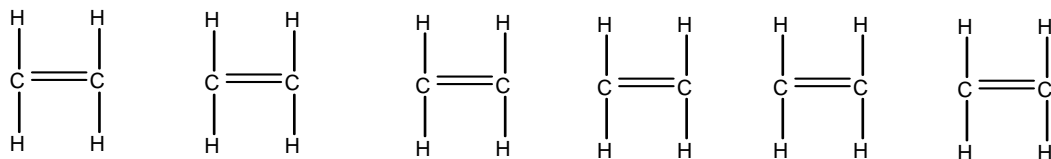
homopolymer

If there are two repeating unit it is a copolymer

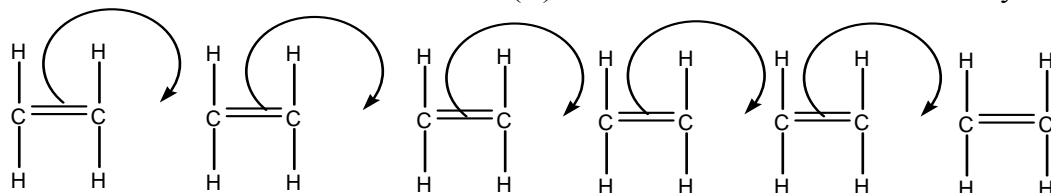


copolymer

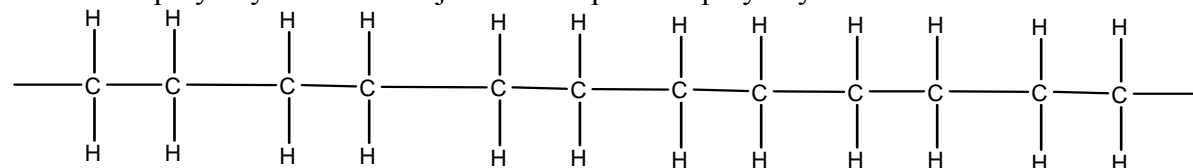
Another way to classify polymers is how they chemically combine. The two classifications are addition polymers and condensation polymers. Addition polymers are typically homopolymers with the most common being polyethylene. Here are ethylene monomers ready to become polymers (an initiator molecule starts the reaction)



One of the bonds between the carbons (C) breaks and attaches to the next ethylene molecule



to become polyethylene. This is just a small piece of polyethylene.



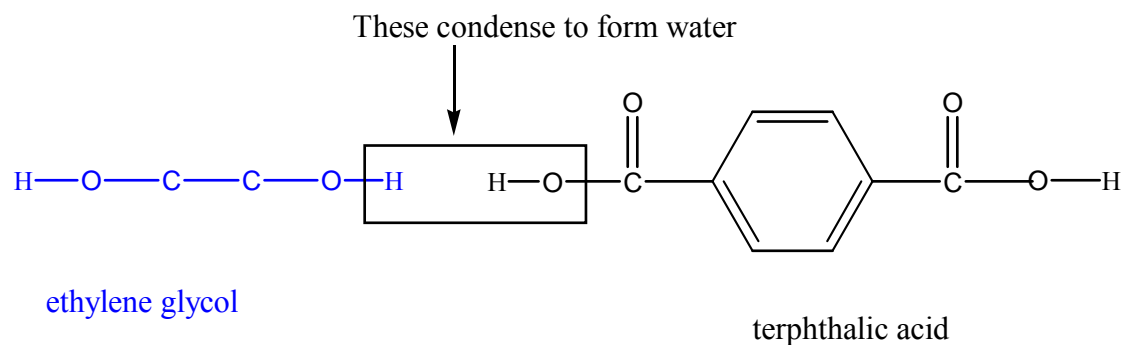
A common initiator for addition polymers is benzoyl peroxide which is a common ingredient in acne medicine. The benzoyl peroxide when heated forms 2 free radicals which are molecules with an unpaired electrons. These are very reactive and potentially harmful to biological systems. The free radicals then start the addition reaction.

Below is a table of common addition polymers

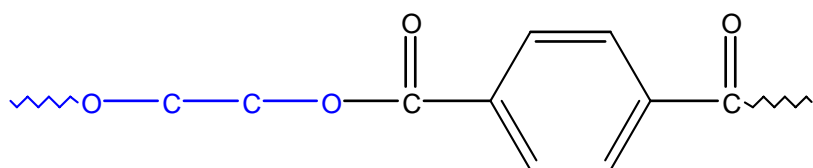
Recycling code	Name	Monomer by itself	Monomer part of polymer
2 4	high density polyethylene (HDPE) low density polyethylene (LDPE)	ethylene (polyethylene)	
5	PP propylene (polypropylene)		
6	PS styrene (polystyrene)		
3	vinyl (V) vinyl chloride (polyvinyl chloride or PVC)		
3	vinyl (V) vinylidene chloride (polyvinylidene chloride) (Saran)		
no code	tetrafluoroethylene (polytetrafluoroethylene or PTFE or teflon)		
no code	acrylonitrile (polyacrylonitrile or Orlon)		
	vinyl acetate (polyvinyl acetate in white glue and chewing gum)		

There are two very common condensation polymers polyesters and polyamides

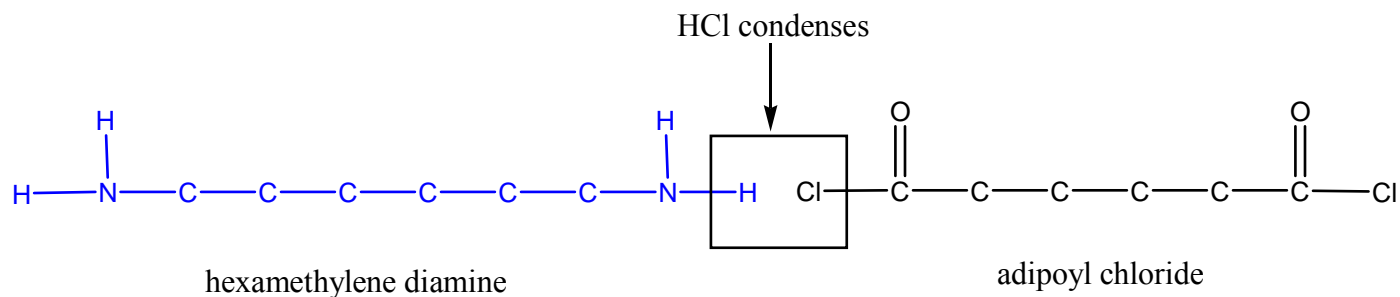
Polyester



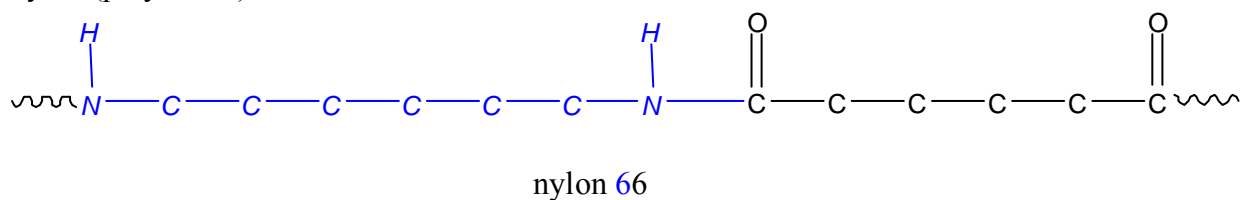
Polyethylene terephthalate which is a polyester It has a recycling code which is 1 (PETE)



Polyamides



Nylon (polyamide)



It is 66 because there are 6 blue carbons and 6 black carbons. There are many other types of "nylons" depending on the number of carbons.

