

CHEMISTRY 410

LABORATORY MANUAL



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Chemistry 410: "Chemistry for Health Sciences" Laboratory Manual

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Course Safety Information

Laboratory Course Safety Information

Safety is our primary concern in the chemistry laboratory. We will be working with fire, toxic chemicals, glassware, and other hazards. It is your responsibility to work in a way that keeps you and your classmates safe. Be aware of laboratory dangers and work safely by following these guidelines.

Laboratory Guidelines on Personal Conduct and Safety:

1. ALWAYS wear safety glasses or goggles in the laboratory.
2. Come to lab prepared to perform the experiment. Complete required pre-lab assignments.
3. Arrive for lab on time, so that you may hear important safety information.
4. Do not enter the laboratory until your instructor arrives and allows you to enter.
5. Never work alone in the laboratory.
6. Wear sensible clothing in the lab. This includes fully closed shoes, long pants or skirts, and fitted clothing. Loose, flowing clothing and nylon or baggy jackets are a hazard and should not be worn. Shorts and short skirts do not adequately protect your legs from splashes or spills and may not be worn. Lab coats or aprons are a good idea to protect your clothing. Protect (and cover) your skin. Tie back long hair.
7. Keep the laboratory clean and avoid clutter. Place all things brought with you to lab (coats, bags, etc.) in the cubbyholes or under your lab bench.
8. Notify the instructor immediately of all spills. The instructor will tell you how it may be safely cleaned up.
9. Never leave an experiment unattended.
10. Never perform unauthorized experiments. Any changes in procedures must be approved by your instructor.
11. Notify the instructor immediately of all injuries or burns, no matter how minor.
12. No eating, drinking, smoking, or chewing gum is allowed in the laboratory. Tasting chemicals or mouth-pipetting is also not allowed.
13. Speak in normal conversational tones in the laboratory.
14. No horseplay and no running.
15. At the completion of your lab work, wash your bench top down, be sure gas jets are off, remove all solid matter from your sink, and put all equipment away.
16. Some chemicals and latex gloves can cause allergic reactions. Always be aware of this possibility and avoid contact with chemicals or inhalation of their vapors whenever possible.
17. Notify the lab instructor of any of the following:
 - You wear contact lenses.
 - You have a medical condition or concern that the instructor should be aware of.
 - You need special accommodations to work in the lab.

Laboratory Hazards and Proper Use of Equipment:

1. Fire in the laboratory

1. Gas burners are a constant burn hazard. Use caution around them. Never leave a lighted burner unattended. Always turn off gas jets when finished and check before leaving the lab so that gas fumes do not build up in the lab.
2. If your clothes catch fire, drop and roll on the floor to smother the flames. If you see someone with clothes on fire, use the fire blanket to smother the flames. The emergency shower may be used, if you are close enough. DO NOT use the emergency shower if clothes that are burning are doused with organic solvents. Notify the instructor immediately to arrange for additional first aid.
3. Small burns should be put under cold running water. Notify your instructor and ask for a bandage. Larger burns should be immediately brought to the instructor's attention, so that additional medical assistance may be obtained.
4. If you have a very small fire in a small container, smother it with a towel, watch glass or book (but not your hand).
5. If you have a waste paper basket - sized or sink fire, use one of the fire extinguishers in the lab.

6. In case of a large fire involving the lab itself, the room and building should be evacuated according to the following procedure:

- Make others aware of the problem. Yell fire, pull the fire alarm lever to alert others.
- Use common sense, if you can turn things off (particularly gas jets and electrical equipment) on the way out, do so.
- Leave the room and get out of the building. Assemble in front of Building 7 so that your instructor may take roll to make sure everyone is out of the building.

2. Glassware

1. Soft glass containers, such as bottles, funnels, and graduated cylinders should never be heated directly with a flame. They may shatter.
2. Check beakers, flasks and test tubes (which can be heated) prior to heating for chips and scratches. Chipped and scratched glassware can easily break. It should not be heated.
3. Hot glass stays hot for long time and you can't see the difference. Place hot glassware on an insulated metal screen (to warn others that it's hot). Let it cool for a considerable time before picking it up, and then, do so cautiously.
4. If you break glass, never touch the broken pieces. Notify the instructor and sweep up the pieces using the broom and dust pan. Use tongs or tweezers to pull glass out of the sink. Dispose of the glass in the designated container, not the trash can.

3. Handling Chemicals

1. Read labels carefully. Be certain that you are using the correct substance.
2. Use only what is needed. Never return chemicals to their original containers. Share the extra with another student or ask the instructor how to dispose of it.
3. Leave chemicals in their proper places. Do not take the reagents to your workspace. Hold the bottle stopper in your hand while you pour (don't put the stopper onto the bench). Replace tops on bottles securely, and clean up any spills that may have occurred.
4. Label all chemicals that you take to your workspace. Grease pencils and labeling tape are available.
5. When mixing liquid chemicals with water, always add the concentrated chemicals to the water. Slowly add the chemical while mixing. Heat is generated when acid is added to water. Slowly mix in acid to water to avoid the solution boiling out of the container.
6. Use caution when heating liquids. Small quantities of liquid in test tubes can be heated safely in a beaker of boiling water. If you heat a test tube directly in a flame, you should:
 - use a test tube holder,
 - point the mouth of the test tube away from people,
 - move the test tube in and out of the flame slowly, heating the side of the test tube rather than the bottom to avoid "bumping" - the ejection of the liquid out of the tube.
7. Never heat flammable liquids with a flame. Use a hot plate or hot-water bath for this.
8. Exercise great care if you are directed to check for the odor of a substance. Use your hand to waft some of the vapor toward your nose. Never inhale deeply.
9. Chemical spills on the skin or clothing must be considered serious. If acids, bases (alkalines), or other corrosive materials come into contact with the skin, wash them off immediately with liberal amounts of water. Seek assistance from the lab instructor. If anything splashes in your eyes, use the eyewash at once and tell the lab instructor.
10. Carry chemicals carefully. Use two hands, one under the bottle, one on the neck or handle. Alert people that you are bringing the chemical through a crowded area.
11. When pouring liquid chemicals, pour from a smaller mouth to a bigger mouth vessel or use a funnel. Never return chemicals to a bottle, you may contaminate the entire amount.
12. If you spill liquid chemicals, you must clean up the spills. Follow your instructor's directions.
13. Always return bottle tops to the proper bottle and secure them firmly. Never pick up a bottle by its cap as a previous user may have put the cap on loosely.
14. Whenever toxic gases are evolved in an experiment, the work must be carried out in one of the fume hoods. Such gases include oxides of nitrogen and sulfur (NO_x, SO_x), the halogens (Cl₂, Br₂, I₂) and hydrogen sulfide (H₂S).

15. Mercury and mercury spills (from broken thermometers, etc.) must be thoroughly cleaned up immediately. Notify your instructor. They will assist you with proper procedures. Never put broken thermometers containing mercury in any of the disposal containers in the lab.
16. Organic liquids must be handled with care. Some organic vapors are respiratory irritants, others are poison. Some of the liquids can be absorbed through the skin. Many organic liquids are flammable. Ask before disposing of organic compounds. Many are disposed on in a container labeled ORGANIC WASTE found in one of the ventilation hoods.

Disposal of Chemicals and Other Materials

1. Waste solids, such as paper towels, matches, filter papers and litmus paper should be discarded in the paper waste cans. Matches should be wet prior to dumping.
2. Your instructor will direct you in disposal of chemicals used in the experiments. Labeled waste containers will be provided for disposal of chemicals that may not be put down the drain or in the paper waste cans. Do not dump chemicals down the drain or into the waste cans unless instructed to.
3. Acids and bases must be neutralized before they are flushed down the drain, with lots of water.
4. Broken or disposable glassware should be placed in the container labeled for glass waste.

Using Electrical Equipment in the Laboratory

Electricity and liquids are a dangerous combination! Do not allow water, aqueous solutions, or other liquids to come into contact with electrical cords or electrical conductors. If your electrical equipment crackles, snaps, or begins to give off smoke, do not attempt to disconnect it! Call your instructor immediately.

Ventilation

Our laboratories are ventilated through the four ventilation hoods (also called fume hoods) at the front of the room. The fan noise is easily noticed. Should these fans stop, the room should be evacuated. Turn off all gas burners, water faucets, and electrical equipment and leave the room. Your instructor will give further instructions. For certain experiments, you will be asked to work in one of the hoods. If any chemicals are being stored in the hood, be sure to ask your instructor if it is safe to work in that hood with the chemicals there.

Laboratory Safety Equipment

In order to work safely in the lab, be certain that you are aware of the location of all of the following items in the laboratory:

- Eyewash station
- Safety shower
- Fire blanket
- Chemical Stockroom
- First aid kit
- Ventilation hoods
- Instructor's desk

Licensing

A detailed breakdown of this resource's licensing can be found in [Back Matter/Detailed Licensing](#).

CHAPTER OVERVIEW

1: Measurement and Significant Figures

[1.1: Measurement and Significant Figures Lab Procedure](#)

[1.2: Measurement and Significant Figures Pre-laboratory Assignment](#)

[1.R: Measurement and Significant Figures \(Report\)](#)

Adapted from

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1.1: Measurement and Significant Figures Lab Procedure

Learning Objectives

- To use a metric ruler to measure the dimensions of regular geometric shapes, and to use these measurements to determine the areas of the shapes.
- To measure the volume of a sample of water using a graduated cylinder and a beaker in order to compare their precision.
- To measure the mass of an item using a triple-beam balance and an analytical (electronic) balance in order to compare their precision; also, to determine the mass of a powder by weighing by difference.

Background

Our knowledge of chemistry and chemical processes largely depends on our ability to obtain correct information about matter. Often this information is quantitative, in the form of measurements. In this lab, students will be introduced to some common measuring instruments so that they can practice making measurements, and to learn about instrument precision. In Part A of this lab, a metric ruler will be used to measure length in centimeters (cm). In Part B, a beaker and a graduated cylinder will be used to measure liquid volume in milliliters (mL). In Part C, an electronic balance will be used to measure mass in grams (g).

Since all measuring devices are subject to some error, it is impossible to make exact measurements. Scientists record all the digits of a measurement that are known exactly, plus the first one that is uncertain. These digits are collectively referred to as significant digits. Digital instruments, such as an electronic balance, are designed to limit themselves to the correct number of significant digits, and their readings are properly recorded as given. However, when using analog instruments such as rulers and thermometers, the experimentalist is responsible for determining the correct number of significant figures. These instruments are properly read to one place beyond the graduations of the scale.

The system of measurement used in science and the healthcare field is the metric system. Within the metric system, a variety of units can be employed to measure the same fundamental quantity; for example, energy could be expressed within the metric system in units of ergs, electron-volts, joules, and two kinds of calories. Therefore, a more basic set of units the Systeme Internationale (SI) units that are now recognized as the standard for science and for technology of all kinds. In principle, any physical quantity can be expressed in terms of base units (Table 1), with each base unit defined by a standard described in the NIST Web site.

Table 1: SI Units

Property	Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Volume	liter	L
Time	second	s
Temperature (absolute)	kelvin	k
Amount of substance	mole	mol

Owing to the wide range of values that quantities can have, it has long been the practice to employ prefixes such as milli and kilo to indicate decimal fractions and multiples of metric units. As part of the SI standard, this system has been extended and formalized (Table 2).

Table 2: Prefixes used to scale up or down base units

Prefix	Abbreviation	Multiplier
giga	G	10^9
mega	M	10^6
kilo	k	10^3
deci	d	10^{-1}
centi	c	10^{-2}
milli	m	10^{-3}
micro	μ	10^{-6}
nano	n	10^{-9}
pico	p	10^{-12}

Measured and Exact Numbers

Counting is the only type of measurement that is free from uncertainty, provided the number of objects being counted does not change while the counting process is underway. The result of such a counting measurement is an example of an **exact number**. If we count eggs in a carton, we know exactly how many eggs the carton contains. The numbers of defined quantities are also exact. By definition, 1 foot is exactly 12 inches, 1 inch is exactly 2.54 centimeters, and 1 gram is exactly 0.001 kilogram. Quantities derived from measurements other than counting, however, are uncertain to varying extents due to practical limitations of the measurement process used.

Significant Figures in Measurement

The numbers of measured quantities, unlike defined or directly counted quantities, are not exact. To measure the volume of liquid in a graduated cylinder, you should make a reading at the bottom of the meniscus, the lowest point on the curved surface of the liquid, Figure 1.

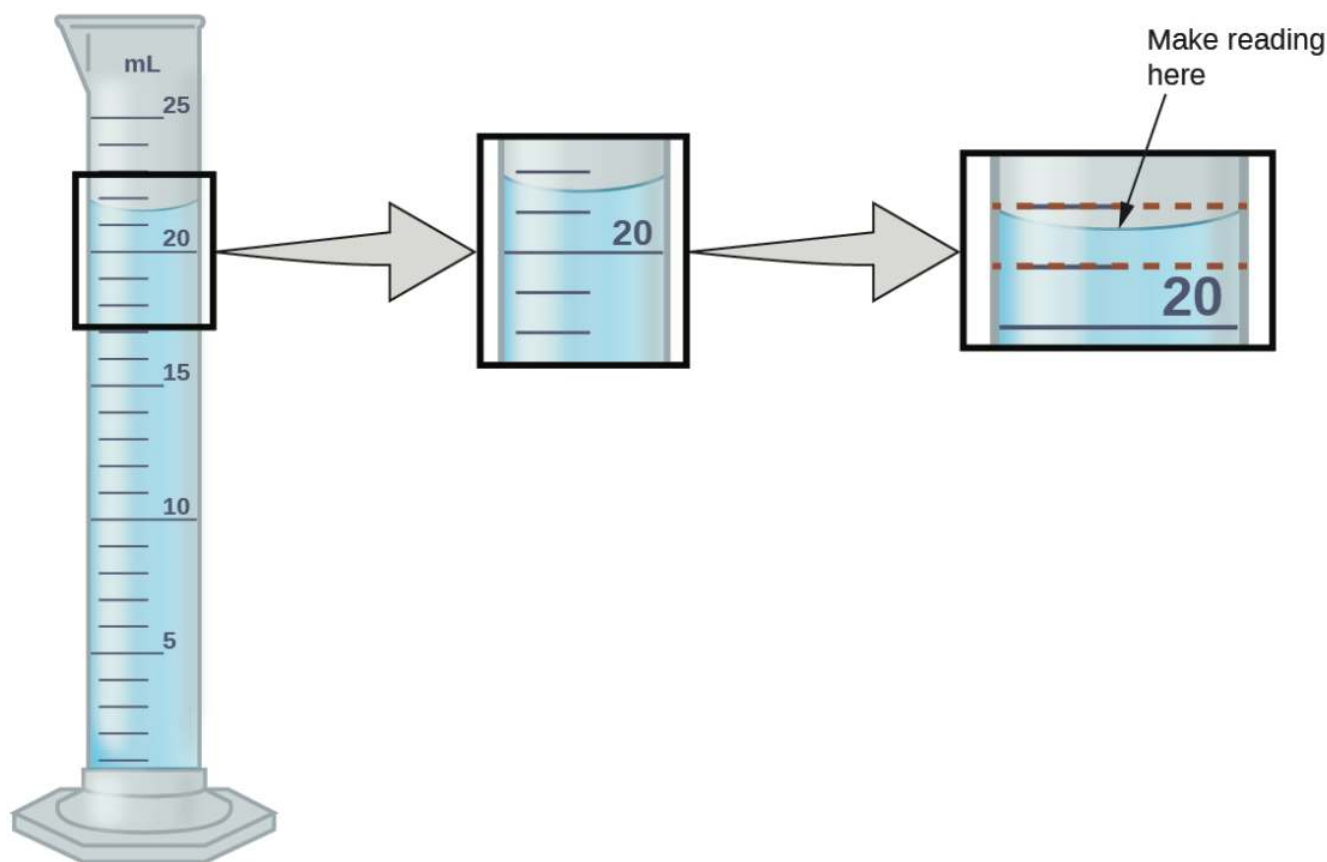
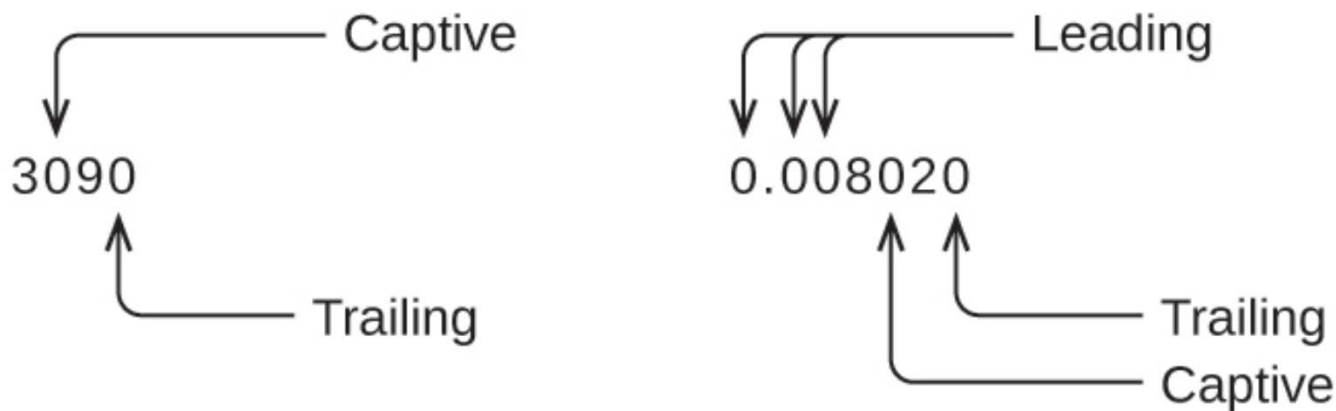


Figure 1.1.1: To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus.

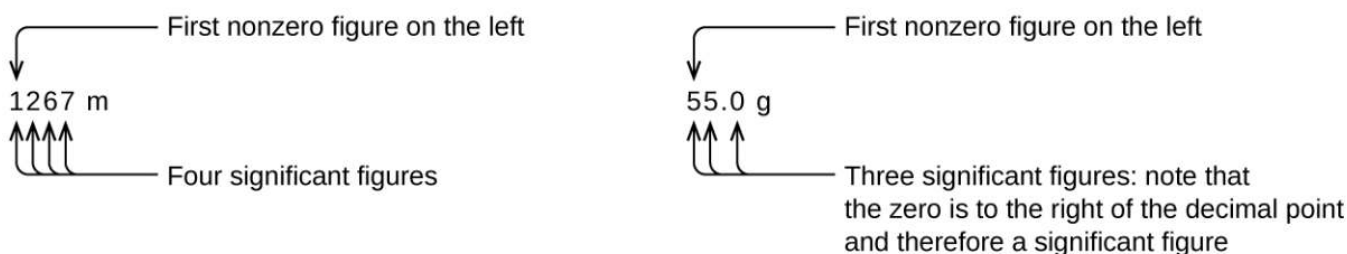
Refer to the illustration in Figure 1. The bottom of the meniscus in this case clearly lies between the 21 and 22 markings, meaning the liquid volume is certainly greater than 21 mL but less than 22 mL. The meniscus appears to be a bit closer to the 22-mL mark than to the 21-mL mark, and so a reasonable estimate of the liquid's volume would be 21.6 mL. In the number 21.6, then, the digits 2 and 1 are certain, but the 6 is an estimate. Some people might estimate the meniscus position to be equally distant from each of the markings and estimate the tenth-place digit as 5, while others may think it to be even closer to the 22-mL mark and estimate this digit to be 7. Note that it would be pointless to attempt to estimate a digit for the hundredths place, given that the tenths-place digit is uncertain. In general, numerical scales such as the one on this graduated cylinder will permit measurements to one-tenth of the smallest scale division. The scale in this case has 1-mL divisions, and so volumes may be measured to the nearest 0.1 mL.

This concept holds true for all measurements, even if you do not actively make an estimate. If you place a quarter on a standard electronic balance, you may obtain a reading of 6.72 g. The digits 6 and 7 are certain, and the 2 indicates that the mass of the quarter is likely between 6.71 and 6.73 grams. The quarter weighs about 6.72 grams, with a nominal uncertainty in the measurement of ± 0.01 gram. If we weigh the quarter on a more sensitive balance, we may find that its mass is 6.723 g. This means its mass lies between 6.722 and 6.724 grams, an uncertainty of 0.001 gram. Every measurement has some uncertainty, which depends on the device used (and the user's ability). All of the digits in a measurement, including the uncertain last digit, are called significant figures or significant digits. Note that zero may be a measured value; for example, if you stand on a scale that shows weight to the nearest pound and it shows "120," then the 1 (hundreds), 2 (tens) and 0 (ones) are all significant (measured) values.

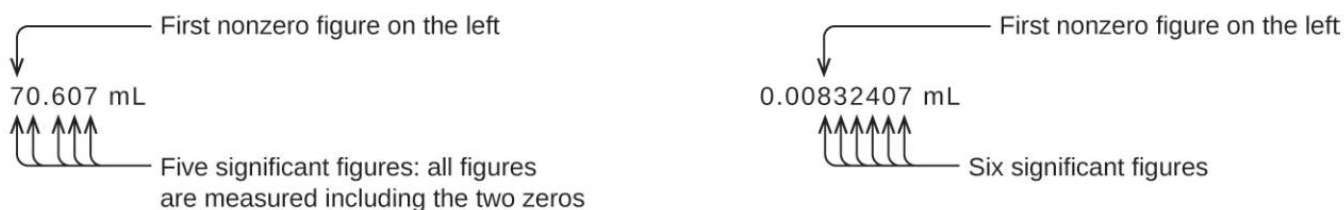
Whenever you make a measurement properly, all the digits in the result are significant. But what if you were analyzing a reported value and trying to determine what is significant and what is not? Well, for starters, all nonzero digits are significant, and it is only zeros that require some thought. We will use the terms "leading," "trailing," and "captive" for the zeros and will consider how to deal with them.



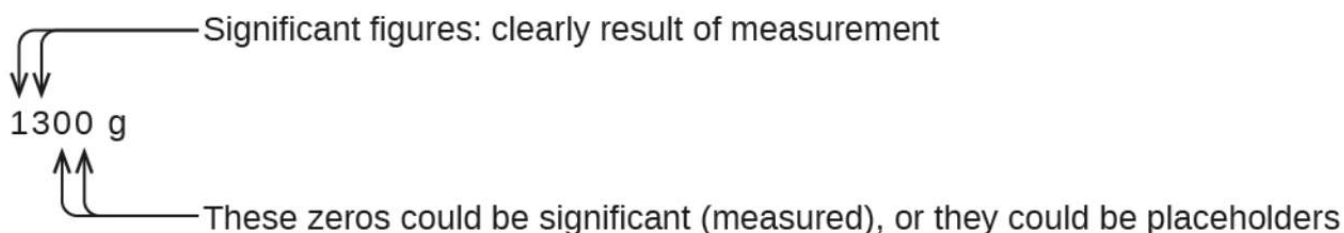
Starting with the first nonzero digit on the left, count this digit and all remaining digits to the right. This is the number of significant figures in the measurement unless the last digit is a trailing zero lying to the left of the decimal point



Captive zeros result from measurement and are therefore always significant. Leading zeros, however, are never significant—they merely tell us where the decimal point is located.



The leading zeros in this example are not significant. We could use exponential notation (as described in Appendix B) and express the number as 8.32407×10^{-3} ; then the number 8.32407 contains all of the significant figures, and 10^{-3} locates the decimal point. The number of significant figures is uncertain in a number that ends with a zero to the left of the decimal point location. The zeros in the measurement 1,300 grams could be significant or they could simply indicate where the decimal point is located. The ambiguity can be resolved with the use of exponential notation: 1.3×10^3 (two significant figures), 1.30×10^3 (three significant figures, if the tens place was measured), or 1.300×10^3 (four significant figures, if the ones place was also measured). In cases where only the decimal-formatted number is available, it is prudent to assume that all trailing zeros are not significant.



When determining significant figures, be sure to pay attention to reported values and think about the measurement and significant figures in terms of what is reasonable or likely when evaluating whether the value makes sense. For example, the official January 2014 census reported the resident population of the US as 317,297,725. Do you think the US population was correctly determined to the reported nine significant figures, that is, to the exact number of people? People are constantly being born, dying, or moving into or out of the country, and assumptions are made to account for the large number of people who are not actually counted. Because of these uncertainties, it might be more reasonable to expect that we know the population to within perhaps a million or so, in which case the population should be reported as 3.17×10^8 people.