Activity Procedures

Packing Peanuts:

1. Place a starch packing peanut in a beaker of water (H_2O) and observe. Record on below.
2. Place some polystyrene packing peanut in a small beaker of acetone (fingernail polish remover for natural nails) and record observations.
3. Place some polystyrene packing peanut in a small beaker of ethyl acetate (fingernail polish remover for acrylic nails – another polymer!) and record observations.

Super Slurper:

1. Place a small amount (~1 tsp) of sodium polyacrylate in a beaker of water and observe.
2. When you are finished with this step pour some table salt ($NaCl$) in the same beaker and record observations.

Edible Plastic: (but do not eat it!!)

1. Place the following in a 100mL flask.
 - 2.5 grams cornstarch or 2.5 grams potato starch
 - 2 milliliters of glycerin solution (half glycerin and half water) and food color as desired
 - 3 milliliters of dilute hydrochloric acid (0.1 mol/L HCl solution)
 - 20 milliliters of distilled water (25 mL if one uses potato starch)
2. Now heat this mixture for 15 minutes at a temperature of 105 - 110°C.
3. Remove from heat and add 3 milliliters of dilute sodium hydroxide solution (4 g $NaOH$ /L solution)
4. Mix thoroughly and then pour onto a piece of acrylic glass (Plexiglas). Wait to cool. Record observations.

Gluep:

1. Make a 4% borax solution by mixing $\frac{1}{4}$ cup borax in a quart of warm water (already made). Stir well. All the borax need not dissolve. Just use the liquid solution.
2. Mix equal volumes of white glue and water (already made). Add food color as desired. Stir well.
3. Mix the glue solution and the borax solution in a ratio of 3:1. For example mix 30 milliliters (2 tablespoons) of the glue solution and 10 milliliters (2 teaspoons) of the borax solution. Knead well.

Congratulation, you have made gluep! You can make colored gluep by adding some food color to the glue solution before you add the borax solution. It will dry out and become crumbly but you can save it in a zippered plastic bag. If it gets stuck on fabric, soak it in vinegar.

The science: White glue is a polymer of polyvinyl acetate. The chains of the polymer are independent of each other and can slide by each other so the glue flows. Borax makes links between the independent chains so the chains are no longer independent but now are all connected. Now the glue can no longer flow and it solidifies into gluep. This is cross linking the polymer. Charles Goodyear discovered that you can crosslink natural latex rubber using sulfur. This made the natural much harder and as a result it could be use for many more purposes including tires. That process is call vulcanizing.

Slime: 1. Mix 30 milliliters (2tbs) of 4% polyvinyl alcohol solution with 5 milliliters (1 tsp) of 4% borax solution. Add food color as desired. Knead well. Adjust as necessary. Record observations.

Polyvinyl Alcohol Fibers:

1. Place 4% polyvinyl alcohol solution in a beaker to a depth of about inch (2.5 centimeters). Add food color.
2. Tip the beaker slightly and gently float an equal amount of acetone.
3. Using tweezers pick up the material at the interface of the layers. Place the strands on some plastic to dry. Put some dried strands in water to see what happens. Record observations.

Nylon: Class Project

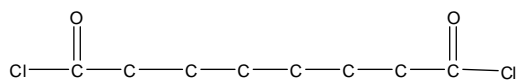
1. Place 5 mL of solution A in a small beaker. Solution A contains 0.5M hexamethylenediamine solution.
2. Place 5 mL of solution B in a second small beaker. Solution B contains 0.25M adipoyl chloride solution .
3. Float solution A on top of solution B. Wearing gloves use tweezers to grab the nylon fibers at the solutions' interface. How long a fiber can you get? Roll them up on the wood splints. Record observations.

Plexiglas: Optional

1. Place about 10 milliliters of methylmethacrylate in a test tube provided on the cart.
2. Have your instructor place about 0.05 grams of benzoyl peroxide in the tube and mix.
3. You can place an item to be encased in Plexiglas in the tube now.
4. Put the test tube in a 90°C water bath for about 10 minutes.
5. Transfer tube to a 50°C water bath until the Plexiglas hardens.

Observations and questions:

1. What happened to each packing peanut when you put them in water?
2. What happened to each packing peanut when you put them in acetone?
3. What happened to each packing peanut when you put them in ethyl acetate?
4. What happened when you put the sodium polyacrylate in water?
5. What happened when you put the sodium chloride in the beaker with sodium polyacrylate in water?
6. How did your edible plastic turn out?
7. Which did you like better, gluep or slime? Why?
8. What did the borax do to solidify the gluep and slime?
9. What did the crude nylon look like? How long was your fiber?
10. Is nylon a homopolymer or copolymer?
11. Is Plexiglas a homopolymer or copolymer?
3. Is nylon a condensation or addition polymer?
4. Is Plexiglas a condensation or addition polymer?
5. Nylon 66 is made by reacting hexamethylenediamine and adipoyl chloride. Draw a piece of nylon 68 which is made by reacting hexamethylenediamine and sebacoyl chloride.



sebacoyl chloride

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Lab 9 ACIDS, BASES, and pH

Acids and bases are present in virtually all water solutions and influence our everyday considerably.

I. ACIDS are compounds which produce the H^+ ion in water. The formula for almost always have an H in them. For this class, we will only consider those acids that have H in the formula. Acids can be classified as inorganic or organic. There are enumerable examples of each.

A. Inorganic Acids are compounds that have H^+ as the cation combined with some anion. The formula of inorganic acids always start with H and always have acid in their names.

Examples: HCl – hydrochloric acid – also known as muriatic acid in pool supply and hardware stores
HNO₃ – nitric acid – one component of acid rain
H₂SO₄ – sulfuric acid – another component of acid rain, mine run-off, and also battery acid
H₂CO₃ – carbonic acid – naturally present in rain caused by CO₂ in air dissolving in water
H₃PO₄ – phosphoric acid – present in cola drinks and used as a rust and stain remover

B. Organic acids are compounds that have the $-CO_2H$ moiety at one end of a carbon chain. The H of the $-CO_2H$ is the one that produce H^+ in water. Other H's in organic compounds do not produce H^+ in water (there are a few exceptions which we will not consider)

Examples: HCO₂H formic acid – used as a chemical weapon by many ants
CH₃CO₂H acetic acid – the acid in vinegar
CO₂H citric acid – the acid in citrus
CO₂H lactic acid – the acid in milk
CO₂H ascorbic acid – vitamin C
CO₂H salicylic acid – in acne face wash and cosmetics and used to make aspirin

II. BASES are compounds that either produce OH⁻ in water or react with H^+ . Classifications include metal hydroxides, metal oxides, ammonia, amines, and inorganic salts which contain any anion except Cl⁻, Br⁻, I⁻, NO₃⁻, SO₄²⁻, ClO₃⁻

A. Metal hydroxide base have a metal cation combined with hydroxide anion, OH⁻.

Examples: NaOH sodium hydroxide – known as lye in the grocery store; used in Drano
KOH potassium hydroxide
Ca(OH)₂ calcium hydroxide – known as slaked lime at the hardware store
Fe(OH)₃ iron (III) hydroxide – wet rust

B. Metal oxides are just like metal hydroxides but water has been removed.

Example: CaO calcium oxide - remove H₂O from Ca(OH)₂; known as quick lime.

C. Ammonia has the formula NH₃. (sometimes called ammonium hydroxide with a formula NH₄OH) It is used in many cleaners especially window cleaners.

D. Amines are organic forms of ammonia. A carbon atom replaces one or all of the H's on NH₃. They tend to be stinky. They give fish that fishy smell. Present in proteins in the form of amino acids. Also form a class of compounds extracted from plants called alkaloids.

E. Basic salts: Salts are defined as ionic compounds which the cation is not H^+ (that would make it an acid) and the anion is not OH⁻ (that would make it a base). Basic salts are salts that contain any anion except Cl⁻, Br⁻, I⁻, NO₃⁻, SO₄²⁻, ClO₃⁻ (a few more exist but we won't consider them).

In basic salts, the anion reacts with the H^+ form an acid.

Examples: Na₂CO₃ sodium carbonate – known as washing soda in the store. Use for laundry.
CaCO₃ calcium carbonate – known as limestone, used in antacids such as Tums
Na₃PO₄ sodium phosphate – known as TSP in stores which is a cleaning material
NaHCO₃ sodium bicarbonate or sodium hydrogen carbonate - baking soda

REACTIONS

ACIDS react with BASE!!

Examples: $\text{HCl} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaCl}$

$2 \text{HCl} + \text{CaO} \rightarrow \text{H}_2\text{O} + \text{CaCl}_2$

$\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}$

$\text{HCl} + \text{CH}_3\text{NH}_2 \rightarrow \text{CH}_3\text{NH}_3\text{Cl}$

$2 \text{HCl} + \text{CaCO}_3 \rightarrow \text{H}_2\text{CO}_3 + \text{CaCl}_2$ (H_2CO_3 decomposes to make CO_2 and H_2O)

$2 \text{HCl} + \text{NaHCO}_3 \rightarrow \text{H}_2\text{CO}_3 + 2\text{NaCl}$ (H_2CO_3 decomposes to make CO_2 and H_2O)

pH

pH is a way to measure and determine if a solution is acidic or basic (another term for a basic solution is alkaline). The pH scale is a numerical scale with $\text{pH} = 7$ is considered neutral which means that it is neither acidic or basic. It is neither acidic nor basic because the concentration of H^+ equals the concentration of OH^- . We write this mathematically as $[\text{H}^+] = [\text{OH}^-]$.

When the pH is below 7 the solution is considered acidic. The lower the pH, the more acidic the solution is. In acidic solutions, the concentration of H^+ is greater than the concentration of OH^- . The more acidic a solution is the lower the pH is and the HIGHER the concentration of H^+ is.

When the pH is above 7, the solution is considered basic or alkaline. The higher the pH, the more basic and the less acidic the solution is. In this case the concentration of H^+ is less than the concentration of OH^- .

pH is a logarithmic scale. What that means is a difference of one pH unit means a tenfold difference in the concentration of H^+ . A solution with a pH equal to 6 has 10 times more concentration of H^+ than a solution with a $\text{pH} = 7$. A pH of 5 is 10 times more concentrated H^+ than a $\text{pH} = 6$. So a solution with a $\text{pH} = 5$ has 100 times more H^+ than a solution with a $\text{pH} = 7$. We will not calculate pH from formulas in this class.

Measuring pH can be done with a pH meter which will give a direct numerical reading. Another older and cheaper way to measure pH is to use indicators. Indicators have different colors depending upon pH. Some indicators just have two colors, other have a range of colors. The two color indicators can only give information about one pH level. For example phenolphthalein is colorless below $\text{pH} = 8.2$ but at $\text{pH} = 8.2$ and above it turns pink. So if you add phenolphthalein to a solution and it turns pink, you know that the pH of the solution is 8.2 or above. If it does not turn pink, then the solution is below $\text{pH} 8.2$. Other indicators can have a variety of colors depending on the pH of a solution. One such indicator is called universal indicator. Another is made from purple cabbage juice. Indicators can also be put on strips of paper. These are used to test the pH of swimming pool, Jacuzzis, and soils.

Buffers

Buffers are solutions that resist a large change in pH. They contain molecules which can neutralize both H^+ and OH^- making the solution very resistant to changes in pH. That is good for living organisms. Living organisms are highly buffered because any large change in pH of a living organism would be lethal.

Lab Procedure: pH

1. In the center bench is a set of test tubes that contain a few squirts of purple cabbage juice indicator and the following pH values: 2, 3, 4, 5, 6, 7, 8, 9, 10, 11, 12. These are called standards and are to be used to compare to the colors of your solutions in procedure 2 to determine the pH of those solutions. Write down the colors of the standards in the lab report
2. In the 50 mL beakers, put about 10 mL of each of the solutions to be tested. These are variety of every day consumer products such as vinegar, ammonia, glass cleaner, baking soda in water, borax in water, TSP in water, glass cleaner in water, liquid soaps in water, sprite, club soda, white wine, lemon juice and anything else made available. If the material is solid, dissolve some in a small amount of water in a larger beaker. Record your choices on table in the lab report.
3. Put a few squirts of cabbage juice into the solutions to be tested. Compare the color to the color of the standard solutions. By comparison, determine the pH of each solution tested. Record on table in the lab report.
4. Use a pH meter to measure the pH of each of these solutions and record on table in the lab report.

Lab procedure: Buffers

1. Put 25 mL of distilled water in a 50 or 100mL beaker, 25 mL of 0.1M NaCl solution in another beaker, and 25 mL of pH 6 buffer in a third beaker.
2. Place the pH probe in the beaker with distilled water. Measure and record the pH.
3. Now add one drop of 0.1 M HCl solution to the beaker keeping the pH probe in the solution. Swirl to mix; then measure and record the pH.
4. Now add 4 more drops of 0.1M HCl and record the pH. Add 5 more drops and record the pH.
5. Remove and rinse the pH probe with distilled water and put it in the 0.1M NaCl solution. Measure and record the pH.
6. Repeat steps 3 and 4.
7. Remove and rinse pH probe with distilled water and put it in the pH = 6 buffer solution. Measure and record the pH.
8. Repeat steps 3 and 4.
9. Pour the solutions down the sink, rinse the beakers and pH probe.

Standards

pH	2	3	4	5	6	7	8	9	10	11	12	13
Indicator color												

Household materials	Indicator color	pH using indicator	pH using meter

Distilled water	pH =	0.1 M NaCl solution	pH =	Buffer solution	pH =
+ 1 drop HCl	pH =	+ 1 drop HCl	pH =	+ 1 drop HCl	pH =
+ 2 drops HCl	pH =	+ 2 drops HCl	pH =	+ 2 drops HCl	pH =
+ 3 drops HCl	pH =	+ 3 drops HCl	pH =	+ 3 drops HCl	pH =
+ 4 drops HCl	pH =	+ 4 drops HCl	pH =	+ 4 drops HCl	pH =
+ 5 drop HCl	pH =	+ 5 drop HCl	pH =	+ 5 drop HCl	pH =

7. How does water solution differ from the buffer solution when adding HCl?

8. Is the sodium chloride solution a buffer? How can you tell?

Lab 10

A COPPER REACTION CYCLE

This is way cool!!

SAFETY PRECAUTIONS: Concentrated nitric acid is hazardous. It produces skin burns and the vapor is a lung irritant. Handle it in the hood while wearing safety glasses. A dropping pipette is useful for transferring small amounts of acid. Rinse your hands after using the nitric acid. The reaction with nitric acid and copper should be done in the hood. NaOH is corrosive and should be handled only when wearing safety glasses. Methanol and acetone are flammable. Use them in the hood, keep them away from flames, and dispose of them in the labeled container in the hood.