Accuracy and Precision

Scientists typically make repeated measurements of a quantity to ensure the quality of their findings and to know both the precision and the accuracy of their results. Measurements are said to be precise if they yield very similar results when repeated in the same manner. A measurement is considered accurate if it yields a result that is very close to the true or accepted value. Precise values agree with each other; accurate values agree with a true value. These characterizations can be extended to other contexts, such as the results of an archery competition (Figure 2).

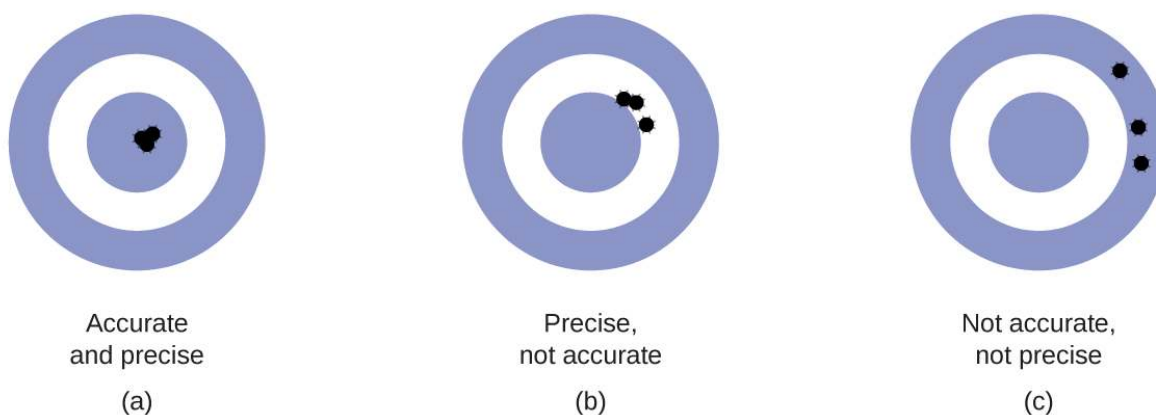
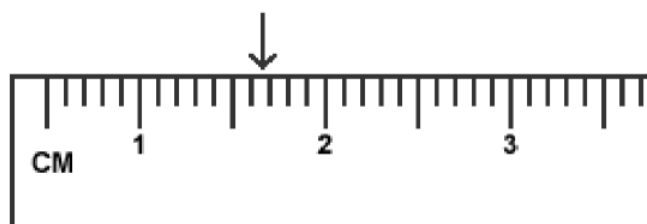


Figure 1.1.2: (a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.

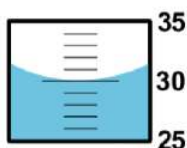
Part A: Measuring Length



The ruler markings are every 0.1-centimeter. The correct reading is 1.67 cm. The first 2 digits 1.67 are known exactly. The last digit 1.67 is uncertain. You may have instead estimated it as 1.68 cm.

Part B: Measuring Volume of a Liquid

When measuring liquid volumes, the graduated scale must be read from the lowest point of the curved surface of the liquid – the liquid meniscus.



The graduated cylinder markings are every 1-milliliter. The correct reading is 30.0 mL. The first 2 digits 30.0 are known exactly. The last digit 30.0 is uncertain. Even though it is a zero, it is significant and must be recorded.

Part C: Measuring Mass

The mass of an object is the amount of matter present in the object, whereas weight is the force of gravity on an object. We must use mass for accurate measurements of how much matter we have. We still call the process of obtaining an accurate mass "weighing". Accurate weighing is usually done with a single pan balance. The accuracy of a measured value, such as a melting point, may be evaluated by a calculation of percent error. Percent error is a common way of reporting how close a measured experimental value (EV) is to the true value (TV):

$$\text{Percent Error} = \frac{|EV - TV|}{TV} \times 100 \quad (1.1.1)$$

Accurate measurements will typically have low percent errors of <5%.

Experimental Procedure

Materials and Equipment

Metric ruler*, shape sheet, electronic balance, large test tube, 100-mL beaker, 100-mL graduated cylinder, triple-beam balance, 250-mL Erlenmeyer flask, electronic balance, sugar, small watch glass.

Safety

Always wear your safety goggles!

Part A: Measuring the Dimensions of Regular Geometric Shapes

1. Obtain a ruler from the instructor's desk.
2. Obtain a "three shape object" from the fume hood, and then use the ruler to measure the dimensions in centimeter to three decimal place. Measure the length and width of the rectangle, and the diameter of the circle. Record these measurements on your report form.
3. When finished, return the ruler.
4. Use your measurements to calculate the areas of the assigned geometrical shapes.

$$\begin{aligned} \text{Area of a rectangle} &= \text{length} \times \text{width} \\ \text{Area of a circle} &= \pi r^2 \\ (\text{where } r = \text{radius} &= 1/2 \times \text{diameter}) \end{aligned} \quad (1.1.2)$$

Part B: Measuring the Volume of a Sample of Water

1. Obtain a large test tube. Fill this test-tube to the brim with tap water, then carefully transfer it to a 50-mL beaker (obtain from your bin). Note that if your 50-mL beaker has no scale markings on it, you will need to get a replacement from the stockroom or your instructor and swap it for one that does. Measure and record the volume of water in the beaker.
2. Again, fill the same test-tube to the brim with tap water, then carefully transfer it to a 50- mL or 100-mL graduated cylinder (obtain from your bin). Measure and record the volume of water in the graduated cylinder. Do these measured volumes have the same number of significant figures?

Part C: Measuring the Mass of Solids

Comparing the Precision of two types of Balances

1. Use a triple-beam balance to obtain the mass of a 250-mL Erlenmeyer flask (obtain from your bin). Now use an electronic balance to obtain the mass of the same Erlenmeyer flask. Do these measured masses have the same number of significant figures? Be sure to record your measured masses on your report form.

Weighing by Difference

1. Using the electronic balance again, obtain the mass of a 50-mL beaker. If you already used this same beaker in Part B, make sure that you carefully dry it before weighing it. Add two spoonfuls of salt to this beaker, using a scoopula. Do not do this over

the balance! Then obtain the new combined mass of both the beaker and the salt. Be sure to use the same electronic balance as before.

2. When finished, dispose of the salt used in the waste container under the fume hood. Use your two measurements to determine the mass of salt (only) weighed out

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1.2: Measurement and Significant Figures Pre-laboratory Assignment

In Part A of this lab, you will measure the dimensions (length, width, diameter) of several geometric shapes.

Questions

1. Using a ruler, you measure the length of a rectangle to be 12.75 cm and the width to be 3.64 cm. Calculate the area of this rectangle (show work), reporting your answer to the correct number of significant figures.
 1. What is the formula for the area of a circle?
2. In Part B of this lab, you will measure the volume of a sample of water, in milliliters (mL). What two measuring instruments will you use to measure the water volume? Consider the following two volume measurements: 57.7 mL and 57.68 mL. Which of these is the more precise measurement, and why?
3. In Part C of this lab, you will measure the mass of several different items, in grams (g). What are the two types of balances you will use to measure mass?
4. An empty beaker has a measured mass of 29.456 g. When some salt is added to the beaker, the combined mass is 36.176 grams. Calculate the mass of the salt only (show work), reporting your answer to the correct number of significant figures

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1.R: Measurement and Significant Figures (Report)

Part A: Measuring the Dimensions of Regular Geometric Shapes

Experimental Data

Shape	Dimensions	Measurement	Significant Figures
Circle	Diameter		
Rectangle	Length		
	Width		

Data Analysis

Perform the conversions indicated below. Show your work, and report your answers in scientific notation.

- Convert the measured rectangle length to mm.
- Convert the measured circle diameter to m.
- Calculate the areas of your rectangle and circle in cm^2 . Show your work, and report your answers to the correct number of significant figures.
 - Area of rectangle
 - Area of circle
- Convert the area of your circle to mm^2 . Show your work, and report your answer in scientific notation.

Part B: Measuring the Volume of a Sample of Water

Experimental Data

Measuring Device	Volume Measurement	Significant Figures
50 mL Beaker		
50 mL Graduated Cylinder		

Data Analysis

- Compare your volume measurements in the table above. Which instrument, the beaker or the graduated cylinder, provides the more precise measurement? Explain.
- Convert the volume of water obtained using the graduated cylinder to cm^3 . Show your work, and report your answer in scientific notation.

Part C: Measuring the Mass of Solids

Table 1: Mass of an Erlenmeyer Flask

Measuring Device	Mass Measurement	Significant Figures
Triple-Beam Balance		
Electronic Balance		

Table 2: Weighing by Difference

Measuring Device	Mass Measurement	Significant Figures
Mass of empty beaker		
Mass of beaker + sugar		

Data Analysis

1. Compare your mass measurements obtained for the Erlenmeyer flask in Table 1. Which balance, triple-beam or electronic, provides the more precise measurement? Explain.
2. Consider the data you obtained in Table 2.
 1. Calculate the mass of salt weighed out. Show your work.
3. Circle one: When performing the above calculation, significant figures / decimal places are the primary consideration.
4. Perform the conversions indicated below. Show your work, and report your answers in scientific notation.
 1. Convert the mass of the salt weighed out to mg.
 2. Convert the mass of the salt weighed out to kg.

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CHAPTER OVERVIEW

2: Density of Liquids and Solids

[2.1: Density of Liquids and Solids Lab Procedure](#)

[2.2: Density of Liquids and Solids Pre-lab Assignment](#)

[2.R: Density of Liquids and Solids \(Report\)](#)

Adapted from

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2.1: Density of Liquids and Solids Lab Procedure

Learning Objectives

- To determine the density of pure water
- To determine the density of aluminum (applying the technique of water displacement) and to use this value to determine the thickness of a piece of aluminum foil

Background

Density is defined as the mass per unit volume of a substance, and it is a physical property of matter. A physical property can be measured without changing the chemical identity of the substance. Since pure substances have unique density values, measuring the density of a substance can help identify that substance. Density is determined by dividing the mass of a substance by its volume:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \quad (2.1.1)$$

The units of density are commonly expressed as g/cm³ for solids, g/mL for liquids, and g/L for gases.

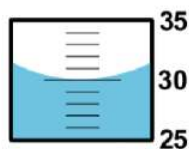
Density is also an intensive property of matter. This means that the value of density is independent of the quantity of matter present. For example, the density of a gold coin and a gold statue are the same, even though the gold statue consists of the greater quantity of gold. This is in contrast to extensive properties, like volume (the amount of space occupied by matter), which depend of the quantity of matter present. The more matter present, the larger the volume.

In Part A of this lab, the mass and volume of distilled water will be measured in order to determine the density of water. Measurements will be performed on three samples of water to improve precision and accuracy. Mass will be measured with an electronic balance, in grams (g), and volume will be measured directly with a graduated cylinder, in milliliters (mL). Recall that when measuring liquid volumes, the graduated scale must be read from the lowest point of the curved surface of the liquid (the meniscus).

The accuracy of the experimentally determined density of water will then be evaluated by comparison to the true, accepted density of water (see values posted in the laboratory).

Measuring the Volume of a Liquid

When measuring liquid volumes, the graduated scale must be read from the lowest point of the curved surface of the liquid – the liquid meniscus.



The graduated cylinder markings are every 1-milliliter. The correct reading is 30.0 mL. The first 2 digits 30.0 are known exactly. The last digit 30.0 is uncertain. Even though it is a zero, it is significant and must be recorded.

In Part B of this lab, the density of aluminum will be determined using aluminum pellets. Again, mass will be measured using an electronic balance, in grams (g). However, since the pellets have irregular shapes, their volume must be measured indirectly using the technique of water displacement (also known as Archimedes Principle). This is because the volume of water that the solid displaces when it is immersed in the water is the same as the volume of the solid itself. The accuracy of this experimentally determined density will also be evaluated by comparison to the true, accepted density of aluminum.

Measuring the Volume of an Irregularly Shaped Solid

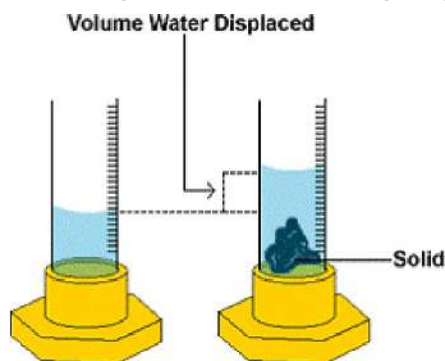


Figure 2.1.1: Volume water displaced = Final volume – Initial volume \ Volume water displaced = Volume of solid \ Note that 1 mL = 1 cm³

The density of aluminum will then be used in an applied problem to determine the thickness of a piece of aluminum foil. The piece of foil used can be considered to be a very flat rectangular box, where

$$\text{Volume of foil} = \text{length} \times \text{width} \times \text{thickness} \quad (2.1.2)$$

The foil volume can be obtained from the measured mass of the foil and the density of aluminum. Thus, if the length and width of the foil rectangle are measured, then the foil's thickness may be calculated.

Experimental Procedure

Materials and Equipment

50-mL graduated cylinder, aluminum pellets, small beaker, aluminum foil, thermometer, electronic balance, distilled water.

Safety

Be especially careful when adding the aluminum to your graduated cylinder, as the glass could break. Tilt the graduated cylinder and allow the pellets to gently slide to the bottom.

Part A: The Density of Water

Using the electronic balance, obtain the mass of your 50-mL or 100-mL graduated cylinder. Make sure it is dry before you weigh it.

1. Add 20-25 mL of deionized water to the graduated cylinder. Precisely measure this volume of water (to two decimal place). Then measure the combined mass using the electronic balance.
2. Repeat Step 1 to obtain a second and third set of mass and volume measurements. Use your thermometer to record the temperature of the water in your graduated cylinder.

Analysis: Subtract the mass of the empty cylinder from each combined mass measurement to obtain three mass measurements of water. Use the three sets of mass and volume measurements to calculate three density values for water. Then take the average of these three density values. Finally, look up the true density of water at the temperature used (posted on the wall in the front of lab), and evaluate the accuracy of your average density value by calculating your percent error.

Part B: The Density of Aluminum and the Thickness of Foil

The Density of Aluminum

1. Using the electronic balance, obtain the mass of a clean, dry small beaker.
2. Obtain a sample of aluminum from the fumehood. Transfer all the pellets to the beaker, and measure the mass of the beaker and pellets. Pour 30-35 mL of water into your 100-mL graduated cylinder. Precisely measure this volume.
3. Carefully add all the aluminum pellets to the water, making sure not to lose any water to splashing. Also make sure that the pellets are all completely immersed in the water. Measure the new volume of the water plus the pellets.
4. When finished, retrieve and dry the aluminum pellets and return them to the fumehood.

Analysis: Use your measured mass and volume (obtained via water displacement) of the aluminum pellets to calculate the density of aluminum. Then look up the true density of aluminum and evaluate your accuracy by calculating your percent error.

The Thickness of Aluminum Foil

1. Now obtain a rectangular piece of aluminum foil from the fumehood. Use the ruler to measure the length and width of the piece of foil.
2. Measure the mass of the foil using the electronic balance.
3. When finished, return the foil to the fumehood and the ruler to the bin on the instructor's desk.

Analysis: Use these measurements along with the density of aluminum to calculate the thickness of the foil.

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2.2: Density of Liquids and Solids Pre-lab Assignment

Questions

1. Circle the correct responses in the following statement: Density is a physical / chemical property of matter and an intensive /extensive property of matter.
2. What devices will you use to measure the mass and the volume of water in Part A of this lab?
3. In Part B of this lab you will perform several measurements in order to determine the density of a metal. Name this metal. Describe the technique you will use to measure the volume of this metal.
4. Consider the tabulated data collected by a student for an unknown metal sample. Use this data to calculate the density of the metal (in g/cm^3). Show your work clearly.

Mass of empty beaker	44.656 g
Mass of beaker + metal sample	124.400 g
Initial volume of water in cylinder	12.7 mL
Final volume of water + metal sample	21.6 mL

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2.R: Density of Liquids and Solids (Report)

Part A: The Density of Water

Experimental Data

	1st water addition	2nd water addition	3rd water addition
Mass of empty cylinder			
Mass of cylinder + water			
Mass of water			
Volume of water			
Density of water			
Average density of water			

Temperature of water: _____

Data Analysis

Look up the true density of water at the temperature recorded. Obtain this value from the chart located in laboratory. Then use this to calculate the percent error in your average density of water. Show your work.

Part B: The Density of Aluminum and the Thickness of Foil

Table 1: The Density of Aluminum

Mass of empty beaker	
Mass of beaker + Al pellets	
Initial volume of water in cylinder	
Final volume of water and Al pellets	
Volume of Al pellets	

Table 2: The Thickness of Aluminum Foil

Mass of Al foil	
Length of Al foil	
Width of Al foil	

Data Analysis

1. Use your measured mass and volume of the pellets (in Table 1) to calculate the density of aluminum, in g/cm^3 . Show your work, and report your answer to the correct number of significant figures.
2. Look up the true density of aluminum at <http://www.chemicool.com>: _____
3. Use this to calculate the percent error in your experimentally determined density value. Show your work.
4. Use your measurements for the aluminum foil (in Table 2) along with the true density of aluminum to calculate the foil thickness, in cm. Show your work, and report your answer in scientific notation. Consider the foil to be a very flat rectangular box, where: $\text{Volume of foil} = \text{length} \times \text{width} \times \text{thickness}$

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CHAPTER OVERVIEW

3: Qualitative Analysis Testing for Cations and Anions

[3.1: Qualitative Analysis Testing for Cations and Anions Lab Report](#)

[3.2: Qualitative Analysis Testing for Cations and Anions Pre-lab Assignment](#)

[3.R: Qualitative Analysis Testing for Cations and Anions \(Report\)](#)

Adapted from

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3.1: Qualitative Analysis Testing for Cations and Anions Lab Report

Learning Objectives

To perform flame tests of metal cations in order to observe their characteristic colors To test for the presence of cation or anion by carrying out qualitative analysis Identify cation or anion present in an unknown sample

Background

Cations and anions play very important roles in health and medicine. Ions such as sodium, potassium, calcium, chloride, phosphate, and bicarbonate are needed to maintain proper hydration and blood pH, and they are needed for the proper function of nerve and muscle tissue. Abnormal levels of these ions in the body can have severe consequences leading to cardiac and neuronal problems, and if not treated can even cause death.

A common blood test that is often ordered by doctors is a chemistry panel (also known by a variety of other names such as Chem12, Chem7, SMA 20, SMAC, CMP, etc.) for a patient. A sample of the patient's blood is sent to a lab where a series of tests are carried out to measure the presence and quantity of certain chemicals, some of which are ions.

A series of chemical reactions is carried out that give distinct results, such as color changes, the formation of gases or insoluble precipitates (a solid), for each ion being tested. In this experiment, we will carry-out flame test and qualitative studies to identify different metal ions present in a sample.

Flame Test

Flame tests are used to identify the presence of a relatively small number of metal ions in a compound. Not all metal ions give flame colors. For Group 1 compounds, flame tests are usually by far the easiest way of identifying which metal you have. For other metals, there are usually other easy methods which are more reliable - but the flame test can give a useful hint as to where to look.



Figure 3.1.1: Colored flames from strontium, cesium, sodium and lithium (from left to right). Picture courtesy of the Claire Murray and Annabelle Baker from the Diamond Light Source.

Qualitative Analysis

When a sample of blood is tested, the technician will record their observations of the results, by comparing their observations to the expected outcomes, they can determine the presence or absence of a given ion in the blood sample. This type of analysis, where only the presence or absence of a particular substance (in this case ions) is determined, is called qualitative analysis.

In qualitative analysis, the ions in a mixture are separated by selective precipitation. Selective precipitation involves the addition of a carefully selected reagent to an aqueous mixture of ions, resulting in the precipitation of one or more of the ions, while leaving the rest in solution. Once each ion is isolated, its identity can be confirmed by using a chemical reaction specific to that ion.

Experimental Procedure

The goal of this laboratory is to introduce you to the process of qualitative analysis by determining the ionic composition of an unknown solution (imagine it is the serum from a patient's blood sample). You will first carry out a series of experiments on solutions containing a particular cation or anion (standard solutions) and record your observations. Then you will repeat these experiments and record your observations on a solution containing unknown ions (the unknown solution). You will then compare

your observations for the unknown with those obtained for the standard solutions and make a qualitative determination as to the ionic composition of the unknown solution.

Safety

Wear your safety goggles and gloves!

Materials

Spot plate, wood splints, small beaker, test tubes, tube rack, Bunsen burner, stirring rod, hot plate, 0.1 M NaCl, 0.1 M KCl, 0.1 M CaCl₂, 0.1 M CuSO₄, 0.1 M Sr(NO₃)₂, 0.1 M (NH₄)₂C₂O₄, 0.1 M NH₄Cl

Part A. Tests for Cations

1. Flame test for sodium, potassium, calcium, strontium, copper

Obtain the standard NaCl, KCl, CaCl₂, CuSO₄, Sr(NO₃)₂ solutions (these are your known solutions). Place the wood splint into the spot plate well with the NaCl solution and then into the flame. Record the color produced by the Na⁺ ion. Repeat the test with the KCl solution. The color (pink-lavender) of K⁺ does not last long, so look for it immediately. Record the color of the 2⁺ 2⁺ K⁺ ion. Repeat the flame test with Ca . Record the color of the Ca flame. Repeat test for the remaining cations and record your observations.

2. Oxalate test for calcium

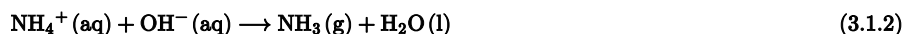
Place 5-6 drops of the standard CaCl₂ solution in a test tube (this is your known solution). Add 1-2 drops of ammonium oxalate solution, (NH₄)₂C₂O₄. Look for a cloudy, white solid (precipitate). If the CaCl₂ solution remains clear, place the test tube in a warm water bath and heat gently for 5 minutes. Look for a precipitate. The net equation for the reaction is:



white precipitate indicates the presence of Ca²⁺.

3. Test for ammonium ion, NH₄⁺

Place 5-6 drops of the standard NH₄Cl solution in test tube (this is your known solution). Add 2 drops of sodium hydroxide (NaOH). You may notice the odor of ammonia. Place the test tube in a hot water bath and gently heat for 5 minutes. Place a strip of red litmus paper across the top of the test tube. The red litmus paper will turn blue as a result of the release of NH₃ (g) from the reaction. Record your observations.



Part B. Tests for Anions

1. Silver nitrate test for chloride

Place 5-6 drops of the known NaCl solution in a test tube. Add 1 drop of AgNO₃ and 1-2 drops of nitric acid (HNO₃). Caution: AgNO₃ stains the skin. Any white solid that remains is AgCl (s) which should not dissolve upon the addition of HNO₃. Record your observations.



white precipitate remains after HNO₃ is added.

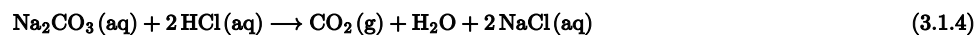
2. Ammonium molybdate test for phosphate

Place 5-6 drops of the Na₃PO₄ solution in a test tube. Add 1-2 drops of HNO₃. Warm the test tube in a hot water bath for 5-10 minutes, after the solutions are warmed add 2 drops of ammonium molybdate solution, (NH₄)₂MoO₄. The resulting yellow precipitate indicates the presence of PO₄³⁻. Record your observations.

3. Carbon dioxide test for carbonate/bicarbonate

Place 5-6 drops of the standard Na₂CO₃ solution in a test tube. While carefully observing the solution, add 1-2 drops of HCl. Watch for the evolution of bubbles of CO₂ gas as you add the hydrochloric acid (HCl). The gas bubbles are formed quickly, and

may be overlooked. If gas bubbles were not observed, add another 2-3 drops of HCl as you watch the solution. Record your observations.



gas bubbles are formed

Part C. Testing a solution for unknown cation and anion

1. Write down the unknown sample number.
2. Repeat each of the tests (A1, A2, A3, B1, B2, and B3) for the unknown solution and write down your observations. For example, when performing the flame test place 8-10 drops of unknown solution into an empty well of a ceramic well plate. Place the wood splint into the solution and then into the flame. Record your observations.
3. Compare these observations to those that you obtained in sections A and B. Determine the cation and anion present in your sample.

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3.2: Qualitative Analysis Testing for Cations and Anions Pre-lab Assignment

Questions

1. Write the chemical symbol or formula for the following:
 - a. Sodium ion
 - b. Potassium ion
 - c. Calcium ion
 - d. Chloride ion
 - e. Phosphate ion
 - f. Bicarbonate ion
 - g. Sodium chloride
 - h. Sodium bicarbonate
 - i. Potassium phosphate
 - j. Calcium bicarbonate
 - k. Potassium bicarbonate
 - l. Calcium phosphate
 - m. Calcium Chloride
 - n. Strontium nitrate
 - o. Copper (II) Nitrate
2. Name one safety precaution that must be followed during this experiment.