Finally, classification of chemical reactions are

- acid/base - H^+ of an acid reacts with OH^- of a base
- redox - electrons transfer from one element to another. Often an element becomes an ion or vice versa.
- precipitation - cations combine with anions to form a solid

PROCEDURE:

1. Cut a length of pure copper wire that weighs about 0.5g and weigh it to the nearest milligram.
2. Coil the wire into a flat spiral and place it in the bottom of a 250mL beaker. **IN THE HOOD CAREFULLY** add 4.0 ml of concentrated nitric acid to the wire in the beaker. The brown gas, N_2O_4 , is a component of smog. Swirl the solution until the copper is completely dissolved.
3. At your desk, add about 100 mL of distilled water and then add 15 ml of 6M NaOH solution to the copper ion solution. This is copper (II) hydroxide or $Cu(OH)_2$.
4. Heat your solution on a hot plate with stirring until it just begins to boil. Stirring is necessary to insure the smooth boiling of the solution. When there is no more apparent reaction, remove the beaker from the hot plate and allow it to settle. You have made copper (II) oxide or CuO . Decant (pour off) the solution from the solid copper oxide. Add 150 ml of hot deionized water, stir and allow to settle again. Decant the solution once again. You have just washed your precipitate.
5. Add 30 ml of dilute sulfuric acid while stirring. Record your observations.
6. In the fume hood, add all at once, 2.0g of 30 mesh zinc metal, stirring until the solution is colorless. Write down your observations. When the evolution of gas has almost ceased, decant the liquid and transfer the copper to a porcelain dish. Wash the product with 5 ml of deionized water, allow it to settle and decant the liquid. Repeat this washing process. Finally wash the product with 5 ml of methanol and decant the methanol. Dry the dish and precipitate on a hot plate.
7. Weigh your smallest beaker and record its mass. Add the washed and dried solid copper from the dish. Weigh the beaker and the copper and record the mass.

1. Exact weight of starting Cu _____
2. The copper metal becomes the pretty blue Cu^{2+} ion in solution.

What kind of reaction is this? _____

3. What is the color of the solid $\text{Cu}(\text{OH})_2$ precipitate? _____
4. Write the balance chemical equation for formation CuO and water from $\text{Cu}(\text{OH})_2$.

5. Observations after the addition of sulfuric acid.

After addition of sulfuric acid, what color is the solution now? _____

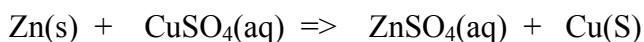
What do you think is causing this color? _____ .

What type of reaction do you think this was? _____

6. Observations after the addition of solid Zn.

What is the gas produced? _____

The reaction produced solid Cu and is described by the equation:



What type of reaction is this? _____

7. Weight of beaker _____

Weight of beaker and Cu. _____

Weight of Cu. _____

8. Calculate the percent of the original copper recovered.

% Cu recovered = (weight of Cu recovered above/weight of original Cu) x 100 = _____

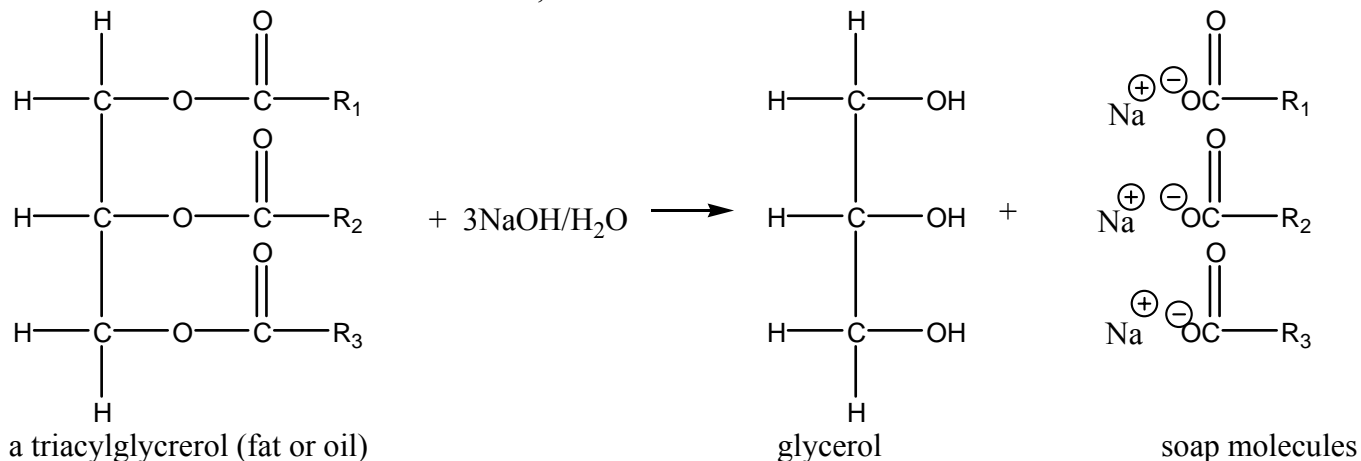
Show calculations here.

Are you a reasonably good chemist? (Did you get all your Cu back?)

Preparation of Soap

Soap is made by a saponification of fats or oils with sodium hydroxide. Fats and oils are chemically called triacylglycerols. Fats and oils are a mixture of a number of different triacylglycerols

Saponification - reaction of a triacylglycerol with sodium hydroxide to make glycerol and the sodium salt of the three fatty acids. The abbreviation R stands for the Rest of a long carbon chains in most common fats and oils the carbon chains have a total 16, 18 or 20 carbons.

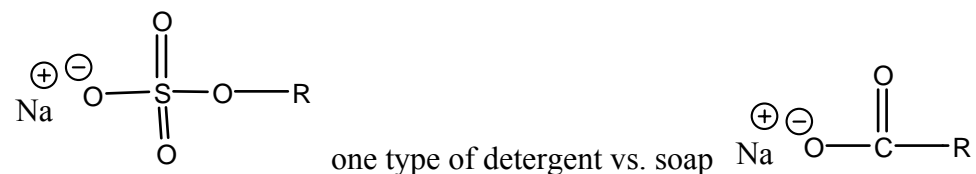


The sodium salts of the fatty acids are soap. All soaps (and detergents) share similar features:

1. One end is long carbon chain which is hydrophobic and attracts hydrophobic grease and dirt molecules.
2. The other end has a hydrophilic group which is attracted to water so the soap and grease go down the drain.

What is the difference between soap and a detergent?

A detergent has something other than a carboxylate group at the end of a long hydrocarbon chain. One type of detergent has a sulfate group at the end.



Why? Because in hard water, water with calcium, magnesium, or iron ions, soap makes a solid precipitate with these ions called soap scum. Detergent remains soluble.

Procedure:

Measure 12 mL of a vegetable oil into a 250 mL Erlenmeyer flask. Add 12 mL of ethyl alcohol (to act as a solvent) add 10 mL of lab 6M sodium hydroxide solution (NaOH). While stirring the mixture constantly with a glass rod, the flask with its contents is heated gently in a boiling water bath, a 600 mL beaker containing about 200 mL of tap water and a few boiling chips can serve as a water bath. DO NOT USE a Bunsen burner! You must use a hotplate. Watch for fires and boil-overs as alcohol is flammable.

After being heated for about 20 min., the odor of alcohol will disappear, indication the completion of the reaction. A pasty mass containing a mixture of the soap, glycerol, and excess sodium hydroxide is obtained. Use an ice-water bath to cool the flask with its contents. To precipitate or "salt out" the soap, add 100 mL of a saturated sodium chloride solution to the soap mixture while stirring vigorously. This process increases the density of the aqueous solution; therefore, soap will float out from the aqueous solution. Filter the precipitated soap with the aid of suction and wash it two times with 20 mL of saturated NaCl each time, then with 10 mL of ice cold water. Observe the appearance of your soap and record your observations below.

Properties of a soap

Emulsifying Properties: Shake 5 drops of mineral oil in a test tube containing 5 mL of water. A temporary emulsion of tiny oil droplets in water will be formed. repeat the same test, but this time add a small piece of the soap you have prepared before shaking. Allow both solutions to stand for a short time. Compare the appearance and the relative stabilities of the two emulsions

Hard Water Reactions: Hard water is water that has moderate amount of minerals dissolved. Minerals are really ionic compounds called salts. The cations Ca^{2+} , Mg^{2+} , and Fe^{3+} cause what is called permanent hardness. To see the effects of hardness on soap, place about 2 mL of the solid soap you have prepared in a 50 mL beaker containing 25 mL of deionized water. Warm the beaker with its contents to dissolve the soap. Pour 5 mL of the soap solution into each of 4 test tubes (nos. 1,2,3,4).

1. Test no. 1 add 2 drops of 7% solution of calcium chloride (7% CaCl_2)

Results

Hooray you made bathtub ring.

2. Test tube no. 2 add 2 drops of a 7% solution of iron(III) chloride (7% FeCl_3)

Results

More bathtub ring.

3. Tube no. 3 add tap water

Results:

Is the tap water hard?

4. pH test: Your skin likes soaps, detergents, and lotions to be slightly acidic with a pH around 6.5
Test soap solution no. 4 with a wide-range pH paper.

What is the approximate pH of your soap solution? _____

Detergent tests:

Add 2 mL of liquid (or solid) detergent in 25 mL deionized water. Warm the beaker with its contents to dissolve and thoroughly mix the detergent. Pour 5 mL of the detergent solution into each of 4 test tubes (nos. 1,2,3,4). Perform the same tests of the detergent as you did on the soaps. Record and compare the detergent tests with the soap tests.

1. Test no. 1 add 2 drops of 7% solution of calcium chloride (7% CaCl₂)

Results and comparison with soap:

Hooray you made bathtub ring.

2. Test tube no. 2 add 2 drops of a 7% solution of iron(III) chloride (7% FeCl₃)

Results and comparison with soap:

More bathtub ring.

3. Tube no. 3 add tap water

Results and comparison with soap:

4. pH test: Your skin likes soaps, detergents, and lotions to be slightly acidic with a pH around 6.5
Test solution no. 4 with a wide-range pH paper.

What is the approximate pH of your detergent solution? _____

Comparison with soap:

5. Draw a soap molecule that has an 18 carbon chain.

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Lab 12

Preparation of Hand Cream

Hand creams are formulated to carry out a variety of cosmetic functions. Among these are softening the skin and preventing dryness; elimination of natural waste products (oils) by emulsification; cooling the skin by radiation thus helping to maintain body temperature. In addition, hand creams must have certain ingredients that aid spreadability and provide body. In many cases added fragrance improves the odor, and in some special cases medications combat assorted ills.

The basic hand cream formulations all contain water to provide moisture and lanolin which helps its absorption by the skin. The latter is a yellowish wax. Chemically, wax is made of esters of long chain fatty acids and long chain alcohols. Lanolin is usually obtained from sheep wool; it has the ability to absorb 25-30 times its own weight of water and to form a fine emulsion. Mineral oil, which consists of high molecular weight hydrocarbons (alkanes), provides spreadability. In order to allow nonpolar (hydrophobic) substances, such as lanolin and mineral oil, to be uniformly dispersed in a polar hydrophilic (hydrophilic) medium, water, one needs strong emulsifying agents (surfactants). An emulsifying agent must have nonpolar hydrophobic portions to interact with the oil and also polar hydrophilic portions to interact with water. A mixture of stearic acid and triethanolamine, through acid-base reaction, yields the salt that has the requirements to act as an emulsifying agent.

Besides the above five basic ingredients, some hand creams also contain alcohols such as propylene glycol (1,2-propanediol), and esters such as methyl stearate, to provide the desired texture of the hand cream.

In this experiment you will prepare each of the hand creams using the combination of ingredients as shown in Table 1.

For each sample in Table 1, mix the ingredients in two beakers. Beaker 1 contains the polar hydrophilic ingredients, and beaker 2 contains the nonpolar hydrophobic substances.

Table 1- Recipes to Prepare Hand Creams

Ingredients	Cream 1	Cream 2	Cream 3	Cream 4	Beaker #
Water	13 mL	13 mL	25 mL	13 mL	1
Triethanolamine	0.5 mL	0.5 mL	0.5 mL	--	1
Propylene glycol	5 drops	5 drops	--	5 drops	1
Stearic acid	2.5 g	2.5 g	2.5 g	2.5 g	2
Methyl stearate	0.25 g	0.25 g	--	0.25 g	2
Lanolin	2 g	2 g	2 g	2 g	2
Mineral oil	2.5 mL	--	2.5 mL	2.5mL	2

- 1) To prepare a sample, put the nonpolar hydrophobic ingredients in a 50 mL beaker (beaker 2) and heat it in a water bath. The water bath can be a 400 mL beaker half filled with tap water and heated with a hot plate. Carefully hold the beaker with crucible tongs in the boiling water until all ingredients melt.
- 2) In the same water bath, heat the 100 mL beaker (beaker 1) containing the polar hydrophilic ingredients for about 5 min. Remove and set it on the bench top. You may color or aroma compounds to this.
- 3) Slowly pour the contents of the 50 mL beaker that holds the molten nonpolar hydrophobic ingredients (beaker 2) into the 100 mL beaker containing polar hydrophilic ingredients (beaker 1). Stir the mixture for 5 min. until you have a smooth uniform paste.

Characterization of the Hand Cream Preparations

1. Test the pH of the hand cream prepared using a wide-range pH paper.

Cream 1 pH =	Cream 2 pH =	Cream 3 pH =	Cream 4 pH =
--------------	--------------	--------------	--------------

2. Rubbing a small amount of the hand cream between your fingers, test for smoothness and homogeneity. Also note the appearance. Record your observations.

Cream 1

Cream 2

Cream 3

Cream 4

3) List ingredients that are hydrophobic.

4) List ingredients that are hydrophilic.

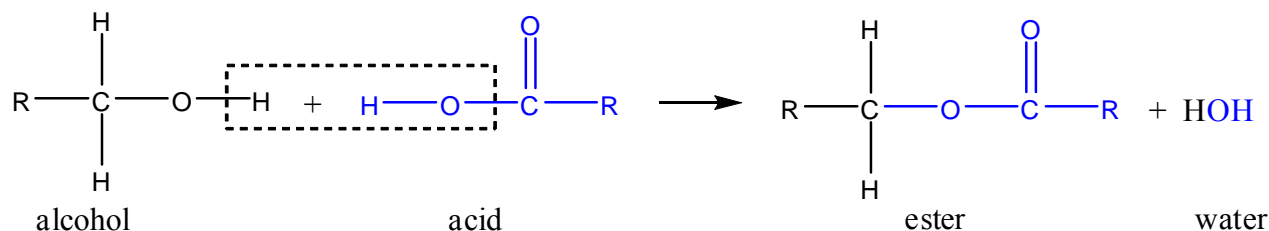
5) Which ingredient do you think is the most critical? Why?

6) Which was your favorite cream? Why?

Lab 13

Synthesis of Various Pleasant Smelling Esters

Chemists can synthesize esters in a number of ways. Today you will make esters by heating an alcohol with a carboxylic acid in the presence of an acid catalyst. The generic equation describing this reaction is



R indicates anything else can be attached here. The name would tell you what is attached

There are two ways to determine that a chemical reaction has occurred. The easiest way is by smell.

Carboxylic acids are typically unpleasant smelling in various ways. Butanoic acid (sometimes called butyric acid) smells like baby barf to me. Alcohols usually smell like a solvent. Esters, however, are very pleasant smelling and are used in fragrances and foods for their smell and flavor. Another piece of evidence for a reaction is water solubility. Small carboxylic acids and small alcohols are very water soluble. Esters are not very water soluble and will separate from water in a second layer much like oil does.

Naming Esters

Esters take their name from both the acid and the alcohol from which they were made. You start with the name of the alcohol and change the anol ending of the alcohol to yl. This becomes the first word of the esters name. Methanol becomes methyl; ethanol becomes ethyl, butanol becomes butyl, etc. Then you change the ic acid part from the carboxylic acid to ate. Acetic acid becomes acetate; propanoic acid becomes propanoate, etc. This becomes the second word of the esters name.