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3.R: Qualitative Analysis Testing for Cations and Anions (Report)

A. Tests for cations

Procedure	Cation tested	Observation
A1. Flame Test	Sodium	
	Potassium	
	Calcium	
	Strontium	
	Copper (II)	
A2. Oxalate test	Calcium	
A3. Ammonium test	Ammonium	

B. Tests for anions

Procedure	Cation tested	Observation
B1. Silver nitrate test	Chloride	
B2. Ammonium molybdate test	Phosphate	
B3. Carbon dioxide test	Carbonate	

C. Testing a solution for unknown Cation & Anions

Procedure	Cation tested	Observation Sample #____
Flame test	Sodium	
	Potassium	
	Calcium	
	Strontium	
	Copper (II)	
Oxalate test	Calcium	
Silver nitrate test	Chloride	
Ammonium molybdate test	Phosphate	
Carbon dioxide test	Carbonate	
Ammonium test	Ammonium	

C3: Based on your observations above, complete the following chart.

Sample # ____ had the following:	Cations:	Anions:
----------------------------------	----------	---------

Data Analysis/Post Lab Questions

1. What is the chemical formula and name of your unknown sample?
2. You detect the presence of calcium ions in a sample using the flame test. What additional test could you do to confirm this finding?

3. You are given a solution that is supposed to contain calcium phosphate. What tests can you perform to confirm the presence of this calcium phosphate? Provide a brief procedure including any observations you are likely to make.
-

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CHAPTER OVERVIEW

4: Calorimetry/Thermochemistry

- [4.1: Calorimetry/Thermochemistry Lab Procedure](#)
- [4.2: Calorimetry/Thermochemistry Pre-lab Assignment](#)
- [4.R: Calorimetry/Thermochemistry \(Report\)](#)

Adapted from

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4.1: Calorimetry/Thermochemistry Lab Procedure

Learning Objectives

- Determine whether a reaction is endothermic or exothermic.
- Determine the best ionic compound to use in a heat pack for treating frostbite based on your experimental results.
- Calculate the average heat capacity of your calorimeter.
- Calculate the specific heat of a metal.

Background

The human body works best within a very narrow temperature range. A temperature drop of as little as 2°C in the body's core causes classic hypothermia symptoms such as mental difficulties and loss of physical coordination. Much more extreme temperature drops in the extremities may be survived, but can lead to frostbite if the flesh freezes. Victims of hypothermia require immediate treatment, and in outdoor situations the treatment is often warmth provided by portable heat sources, such as heat packs.

Heat packs are available that produce warmth through various chemical reactions. Such heat packs are convenient because they only release heat when triggered. One common heat pack contains an internal pouch of water and a solid powder. Once the pouch of water is broken open, there is an exothermic reaction between the water and the powder. These heat packs have limitations. For example, they do not work well in extreme cold because the water within the pack will freeze.

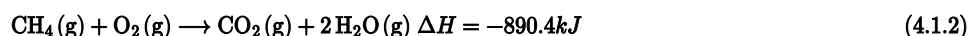
In the coldest environments, heat packs are available that contain only the powder in a resealable waterproof sack. When heat is required, the sack can be opened and any aqueous solution poured inside. The sack is resealed and the reaction produces heat.

Reaction Enthalpy

An important part of chemistry involves studying energy changes that occur during chemical reactions. These energy changes are of fundamental importance in understanding the “driving force” of a chemical reaction. The most common way energy is exchanged between a chemical system and the environment is by evolution or absorption of heat (q). The change in heat energy accompanying a chemical reaction is known as enthalpy change, ΔH . By convention, reactions in which heat is absorbed are labeled endothermic and have positive values of ΔH ; reactions in which heat is released are labeled exothermic and have negative values of ΔH . Cold packs used in athletics are familiar to many sports enthusiasts. In order to derive coldness from the pack, a plastic packet of water is broken inside another packet containing a solid salt such as NH_4NO_3 . In this case, the enthalpy of solution, i.e., heat absorbed when a substance dissolves, is endothermic indicating that heat is absorbed as the salt dissolves. Thus, the enthalpy of solution is designated with a positive sign since energy is absorbed or added.



On the other hand, gas stoves produce heat through the combustion (burning in oxygen) of methane, (see Equation 1). Since heat is produced by the combustion, the reaction is exothermic and enthalpy of combustion (ΔH), i.e., heat released during combustion, must be negative. In fact, ΔH for the methane combustion is -890.4 kJ (see below). It is important to realize that ΔH is related to the coefficients in the balanced equation. Thus, -890.4 kJ of heat is released per every 1 mole of CH_4 that reacts but per every two moles of H_2O that is formed.



The negative sign serves to reinforce the fact that heat is produced rather than absorbed or that this is an exothermic reaction. We can use stoichiometry to calculate the enthalpy of any amount of a reactant.

Heat and Temperature

If an object (such as a pot of water) is positioned to absorb the heat given off during a combustion reaction, then the temperature of the object will change as follows

$$q = m \times c \times \Delta T \quad (4.1.3)$$

Where q = the amount of heat absorbed by the object in Joules

m = the mass of the object being heated in grams

c = the specific heat of the object being heated

ΔT = the change in temperature of the object = final temperature minus the initial temperature = $T_f - T_i$

The specific heat is different for different substances. The specific heat or heat capacity is low for objects which are easily heated or cooled with minimal energy input. Metals usually have low specific heats because their temperature changes very quickly. Water has a high specific heat because it takes a lot of energy to change the temperature of a sample of water. For example:

Example specific heat values

Substance	Specific Heat ($\text{J g}^{-1} \text{C}^{-1}$)
Water	4.18
Air	1.01
Aluminum	0.897
Granite	0.790

Calorimetry

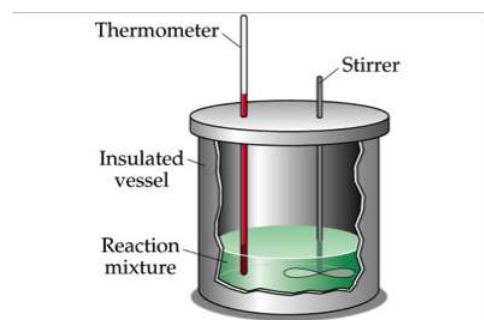


Figure 4.1.1: A simple illustration of a typical calorimeter.

Many experiments in thermochemistry involve a calorimeter. A calorimeter, like the one above, is simply a container that insulates a reaction from the surrounding environment so change in temperature as the reaction proceeds place can be measured accurately measured as independent of the environmental temperature. In lab, you will use two Styrofoam cups as your calorimeter. Ideally any calorimeter would be able to completely maintain the system without losing heat to the surroundings. Unfortunately, no calorimeter is perfect and heat is always lost to the surroundings. It is possible to determine how much heat is lost by the calorimeter by mixing hot water and room temperature water. The heat lost by the hot water is equal to the heat gained by the room temperature water.

$$|-q_{\text{hot}}| = |q_{\text{ambient}}| + |q_{\text{calorimeter}}| \quad (4.1.4)$$

In order to calculate the heat lost to the calorimeter you will calculate q for both the hot water and the room temperature water (using 4.1.3) and then subtract to find the difference. It is also possible to view how heat can be transferred from one object to another. In a perfect system (where no heat is lost to the calorimeter) the heat lost by the hot object would be gained by the cold object such as:

$$q_{\text{lost}} = -q_{\text{gained}} \quad (4.1.5)$$

which could be rearranged to give:

$$q_{\text{hot water}} = -q_{\text{cold water}} \quad (4.1.6)$$

as well as replacing q with 4.1.4 to give

$$mc(T_f - T_i)_{\text{hot object}} = -mc(T_f - T_i)_{\text{cold object}} \quad (4.1.7)$$

Although usually heat is lost to the calorimeter, as we showed in 4.1.4, we can usually assume this value is exceptionally small and omit it from the equation as we did in 4.1.5, 4.1.6, 4.1.7. In this lab, we will be investigating the endothermic and exothermic qualities of salt solutions by dissolving various salts into water and monitoring the temperature. Then you will calibrate a coffee cup calorimeter by determining its heat capacity. Finally, you will calculate the specific heat of copper by dropping hot pennies into water and monitoring the heat exchange.

Experimental Procedure

Materials and Equipment

Sodium chloride, potassium chloride, calcium chloride, 10-15 pennies, DI water bottle. Equipment: 600 mL beaker, 250 mL beaker, 100 mL beaker, 50 mL graduated cylinder, 4- styrofoam cups, 2 lids with hole, stirring rod, 3 small test tubes, thermometer, hot plate.

Safety

It is important to follow the safety guidelines below while performing this lab. Remember to dry the calorimeter and thermometer between each trial. Use appropriate procedures for hot plate use. Dispose of waste in the Waste container as indicated by your instructor. WEAR SAFETY GOGGLES, CLOSED-TOE SHOES.

Part A: Exothermic and Endothermic Dissolution of Salts

In this section of the procedure, you will observe temperature changes as various salts are dissolved in water.

1. The first salt is NaCl, and the corresponding dissolution reaction is $\text{NaCl(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
2. Record the mass of the empty test tube prior to beginning.
3. You will be recording the temperature using a thermometer. Insert the end of the thermometer into the bottom of a clean, dry test tube. (Use the small test tubes).
4. Fill the test tube approximately 2 cm with distilled water. Record the mass of the test tube and water to determine the mass of water.
5. Record the initial temperature of the water.
6. Record the mass of a second (clean and dry) test tube.
7. Fill the second test tube approximately 1 cm with solid NaCl.
8. Pour the solid NaCl into the water and stir gently with the thermometer.
9. Monitor the temperature. The temperature will increase or decrease away from the initial temperature. Eventually the final temperature will be reached before it begins to return to room temperature. Record the final temperature.
10. The solution should be disposed of in the waste container in the fume hood. Rinse and dry the test tube.
11. Repeat Steps 2-7 for CaCl₂ and (if there is time) KCl.

Part B: Calculating the Heat Capacity of a Calorimeter

1. Obtain a hot plate and plug it in.
2. Stack the two Styrofoam cups together and place inside a 400-mL beaker
3. Place the assembly on a balance. Record the mass.
4. Place 50-mL of tap water in the cup assembly and record mass. Subtract to find the mass of water.
5. Place the cardboard lid on top the Styrofoam cups.
6. Insert a thermometer through hole in lid. Obtain the initial temperature of the “cold” water.
7. Attach a thermometer clamp to the thermometer so that the thermometer is not sitting on bottom of cup.
8. This is our calorimeter.
9. Tare a 150mL beaker.
10. Add ~50mL of water to beaker on scale and record weight as the weight of “hot” water.
11. Place the ~150 mL beaker on the hot plate and heat to ~90 degrees.
12. Using the thermometer, record the temperature of the hot water.
13. Tip the lid of the calorimeter up and using beaker tongs immediately pour the hot water into the calorimeter.
14. Immediately replace the lid and begin recording temperature.

15. Gently swirl the calorimeter and measure temperature every 10 seconds until the temperature is constant for 3 readings. (The temperature will increase to the final temperature before beginning to decrease back to room temperature.) Record the final temperature in the data section.
16. Repeat steps 9-14 for a second trial.

Part C: Calculating the Specific Heat of Copper

1. Obtain the mass of ~ 10-20 copper shots or the weight of a small piece of copper wire. Record the mass in your data table.
2. Add copper shots/copper wire to a large (clean and dry) test tube. Place the test tube in a 400-mL beaker containing ~ 150 mL water. The water level should be above the level of the copper shots/copper wire to ensure they are adequately heated.
3. Heat water to ~ 95-100 °C. Record this as the initial temperature of the copper shots/copper wire. Allow the copper shots/copper wire to heat for ~2-3 minutes.
4. Add your calorimeter to the balance and tare it. Add ~40 mL of water. Record the mass of the water in the table.
5. Set up your calorimeter with thermometer as described earlier.
6. Record the initial temperature of the water.
7. Swiftly but carefully remove the test tube from the hot water bath with a pair of test tube tongs. Dump the copper shots/copper wire into the calorimeter and immediately cover with the lid. *Be careful that the hot water on the outside of the test tube does not drip onto your hand or into the calorimeter!
8. Monitor the temperature by taking temperature readings every 10 seconds. When the temperature is consistent for 3 straight measurements (or the temperature begins to cool) record the final temperature of the copper shots/copper wire and water.
9. Repeat steps 1-8 for a second trial. (Be sure to dry the copper shots between trials).
10. Use the data from your table to calculate the specific heat of copper (copper shots/copper wire).

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4.2: Calorimetry/Thermochemistry Pre-lab Assignment

Questions

1. Describe exothermic and endothermic process.
2. List two reagents that will be used in Part A of this experiment.
3. What is calorimetry?

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4.R: Calorimetry/Thermochemistry (Report)

Part A: Exothermic and Endothermic Dissolution of Salts

Experimental Data

	NaCl	CaCl ₂	KCl
Equation for dissolution of salt			
Mass of test tube			
Mass of test tube and water			
Mass of water			
Initial temperature of water			
Mass of test tube			
Mass of test tube and salt			
Mass of the salt			
Final temperature			
Is this exothermic or endothermic dissolution?			
Calculate q for the dissolution of the salt			
Calculate the J/g salt			
Would this salt be most useful for ice or heat packs?			

Part B: Calculating the Heat Capacity of a Calorimeter

Show your calculations

	Trial 1	Trial 2
Mass of calorimeter (Coffee cups and lid)		
Mass of "cold" water		
Initial temperature of "cold" water		
Mass of hot water		
Initial temperature of hot water		
Final temperature of the system		
Specific heat of water	4.184 J/g °C	4.184 J/g °C
q Lost by hot water		
q Gained by cold water		
Heat capacity of calorimeter		
Average heat capacity of calorimeter		

Trial 1		Trial 2	
Time	Temperature	Time	Temperature
0		0	
10		10	
20		20	
30		30	
40		40	
50		50	
60		60	
70		70	

Part C: Calculating the Specific Heat of Copper

Show your calculations

	Trial 1	Trial 2
Mass of copper shots used		
Initial temperature of copper shots/copper wire		
Final temperature of the system		
Mass of water in calorimeter		
Initial temperature of water in calorimeter		
Specific heat of water	4.184 J/g °C	4.184 J/g °C
q Gained by water		
Specific heat of copper (Assuming no heat lost to calorimeter)		
Average specific heat of copper		
Specific heat of copper (Using calorimeter specific heat from Part B)		
Average specific heat of copper		
If the specific heat of Cu is 0.386 J/g °C, what is the % error from the above row?		

Trial 1		Trial 2	
Time	Temperature	Time	Temperature
0		0	
10		10	
20		20	
30		30	
40		40	
50		50	
60		60	
70		70	

Post-Lab Questions

1. According to your results, what salt in Part A would have been the best choice for use in a heat pack? In a cold pack? Defend your choice.
2. Why is it not possible to reuse a heat pack in term of the chemistry?
3. Compare the specific heat of water to the specific heat of metal in the table provided. Which would heat up faster (with less energy required)
4. Why would metal make a poor ingredient in a heat pack?
5. If there was a delay between measuring the initial temperature of a hot object and its transfer to the calorimeter, how would the heat capacity of the object be affected? How would the calculation of the heat capacity of the calorimeter be affected (too high, too low, or no affect)?
6. If hot water from the test tube in Part C had accidentally dripped into the calorimeter, how would the calculation of the specific heat of the calorimeter be affected (too high, too low, or no affect)?

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CHAPTER OVERVIEW

5: Single and Double Replacement Reactions

[5.1: Single and Double Replacement Reactions Lab Procedure](#)

[5.2: Single and Double Replacement Reactions Pre-lab Assignment](#)

[5.R: Single and Double Replacement Reactions \(Report\)](#)

Adapted from

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“LibreTexts Online Chemistry Lab Manual: Types of Chemical Reactions (Experiment)” by Santa Monica College is licensed under CC BY-NC-SA 3.0 US

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5.1: Single and Double Replacement Reactions Lab Procedure

Learning Objectives

- Perform and observe the results of a variety of single and double replacement reactions.
- Become familiar with some of the observable signs of these reactions
- Identify the products formed in each of these reactions.
- Write balanced chemical equations for each of the double replacement reactions studied.

Background

During a chemical reaction both the form and composition of matter are changed. Old substances are converted to new substances, which have unique physical and chemical properties of their own. Some of the observable signs that a chemical reaction has occurred include:

- Metallic deposit
- Appearance of bubbles
- Temperature change occurs
- Color change occurs
- Precipitate (cloudy, tiny particles) appears

Note that there are other observable signs for chemical reactions, but these are most likely to be seen in this lab

Single Replacement Reaction

In Part A of this lab we will examine single replacement reactions. This is one type of oxidation-reduction reaction, or **redox reaction**, because it occurs via a transfer of electrons.

Single replacement reactions have the general form: $A + BC \rightarrow B + AC$

Here, A is an element and BC is usually an aqueous ionic compound or an acid (consisting of B⁺ and C⁻ aqueous ions). Element A replaces element B in the compound BC; this results in the formation of a new element B and a new ionic compound or acid, AC. If the new element B is a metal, it will appear as a metallic deposit. If it is a gas, it will appear as bubbles.

✓ Example 5.1.1

magnesium metal + aqueous aluminum chloride

Since Mg is more active than Al, a single replacement reaction will occur. The predicted products are aluminum metal and aqueous magnesium chloride

Reaction Equation



✓ Example 5.1.2

Aluminum Metal + Aqueous Magnesium Chloride

Since Al is not more active than Mg, a replacement will NOT occur.

Reaction Equation



Double Replacement Reactions

All double replacement reactions have the general form:



Reactions that can be classified as double replacements include precipitation reactions, neutralization reactions and gas forming reactions.

Precipitation Reactions

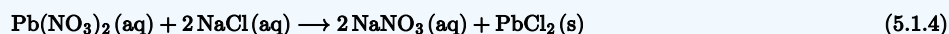
Here AB and CD are usually aqueous ionic compounds (or acids) consisting of aqueous ions (A^+ and B^- , C^+ and D^-). When a double replacement reaction occurs, the cations and anions switch partners, resulting in the formation of two new ionic compounds AD and CB, one of which is in the solid state. This solid product is an insoluble ionic compound called a precipitate. To determine whether a product ionic compound will be soluble or insoluble, consult the Solubility Rules provided at the end of the Background section. Note that if both of the predicted products are soluble, a precipitation reaction will not occur.

✓ Example 5.1.3

aqueous lead (II) nitrate + aqueous sodium chloride

The predicted products are lead (II) chloride (insoluble) and sodium nitrate (soluble). Since one of the predicted products is insoluble, a precipitation reaction will occur.

Reaction Equation



Neutralization Reactions

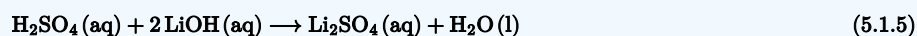
Here AB is an acid (consisting of H^+ and X^- aqueous ions) and BC is a base (consisting of M^+ and OH^- ions). When a double replacement reaction occurs, the cations and anions switch partners, resulting in the formation of water and a new ionic compound (or salt), which is usually soluble. Neutralization reactions are exothermic, and are generally accompanied by a noticeable release of heat.

✓ Example 5.1.4

sulfuric acid + aqueous lithium hydroxide

The predicted products are water and lithium sulfate.

Reaction Equation



Gas Forming Reactions

In these reactions one of the products (AD or CB) after the double replacement is in the gaseous state, such as hydrogen sulfide (H_2S) or ammonia (NH_3). One of the products could also be carbonic acid (H_2CO_3) or sulfurous acid (H_2SO_3). Both carbonic acid and sulfurous acid are unstable and will decompose to form carbon dioxide and sulfur dioxide gases, respectively:

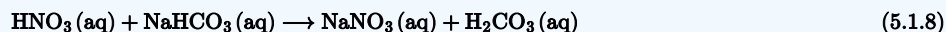


✓ Example 5.1.5

nitric acid + aqueous sodium bicarbonate

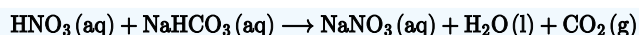
The predicted products are carbonic acid and sodium nitrate. However carbonic acid decomposed to carbon dioxide and water.

Reaction Equation



Decomposes

Final equation:



(5.1.9)

Writing Equations for Reactions

Write chemical formulas for each reactant and place a yield arrow ($\rightarrow$) after the last reactant.

Identify the reaction type – precipitation, neutralization or gas forming. If you determine that a reaction will occur, write the correct formula(s) of the products after the arrow.

If you determine that a reaction will not occur, write “no reaction” after the arrow. Balance the equation (to ensure mass conservation). Be sure to include the physical states of all reactants and products in your final equation.

Review: Chemical Formulas and Names of Ionic Compounds

Ionic compounds are formed when positive ions (cations) and negative ions (anions) are attracted to each other via strong electrostatic forces. This attraction is called ionic bond. The basic rules for writing the chemical formulas of ionic compounds:

- Determine the formulas and charges on the cation and anion involved in the compound.
- Combine the ions in a ratio that results in the formation of a neutral ionic compound.
- The charge of all the cations must equal the total charge of all the anions in the compound. The number of each element present in the compound becomes a subscript in the chemical formula.

✓ Example 5.1.6

Write the formula for iron (III) chloride

The basic rules for writing names of ionic compounds:

- Both the cation and anion must be named.
- Always name the cation first, then the anion.

Solution

First identify the cation and the anion in this compound – iron (III) – Fe cation and chloride Cl - 3+ anion. For a neutral compound, one Fe is needed for every 3 Cl - . The formula of the compound is FeCl₃.

✓ Example 5.1.7

Name the ionic compound Al(NO₃)₃. First identify the cation and anion in this compound.

Solution

Cation = Al³⁺ : aluminum

Anion = NO₃⁻ : nitrate

The name of this compound is aluminum nitrate.

Experimental Procedure

Safety

- Be especially cautious when using 6M HCl, 3M H₂SO₄ and 6M NaOH as they can burn your skin.
- Be aware that skin discoloration will result from contact with AgNO₃. If you feel any tingling sensations or see any color changes on your skin, flush with water immediately for a minimum of 15 minutes. Inform your instructor of any chemical contact as soon as possible.
- WEAR SAFETY GOGGLES, CLOSED-TOE SHOES

Materials and Equipment

Solids: sodium bicarbonate, copper, zinc, magnesium

Solutions: 6 M sodium hydroxide, 3 M sulfuric acid, 6 M hydrochloric acid, 1 M sodium chloride; all other solutions are 0.1 M and include sodium chloride, iron (III) chloride, ammonium hydroxide, sodium carbonate, cobalt (II) chloride, sodium phosphate, copper (II) sulfate, potassium nitrate, barium chloride, silver nitrate, copper (II) sulfate, zinc nitrate, aluminum sulfate. Equipment: 9 small test tubes, test tube rack, large beaker, 2 small beakers.

Part A: Single Replacement Reactions

Use the medium sized test tubes; place them in a large beaker. Use clean test tubes that have been rinsed with distilled water. The test tubes do not have to be dry. Use 3-mL quantities of all solutions. A good estimate is to measure out 3 mL of water in a test tube and use it as a comparison.

Place one piece of metal in the test tube first, and then add the solution. The metal should be completely immersed in the solution used. If results are not obtained immediately, give the reaction some time. Some reactions take longer than others.

Perform the following reactions, and record your observations for each on the data sheet. All waste is to be disposed of in the container in the hood!

1. Zinc metal + hydrochloric acid
2. Copper metal + aqueous silver nitrate
3. Copper metal + aqueous zinc nitrate
4. Magnesium metal + hydrochloric acid

Part B: Double Replacement Reactions

Use approximately 3-mL quantities of all solutions. A good estimate is to measure out 3 mL of water in a test tube and use it as a comparison.

Perform the following reactions, and record your observations for each on the data sheet. If results are not obtained immediately, give the reaction some time. Some reactions take longer than others. All waste is to be disposed of in the container in the hood!

1. Aqueous sodium chloride + aqueous silver nitrate
2. Aqueous sodium phosphate + aqueous copper (II) sulfate
3. Hydrochloric acid + solid sodium bicarbonate (just a small scoop)
4. Aqueous barium chloride + sodium sulfate
5. Hydrochloric acid + aqueous sodium hydroxide
6. Aqueous sodium carbonate + aqueous cobalt (II) chloride
7. Aqueous sodium chloride + aqueous potassium nitrate
8. Aqueous iron (III) chloride + aqueous ammonium hydroxide

When finished, complete the data sheet by writing the balanced equation for each reaction.

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5.2: Single and Double Replacement Reactions Pre-lab Assignment

Questions

1. Name two observable signs that a chemical reaction has occurred.
2. Name two chemicals used in this experiment that must be handled with caution.
3. Give an example of a double replacement and single replacement reaction.

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5.R: Single and Double Replacement Reactions (Report)

Experimental Results

Part A: Single Replacement Reaction

For each of the reactions performed:

- record your observations
- predict the names and states of the products formed
- write the balanced equation, including all physical states

Zinc metal + hydrochloric acid

Observations	
Product Names	
Balanced Equation	

Copper metal + aqueous silver nitrate

Observations	
Product Names	
Balanced Equation	

Copper metal + aqueous zinc nitrate

Observations	
Product Names	
Balanced Equation	

Magnesium metal + hydrochloric acid

Observations	
Product Names	
Balanced Equation	

Data Analysis Questions

1. In reactions a-e, which (if any) did not result in a reaction? Explain.

Part B: Double Replacement Reactions

- predict the reaction type (precipitation, neutralization or gaseous) -- record your observations
- predict the names and states of the products formed
- write the balanced equation, including all physical states.

Aqueous sodium chloride + aqueous silver nitrate

Observations	
Product Names	
Reaction Type	
Balanced Equation	

Hydrochloric acid + aqueous sodium hydroxide

Observations	
Product Names	
Reaction Type	
Balanced Equation	

Hydrochloric acid + solid sodium bicarbonate

Observations	
Product Names	
Reaction Type	
Balanced Equation	

Aqueous iron (III) chloride + aqueous ammonium hydroxide

Observations	
Product Names	
Reaction Type	
Balanced Equation	

Aqueous barium chloride + sodium sulfate

Observations	
Product Names	
Reaction Type	
Balanced Equation	

Aqueous sodium phosphate + aqueous copper(II) sulfate

Observations	
Product Names	
Reaction Type	
Balanced Equation	

Aqueous sodium carbonate + cobalt(II) chloride

Observations	
Product Names	
Reaction Type	
Balanced Equation	

Aqueous sodium chloride + aqueous potassium nitrate

Observations	
Product Names	
Reaction Type	
Balanced Equation	

Post-Lab Questions

- Describe each of the following compounds as ionic, covalent, or acid.
- Use IUPAC nomenclature rules to write the chemical formula or the chemical name.

Name	Formula	Description
Aluminum Oxide		
Tin(IV) sulfate		
Silver sulfite		
Carbon tetrachloride		
	P ₂ O ₅	
	NO ₂	
	H ₂ SO ₄	
	PbCl ₂	
	NBr ₃	
	AlPO ₄	

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CHAPTER OVERVIEW

6: Mole Ratios and Reaction Stoichiometry

[6.1: Mole Ratios and Reaction Stoichiometry Lab Procedure](#)

[6.2: Mole Ratios and Reaction Stoichiometry Pre-lab Assignment](#)

[6.R: Mole Ratios and Reaction Stoichiometry \(Report\)](#)

Adapted from

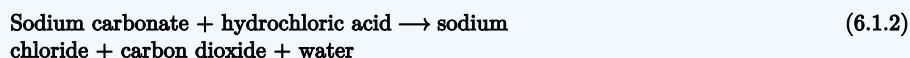
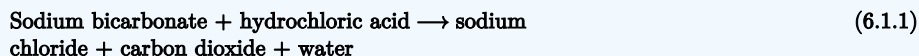
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6.1: Mole Ratios and Reaction Stoichiometry Lab Procedure

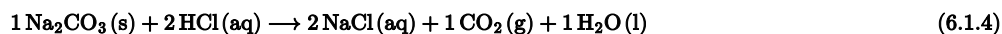
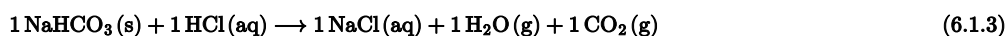
Learning Objectives

To experimentally determine the mole-to-mole ratios between the underlined reactants and products in the following two double displacement “gas forming” reactions:



Background

The easiest way to obtain the mole-to-mole ratios would be to simply balance the chemical equations for these reactions. This would be considered a theoretical approach to the problem. The balanced equations for reactions 6.1.1 and 6.1.2 are:



In reaction 6.1.3, the balancing coefficients indicate that there is a 1:1 mole ratio between reactant NaHCO₃ and product NaCl. This means that for every 1 mole of sodium bicarbonate that reacts, 1 mole of sodium chloride should be produced. However, in reaction 6.1.4, the balancing coefficients indicate that there is a 1:2 mole ratio between reactant Na₂CO₃ and product NaCl. In this case, for every 1 mole of sodium carbonate that reacts, 2 moles of sodium chloride should be produced.

To determine these mole-to-mole ratios experimentally, a quantitative analysis of both reactions is required. Specifically, a pre-weighed mass of sodium bicarbonate/carbonate will be allowed to react with a slight excess of hydrochloric acid. The sodium chloride product will then be carefully retrieved, dried and weighed at the end of the reaction. This mass of collected sodium chloride is called an experimental yield. Both the mass of sodium bicarbonate/carbonate reactant used and the mass of sodium chloride product can be converted to mole quantities via their respective molar masses. Finally, dividing both the reactant and product moles by the lower of the two values should yield the simplest whole number mole-to-mole ratio of reactant and product.

While an experimental product yield is obtained by actually performing a reaction in lab, a theoretical yield is the maximum mass of product that could be obtained from a reaction provided that no errors occur. Theoretical product yields can only be determined by performing a series of stoichiometric calculations. Thus, using this method, theoretical yields of sodium chloride will be calculated for reactions A and B. Once obtained, the percent yield of sodium chloride can be determined for both reactions, where

$$\text{Percent Yield} = \frac{\text{Experimental Yield}}{\text{Theoretical Yield}} \times 100 \quad (6.1.5)$$

Good experimental practices in the lab (with minimum error) generally result in a high percent yield, where the experimental yield closely approaches the theoretical yield.

Experimental Procedure

Materials and Equipment

Solid sodium bicarbonate (NaHCO₃), solid sodium carbonate (Na₂CO₃), 6 M hydrochloric acid (HCl), electronic balance, evaporating dish, watch glass (to fit as cover for the evaporating dish), stand and ring clamp, wire gaze, dropper pipette, small beaker, and Bunsen burner.

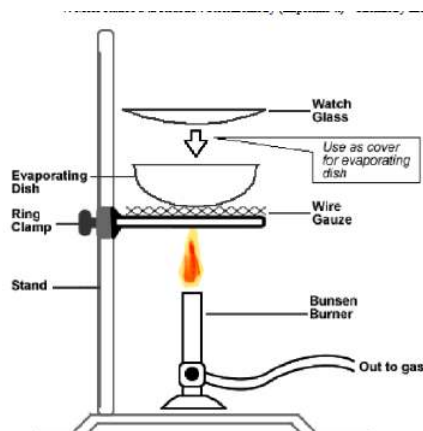
Safety

- Be especially careful when handling the 6 M HCl (aq), as it can cause chemical burns to the skin. If any acid spills on you, rinse immediately under running water for up to 15 minutes and report the accident to your instructor.
- Acid spills may also be neutralized using the sodium bicarbonate solution in the fume hood. Also, be sure to exercise appropriate caution when using the Bunsen burner and handling hot equipment.

- ALWAYS WEAR YOUR GOGGLES!!!

Procedure

1. Measure and record the mass of your clean dry evaporating dish + watch glass (assembled together with the watch glass acting as a cover on top of the evaporating dish).
2. Carefully add 0.3 – 0.4 g of solid sodium bicarbonate (NaHCO_3) to the evaporating dish. Do not do this over the balance! Then measure and record the mass of the evaporating dish + watch glass + NaHCO_3 .
3. Obtain about a 5-mL quantity of hydrochloric acid (HCl) in your 10 mL graduated cylinder. Then using your dropper pipette, add the drop by drop to the sodium bicarbonate in the evaporating dish. The reaction will be evident by the bubbling that takes place. Gently mix the reactants after every 4-5 drops of HCl . Continue adding HCl until the bubbling stops and all of the NaHCO_3 is dissolved – this indicates that the reaction is complete.
4. Assemble the stand, ring clamp and wire gauze apparatus for heating as shown in the figure below. Cover the evaporating dish with the watch glass and place it on the wire gauze.



- 1.
5. Gently heat the solution in the covered evaporating dish with a Bunsen burner flame in order to remove the water generated in the reaction (as well as any excess HCl present). The flame should be adjusted to a lower temperature and wafted under the evaporating dish constantly. Continue heating until the contents are completely dry. Note that the watch glass cover should also be dry!
6. After allowing the evaporating dish to cool to room temperature, measure and record the mass of the evaporating dish + watch glass + residue (NaCl).
7. Repeat steps 1 to 6 with a 0.3 – 0.4 g sample of sodium carbonate (Na_2CO_3).
8. The waste from this experiment may be disposed of waste container in the hood.
9. Analysis: Experimental Mole-to-Mole Ratios – Convert the initial mass of sodium bicarbonate (or carbonate) reactant to moles (via its molar mass). In a similar manner, convert the final mass of collected sodium chloride product to moles (via its molar mass). Finally, obtain the simplest whole number mole-to-mole ratio by dividing both the reactant and product moles by the lower of the two values.
10. Analysis: Percent Yields – Calculate the theoretical yield of NaCl for both reactions 3 and 4 via standard mass-to-mass stoichiometry. Use your masses of sodium bicarbonate/carbonate reactants weighed out in lab as the starting point and the mole ratios from the balanced equations for these calculations. Then determine your percent yield for each reaction using the calculated theoretical yields along with your experimental yields of NaCl , obtained in lab. See the Theory Section for the equation.

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6.2: Mole Ratios and Reaction Stoichiometry Pre-lab Assignment

Reactions

Sodium bicarbonate + hydrochloric acid $\rightarrow$ sodium chloride + carbon dioxide + water (6.2.1)

Sodium carbonate + hydrochloric acid $\rightarrow$ sodium chloride + carbon dioxide + water (6.2.2)

$1 \text{ NaHCO}_3 (\text{s}) + 1 \text{ HCl} (\text{aq}) \rightarrow 1 \text{ NaCl} (\text{aq}) + 1 \text{ H}_2\text{O} (\text{g}) + 1 \text{ CO}_2 (\text{g})$ (6.2.3)

$1 \text{ Na}_2\text{CO}_3 (\text{s}) + 2 \text{ HCl} (\text{aq}) \rightarrow 2 \text{ NaCl} (\text{aq}) + 1 \text{ CO}_2 (\text{g}) + 1 \text{ H}_2\text{O} (\text{l})$ (6.2.4)

Questions

- Calculate the molar mass for the following compounds:
 - NaHCO_3 :
 - Na_2CO_3 :
- Your goal in this lab is to experimentally verify the mole-to-mole ratios between a certain reactant and a certain product in both reactions. Identify the two substances in Reaction 6.2.4.
- In Reaction 6.2.3, you will react a pre-weighed sample of sodium bicarbonate with acid. In Reaction 6.2.4, you will use sodium carbonate instead of the bicarbonate. Name the "container" that you will perform both reactions in.
- What is the purpose of the watch glass?
- How will you know when enough acid has been added to the sodium bicarbonate (in Reaction 6.2.3) or sodium carbonate (in Reaction 6.2.4), and that the reactions are complete?
- After mixing the reactants together, you will then use a Bunsen burner to heat the contents of the reaction "container". When should you stop heating?
- Once heating is complete, you will then weigh the substance that remains in the reaction "container". What is the name of this substance?
- The mass of this substance could be described as your (circle one): theoretical yield/experimental yield/percent yield

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6.R: Mole Ratios and Reaction Stoichiometry (Report)

Reaction A: Sodium Bicarbonate and Hydrochloric Acid

Experimental Data

Mass of evaporating dish + watch glass (g)	
Mass of evaporating dish + watch glass + Sodium bicarbonate	
Mass of sodium bicarbonate used	
Mass of evaporating dish + watch glass + sodium chloride	
Mass of sodium chloride collected (experimental yield)	

Data Analysis

- Use your data to determine the experimental mole-to-mole ratio between sodium bicarbonate and sodium chloride. Show your work for each step.
 - Convert the mass of sodium bicarbonate used to moles.
 - Convert the mass of sodium chloride collected to moles.
- Divide both of your results from the preceding two steps by the lower mole value to determine the simplest mole-to-mole ratio between sodium bicarbonate and sodium chloride.
 - Simplest mole ratio before rounding _____ moles NaHCO_3 : _____ moles NaCl
 - Simplest whole number ratio after rounding _____ moles NaHCO_3 : _____ moles NaCl
- Determine your percent yield of sodium chloride in reaction 3. Show your work for each step.
- Use mass-to-mass stoichiometry, calculate the theoretical yield of NaCl in reaction 3. Use your initial mass of sodium bicarbonate reactant as starting point, along with the relevant mole ratio from the balanced equation to perform this calculation.
- Is your percent yield here for reaction 3 greater than or less than 100%? Give one possible source of error that could explain the percent yield you obtained.

Reaction B: Sodium Carbonate and Hydrochloric Acid

Experimental Data

Mass of evaporating dish + watch glass (g)	
Mass of evaporating dish + watch glass + Sodium bicarbonate	
Mass of sodium bicarbonate used	
Mass of evaporating dish + watch glass + sodium chloride	
Mass of sodium chloride collected (experimental yield)	

Data Analysis

- Use your data to determine the experimental mole-to-mole ratio between sodium carbonate and sodium chloride. Show your work for each step.
 - Convert the mass of sodium carbonate used to moles.
 - Convert the mass of sodium chloride collected to moles.
- Divide both of your results from the preceding two steps by the lower mole value to determine the simplest mole-to-mole ratio between sodium carbonate and sodium chloride.
 - Simplest mole ratio before rounding: _____ moles Na_2CO_3 _____ moles NaCl
 - Simplest whole number mole ratio after rounding: _____ moles Na_2CO_3 _____ moles NaCl

3. Determine your percent yield in reaction 4. Show your work for each step. Write the balance equation for reaction 4 the reaction between sodium carbonate and hydrochloric acid.
4. Using mass-to-mass stoichiometry, calculate the theoretical yield of NaCl for reaction 4. Use your initial mass of sodium carbonate reactant as a starting point, along with the relevant mole ratio from the balanced equation to perform this calculation.
5. Is your percent yield here for reaction 4 greater than or less than 100%? Give one possible source of error that could explain the percent yield you obtained.

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CHAPTER OVERVIEW

7: Osmosis and Dialysis

[7.1: Osmosis and Dialysis Lab Procedure](#)

[7.2: Osmosis and Dialysis Pre-lab Assignment](#)

[7.R: Osmosis and Dialysis \(Report\)](#)

Adapted from

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“Biology OER: Osmosis and Diffusion” by The OpenLab at City Tech is licensed under CC BY-NC-SA 3.0

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7.1: Osmosis and Dialysis Lab Procedure

Learning Objectives

- To introduce the concept of solute concentration and osmosis and dialysis.
- Using different solute concentrations observe the process of osmosis.
- Use dialysis to separate out glucose and chloride ions from a mixture of chloride ions, glucose, and starch.

Background

Health care professionals work with a variety of solutions, the most common type probably being intravenous (IV) solutions. A solution is a homogeneous mixture of one or more solutes dissolved in a solvent. In a solution, the substance that is present in the greater amount is called the solvent, while the substance (or substances) that is present in the smaller amount is called the solute. Although most common solutions are liquids, and the most common solvent is water, solutions can be made from solvents and solutes that are liquids, solids or gases. Body fluids are examples of liquid solutions while the air we breathe is an example of a solution containing gaseous components.

Diffusion

Diffusion is the process in which a substance moves from an area of high concentration to an area of lower concentration. It is important for membranes to be semi-permeable. This difference in concentration is referred to as a concentration gradient. This movement does not require any external energy, but uses the free energy intrinsic to the system.



Diffusion

Figure 7.1.1: Concentrated dye diffuses along the concentration gradient until reaching equilibrium (no net movement).

If membranes were universally permeable, they would not legitimately serve their purpose as membranes; certain substances need to be kept out of a cell, and others kept in.

If membranes were not at all permeable, there would be no interface between the cell and its environment—effectively starving the cell. Membranes, being selectively permeable, allow in nutrients and other necessary substances, and also provide for the purging of cell waste.

Osmosis

Osmosis is a special case of diffusion. Instead of observing the net change in solute, osmosis follows the net movement of solvent across a semipermeable membrane. Since a semi-permeable membrane permits specific things to pass through, some solutes are partitioned.

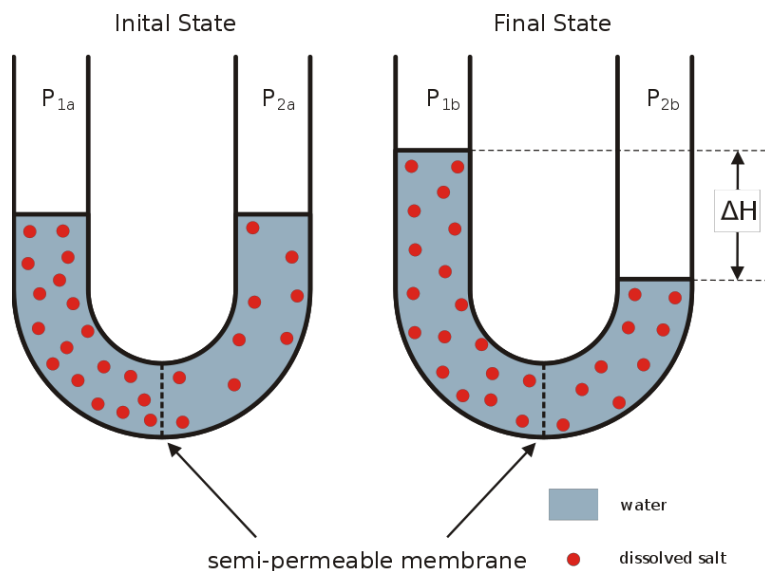


Figure 7.1.2: A semi-permeable membrane allows the solvent to pass but not this red salt molecule. The water moves along the concentration gradient (of water). This movement of water causes an osmotic pressure.

A cell lacking a cell wall is affected greatly by the tonicity of the environment. In a hypertonic solution where the concentration of dissolved solute is high, water will be drawn out of the cell. In a hypotonic solution where the concentration of dissolved solute is lower than the interior of the cell, the cell will be under great osmotic pressure from the environmental water moving in and can rupture.

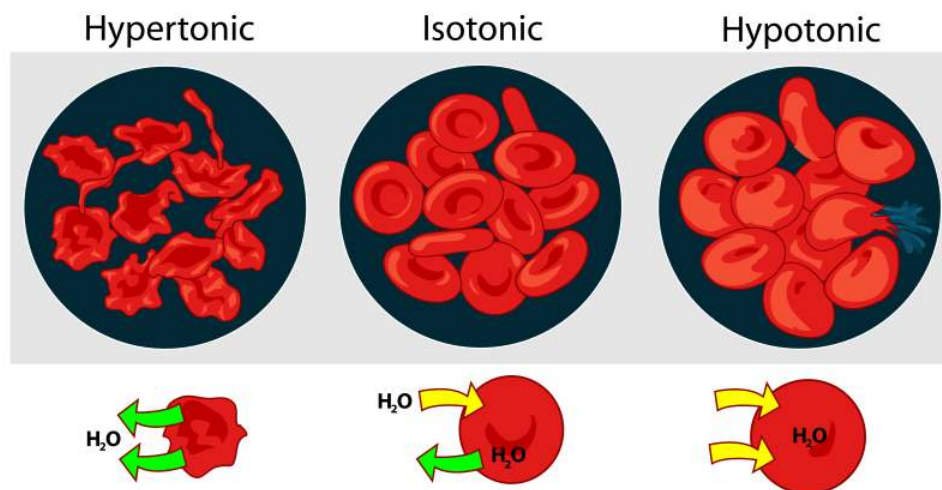


Figure 7.1.3

Plants have rigid cell walls composed of cellulose. These cell walls permit for maintenance of cellular integrity when the external environment is hypotonic (less dissolved substances). In this situation, the water moves into the cell. Without the cell wall, the cell would burst open from the excessive water pressure entering the cell. This state of swelling is referred to as turgid, resulting from turgor pressure.

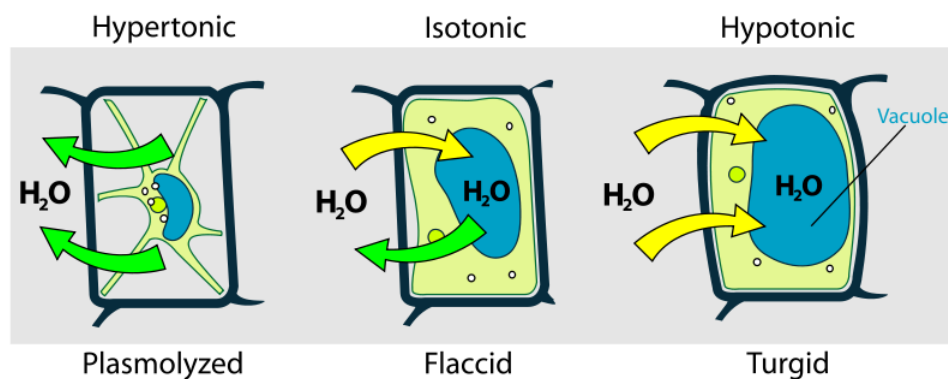


Figure 7.1.4: Cell walls of a plant retain the shape of the cell despite the state of external tonicity.

When the exterior environment is hypertonic, (greater amount of dissolved substances), the reverse condition occurs whereby the cellular fluid exiting the cell reduces the size of the cytoplasm. This condition is referred to as plasmolysis.

Dialysis

Dialysis is the separation of colloids from dissolved ions or molecules of small dimensions, or crystalloid, in a solution. A colloid is any substance that is made of particles that are of an extremely small size: larger than atoms but generally have the size of 10^{-7} cm ranging to 10^{-3} cm. A crystalloid is a substance that has some or all of the properties of a crystal or a substance that forms a true solution and diffuses through a membrane by dialysis. Dialysis is a process that is like osmosis. Dialysis is possible because of the unequal rates of diffusion through a semipermeable membrane.

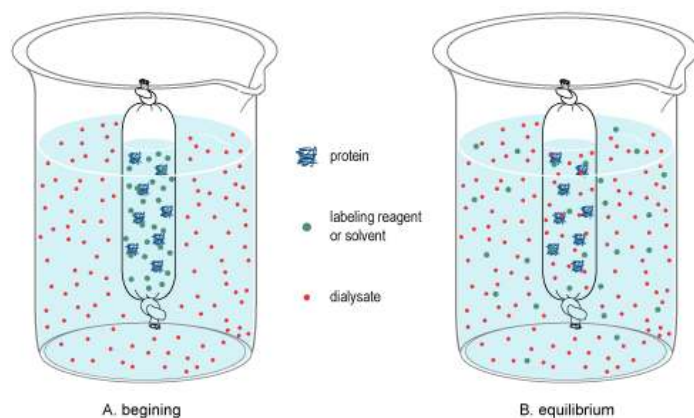


Figure 7.1.5: Graphic showing the diffusion of solutes across a membrane during dialysis. Image used with permission from wikipedia (Potcherboy)

This experiment will introduce you to the concept of solute concentration, osmosis and dialysis. You will be given solutions with different solute concentrations which you will use to observe the process of osmosis. Finally, dialysis is used to separate out glucose and chloride ions from a mixture of chloride ions, glucose, and starch.

Experimental Procedure

Safety

WEAR YOUR SAFETY GOGGLES

Materials: 3 dialysis tubing, three 250 mL beakers, pipette, distilled water, thread, 10% (m/v) sodium chloride solution, and 30% sodium chloride solution.

Part A. Osmosis

1. You will be given three pieces of dialysis tubing of equal length (8-10 cm). Soak them in a beaker with water for about 10 min to hydrate them. Tie a knot at the end of each bag and fill half with 10% (m/v) sodium chloride solution using a pipette. Gently squeeze out any air above the liquid and close bag with another knot. Rinse bags with tap water, pat dry using a paper towel to remove any liquid from the outside of the bag and weigh the bag. Record the weight of each bag. **DO NOT MIX UP BAGS!**
2. Label three labeled beakers and fill with (a) 50 mL of distilled/deionized water, (b) 50 mL of 10 % (m/v) sodium chloride, and (c) 50 mL of 30% sodium chloride. Be sure to keep track of the bags placed in each solution by labeling the beakers. Allow the bags to remain undisturbed in each solution for 30 min. Gently swirl the beaker every 5 minutes. At the end of the 30 minutes quickly remove each bag, rinse the bag well with distilled water, gently pat the bag completely dry to remove any liquid from the outside the bag and reweigh the bag. Record your results for each bag.