Example: If you react hexanol with formic acid, you make hexyl formate.

Procedure:

1. Prepare a hot water bath by filling your 600 mL beaker about half full of water and place it on a hot plate. Maintain the water at a temperature between 60 - 70°C.
2. Obtain 6 clean, dry test tubes. Mix the alcohols and carboxylic acids following the table. Use 6 mL of the alcohol and 2 mL of the acid. If the acid is solid, use about 1 gram.
3. Now add about 15 drops of concentrated sulfuric acid (H₂SO₄) to each test tube. CAUTION: Very dangerous acid. Will burn your skin very quickly. Avoid contact. Neutralize spills on your desk and wipe up. Rinse off drips on the side of the bottle. SKIN CONTACT: Rinse immediately with lots of water and let your instructor know!!!!
4. Place the test tubes in the hot water bath for ten minutes. If you still smell the acid or alcohol, raise the temperature to 80°C and heat 10 more minutes.
5. Carefully (keep away from your face and waft the smell towards your nose) smell your test tubes and report what you smelled and any other observations. Dispose in the container in the hood.

Record your data in the following table and draw the Lewis dot structures for each acid and alcohol. Many of them are in your text. You can find the others on the Internet or from your instructor. Then complete the subsequent table with names and structures of the esters that you made.

Table of mixing

Tube #	Alcohol (structure)	Acid (structure)	Odor of ester
1	methyl alcohol	salicylic acid (2-hydroxybenzoic acid)	
2	ethyl alcohol	butanoic acid	
3	pentyl alcohol	acetic acid	
4	isoamyl alcohol (3-methyl-1-butanol)	acetic acid	
5	pentyl alcohol	propanoic acid	
6	benzyl alcohol	acetic acid	
7	isobutyl alcohol (2-methyl-1-propanol)	formic acid	
8	ethanol	acetic acid	

Complete the table

Tube #	Name of ester	Structure of ester
1		
2		
3		
4		
5		
6		
7		
8		

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Lab 14

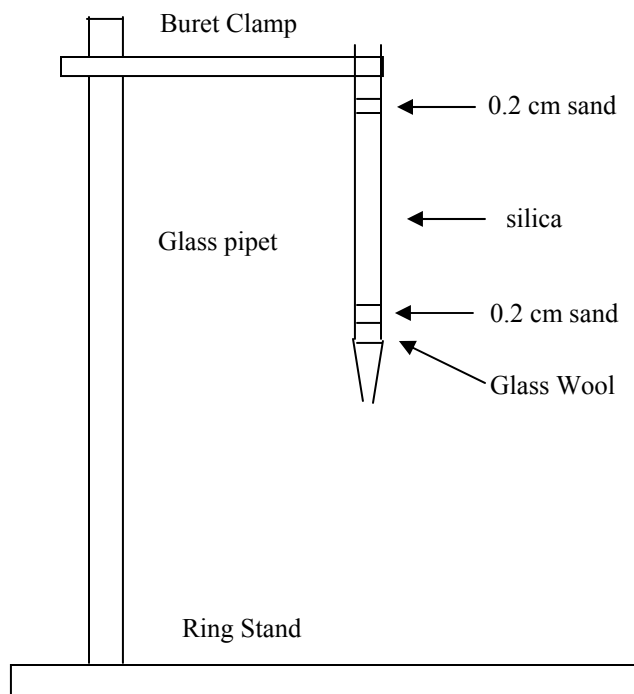
Chromatography Of Plant Pigment Extracts

Extraction of plant pigments

1. Obtain some chlorophyll containing leaves or grass. Put them in your mortar and add 50 mL of a 4:1 mixture of ligroin: acetone. Grind until your liquid is **dark** green, the darker the better. You may need to get some more leaves. If your liquid evaporates, add a little more.
2. Decant the solution and then either centrifuge or filter to remove all of the remaining solids.
3. Dispose of the grass contain solvent in the waste beaker. All ligroin and acetone should be put in the non-halogenated waste container in the hood.

Separation by chromatography

1. Assemble the apparatus pictured.
2. Soak the glass wool in ligroin before putting in the pipette.
3. Use small spatulas to place the sand and silica in the pipette. **DO NOT BREATHE THIS DUST!!**
4. Fill the pipette with ligroin and note the drip rate. About 1 drop per second is good. More than three times faster or slower means you need to reassemble your apparatus.
5. Fill your pipette with ligroin and when it falls to the top layer of sand, quickly fill the pipette with your plant extract.
6. Keep filling your pipette with ligroin, never allowing it to fall below the top layer of sand.
7. Pigments should begin to separate. One remains virtually stationary and one moves down the pipette. When those pigments are about to exit the pipette, place a clean container under the pipette and collect this fraction.
8. Now start refilling the pipette with acetone in place of ligroin.
9. The other pigments should begin to move down the pipette. When they reach the bottom, collect them in a clean container.
10. Place each of the pigments in the UV box and observe. Write down your observations. **DO NO LOOK DIRECTLY AT UV LIGHT!!!**



Questions

1. a. In the chromatography procedure what was the mobile phase(s)?

b. What was the stationary phase?
2. What color pigments are attracted to the ligroin?
3. What color pigments are attracted to the acetone?
4. What do you think would have happened if you had added acetone to the pipette first and then ligroin?

5. Acetone mixes easily with water. It can be considered hydrophilic. Ligroin does not mix with water and it can be considered hydrophobic. Which color pigments do you think are more hydrophobic, the greens or the yellows? Explain your reasoning.

6. For the green pigment, what colors of the spectrum do you think are absorbed? What colors are reflected.
7. For the yellow pigment, what colors of the spectrum do you think are absorbed? What colors are reflected.

8. Which pigment, yellow or green, is the cause of the red fluorescence?

The spectrum of the pigments (optional)

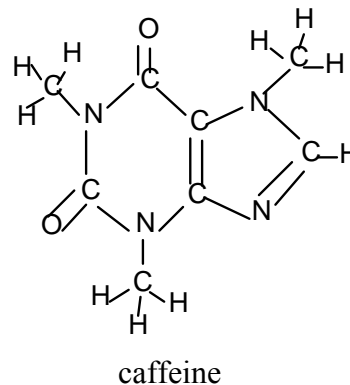
Your instructor will show you how to use the spectrophotometers.

9. Determine the absorption spectrum of the yellow pigment to come out of the pipette from 400 to 700 nm by measuring the absorption of light every 10 nm. Use ligroin as your blank.
10. Determine the absorption spectrum of the green pigment to come out of the pipette from 400 to 700 nm by measuring the absorption of light every 10 nm. Use acetone as your blank.
11. Graph all of your spectral data on the graph paper provided. Put wavelength on the x axis and absorption on the y axis

Lab 15

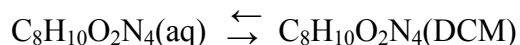
Extraction and Identification of Caffeine

Caffeine is a naturally occurring alkaloid present in many plants. It is in the purine family of chemical compounds, the same purine family as adenine and guanine found in nucleic acids (DNA and RNA). Its chemical formula is $C_8H_{10}O_2N_4$ and the structure is diagramed at the right. Caffeine is found in coffee, tea, and chocolate and is added to many carbonated beverages. Caffeine is more soluble in hydrophobic liquids than in water. You will extract the caffeine from a beverage using hydrophobic dichloromethane. After crystallization, you will identify the crystals as caffeine using melting point analysis.