Part B. Identification Tests for Chloride, Glucose and Starch

Materials: 10% NaCl solution, AgNO₃ solution, 10% glucose solution, 1% starch solution, Benedict's reagent, DI water, test tubes.

1. Chloride test: (wear gloves for this part of the experiment!) Place 5-6 drops of 10% NaCl in a test tube and 5-6 drops of deionized (DI) water in another test tube. The test tube with water will be the control or comparison sample. Test for Cl⁻ by adding 1 drop of AgNO₃ to each. Record your observations. Compare the results for the NaCl solution with the results of the water sample.
2. Glucose test: Place 1.0 mL of 10% glucose solution in a test tube. Place 1.0 mL of DI water into another test tube. Add 1.0 mL of Benedict's reagent to the glucose and water. Heat both tubes in a warm water bath for 5 minutes. Record your observations. Remove tube from the water, label them and keep on table for comparison reasons. Note: Each time you carry out the glucose test, the tube containing Benedict's reagent must be heated in a boiling water bath.
3. Starch test: Place 1.0 mL of 1% starch solution in one test tube and 1.0 mL of DI water in another test tube. Add 1 drop of iodine reagent to each. Record your observations. Compare the results for the starch sample with the results for the water sample.

Part C. Dialysis

Materials: 10% NaCl solution, AgNO₃ solution, 10% glucose solution, Benedict's reagent, dialysis tubing, thread, 100 mL beakers, distilled water.

In a test tube combine approx. 1 mL of 10% NaCl, 1 mL of 10% glucose, and 1 mL of 1% starch solution. Soak a dialysis tube in a 150 mL beaker with distilled water for about 10 min to hydrate it. Tie a knot at one end of a 4-5 cm long piece of dialysis tubing. Fill the dialysis bag with the mixture, and tie a knot at the other end, rinse bag with water to remove any of the mixture that may be on the outside of the bag, pat dry and place in a beaker with about 30 mL of distilled water, making sure bag is submerged.

1. Immediately remove 3 mL of the water in the beaker (surrounding the dialysis bag). Divide the sample into 3 test tubes (about 1.0 mL each). Repeat each of the three identification tests that you did in part B on these three samples (one test per sample). A substance is present in the sample if a test for that substance gives the same results as in the identification tests in part B. Record the result as positive (+). If the substance is absent (there is no chemical reaction), record the test as negative (-).
2. Repeat the procedures in part B after 15 minutes of placing the bag in the beaker.
3. Repeat the procedures in part B after 30 minutes of placing the bag in the beaker.
4. After the 30 minutes test, carefully open the bag and remove the contents and perform the three identification tests for chloride, glucose and starch.

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7.2: Osmosis and Dialysis Pre-lab Assignment

Questions

1. What is the difference between osmosis and dialysis?
2. Why is it important for membranes to be selectively permeable?

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7.R: Osmosis and Dialysis (Report)

Experimental Results

Part A: Osmosis

	Distilled H ₂ O	10% NaCl	30% NaCl
Initial weight of dialysis bag (g)			
Final weight of dialysis bag (g)			
Change in weight (g)			

B. Identification Tests for Chloride, Glucose and Starch

Test Performed	Reagent Added	Observation with test solution	Observation with water (control)
Chloride			
Glucose			
Starch			

C. Dialysis

OBSERVATIONS FOR:			
Time	Chloride test	Glucose test	Starch test
0 minutes			
15 minutes			
30 minutes			
Content of bag			

Data Analysis/ Post-lab Questions

- Imagine that the dialysis bag you used in part A is a 'cell' and that the solution inside the bag is intracellular fluid.
 - Which solution is isotonic to the fluid inside the bag?
 - Which solution is hypertonic to the fluid inside the bag?
 - Which solution is hypotonic to the fluid inside the bag?
- What would have happened in each of the three beakers in part A if the bags were filled with distilled water instead of the 30% sodium chloride solution?
- In part C which substances were found in the solution outside the dialysis bag?
- How did these substances in question 3 go into the solution outside the bag?
- What substances were retained inside the bag in part C? Why were they retained?
- 0.9 % (m/v) NaCl and 5% (m/v) glucose solutions are said to be isotonic with respect to the intracellular fluid of red blood cells. A red blood cell placed in these solutions would undergo neither crenation nor hemolysis (i.e. no change in the red blood cell would occur). Given this information predict if a red blood cell would undergo crenation, hemolysis, or no change in the following solutions:
 - 2% (m/v) glucose
 - 2% (m/v) NaCl
 - 0.1% (m/v) NaCl
 - 10% (m/v) glucose

7. Are the solutions given above isotonic, hypotonic or hypertonic with respect to the intracellular fluid of a red blood cell?

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CHAPTER OVERVIEW

8: Titration of Vinegar

[8.1: Titration of Vinegar Lab Procedure](#)

[8.2: Titration of Vinegar Pre-lab Assignment](#)

[8.R: Titration of Vinegar \(Report\)](#)

Adapted from

“LibreTexts Online Chemistry Lab Manual: Proper Use of a Buret” by Oliver Seely (Professor of Chemistry, Emeritus; California State University Dominguez Hills) is in the public domain.

“LibreTexts Online Chemistry Lab Manual: Titration of Vinegar (Experiment)” by Santa Monica College is licensed under CC BY-NC-SA 3.0 US

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8.1: Titration of Vinegar Lab Procedure

Learning Objectives

- To determine the molarity and percent by mass of acetic acid in vinegar

Background

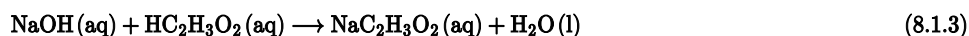
Vinegar is essentially a solution of acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) in water. The concentration of acetic acid in vinegar may be expressed as a molarity (in mol/L):

$$\text{Molarity} = \frac{\text{Moles of Acetic Acid}}{\text{Volume of Vinegar (in L)}} \quad (8.1.1)$$

or as mass percent

$$\text{Mass} = \frac{\text{Mass of Acetic Acid}}{\text{Mass of Vinegar}} \times 100 \quad (8.1.2)$$

In this experiment, a technique known as a titration will be used to determine the concentration of acetic acid in vinegar. A titration involves performing a controlled reaction between a solution of known concentration (the titrant) and a solution of unknown concentration (the analyte). Here, the titrant is an aqueous solution of $\sim 0.1 \text{ M}$ sodium hydroxide (NaOH) and the analyte is vinegar. When mixed, a neutralization reaction occurs between sodium hydroxide and the acetic acid in vinegar:



The sodium hydroxide will be gradually added to the vinegar in small amounts from a buret. A burette is a device that allows the precise delivery of a specific volume of a solution. The NaOH will be added to the vinegar sample until all the acetic acid in the vinegar has been exactly neutralized. At this point the reaction is completed, and no more NaOH is required. This is called the equivalence point of the titration.

In order to know when the equivalence point is reached, an indicator solution called phenolphthalein is added to the vinegar at the beginning of the titration. Phenolphthalein is a pH sensitive organic dye. Phenolphthalein is colorless in acidic solutions like vinegar, and deep pink in basic solutions like sodium hydroxide. At the equivalence point of the titration, just one drop of NaOH will cause the entire solution in the Erlenmeyer flask to change from colorless to a very pale pink.



In acidic solution



In basic solution



equivalence point

As the titration is performed, the following data will be collected:

- The molarity of NaOH (aq) used
- The volume of NaOH (aq) used to neutralize the vinegar
- The volume of vinegar used. Using this data, the molarity and mass percent of acetic acid in vinegar can be determined.

Experimental Procedure

Materials and Equipment

50-mL burette*, 5-mL volumetric pipette*, pipette bulb*, $\sim 0.1 \text{ M}$ NaOH (aq), vinegar, phenolphthalein, burette stand, two 250-mL (or 125 mL) Erlenmeyer flasks, wash bottle with distilled water, funnel

Safety

Be especially careful when handling the sodium hydroxide base (NaOH), as it is corrosive and can cause chemical burns to the skin. If any NaOH spills on you, rinse immediately under running water for up to 15 minutes and report the accident to your

instructor. Make certain to wear safety glasses when handling NaOH when doing the experiment!

Titration Procedure

Your instructor will demonstrate the correct use of the volumetric pipette and burette at the beginning of the lab session. Detailed instructions on how to use a pipette are also found on the last page of this handout. Note that three titrations must be performed.

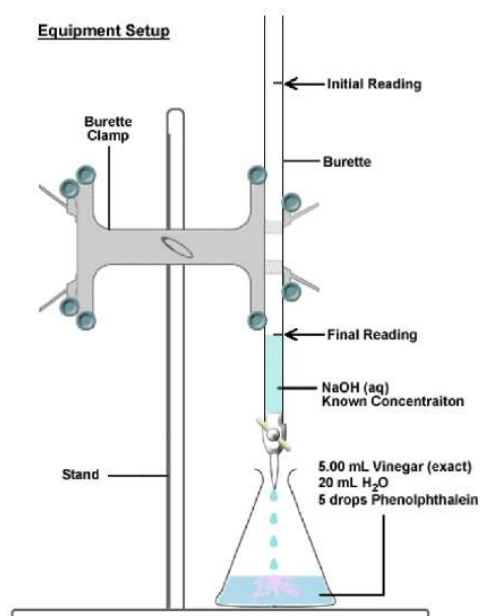
1. Obtain a 50-mL burette from the drying rack, 5-mL volumetric pipette and a pipette bulb from the fume hood.
2. Setting up the buret and obtain the NaOH
3. Rinse the inside of the buret with distilled water. Allow the distilled water to drain out through the tip in order to ensure that the tip is also rinsed.
4. Now rinse the burette with a small amount of NaOH (aq). To do this, add about 5-mL of NaOH (aq) to the burette, then twirl the burette on its side (over the sink) to rinse its entire inner surface. Then allow the NaOH (aq) to drain out through the tip and collect in a small beaker.
5. Fill the buret with NaOH (aq) up to the top, between 0-mL and 5-mL. Use a funnel to do this carefully, below eye-level, and preferably over a large beaker. After this you will need to flush the tip of the burette – your instructor will show you how to do this. Now measure the volume at the level of the NaOH precisely, and record it as the “Initial Burette Reading” on your report. Also record the exact molarity of the NaOH (aq), which is labeled on the stock bottle.

Preparing the vinegar sample

1. The volumetric pipette used in this lab is designed to measure and transfer exactly 5.00 mL of solution. First, rinse the inside of the volumetric pipette with distilled water. Using the pipette bulb, draw the water into the pipette up above the 5-mL mark, then allow it to drain out through the tip. You may want to do this several times for practice. Then perform a final rinse, but this time use vinegar.
2. Now use the volumetric pipette to transfer 5.00-mL of vinegar into a clean 250-mL Erlenmeyer flask (see instructions on page 4). Record this volume of vinegar on your report. Then add about 20-mL of distilled water and 2-3 drops of phenolphthalein to this Erlenmeyer flask.

Performing the titration

1. Begin the titration by slowly adding NaOH (aq) from the burette to the vinegar in the Erlenmeyer flask. Swirl Erlenmeyer flask as you add the base in order to efficiently mix the chemicals. Some pinkness may appear briefly in the flask as the base is added, but it will quickly disappear as the flask is swirled.
2. As the equivalence point is approached, the pink color will become more pervasive and will take longer to disappear. When this occurs, start to add the NaOH (aq) drop by drop. Eventually the addition of just one drop of NaOH (aq) will turn the solution in the Erlenmeyer flask a pale pink color that does not disappear when swirled. This indicates that the equivalence point has been reached. Do not add any more NaOH (aq) at this point. Measure this volume of NaOH (aq) precisely, and record it as the “Final Burette Reading” on your report.
3. Refill your burette with NaOH (aq), and then repeat this procedure for a second sample of vinegar, and then a third sample of vinegar. You do not need to flush the tip of the burette again. Note that if you use less than 25-mL of NaOH (aq) for the second titration, you do not need to refill the burette for the third titration; also, that you will need to clean out and re-use one of your Erlenmeyer flasks for the third titration. You and your partner(s) will take turns performing these titrations.
4. When finished, dispose of your chemical waste as instructed.



Pipeting instructions

1. Get the appropriate amount of the solution you wish to pipet in a clean, dry beaker. Never pipet directly out of the stock bottles of solution. This creates a contamination risk.
2. Insert the tip of the pipet into the beaker of solution so that it is about a quarter inch from the bottom. Be sure not to press the tip against the bottom of the container.
3. If you are right handed, hold the pipet in your right hand, leaving your index finger free to place over the top of the pipet. With your left hand, squeeze the pipet bulb. Press it firmly over the top of the pipet, but **DO NOT INSERT THE PIPET DEEP INTO THE BULB!**
4. Release the pressure on the bulb and allow the solution to be drawn up into the pipet until it is above the volume mark. Do not allow the solution to be sucked into the bulb itself.
5. Quickly remove the bulb and place your index finger firmly over the top of the pipet. Then remove the pipette tip from the beaker of solution.
6. Slowly roll your finger to one side and allow the liquid to drain until the bottom of the meniscus is aligned with the volume mark. With practice, you will be able to lower the liquid very, very slowly.
7. When the bottom of the meniscus is even with the volume mark, press your index finger firmly on the top of the pipet so no liquid leaks out. Touch the tip once to the side of the beaker to remove any hanging drops.
8. To transfer the solution, place the tip of the pipette against the wall of the receiving container at a slight angle. Then allow the liquid to drain from the pipet.
9. When the solution stops flowing, touch the pipet once to the side of the receiving container to remove any hanging drops. **DO NOT** blow out the remaining solution. The pipet has been calibrated to deliver the appropriate amount of solution with some remaining in the tip.

Calculations

Molarity of Acetic Acid in Vinegar

1. First, using the known molarity of the NaOH (aq) and the volume of NaOH (aq) required to reach the equivalence point, calculate the moles of NaOH used in the titration.
2. From this mole value (of NaOH), obtain the moles of HC₂H₃O₂ in the vinegar sample, using the mole-to-mole ratio in the balanced equation.
3. Finally, calculate the molarity of acetic acid in vinegar from the moles of HC₂H₃O₂ and the volume of the vinegar sample used.

Mass Percent of Acetic Acid in Vinegar

1. First, convert the moles of $\text{HC}_2\text{H}_3\text{O}_2$ in the vinegar sample (previously calculated) to a mass of $\text{HC}_2\text{H}_3\text{O}_2$, via its molar mass.
2. Then determine the total mass of the vinegar sample from the vinegar volume and the vinegar density. Assume that the vinegar density is 1.000 g/mL (= to the density of water).
3. Finally, calculate the mass percent of acetic acid in vinegar from the mass of $\text{HC}_2\text{H}_3\text{O}_2$ and the mass of vinegar.

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8.2: Titration of Vinegar Pre-lab Assignment

Questions

1. In this lab, you will perform a titration using sodium hydroxide and acetic acid (in vinegar). Write the balanced neutralization reaction that occurs between sodium hydroxide and acetic acid.
2. During this experiment, a variety of reagents will be used. Is sodium hydroxide the analyte or the titrant?
3. What volume of acetic acid is used?
4. Is the acetic acid the analyte or the titrant?
5. You will add sodium hydroxide to the acetic acid until all the acetic acid is consumed. This is a special point in the titration called the _____ point.
6. An indicator solution is used to indicate when all the acetic acid has been consumed and that the reaction is complete. What is the name of the indicator solution?
7. Is this indicator mixed with sodium hydroxide or acetic acid?
8. How exactly does the indicator let you know when the reaction is complete?

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8.R: Titration of Vinegar (Report)

Experimental Data

	Trial 1	Trial 2	Trial 3
Initial Buret Reading			
Final Buret Reading			
Volume of NaOH used			
Molarity of NaOH used			
Volume of vinegar used			

Data Analysis

- Write the balanced equation for the neutralization reaction between aqueous sodium hydroxide and acetic acid.

The Molarity of Acetic Acid in Vinegar

	Trial 2	Trial 2	Trial 3
Moles of NaOH used in titration			
Moles of acetic acid neutralized in vinegar sample			
Molarity of acetic acid in vinegar			
Average molarity			

- Determine the molarity of acetic acid in vinegar. Show all calculations for one trial below.

The Mass Percent of Acetic Acid in Vinegar

	Trial 1	Trial 2	Trial 3
Mass of HC ₂ H ₃ O ₂ in vinegar sample			
Mass of vinegar sample (assume density = 1.00 g/mL)			
Mass Percent of HC ₂ H ₃ O ₂ in vinegar			
Average Mass Percent			

- Determine the mass percent of acetic acid in vinegar. Show all calculations for one trial below.

Questions

- What was the purpose of the phenolphthalein indicator in this experiment? Be specific.
- Suppose you added 40 mL of water to your vinegar sample instead of 20 mL. Would the titration have required more, less or the same amount of NaOH (aq) for a complete reaction? Explain.
- Consider a 0.586 M aqueous solution of barium hydroxide, Ba(OH)₂ (aq). How many grams of Ba(OH)₂ are dissolved in 0.191 dL of 0.586 M Ba(OH)₂ (aq)?
- What volume (in L) of 0.586 M Ba(OH)₂ (aq) contains 0.466 ounces of Ba(OH)₂ dissolved in it?
- If 16.0 mL of water are added to 31.5 mL of 0.586 M Ba(OH)₂ (aq), what is the new solution molarity?

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CHAPTER OVERVIEW

9: Acid, Bases and pH

[9.1: Acid, Bases and pH Lab Procedure](#)

[9.2: Acid, Bases and pH Pre-lab Assignment](#)

[9.R: Acid, Bases and pH \(Report\)](#)

Adapted from

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9.1: Acid, Bases and pH Lab Procedure

Learning Objectives

- To set up and show how to use a pH indicator
- To determine the pH of common solutions
- To understand pH differences of acids and bases
- To learn to use a laboratory pH meter
- To understand relationship between pH and H⁺ ion concentration

Background

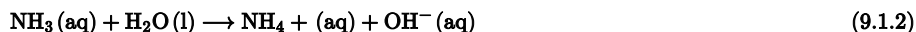
Definition of an acid & a base

An acid is a compound that supplies a hydrogen ion in a reaction. Conversely, a base is a compound that accepts a hydrogen ion in a reaction. Using the reaction with hydrochloric acid in water as an example to demonstrate the definitions of an acid and a base. Hydrochloric acid and water molecules reactants, while chloride ion and hydronium ion are products:



What has happened in this reaction is that the original hydrochloric acid molecule (the acid) has donated a hydrogen ion to the original water molecule (base), which in turn has accepted the hydrogen ion.

Another example is with ammonia (a weak base) in water:



In this reaction, the original water molecule (the acid) has donated a hydrogen ion to the original ammonia molecule (base), which in turn has accepted the hydrogen ion.

pH and Buffers

A pH value is a number, usually between 0 and 14, that represents the acidity, neutrality or basicity of a solution. The “pH” is always written with a lowercase “p” and an uppercase “H”, which stands for “power of hydrogen.” pH values are related to hydrogen ion (H⁺) concentrations. The mathematical relationship between pH and is described by the equation

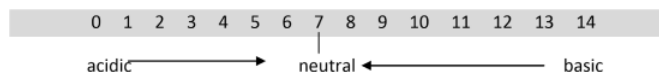
$$\text{pH} = -\log [\text{H}^+] \quad (9.1.3)$$

There is an inverse relationship between pH and H⁺ ion concentration (in brackets, expressed in units of molarity, M). As the H⁺ concentration decreases, the pH value increases, and vice versa. When the pH value is a whole number (e.g. pH 7), the number is equal to the negative exponent of the ion concentration.

$$\text{pH value} = X \rightarrow [\text{H}^+] = 10^{-X} \text{ M} \quad (9.1.4)$$

So for pH 7, the H⁺ ion concentration is 10⁻⁷ M.

The pH values of everyday chemicals typically range from pH 0 to pH 14. Values between 0 and 7 indicate an acidic solution. Values between 7 and 14 indicate a basic solution. A pH of exactly 7 indicates that a solution is neutral, neither acidic nor basic. Pure water is pH 7. Tap water and bottled water may have slightly different pH values, depending upon other components (metal ions, additives). The pH scale is shown below.



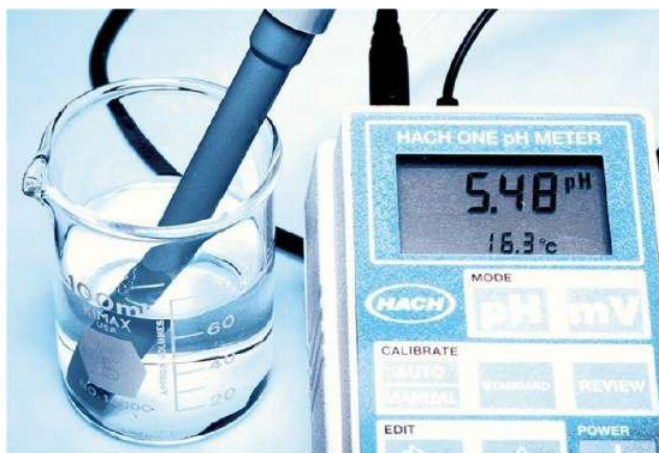
The lower the pH value, the more acidic the solution; the higher the pH value, the more basic the solution. Basic solutions are also called alkaline solutions. It should be noted that the pH scale does extend beyond 0 and 14. Strong laboratory acids typically have pH values less than 0 (negative pH values) and strong laboratory bases typically have pH values greater than 14. Thus, they are considerably more dangerous. For our purposes, we will focus on solutions that have pH values in the 0-14 range.

The concept of pH is widely used in all areas of science including agriculture, biology, engineering and medicine. Many commercial products use pH as an advertisement tool, such as shampoo and water; more recently, food and drink of certain pH has been touted as more healthful.

A pH indicator is a substance that, when a small amount of it is added to a solution of unknown pH, will change its color. This is a way to determine pH of a solution visually.

The indicator used in this lab will be obtained from a natural source, red cabbage. Cabbage indicator yields a particular color depending on the pH of the solution. pH indicators are a good way to easily and quickly show the approximate pH by color when compared to a standard. An everyday example where a pH indicator is used is for testing a water sample from a swimming pool.

While pH indicators are useful for qualitative purposes, when an exact quantitative value is needed, a pH meter is used. A laboratory pH meter typically has a special probe capped with a membrane that is sensitive to H^+ ion concentrations. The meter reading shows an exact pH value of the solution probed.



pH meters are used to measure pH values of water samples, such as determining acidity of rainwater samples. Rain water contains dissolved carbon dioxide that produces a weakly acidic solution. Rain naturally has a pH between 5 and 6. The pH of rain in parts of the U.S. is less than pH 5, which is harmful to aquatic life and human health. This is called acid rain.

Living organisms are very sensitive to the effects of acids and bases in their environment. An excess of H^+ or OH^- can interfere with the functioning of biological molecules, especially proteins. Thus, in order to maintain homeostasis and survive, organisms must maintain a stable internal pH.

A buffer is a solution whose pH resists change on addition of small amounts of either an acid or base. To be a good buffer, a solution should have a component that acts as a base (reacts with H^+) and a component that acts as an acid (reacts with OH^-).

The buffering capacity of solution is tested by adding small amounts of acid (for example, HCl) and base (for example, NaOH) and checking the pH after each addition. If the pH changes only slightly (≤ 1.0 pH unit), the solution is a good buffer. Eventually its buffering capacity will be exhausted, however, and the pH will change dramatically.

Experimental Procedure

Materials and Equipment

400-mL beaker, ring stand, wire gauze, Bunsen burner, large test tubes, dropper pipet, stirring rod, wash bottle with distilled water, pH paper, pH meter, 0.1 M acetic acid, 0.1 M sodium acetate, 0.1 M hydrochloric acid, 0.1 M sodium hydroxide, cabbage solution, buffer solution.

Safety

Exercise appropriate caution when using the Bunsen burner. Personal Protective Equipment (PPE) required: safety goggles, closed-toe shoes, gloves.

Part A: Preparing pH Indicator and pH Standards

1. Tear a few leaves of red cabbage into small pieces and place the leaves into a 500-mL beaker. Add about 500-mL of distilled water to this beaker. Make sure that all of the leaf pieces are completely submerged.
2. Gently boil the mixture on heating plate until it appears dark purple in color (5-10 min). Turn off the heat and allow to cool (5 min). {ALREADY PREPARED- OBTAIN FROM THE FUMEHOOD.}
3. Work in groups clean 13 test tubes and label 1-13.
4. Add 3 mL of each buffer solution in a separate tube to create a set with pH from 1-13.
5. Add 2 mL cabbage indicator solution to each test tube and mix, record the colors of the pH standards. SAVE PREPARED STANDARD FOR COMPARISON

Part B: Qualitative Analysis for pH Values of Everyday Chemicals

Materials: cabbage solution, shampoo, soda, mouth washFs, bleach, laundry detergent, lemon juice, vinegar, tap water, 0.1 M HCl, 0.1M NaOH

1. Each group will bring a small beaker to the fume hood and take ~50 mL of the cabbage indicator back to their bench.
2. Obtain 10 large test tubes (clean, but may be wet). Label each test tube with the solutions to be tested.
3. Pour about 3-mL of each solution into the appropriately labeled test tube.
4. Using a dropper pipet, add an equal volume of cabbage indicator solution. If necessary, stir to mix with a clean stirring rod (rinse with distilled water between uses).
5. Record the resulting color of the sample after mixed with the cabbage indicator. Compare this color with pH standards you prepared in part A to determine the pH of the sample. The color may be between the pH standard colors (e.g. green-blue instead of green or blue alone). For these, record the pH to 0.5 values (e.g. pH = 9.5 instead of 9 or 10).

Note

Do not discard the contents in these test tubes as they will be used in the next section.

Part C: Quantitative Analysis for pH Values of Everyday Chemicals

1. Plug the probe into one port on the side of the pH meter. Plug the AC adapter into the other side of the pH meter; plug the adapter into an electrical outlet. You should see a pH value reading.
2. Prepare the probe to make pH measurements: remove from the storage bottle and thoroughly rinse the lower section of the probe with distilled water/wash bottle. **Note:** Do not completely submerge the probe. The handle is not waterproof.
3. Use the same ten test tubes containing samples from Part B. Or, complete steps 1 and 2 of Part B to obtain ten samples for analysis. Insert the pH probe directly into each test tube. SPECIAL CARE IS NEEDED WHEN INSERTING THE PROBES INTO THE TEST TUBES. The probes must NOT touch the glass rim of the test tubes or the pH blub can easily be broken and the probe destroyed.
4. Record the pH value (to 0.01 pH) shown on the pH meter screen.
5. After each pH measurement, the probe must be thoroughly rinsed with distilled water.
6. When you are finished making measurements, rinse the probe with distilled water. Slide the cap onto the probe, and then screw the cap onto the storage bottle so the tip of the probe is immersed in the storage solution.

Part D: Effect of Buffers on pH

1. Obtain 4 large test tubes. Label the test tubes A, B, C, and D.
2. Add 10-mL of distilled water to tubes A and C.
3. Add a 5-mL quantity of both 0.1 M $\text{H}_2\text{C}_2\text{H}_3\text{O}_2$ (acetic acid) and 0.1 M $\text{NaC}_2\text{H}_3\text{O}_2$ (sodium acetate) to tubes B and D. This mixture of acetic acid and sodium acetate is a buffer solution. Stir to mix completely.
4. Using pH paper, determine the pH of the contents of each test tube (A-D). Use the stirring rod to dab a small drop of the solution to be tested onto a piece of pH paper. Then compare the color obtained to the pH scale to determine the pH value. Record these pH values.
5. Add 5 drops of 0.1 M HCl (hydrochloric acid) to test tubes A and B. Record the pH using pH paper.
6. Add 5 drops of 0.1 M NaOH (sodium hydroxide) to test tubes C and D. Record the pH using pH paper.

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9.2: Acid, Bases and pH Pre-lab Assignment

Questions

1. What is a pH indicator?
2. A solution that has a pH of 10 is considered?
3. A solution with a concentration of $1.2 \times 10^{-5}\text{M}$ has a pH of?

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9.R: Acid, Bases and pH (Report)

Experimental Results

Part A: Color of Red Cabbage Indicator With pH Standards

pH Standard	Color with Cabbage Indicator
1	
2	
3	
4	
5	
6	

pH Standard	Color with Cabbage Indicator
8	
9	
10	
11	
12	
13	

pH Standard	Color with Cabbage Indicator
7	

Parts B and C: pH of Everyday Chemicals

Chemical	Color with indicator	Qualitative pH	Acidic, basic, or neutral?	Quantitative pH
Soda				
Shampoo				
Mouth wash				
Bleach				
Laundry detergent				
Lemon juice				
Vinegar				
Tap water				
0.1 M HCl				
0.1 M NaOH				

Part D: Effect of Buffers on pH

Test Tube	Content	Initial pH	Chemical Added	New pH	pH Change
A	Water			HCL	
B	Buffered solution (acetic acid and sodium acetate)			HCl	
C	Water			NaOH	
D	Buffered solution (acetic acid and sodium acetate)			NaOH	

Post-Lab Questions

1. What is an acidic solution? What is a basic solution?
2. What is an alkaline solution?
3. What is a pH indicator? What are common uses of pH indicators?
4. Write the mathematical equation that relates pH value and H^+ ion concentration:
5. Circle correct choice: Acids have (high OR low) pH, and (high OR low) H^+ ion concentration. Bases have (high OR low) pH, and (high OR low) H^+ ion concentration.
6. When the H^+ ion concentration is expressed in brackets $[H^+]$, what are the units of the for H^+ ion concentration?
7. Does a solution with pH 10 have equal, less or more H^+ ions than of a solution with a pH 6? Calculate the $[H^+]$ for both solutions, include units in your answer:
 1. For pH 10, $[H^+] =$ _____
 2. For pH 6, $[H^+] =$ _____
8. The two methods of determining pH values (pH indicator versus pH meter) should show similar pH values for those solutions. What was different?

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CHAPTER OVERVIEW

10: Properties of Alcohol, Aldehydes and Ketones

[10.1: Properties of Alcohol, Aldehydes and Ketones Lab Procedure](#)

[10.2: Properties of Alcohol, Aldehydes and Ketones Pre-lab Assignment](#)

[10.R: Properties of Alcohol, Aldehydes and Ketones \(Report\)](#)

Adapted from

“LibreTexts Online Chemistry Lab Manual: Aldehydes and Ketones” by Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, Adam Hahn, and ChemPRIME at Chemical Education Digital Library (ChemEd DL) is licensed under CC BY-NC-SA 3.0 US

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10.1: Properties of Alcohol, Aldehydes and Ketones Lab Procedure

Learning Objectives

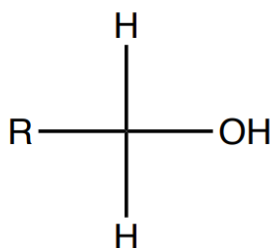
- Study the chemical properties of alcohols, aldehydes and ketones.
- Perform solubility test in water, and organic solvents
- Test the reactivity with various chemical reagents.

Background

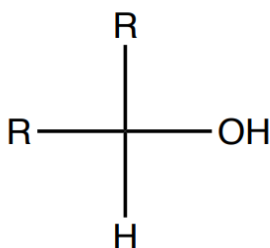
Structure of Alcohols

Alcohols are organic compounds containing a hydroxyl (R-OH) functional group bonded to a carbon atom that is not bonded to a carbonyl carbon (C=O). If the hydroxyl group is bonded to an aromatic ring (benzene ring), a class of compounds called phenols are formed, which have properties different than regular alcohols, and will not be used in this experiment.

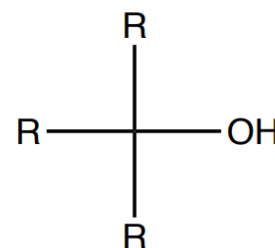
Alcohols can be separated into three subclasses, primary (1°), secondary (2°) and tertiary (3°) based on the number of alkyl (R-) groups attached to the carbon atom with the hydroxyl (OH) group attached.



Primary alcohol



Secondary alcohol



Tertiary alcohol

Physical Properties of Alcohols

The most common alcohols are colorless liquids at room temperature. Their melting and boiling points are considerably higher than those of alkanes, alkenes, and alkynes of similar size due to the ability to hydrogen bond. The ability to form hydrogen bonds also makes low molar mass alcohols soluble in water. However, as the alkane portion of the molecule increases in size the solubility decreases, because the hydrogen bonds formed by the alcohol group cannot counteract the nonpolar alkane part.

Chemical Properties of Alcohols

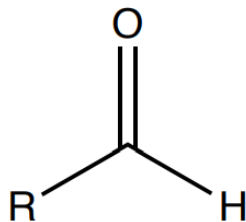
Chemically alcohols undergo two main categories of reactions: oxidation and dehydration. There are several common oxidizing agents (compounds that will oxidize organic molecules). These include:

- Potassium Permanganate
- Chromic Acid
- Tollens reagent
- Fehling's/Benedict's reagent

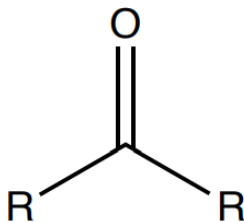
Due to the complex reaction mechanisms and balancing required, we will focus on the organic compounds formed in the reactions only. Depending on the classification of alcohol (primary, secondary, tertiary) they will be oxidized to different classes of molecules. The reaction results in the replacement of the OH group by either an Aldehyde, Carboxylic Acid or Ketone group.



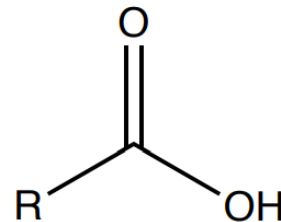
Tertiary alcohol $\xrightarrow{[O]}$ NR (10.1.3)



Aldehyde



Ketone



Carboxylic acid

Structure of Aldehydes and Ketones

Aldehydes and Ketones are organic compounds containing a carbonyl carbon functional group. Carboxylic Acids and Esters also contain a carbonyl carbon, and will be explored in a future experiment. The carbonyl carbon is a polar group with the carbon having a slight excess of positive charge and the oxygen atom having a slight excess of negative charge.

Chemical Properties

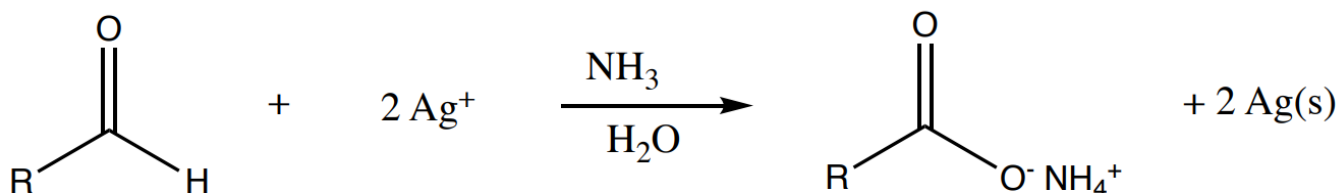
Aldehydes and ketones are created by the mild oxidation of primary and secondary alcohols. One such method to oxidize alcohols is with copper (II) oxide. Upon heating, copper wire (Cu) in an open flame leads to the formation of copper (II) oxide. The copper (II) oxide is then reacted with an alcohol to form an aldehyde or ketone, copper (I) oxide and water are products of the reactions.

Primary alcohol $\xrightarrow{[O]}$ Aldehyde (10.1.4)

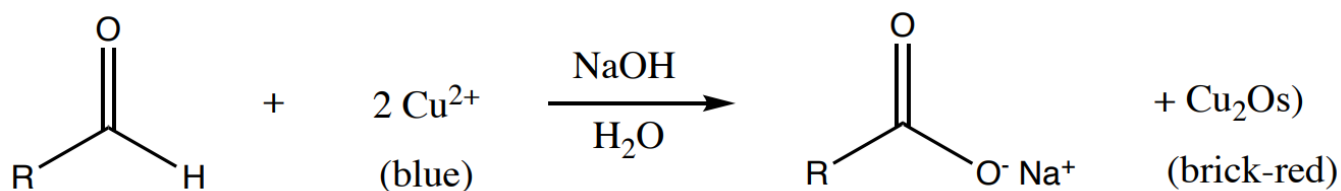
Secondary alcohol $\xrightarrow{[O]}$ Ketone (10.1.5)

Chemically aldehydes and ketones both contain a carbonyl carbon and thus have similar chemical reactivities. However, aldehydes are more susceptible to oxidation because of the hydrogen atom attached to the carbonyl group. This is the basis for distinguishing between these two classes of compounds. Several tests are useful for differentiating between aldehydes and ketones.

The first test is referred to as the Tollens' or the Silver Mirror test. The Tollens test uses Ag^+ ions to oxidize the aldehyde, as a result the silver ions are reduced to silver metal and either plated out on the surface of the reaction container, or forms a dark silver metal precipitate.

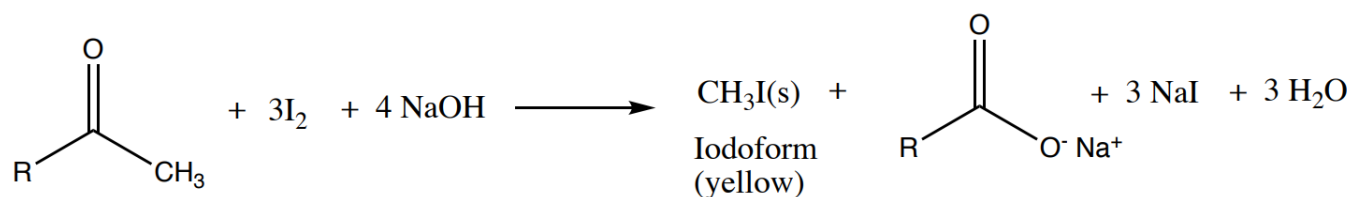


A second test for differentiating aldehydes and ketones is the Fehling's test. The Fehling's reagent uses Cu^{2+} ions to oxidize the aldehyde. The copper ions are reduced to form copper (I) oxide a reddish precipitate.



A stronger oxidant such as chromic acid which was used previously to oxidize alcohols can also be used to oxidize aldehydes. The chromic acid in the Bordwell-Wellman reagent is an orangeyellow solution that is reduced to Cr^{3+} which is a green color.

The last test is very specific for methyl ketones which are reacted with iodine (I_2) in the presence of a strong base to form iodoform (iodomethane) which is insoluble in water and forms yellow crystals. The resulting iodoform has an odor similar to chloroform.



Experimental Procedures

Safety

WEAR YOUR SAFETY GOGGLES and GLOVES

Materials

Test tubes, test tube rack, disposable pipettes, ethanol, 1-butanol, 2-butanol, t-butyl alcohol, cyclohexanol, 1-hexanol, octanol.

Properties of alcohol

Solubility Tests - Water

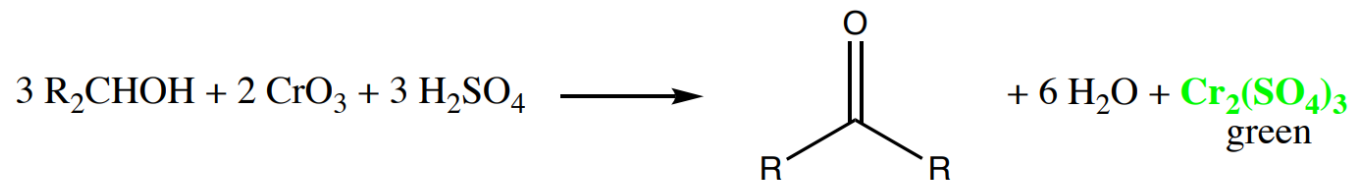
Test the solubility of each of the listed substance with water by adding 10 drops of the substance to be tested to 3 mL of water in a test tube. Mix each of the test tubes vigorously for 15 seconds. Wait 30 seconds. Record your observations. Note which pairs are miscible and which are not.

Repeat the above test on your unknown sample. Dispose of the solutions in the waste bottle indicated by your instructor.

Oxidation of Primary and Secondary Alcohols

Jones Oxidation

Materials: test tubes, tube rack, DI-water, ethanol, 1-butanol, 2-butanol, t-butyl alcohol, cyclohexanol, Jones reagent

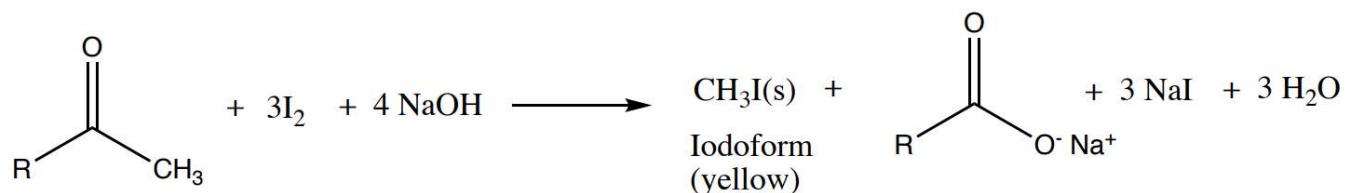


In a test tube place 15 mg or 10 drops of each alcohol and add 2 drops of Jones reagent (chromic acid in sulfuric acid) to the tube. Observe the color change as you add the chromic solution to the test tube. A color change within 15 seconds indicates a positive test for a primary or secondary alcohol. No color change indicated no reaction has occurred. Note that aldehydes also give a positive test, but tertiary alcohols do not. The Jones reagent will already be prepared for you. Record your observations.

Repeat the above test on your unknown sample. Record your observations.

Iodoform Test for Methyl Ketones

Materials: test tubes, tube rack, 3 M sodium hydroxide, warm water bath, iodine solution, acetone, butanal, 2-butanone, cyclohexanone, benzaldehyde, stoppers



Obtain and label six test tubes place 2 mL of water in each of the six separate test tubes. Add 4 drops or 20 mg of solid to the test tubes. Add 20 drops of 3 M NaOH to each, slowly add 40 drops of iodine solution. Stopper the test tube and shake vigorously, the formation of a yellow solid precipitate indicates a positive test. Record your observations.

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10.2: Properties of Alcohol, Aldehydes and Ketones Pre-lab Assignment

Questions

1. What are two safety precautions that must be followed while completing this experiment?
2. What is the evidence of a positive result in the Chromic Acid test?
3. What is the evidence of a positive result in the Iodoform test?

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10.R: Properties of Alcohol, Aldehydes and Ketones (Report)

Experimental Results

Solubility Test – Water

Mixture	Observation	Miscible/Immiscible	
		30 sec	5 mins
1-butanol			
2-butanol			
Ethanol			
2-methyl-2-propanol (tbutyl alcohol)			
Cyclohexanol			
Unknown _____			

1. Which compounds were completely soluble in water:
2. Which compounds were insoluble in water:
3. Comment on any trends noticed in your observations of solubility and insolubility in water.

Jones Oxidation for Primary and Secondary Alcohols

Mixture	Observation	Write the complete reaction
Control (H ₂ O)		
1-butanol		
2-butanol		
Ethanol		
2-methyl-2- propanol (t-butyl alcohol)		
Cyclohexanol		
Unknown _____		

1. What was the purpose of the control?
2. What is the evidence of a positive test?
3. Did any of the compounds tested behave unexpectedly? Explain. Why might this have occurred?

Properties of Aldehydes and Ketones

Solubility Test – Water

Compound	Structure	Solubility Obs.	Miscible/Immiscible
Butanal			
Benzaldehyde			
Acetone			
2-butanone (methyl ethyl ketone)			
Cyclohexanone			

1. What trends were observed in the solubility of aldehydes and ketones in water?

Oxidation Tests

Compound	Jones	Iodoform
Butanal		
Benzaldehyde		
Acetone		
2-butanone (methyl ethyl ketone)		
Cyclohexanone		

1. For the Jones test did any compounds tested not react as expected? Explain.

2. For the Iodoform test did any compounds tested not react as expected? Explain.

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CHAPTER OVERVIEW

11: Making Soap Saponification

[11.1: Making Soap Saponification Lab Procedure](#)

[11.2: Making Soap Saponification Pre-lab Assignment](#)

[11.R: Making Soap Saponification \(Report\)](#)

Adapted from

“LibreTexts Online Chemistry Lab Manual: Making Soap - Saponification (Experiment)” by Santa Monica College is licensed under CC BY-NC-SA 3.0 US

“The Basics of General, Organic, and Biological Chemistry v. 1.0: 17.2 Fats and Oils ” by David W. Ball, John W. Hill, and Rhonda J. published by Saylor Academy is licensed under CC BY-NC-SA 3.0

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11.1: Making Soap Saponification Lab Procedure

Learning Objectives

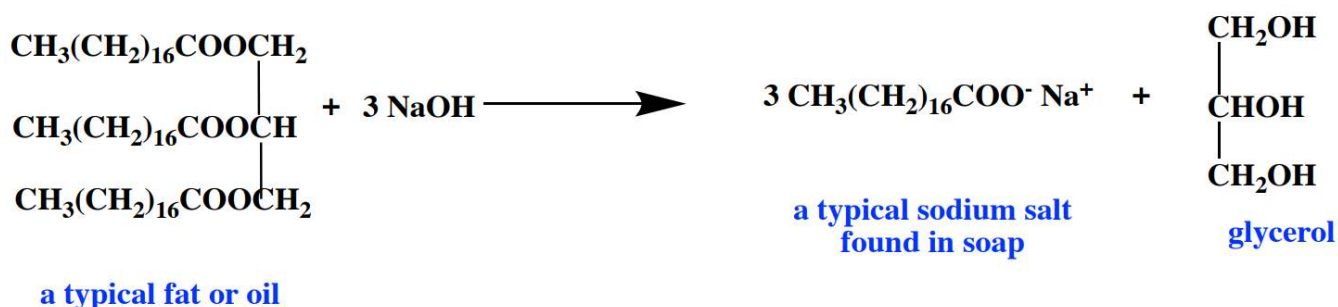
- The objective of this laboratory is to make lye soap via the saponification reaction.

Background

Soap making has remained unchanged over the centuries. The ancient Roman tradition called for mixing rain water, potash and animal tallow (rendered form of beef or mutton fat). Making soap was a long and arduous process. First, the fat had to be rendered (melted and filtered). Then, potash solution was added. Since water and oil do not mix, this mixture had to be continuously stirred and heated sufficiently to keep the fat melted. Slowly, a chemical reaction called saponification would take place between the fat and the hydroxide which resulted in a liquid soap. When the fat and water no longer separated, the mixture was allowed to cool. At this point salt, such as sodium chloride, was added to separate the soap from the excess water. The soap came to the top, was skimmed off, and placed in wooden molds to cure. It was aged many months to allow the reaction to run to completion.

All soap is made from fats and oils, mixed with alkaline (basic) solutions. There are many kinds of fats and oils, both animal and vegetable. Fats are usually solid at room temperature, but many oils are liquid at room temperature. Liquid cooking oils originate from corn, peanuts, olives, soybeans, and many other plants. For making soap, all different types of fats and oils can be used – anything from lard to exotic tropical plant oils.

Saponification Reactions



Soap can emulsify fats and oils by forming micelles around oil droplets. The soap molecules surround an oil droplet so that their nonpolar tails are embedded in the oil and their charged “head” groups are on the exterior of the droplets, facing the water. If the oil droplets are small enough and if there are enough soap molecules to surround them, the oil droplets become dispersed in the water and can then easily be washed away. Therefore, using lots of soap, hot water, and agitation can help clean greasy dishes. Hot water can melt solid fats, and agitation can help break up the fats and oils into smaller droplets.

Using lots of soap makes it more likely that there will be enough soap molecules to surround and emulsify all of the fat droplets. Soaps are less effective in hard water, which is water that contains a significant concentration of magnesium, calcium and iron ions. These ions form precipitates with soap molecules, and this precipitate is often seen as a gray line on a bathtub or sink and is often called “soap scum”. Since soap forms a precipitate with these ions, it means that many of the soap molecules are no longer present in the solution. Therefore, soap will form fewer suds in hard water. “Soft water” is water that contains very few or no ions that precipitate with soap. Soap will therefore be much more effective in soft water than in hard water.

Detergents are similar to soaps in that they have a charged head group and a long nonpolar tail group, but they are not prepared from natural fats or oils. Detergents are useful because they do not form precipitates with magnesium, iron or calcium ions, which means that they work in both soft and hard water. Shown below is a typical detergent molecule, sodium lauryl sulfate (which you may recognize from ingredient lists of shampoos or other cleaning products):

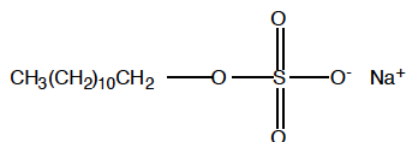


Figure 11.1.1: Sodium Lauryl Sulfate (a non-biodegradable detergent)

After detergents started being widely used, it was discovered that they were not broken down in sewage treatment plants. Many streams and lakes became contaminated with detergents and large amounts of foam appeared in natural waters. Biodegradable detergents were then developed. Shown below is an example of a biodegradable detergent, sodium laurylbenzenesulfonate.