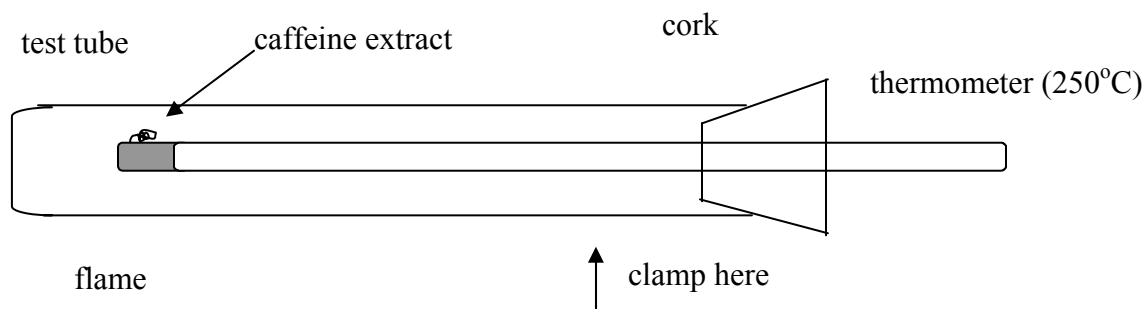
Procedure

1. Into a separatory funnel, add 50.00 mL of caffeine containing beverage, 1.00 mL of 20% aqueous solution of sodium carbonate, and 6.00 mL of dichloromethane (DCM) sometimes referred to methylene chloride. Swirl for 5 min. using the technique demonstrated in class.
2. Transfer the lower, dichloromethane layer to your smallest beaker. Be sure this beaker is dry and has been weighed. Record the mass of the empty beaker. Add another 6.00 mL of DCM to the separatory funnel and swirl for 5 minutes, and transfer the lower layer to same beaker. If you have any water floating on the top, show your instructor. During this extraction the following equilibrium is established



With the position of equilibrium favoring the caffeine dissolved in DCM. By removing the product ($C_8H_{10}O_2N_4(DCM)$), and then adding more DCM the equilibrium shifts to the right according to Le Chatlier and virtually all of the caffeine is removed from the aqueous phase with the second extraction.

2. UNDER THE HOOD. Place the beaker with DCM and caffeine on a hot plate. On low to medium low heat, evaporate the DCM with the hood sash closed. Watch closely and remove the beaker from the hot plate when the liquid level is as thick as a penny. DO NOT OVER HEAT. The remaining DCM will evaporate off in a few minutes. When all the DCM is gone and the caffeine has crystallized, weigh the beaker and contents on the same balance that you used previously. Record the mass of the beaker and caffeine.
3. Look at you caffeine crystals using a hand lens and describe and sketch them.
4. Determine the melting point by assembling the apparatus in the diagram below. Heat the apparatus using a Bunsen burner and observe when the solid melts. Quickly note the temperature and record it.



Chem 11 Lab 15 Caffeine Data: Name _____

What is the melting point of pure caffeine? How close and pure is your extract?

1. type of beverage _____

2. mass of dry empty beaker _____

3. mass of beaker with caffeine extract _____

4. mass of caffeine extract
subtract mass of data 2 from data 3 _____

5. percent caffeine (weight/volume)
divide grams caffeine by mL of beverage _____

6. Go on the internet and find the range of caffeine
in different beverages. Record this range of
percent caffeine for your beverage _____

7. Experimental melting point _____

8. Go on the internet and find the value
of caffeine's melting point _____

Sketch of crystals

Determination of the Identity of Eight Unknown Solids

The purpose of this experiment is to determine the identity of eight unknown solids. The unknowns are sucrose, sodium bicarbonate, starch, sodium chloride, salicylic acid, ascorbic acid, calcium carbonate and calcium sulfate.

The first part of the experiment determines some chemical and physical properties of the eight substances. These two properties will differentiate between many of the compounds. The physical properties to be determined are the state and color of the substance and its solubility in water. The chemical properties are the acid/base properties (using red cabbage as the acid/base indicator), the reaction with iodine, I_2 , the reaction with hydrochloric acid, HCl, and the reaction with silver nitrate, $AgNO_3$. You will test the 8 substances and observe their properties and reactions. These tests provide you with necessary information to identify these substances even when they are given to you as unknown materials. When chemists use known materials this way, they call them standards.

Procedure: .Knowns (standards):

Solubility: To test for water solubility, place a small quantity (about the size of $\frac{1}{2}$ a pea) of each of eight known substances into eight clean labeled test tubes. Be sure to record the test tube number and sample name. Add approximately 2 ml of deionized water to each sample in the test tube and stir or shake thoroughly. Be sure that you know what substance is in each test tube. Add deionized water to each sample stir thoroughly. Is the substance soluble in water? It is crucial that you do not use too much of the sample for the solubility test as too much sample will look as if it is insoluble when actually it is soluble. Record the results in the chart below. About half are not soluble.

Chemical Tests: For each chemical test use a portion of your original sample of knowns that you dissolved for the solubility test and put some of the sample in the micro test wells. Observe carefully note such things as bubbles, cloudiness (this means a solid has formed), colors, time for changes that occurred, etc. For the substances that did not dissolve still use the solution but make sure that you include some of the solid in your sample. Use your pipette to make the transfers from test tube to micro test wells but be sure that you rinse the pipette several times with distilled water between samples. Observe very carefully any results you obtained and record them on the chart below

pH test: Add about $\frac{1}{2}$ mL of cabbage juice indicator to about $\frac{1}{2}$ mL of each sample. Record results.

I_2 test: Add 2 drops of iodine solution to small portion of each sample. Reaction with Ag^+ ion: Add a few drops of silver nitrate to small portion of each sample..

Reaction with HCl: Add HCl to the rest of your sample. Note any bubbles that form or if the substance dissolves.

Unknowns: Repeat the above procedure using the unknowns. Record your results on the chart below. Compare the data and identify the unknowns.

Data Table

Knowns	Solubility	pH	Iodine	AgNO ₃	HCl
Sucrose (table sugar)					
Sodium bicarbonate (baking soda)					
Starch					
Sodium chloride (table salt)					
Salicylic acid					
Ascorbic acid (vitamin C)					
Calcium carbonate (Tums, limestone)					
Calcium sulfate (gypsum, sheetrock)					
Unknowns					
1					
2					
3					
4					
5					
6					
7					
8					

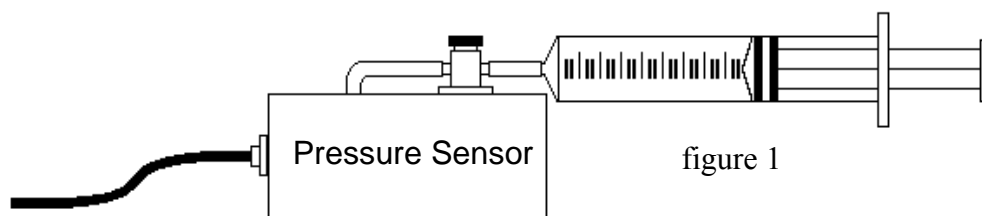
Lab 17: Pressure, Volume and Temperature Relationships in Gases

In this lab, you will experimentally observe the effects of changing volume and temperature on the pressure of gases. Remember, gas atoms and molecules are in constant motion like a mosh pit. The higher the temperature, the faster they move. With all this motion, there are many collisions. The collisions of gas atoms and molecules with another object cause pressure. More collisions lead to higher pressure.