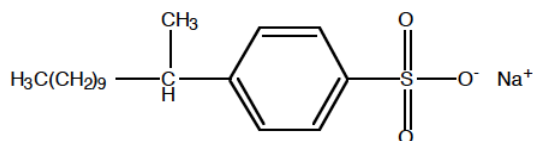


Figure 11.1.2: Sodium Laurylbenzenesulfonate (a biodegradable detergent)

Many commercial detergents also contain phosphate compounds. This can be a problem, because phosphate is a nutrient for plants. Too much phosphate in a pond, lake, or stream accelerates the growth of algae, which consumes too much of the dissolved oxygen in the water. This disturbs the ecosystem in the pond, and some organisms will die. Therefore, you will see some detergents these days that are labeled “phosphate free”. These are better for the environment than phosphate-containing detergents.

In this experiment, you will make soap from a fat or an oil by heating it with sodium hydroxide. You will precipitate the soap by adding it to a concentrated salt solution, and then you will collect the solid soap using vacuum filtration. You will then test the soap you made for its pH and foaming ability. You will test to see how well it emulsifies oil and you will also test its behavior in hard water. You will carry out the same tests on a commercially prepared soap solution and on a commercially prepared detergent solution, and you will compare your soap to the commercial soap and detergent.

Experimental Procedure

Safety

- Be sure to exercise caution when dispensing the 9 M NaOH. If the chemical comes into contact with your skin, immediately rinse with water for a minimum of fifteen minutes and notify your instructor.
- Wear safety goggles, closed-toe shoes and gloves

Materials and Equipment

Materials: olive oil, vegetable oil or coconut oil, 9 M sodium hydroxide solution, food coloring, assorted fragrances(optional), stearic acid, 0.5 M calcium chloride solution, 0.5 M Iron (III) chloride solution, and 0.5 M magnesium chloride solution.

Equipment: tall 250 mL beaker, glass stirring rod, glass pipets and pipet bulbs, pH paper, test tubes, tube rack, 10 mL graduate cylinder, magnetic stir bar and hot/stir plate.

Synthesis of Soap

1. Preheat either the olive oil, vegetable oil or coconut oil to 35°C in a 150-mL beaker. Pour 10 mL of the warm oil into a tall 250 mL beaker.
2. Prior to beginning the reaction, choose your fragrance (optional). You may choose one of the following: holiday candy, island coconut, lavender, cinnamon, vanilla.
3. Add 1-2 drops of desired fragrance, using the pipet provided at front bench; do not mix fragrances.
4. Add 3 mL of 20 % sodium hydroxide solution to the beaker. This is approximately two full dropper squirts.
5. Use the glass stirring rod to mix. You must stir for 20-45 minutes; you may choose to take turns with your lab partner. The mixture will slowly become smoother and opaquer; it should thicken to a pudding-like consistency.
6. After approval by your instructor, add 2-3 drops of desired food coloring. Stir.

7. Add a dash (approximately 1/8 teaspoon or the tip of a spatula) of stearic acid. This will serve as a hardener for the liquid soap. Stir.
8. Pour into chosen mold shape. Label with your names and lab section number.
9. After pouring into the mold (small beaker), the process will continue on its own. The soap will heat up and liquefy again, then cool off slowly, harden and dry. So, the soap must be left undisturbed for at least 12 hours. You will pick up your finished soap in lab next week.

Properties of Synthesized Soap

In this portion of the experiment you will be comparing the properties of the soap you made with a commercial soap and a detergent. You will start out by making three separate soap solutions, as follows:

1. Add about 40 mL of distilled water to 1 g of the soap you prepared. Heat the beaker on a hot plate at medium setting and gentle stir until the soap dissolves. Cool to room temperature. In the data table, describe the appearance of the soap solution.
2. Mix 1g of commercial soap with 40 mL of warm distilled water. Stir the solution gentle and record the appearance of the soap solution.
3. Add 1 g of commercial detergent to 50 mL of warm distilled water and stir the solution gently. Record the appearance in the data table.
4. Remove 5 mL of each solution into three different test tubes and measure the pH of the soap solution along with the pH of distilled water as a reference. Record the pH of the solutions in your data table. Keep the solution for the next step.
5. To the test tubes used in step 4 add 5 drops of oil. Stopper and shake each of the tubes continuously for 10 seconds. Allow the test tube to sit for 5 minutes and describe the observation of this emulsification test in the data table.
6. Place three test tubes in a test tube rack and label them. Pour about 4 mL of the soap solution into each test tube.
7. Add 10 drops of calcium chloride solution to the first tube, 10 drops of iron (III) chloride to the second tube and 10 drops of magnesium chloride to the third tube. Stopper and shake each tube to mix the contents. Describe the color and appearance of the mixture in each test tube.

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11.2: Making Soap Saponification Pre-lab Assignment

Questions

1. What are soaps made up of?
2. Identify the primary safety hazard in this experiment?
3. Describe what occurs when soap can emulsify fats and oils.

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11.R: Making Soap Saponification (Report)

Experimental Data

Part A: Synthesized Soap

1. Describe the color, texture and appearance of synthesized soap:
2. Does it smell like any soap that you have used?

Part B: Properties of Soaps and Detergents

	Synthesized Soap	Commercial Soap	Commercial Detergent
pH			
Emulsification test			
Ca ²⁺			
Fe ³⁺			
Mg ²⁺			

Post-lab Questions

1. The saponification reaction occurs between an ester and a base, shown in the figure in the procedure. In the reaction you performed, what is the ester and what is the base?
2. The base used in the saponification reaction must always contain a hydroxide ion. What bases are most commonly used for this reaction?
3. The products of the reaction are glycerol and a crude soap. The chemical formula of the soap is $\text{CH}_3(\text{CH}_2)_{14}\text{COONa}^+$. Draw the line-angle structure. Circle the portion of the molecule that is water-soluble and box the portion of the molecule that is fat-soluble. Why is the circle portion water-soluble? Why is the box portion fat-soluble?
4. During lab section, why did the saponification reaction require the long period of stirring?
5. Do you think that the type of fat used will make a difference in the product? Why or why not?

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CHAPTER OVERVIEW

12: Synthesis of Aspirin

[12.1: Synthesis of Aspirin Lab Procedure](#)

[12.2: Synthesis of Aspirin Pre-lab Assignment](#)

[12.R: Synthesis of Aspirin \(Report\)](#)

Adapted from

“LibreTexts Online Chemistry Lab Manual: Synthesis of Aspirin (Experiment)” by James Chickos, David Garin, and Valerian D'Souza. University of Missouri– St. Louis; Chemistry) and is licensed under CC BY-NC-SA 3.0 US

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12.1: Synthesis of Aspirin Lab Procedure

Learning Objectives

- Synthesize Aspirin

Background

Over history, many compounds obtained from nature have been used to cure ills or to produce an effect in humans. These natural products have been obtained from plants, minerals, and animals. In addition, various transformations of these and other compounds have led to even more medically useful compounds. During this semester, you will have an opportunity to isolate some pharmacologically active natural products and to synthesize other active compounds from suitable starting materials.

Analgesics are compounds used to reduce pain, antipyretics are compounds used to reduce fever. One popular drug that does both is aspirin. The Merck Index, which is an encyclopedia of chemicals, drugs and biologicals, lists the following information under aspirin: acetylsalicylic acid; monoclinic tablets or needle-like crystals; mp 135 °C (rapid heating); is odorless, but in moist air it is gradually hydrolyzed into salicylic and acetic acids; one gram dissolves in 300 mL of water at 25 °C, in 100 mL of water at 37 °C, in 5 mL alcohol, in 17 mL chloroform.

An aspirin tablet contains a small amount of aspirin (usually 300-400 mg) in a starch “binder” and sometimes contains other ingredients like caffeine and buffers. When aspirin is ingested, it is broken down to salicylic acid by the basic conditions in the small intestine. It is then absorbed into the bloodstream.

In this experiment, the salicylic acid is the limiting reactant and the acetic anhydride is in excess. After the reaction heating period is over, the excess unreacted acetic anhydride will be hydrolyzed by the addition of water to the mixture: water reacts with acetic anhydride to form 2 molecules of acetic acid.

The solid synthesized aspirin will be collected using vacuum filtration. Any other reaction ingredients that are soluble (this includes acetic acid, phosphoric acid, and water) will pass through the filter paper.

The collected aspirin will be tested for its purity using FeCl_3 (aq). Iron (III) ion reacts with phenols to form a purple complex. Salicylic acid contains a phenol group, but acetylsalicylic acid does not. Therefore, if you add FeCl_3 to an aspirin sample and you see a purple color, it means that there is still some salicylic acid present and the sample is impure.

The collected aspirin will then be purified by recrystallization. In this purification method, the crude aspirin will be dissolved in a small amount of warm ethanol. Water will then be added and the solution will be cooled slowly and then chilled. The acetylsalicylic acid will recrystallize, and the solid impurities (unreacted salicylic acid) should remain dissolved in the solution. The solid aspirin will again be collected using vacuum filtration and tested for purity. This aspirin should be purer than the original aspirin.

The final product will be dried and weighed and the theoretical and percent yields will be calculated.

Experimental Procedure

Materials and Equipment

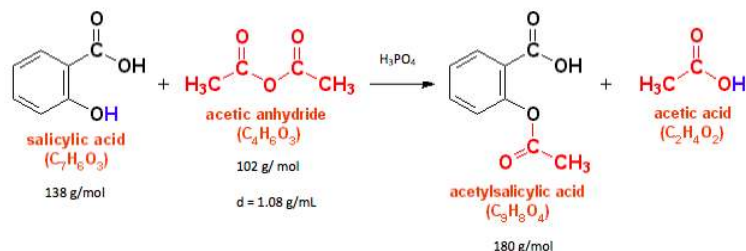
125-mL Erlenmeyer flask, 400-mL beaker, hot plate or Bunsen burner, ice, salicylic acid, acetic anhydride, 5- or 10-mL graduated cylinder, stirring rod, 600-mL beaker, dropper, 85% phosphoric acid (H_3PO_4), iron (III) chloride solution, ethanol, Buchner filtration apparatus, filter paper, spatula, watch glass, thermometer, ring stand and ring clamp, capillary tube, latex tubing/rubber band.

Safety

- Acetic anhydride is irritating to the nose and sinuses. Keep this compound under the hood at all times, and avoid breathing the vapors. Keep the acetic anhydride bottle capped after taking out the reagent.
- The aspirin that you make in this lab is NOT pure enough to be taken internally! Do not ingest the aspirin! Wear your safety goggles and gloves, especially when handling concentrated phosphoric acid.
- Waste Disposal:

- All waste must be placed in the organic waste containers in the fume hood.
- Be especially careful when handling the phosphoric acid and acetic anhydride.

SYNTHESIS OF ASPIRIN (acetylsalicylic acid)



1. Weigh out about 1 gram of salicylic acid on a piece of weighing paper. To do this, first weigh a piece of weighing paper. Place some salicylic acid on the weighing paper and weigh again. Add or remove solid until you have about 1 gram of it on the paper. Record the mass of the weighing paper plus the solid. Subtract to determine the mass of the salicylic acid. Place this solid into a 125-mL Erlenmeyer flask.
2. In the hood, measure out 3.0 mL of acetic anhydride in a small graduated cylinder and add it to the flask.
3. Add 3 drops of concentrated (85%) phosphoric acid. This will be the catalyst for the reaction.
4. Start heating the reaction, and stir gently with a stirring rod. Once the water bath starts boiling, start timing the reaction. When the mixture has been allowed to react at 100°C (the temperature of boiling water) for 15 minutes, you can consider the reaction to be complete. During the heating time, put 3 mL of water in each of two test tubes and chill these two test tubes in an ice bath. (These tubes of cold water will be used to rinse the solid aspirin after you collect it on the filter paper.)
5. After the reaction has gone on for 15 minutes at 100°C , add 1 mL of deionized water to the flask. This water will react with any excess acetic anhydride, converting it to acetic acid.
6. Remove the flask from the water bath. Add 9-10 mL of deionized water to the flask and swirl it around to mix it. As the flask cools, crystals of aspirin will start to form. When you see crystals, put the flask in an ice bath for 10 minutes. If acetylsalicylic acid does not begin to crystallize out, scratch the walls of the flask with a glass rod.

Vacuum filter the product using a Buchner funnel (see figure below for setup). You can use some of the filtrate to rinse the Erlenmeyer flask if necessary. After the liquid has gone through the funnel, pour one of the 3-mL portions of chilled rinse water over the sample to rinse it. When this water has gone through the funnel, rinse the solid with the second chilled portion of water. Leave your aspirin on vacuum for several minutes – this will draw air through the sample and will help dry the aspirin.

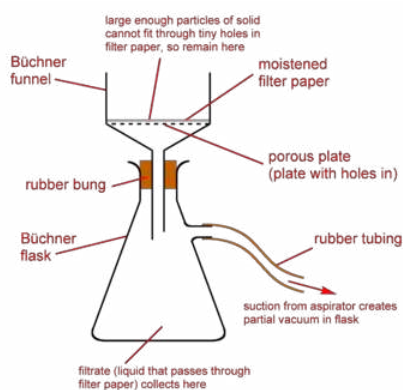


Figure 12.1.1: Collecting the crude Aspirin. Apparatus for vacuum filtration Image from: Wikipedia

Recrystallization of Aspirin

1. Set aside a small amount of the crude aspirin you obtained– you will test its purity later. (You will need enough to test its melting point and to test its reactivity with $FeCl_3$.)

2. Transfer the rest of the crude aspirin to a 50-mL Erlenmeyer flask. Add 4 mL of ethanol and warm the flask on a hot plate until all of the solid dissolves. Immediately remove the flask from the heat and slowly add 13 mL of cold water. Crystals should form. Chill this solution in an ice-water bath, and collect the crystals using vacuum filtration as you previously did.
3. Carefully lift the filter paper with the crystals on it and place it on a clean watch glass. Leave this aspirin in your bin to dry until the next laboratory period.
4. At the following laboratory period (when your aspirin is dry), weigh the aspirin on a piece of weighing paper or a weighing boat. (You will need to scrape it off of the filter paper.)

Ferric Chloride Test for Salicylic Acid

1. For the FeCl_3 test, the samples do not have to be dry. To do the test, get 4 test tubes. Place 1 mL of ethanol and 2 drops of FeCl_3 solution in each tube. Add a few crystals of salicylic acid to one test tube. Add a few crystals of your crude aspirin product to the second tube. In the third tube, place a few crystals of the recrystallized aspirin. Don't add anything else to the fourth tube – it will be your "blank". Shake each of the tubes and record your observations.

Melting point

1. Weigh the product, determine its melting point (lit mp 135-136 °C) and calculate the percentage yield. Dispose your sample in waste container under the fume hood.

Melting point instructions

1. Obtain a capillary tube from your instructor. Press the open end of the capillary tube into the powder. Then turn the tube over and tap the tube lightly against the lab bench to allow the powder to fall into the sealed end. Repeat until you have a depth of about 2-mm of solid in the tube.
2. Assemble your equipment as shown in the diagram in Figure 1.
3. Use a large 400-mL beaker half-filled with tap water for the hot water bath. Use a small piece of latex tubing like a rubber band to attach the capillary tube to your thermometer. The sealed end should be close to the bulb of the thermometer. Place a slotted stopper around the thermometer, and using a clamp, suspend it in the water bath. Slowly heat the water (about 3 degree/minute), using a hot plate. Carefully watch the solid. When the solid melts, note the temperature. Stir the bath continuously and watch your sample carefully.
4. The melting point is the temperature at which liquid first appears. Record your melting point.
5. Your aspirin sample and the capillary tube (together) should be disposed of in the labeled waste container provided when you are finished.

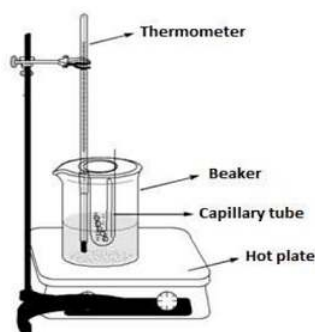


Figure 12.1.2: Melting point assembly

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12.2: Synthesis of Aspirin Pre-lab Assignment

Questions

1. Calculate the molar mass of salicylic acid and acetylsalicylic acid. (Show your calculations in order to receive full credit)
2. Calculate the theoretical yield of acetylsalicylic acid from 2.01 g of salicylic acid assuming excess acetic anhydride.
3. Give two safety precautions that must followed during this experiment

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12.R: Synthesis of Aspirin (Report)

Experimental Data

Mass of flask	
Mass of flask and salicylic acid	
Mass of watch glass	
Mass of watch glass and crude aspirin product	
Mass of watch glass and purified aspirin product	

Calculations

1. Mass of salicylic acid _____
2. Theoretical yield of aspirin _____ (show your work)
3. Mass of purified aspirin _____
4. Percent yield _____ (show your work)
5. Melting point (o C) of aspirin product (Optional) _____

Ferric Chloride Test

Tubes	Reagents	Observations
1	Reference (Ethanol)	
2	Salicylic acid	
3	Crude Aspirin	
4	Recrystallized aspirin	

1. Why is the aspirin washed with cold water?
2. According to your results from the ferric chloride test, what can you say about the purity of your aspirin? Be as specific as possible.
3. Aspirin that has been stored for a long time may give a vinegar-like odor and give a purple color with FeCl_3 . What reaction would cause this to happen?

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CHAPTER OVERVIEW

13: Detection of Carbohydrates

[13.1: Detection of Carbohydrates Lab Procedure](#)

[13.2: Detection of Carbohydrates Pre-lab Assignment](#)

[13.R: Detection of Carbohydrates \(Report\)](#)

Adapted from

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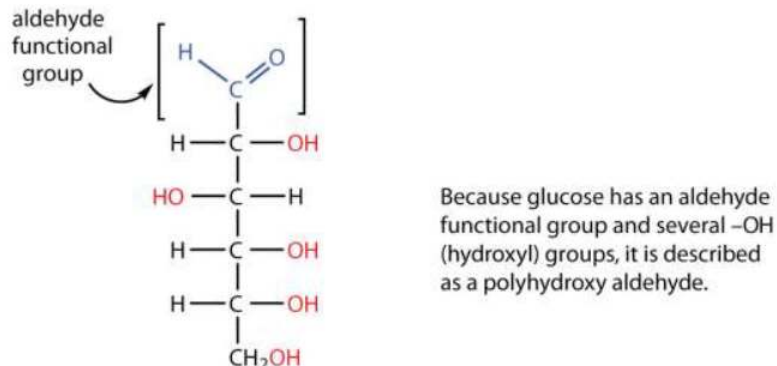
13.1: Detection of Carbohydrates Lab Procedure

Learning Objectives

- Distinguish between reducing and non-reducing sugars using the Benedict's test.

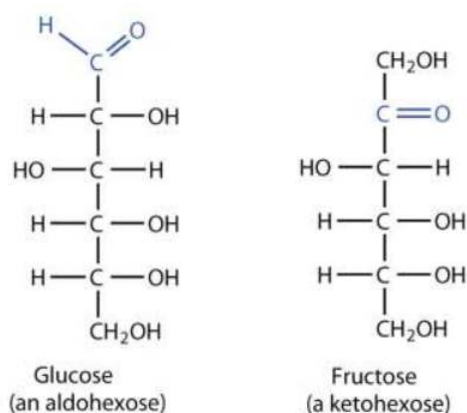
Background

All carbohydrates consist of carbon, hydrogen, and oxygen atoms and are polyhydroxy aldehydes or polyhydroxyl ketones are compounds that can be broken down to form them. Examples of carbohydrates include starch, fiber, the sweet-tasting compounds called sugars, and structural materials such as cellulose.



Classification of Monosaccharides

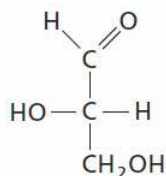
The simplest carbohydrates—those that cannot be hydrolyzed to produce even smaller carbohydrates—are called monosaccharides. Two or more monosaccharides can link together to form chains that contain from two to several hundred or thousand monosaccharide units. Prefixes are used to indicate the number of such units in the chains. The most important naturally occurring monosaccharides contain three to six carbon atoms per molecule. Monosaccharides of specific sizes may be indicated by names composed of a stem denoting the number of carbon atoms and the suffix -ose. For example, the terms triose, tetrose, pentose, and hexose signify monosaccharides with, respectively, three, four, five, and six carbon atoms. Monosaccharides are also classified as aldoses or ketoses. Those monosaccharides that contain an aldehyde functional group are called aldoses; those containing a ketone functional group on the second carbon atom are ketoses.



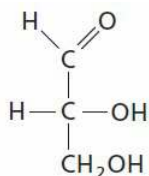
Although a variety of monosaccharides are found in living organisms, three of these hexoses are particularly abundant: D-glucose (found in fruit juices, honey and corn syrup), D-galactose (released when lactose is hydrolyzed), and D-fructose (found in honey, sweet fruits).

D- and L-glyceraldehyde provide reference points for designating and drawing all other monosaccharides. Sugars whose Fischer projections terminate in the same configuration as Dglyceraldehyde are designated as D sugars; those sharing the same

configuration as Lglyceraldehyde are designated as L sugars.



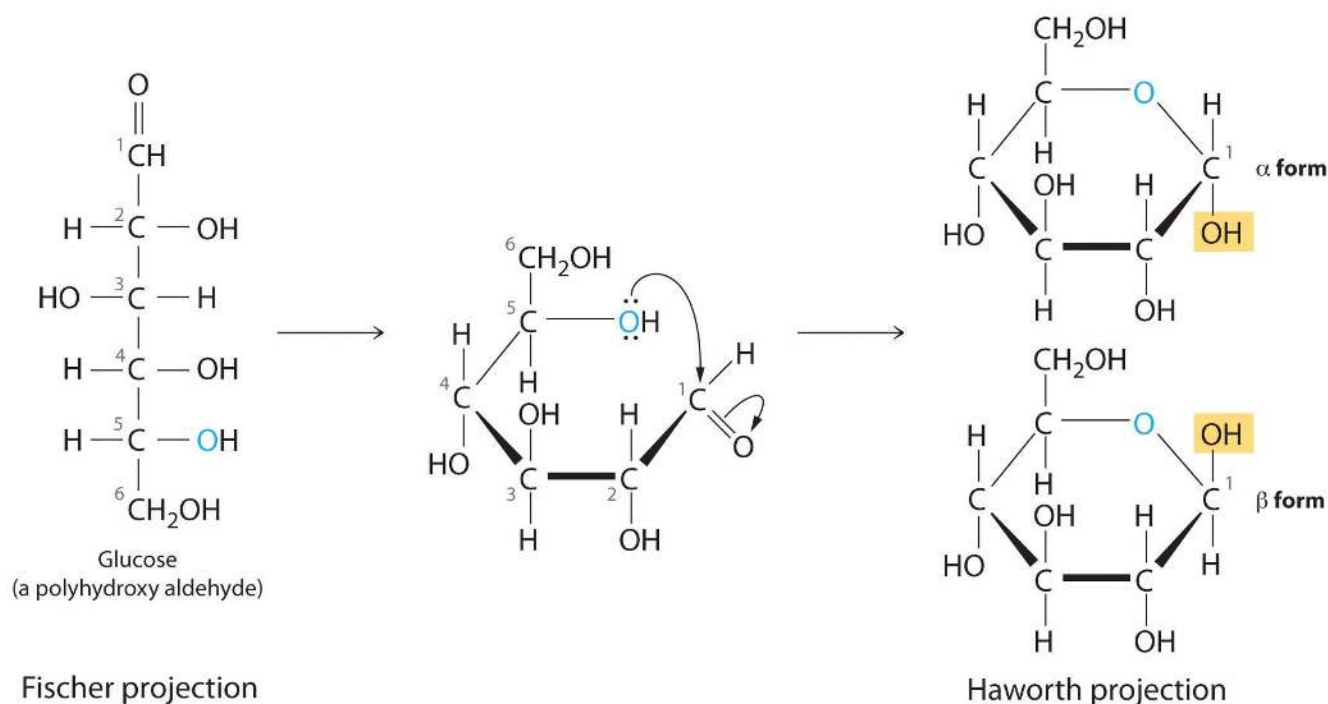
L-Glyceraldehyde



D-Glyceraldehyde

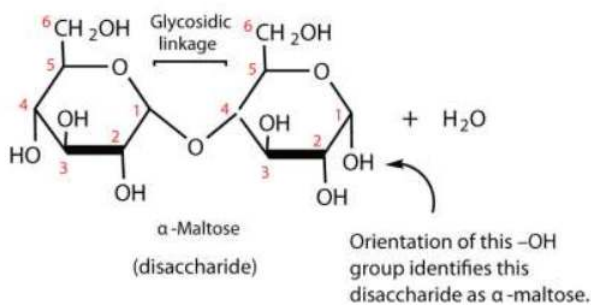
Haworth Structures

Glucose tends to exist in the cyclic form, which is depicted by Haworth structure. The structure shown on top with the OH group on the first carbon atom projected downward, represent what is called the alpha (α) form. The structures on the bottom, with the OH group on the first carbon atom pointed upward, is the beta (β) form.