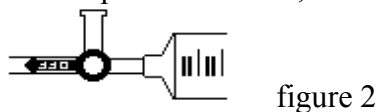
PROCEDURE:

Part I : Pressure and Volume Relationship at Constant Temperature

1. Get a Lab Pro and connect the power transformer to it and then connect the pressure sensor to channel 1 and the stainless steel temperature probe to channel 2.
2. Connect the Lab Pro to the cable coming out of rear of the computer Plug in power transformer and turn on computer.
3. Open **Logger Pro 2.2.1 program** A dialogue box should open. Hit scan. Should show Lab Pro - USB Hit OK.
4. Select FILE, OPEN, Chemistry with Computers, EXP 06 Boyle's Law.
5. Select Exp06 Pressure Sensor. Do NOT select Gas-Pressure sensor!!
6. Select Setup, Sensors, Port 1, Pressure-Pressure sensor. In calibration, change Units to Atm.
7. Select DATA, Column Options, Pressure and set decimal places to 2.
8. At the top of the y axis click on the number (250) and enter 3 and hit enter.
9. Connect the syringe to the pressure sensor as shown in figure 1



10. To begin the experiment, click on Collect, when ready to collect data points, click on keep, enter volume setting, and press enter.
11. When finished collecting data, click on Stop.
12. For your first data point, open the valve's side arm to allow air to enter and exit. Align the blue handle with the arm that leads to the pressure sensor, as shown figure 2



Move the piston until the front edge of the black ring is at the 10.0 mL mark. Then close the side arm of the pressure sensor valve by aligning the blue handle with the side arm, as shown in figure 3.

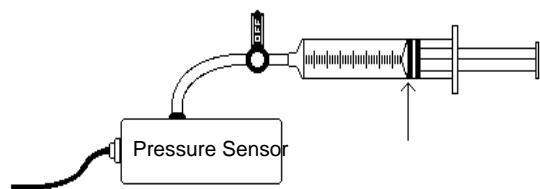


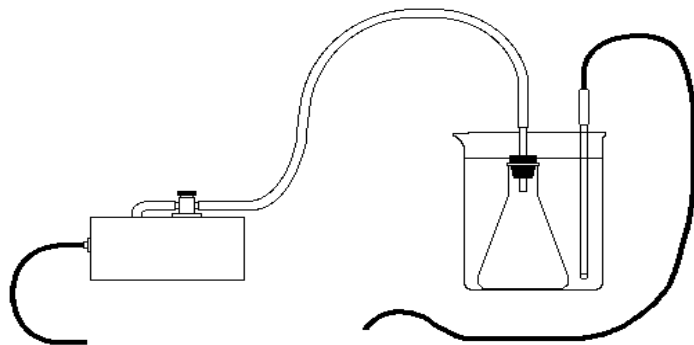
figure 3

Click on keep, type in 10.0 mL and press enter.

13. Be sure to also record your data on the chart in the Lab Report section.
14. For your second data point, move the piston until the front edge of the inside black ring (see Figure 3) is at the 5.0-mL line on the syringe. Hold the piston in this position until the pressure value stabilizes. When the pressure reading has stabilized, click on keep, type in 5.0 mL and press enter.
15. Now move the piston to the 7.5 mL line, and collect another data pair. Again enter / record the pressure and volume values.
16. Repeat this procedure for volumes of 12.5, 15.0, 17.5, and 20.0 mL.
17. Hit print to print the graph, one copy for each person.

Part II : Pressure-Temperature Relationship at Constant Volume

1. Select FILE, OPEN, Chemistry with Computers, EXP 07.
2. Select Exp 07, Pressure-Stnlss. You need not save exp 6 once it is printed
3. Select Setup, Sensors, Click on the ch 1 icon, Pressure-Pressure sensor. Change Units to Atm. Hit OK
4. Select DATA, Column Options, Pressure and set decimal places to 2.
5. At the top of the y axis click on the number (250) and enter 3 and hit enter.
6. On the temperature axis, click on 100 and enter 105 and hit enter. Click on 0 and enter -5 and hit enter.
7. Use large beakers on a hot plate to prepare 3 water baths at approximately 30°C, 60°C and 90°C. You can save time by using the hot tap water which is about 60°C. Also prepare an ice bath at 0°C. **CAUTION:** *Do not burn yourself or the probe wires with the hot water or hot plate.*
8. Connect a piece of plastic tubing to the end opening of the 3-way valve of the pressure sensor. align the blue handle with the arm that leads to the pressure sensor, then open the side arm of the pressure sensor valve to allow air to enter and exit.
9. Insert a 1-hole stopper fitted with a glass tube into a 125-mL Erlenmeyer flask. Twist the stopper to ensure a tight fit. Firmly attach the plastic tubing from the pressure to the glass tube in the stopper. Make sure there are no leaks!
10. Close the side arm of the pressure sensor valve. Align the blue handle with the side arm. The air sample to be studied is now confined in the flask.
11. Submerge the flask and the temperature probe in the ~ 0°C water bath shown in figure 4 figure 4



12. To begin the experiment, click on Collect,
13. Monitor the pressure and temperature readings. When the temperature and pressure readings have both stabilized, record the data and click on Keep

14. Be sure to also record your data on the chart in the Lab Report section.
15. Repeat this procedure using the 30°C, 60°C and 90°C water baths.
16. When finished collecting data, click on Stop. Save your data to a disk for printing and print a copy of the graphs for both experiments for each member of your team.

LAB 17: P,V, and T Relationships LAB REPORT

Lab period _____ Name _____

Partners: _____

Part I :

1. Fill in the table below with your experimental values (including units) for pressure and volume.

Volume	Pressure	P x V	P/V

2. In gas property relationships, you have seen directly proportional and inversely proportional.

Conceptually directly proportional means if V goes up , then P goes up by the same proportion. Mathematically, the data would follow the equation $P = kV$, where k is a proportionality constant. Rearranging this equation algebraically, $P/V = k$ with k being a constant. So if your column P/V is constant or nearly so (depending on how good a chemist you are), then your data indicate a direct relationship. If PV is not constant, then not a direct relationship is indicated.

Are your P/V values approximately constant? Yes or No Circle one!
Then your P/V data indicate what type of relationship, Direct or Not Direct? Circle one!

Conceptually inversely proportional means if V goes up , then P goes down by the same proportion. Mathematically, the data would follow the equation $P = k/V$, where k is another proportionality constant. Rearranging this equation algebraically, $PV = k$ with k being a constant. So if your column PV is constant or nearly so (depending on how good a chemist you are) then your data indicate an inverse relationship. If PV is not constant, then not an inverse relationship is indicated.

Are your PV values approximately constant? Yes or No Circle one!
Then your PV data indicate what type of relationship, Inverse or Not Inverse? Circle one!

From your results, which relationship is more likely, direct or inverse? _____

3. Attach a graph of your pressure-volume data.

Lab 17:

4. Conclusion: Pressure and Volume are _____ proportional when _____ is constant.

5. Use P, V and k to write a mathematical equation showing their relationship. _____

6. Complete the table:

If the Volume ...	Then the Pressure ...
Increases by a factor of 2	_____ by a factor of _____
Decreases by a factor of 2	_____ by a factor of _____
Increases by a factor of 3	_____ by a factor of _____
Decreases by a factor of 3	_____ by a factor of _____

Part II :

1. Fill in the table below with your experimental values (with units) for pressure and temperature. Remember temperature needs to be converted to Kelvins which is Celsius + 273

Pressure (atm)	Temperature (°C)	Temperature (K)	P x T(Kelvins)	P/T(Kelvins)

2. Is PT or P/T approximately constant?

Is pressure and temperature directly or inversely proportional? _____. Precisely and concisely explain how you reached you conclusion.

3. Attach a graph of your pressure-temperature data.

4. Conclusion: Pressure and temperature are _____ proportional when the temperature is measured in _____ .and when _____ is constant.

5. Use P, T and k to write a mathematical equation showing there relationship. _____

6. Explain the relationship between pressure and temperature **using the concepts of collisions of molecules and changing volumes**.

7. Complete the table:

If the Temperature ...	Then the Pressure ...
Increases by a factor of 2	_____ by a factor of _____
Decreases by a factor of 2	_____ by a factor of _____
Increases by a factor of 3	_____ by a factor of _____
Decreases by a factor of 3	_____ by a factor of _____