Disaccharides

Disaccharides ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) are sugars composed of two monosaccharide units that are joined by a carbon–oxygen–carbon bond known as a glycosidic bond. Examples of disaccharides are maltose, sucrose (table sugar) and lactose (milk sugar).

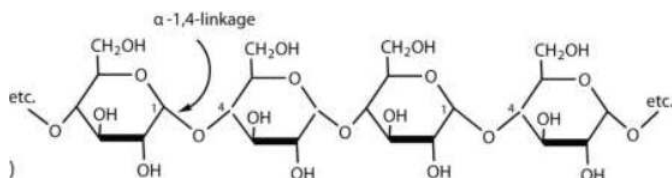


Polysaccharides

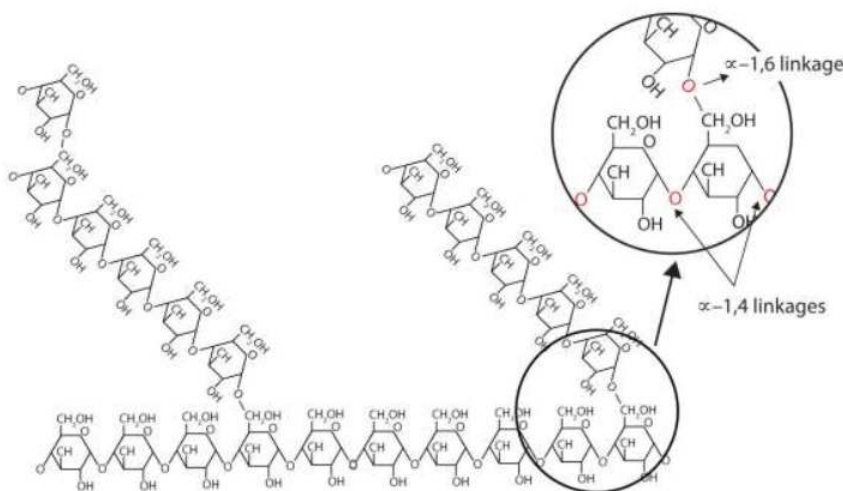
Polysaccharides are very large polymers composed of thousands of monosaccharides joined together by glycosidic bonds. The three most abundant polysaccharides are starch, glycogen, and cellulose.

Starch is the most important source of carbohydrates in the human diet and accounts for more than 50% of our carbohydrate intake. Starch is a mixture of two polymers: amylose and amylopectin. Natural starches consist of about 10%–30% amylose and 70%–90% amylopectin.

Amylose is a linear polysaccharide composed entirely of D-glucose units joined by the α -1,4- glycosidic bonds.

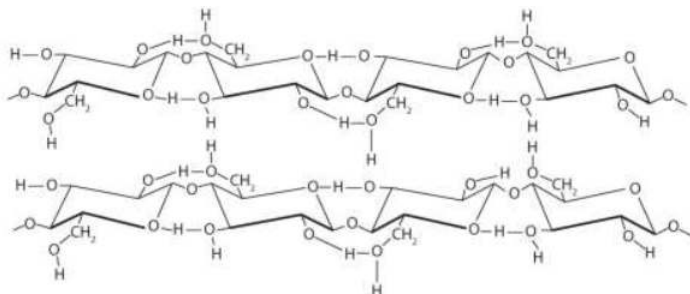


Amylopectin is a branched-chain polysaccharide composed of glucose units linked primarily by α -1,4-glycosidic bonds but with occasional α -1,6-glycosidic bonds, which are responsible for the branching.



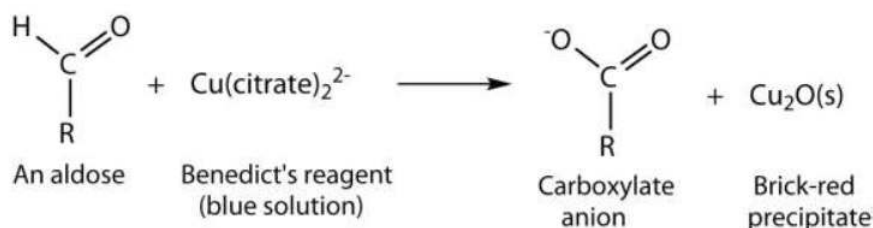
Cellulose

Cellulose, a fibrous carbohydrate found in all plants, is the structural component of plant cell walls. Because the earth is covered with vegetation, cellulose is the most abundant of all carbohydrates, accounting for over 50% of all the carbon found in the vegetable kingdom. Cotton fibrils and filter paper are almost entirely cellulose (about 95%), wood is about 50% cellulose, and the dry weight of leaves is about 10%–20% cellulose.



Benedict's Test for Reducing Sugars

The chemical behavior monosaccharides are determined by their functional groups. An important reaction of monosaccharides is the oxidation of the aldehyde group, one of the most easily oxidized organic functional groups. Aldehyde oxidation can be accomplished with any mild oxidizing agent, such as Tollens' reagent or Benedict's reagent. With Benedict's reagent, complexed copper (II) ions are reduced to copper (I) ions that form a brick-red precipitate of copper (I) oxide.



Any carbohydrate capable of reducing either Tollens' or Benedict's reagents without first undergoing hydrolysis is said to be a reducing sugar.

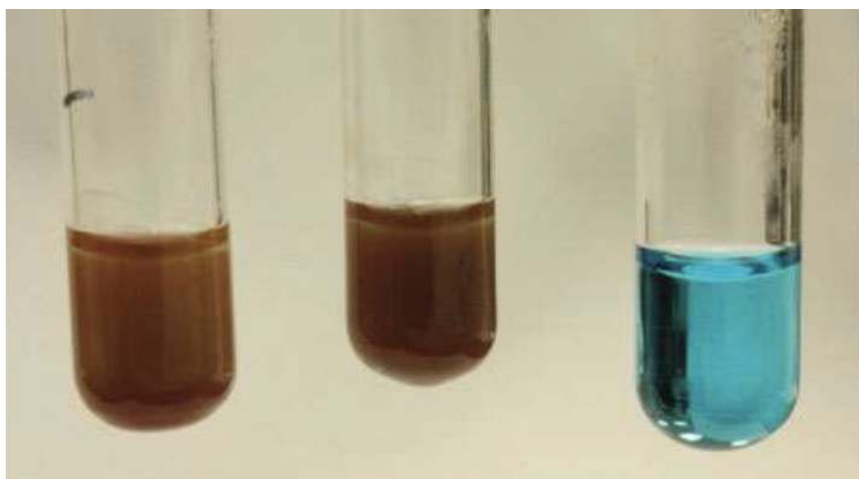


Figure 13.1.1: Benedict's test was performed on three carbohydrates, depicted from left to right: fructose, glucose, and sucrose. The solution containing sucrose remains blue because sucrose is a nonreducing sugar.

Experimental Procedure

Question

Are there simple reducing sugars in my juice?

Diabetes mellitus is a disease that refers to the inability of the cells to take in glucose. The word diabetes refers to urination and mellitus refers to sweetness. Since the cells of diabetics cannot remove glucose from the blood, there is an excess of glucose circulating that is eliminated in the urine. The traditional method of diagnosing someone with diabetes mellitus was to taste the sweetness of the patient's urine. Let's use Benedict's test for the detection process instead of the unhygienic alternative.

Make a hypothesis and ask what we would predict from a Benedict's test if testing a urine sample of someone with diabetes mellitus.

Materials and Equipment

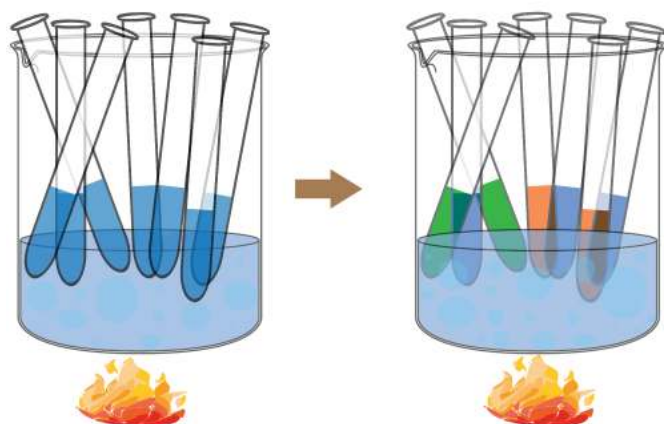
Test tubes, test tube rack, 400-mL beaker, droppers, hot plate, 5- or 10-mL graduated cylinder, Benedict's reagent, apple juice, 2% carbohydrate solutions: glucose, fructose, sucrose, lactose and starch

Benedict's Test for Reducing Sugars Procedure

1. Obtain 7 test tubes and number them 1-7.
2. Add to 10 drops of the materials to be tested (see Table 1) to each tube. Your instructor may ask you to test some additional materials. If so, include additional numbered test tubes.
3. Indicate in the table whether the sample you are testing is positive control, a negative control or experimental.
4. Before you begin the heating of the samples, use predict the color change (if any) for each sample. (use the sample type to aid in your prediction)
5. Add 40 drops (or 2 ml) Benedict's solution to each tube.
6. Place all of the tubes in a boiling water bath for 3 min or until a noticeable color change and observe colors during this time.
7. After 3 min, remove the tubes from the water bath and let them cool to room temperature. Record the color of their contents in your data table.

Iodine Test for Starch Procedure

Boil solutions until color changes



Carbohydrates that are used for energy storage are not reducing sugars since they are polymers that lack free aldehydes. Plant cells store energy in the form of starches like amylose and amylopectin. Since these molecules are larger than monosaccharides or disaccharides, they are not sweet to the taste and are not very soluble in water. An iodine/potassium iodine (KI I₂) staining distinguishes starch from monosaccharides, disaccharides, and other polysaccharides. The basis for this test is that starch is a coiled polymer of glucose — iodine interacts with these coiled molecules and becomes bluish black. Iodine does not react with other carbohydrates that are not coiled, and remains yellowish brown. Therefore, a bluish black color is a positive test for starch, and a yellowish-brown color (i.e., no color change) is a negative test for starch. Notably, glycogen, a common energy storage polysaccharide in animals, has a slightly different structure than does starch and produces only an intermediate color reaction.

1. Obtain 7 test tubes and number them 1-7.
2. Hypothesis Testing: Indicate in the table if the sample is experimental or control. Predict your expected color changes for each sample.
3. Add to each tube the materials to be tested as indicated in the Table 2 below.
4. Your instructor may ask you to test some additional materials. If so, include additional numbered test tubes.
5. Add 10 drops of iodine to each tube. This test does NOT require boiling. Record the color of the tubes' contents in the table below.

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13.2: Detection of Carbohydrates Pre-lab Assignment

Questions

1. What are the three most abundant polysaccharides?
2. What type of bond links two monosaccharide to form a disaccharide?
3. What is the difference between D-glucose and L-glucose?

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13.R: Detection of Carbohydrates (Report)

Experimental Data

Table 1: Benedict's Test for Reducing Sugars

Tube	Solution	Type of Sample (unknown, +control, - control)	Hypothesized color	Benedicts Test color
1	10 drops apple juice			
2	10 drops potato juice			
3	10 drops sucrose solution			
4	10 drops glucose solution			
5	10 drops distilled water			
6	10 drops reducing sugar			
7	10 drops starch solution			

Questions

1. In the Benedict's test, which of the solutions is a positive control? Which is a negative control?
2. Which is a reducing sugar, sucrose or glucose?
3. Which contains more reducing sugars, potato juice or apple juice?

Table 2: Iodine Test for Starch

Tube	Solution	Type of Sample (unknown, +control, - control)	Hypothesized color	Iodine Test color
1	20 drops apple juice			
2	20 drops potato juice			
3	20 drops sucrose solution			
4	20 drops glucose solution			
5	20 drops distilled water			
6	20 drops reducing sugar			
7	20 drops starch solution			

Questions

1. In the Iodine test, which of the solutions is a positive control? Which is a negative control?
2. Which is more positive for the iodine test: apple juice or potato juice?
3. What can you infer about the storage of carbohydrates in apples? In potatoes?
4. Describe the half reaction $\text{Cu}^{2+} \rightarrow \text{Cu}^{+}$ as oxidation or reduction.
5. Describe the half reaction $\text{Cu} \rightarrow \text{Cu}^{+}$ as oxidation or reduction.

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CHAPTER OVERVIEW

14: Paper Chromatography of Amino Acids

[14.1: Paper Chromatography of Amino Acids Lab Procedure](#)

[14.2: Paper Chromatography of Amino Acids Pre-lab Assignment](#)

[14.R: Paper Chromatography of Amino Acids \(Report\)](#)

Adapted from

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14.1: Paper Chromatography of Amino Acids Lab Procedure

Learning Objectives

- To use paper chromatography to identify amino acids.
- To identify components of unknown amino acid sample by calculating R_f values

Background

Chromatography is a collective term for a set of analytical techniques used to separate mixtures. Chroma means color and graph means to write or draw. Paper chromatography is an analytical technique used to separate mixtures of chemicals (sometimes colored pigments) using a partitioning method. The paper in this method is called the stationary phase because it does not move and serves as a substrate or surface for the separation. Analytes (substances being analyzed) are separated from each other based on a differential affinity to a solvent. The solvent dissolves and carries the analytes along the matrix of the stationary phase. Since the solvent moves through a wicking action, it is called the mobile phase.

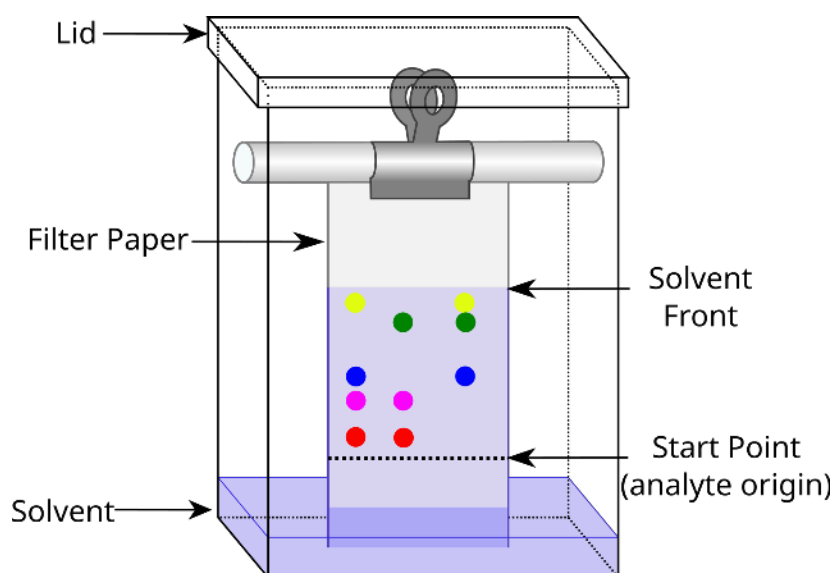


Figure 14.1.1: Credit: Theresa Knott [CC BY-SA 3.0 or GFDL]

The figure above shows a chamber containing mobile phase and a prepared paper stationary phase. A line drawn at the bottom edge of the paper is the starting point/line. The starting line and sample dots must be above the level of the mobile phase when the paper is placed inside the chamber. If the starting line is below the liquid level, the samples will wash out into the mobile phase rather than elute up the stationary phase. Another line is drawn about 10 cm above the bottom edge of the paper. This is the finish line. Its location was chosen for this experiment because when the eluting solution reaches that line, any inks that are mixtures should be clearly separated.

Note

When the solvent front reaches the finish line, the paper should be removed immediately from contact with the mobile phase.

The figure below shows a typical paper chromatogram of the separation of pen inks. There are a few difficulties commonly encountered in the elution process. One problem is that spots tend to spread out as they elute, and can bleed into each other as they proceed up the paper. This can be confusing when interpreting the chromatogram. To avoid this problem, space the spots of sample far apart and make repeated, tiny applications of sample to prevent spreading. Another problem is an uneven solvent front. This can happen if the beaker is nudged – if the mobile phase sloshes inside, the elution trails may travel diagonally, which makes interpretation very difficult. This can also happen if the two edges of the chromatogram are allowed to touch when they are stapled or taped together to form a cylinder.

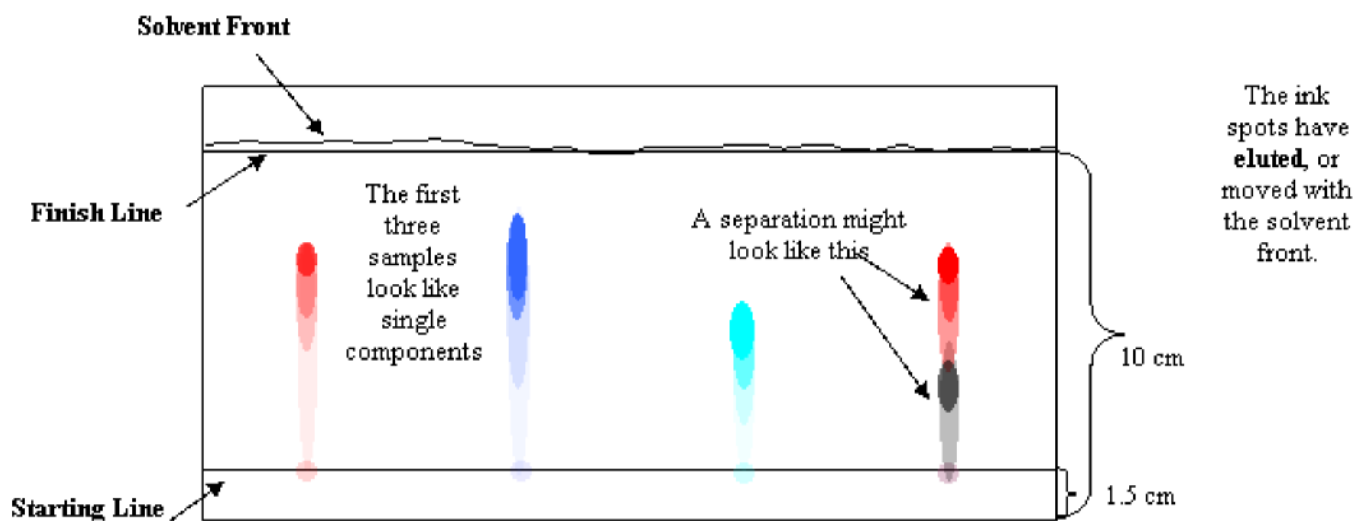


Figure 14.1.2:

A component with a given solubility travels along with the mobile phase at one rate, regardless of what other components are present in the sample. If the red part of purple ink travels at the same rate as pure red ink, and both stop in the same place, the two should be the same red ink. The two red spots should have the same Retention Factor, R_f . The R_f is the distance, D , traveled by the spot divided by the distance traveled by the eluting solution, or Solvent Front, F .

$$R_f = \frac{D}{F} \quad (14.1.1)$$

Comparing the R_f values allows the confirmation of a component in multiple samples because unique components have unique R_f values.

Experimental Procedure

Materials and Equipment

Materials: chromatography paper, amino acids (1% solutions): tyrosine, serine, glutamine, glycine, arginine, glutamic acid and unknown, eluting solution (isopropyl alcohol, 0.5 M NH_4OH ; 0.2% ninhydrin spray).

Equipment: 600-mL beaker, pencil, ruler, evaporating dish, toothpicks, hair dryer, stapler and paper towels.

Safety

Wear safety goggles at all times. Use eluting solution only in the hood. Do not breathe fumes from the eluting solution. Be sure to handle only the dry part of the chromatogram when removing it from the beaker. Wash hands thoroughly if the eluting solution touches your skin.

Part A: Preparation of Chromatography Paper

1. Wash your hands thoroughly to remove excess oils from your skin. Obtain a ruler and a piece of chromatography paper from the fume hood. Handle the paper only on the edges to avoid leaving fingerprints, as these may hinder the elution process.
2. Place the chromatography paper on a sheet of clean notebook paper or paper towel to avoid picking up dirt or contaminants from the bench top. Orient the paper into a "landscape" position and write your name on the top edge of the paper in one corner. Using a pencil and ruler to measure accurately, draw a straight line across the paper, about 1.5 cm above the bottom edge. This is the starting line. At the top of the paper draw a straight line across the paper about 2.0 cm from the top, this is finish line.
3. On the starting line, measure in from one side about 2.5 cm and lightly draw a small "X" centered on the starting line. Draw six more, 1.5 cm apart.
4. In the center of each X, make a small spot of amino acid sample in this order: tyrosine, serine, glutamine, glycine, arginine, glutamic acid and unknown.
5. When you have finished, you should have something that looks like ??? below.
6. Go back over each amino acid spot a second time to ensure there is enough amino acid in the spot.

- Gently curl the paper into a cylinder, with the spots on the outside. Staple the ends together near the top and bottom, taking care that the two edges of the paper do not touch. If they do touch, the eluent will creep on a diagonal, and the spots will run together, or not in straight lines.

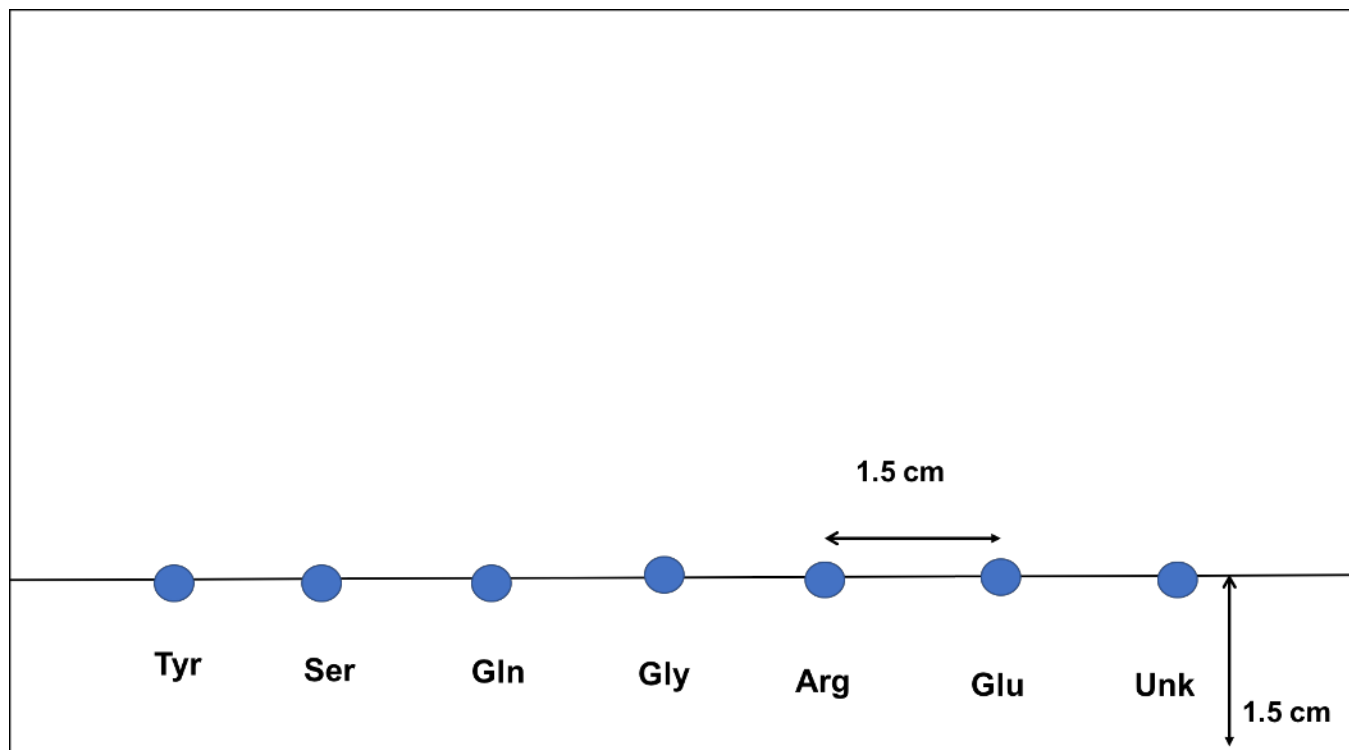


Figure 14.1.3: Prepared chromatography paper.



Figure 14.1.4: Chromatogram in solvent chamber

Part B: Acquisition of Chromatogram

- Take a 600-mL beaker and pour 10-mL of 0.5 M NH_4OH and 20 mL of isopropyl alcohol (eluting solution) into the beaker. Obtain your evaporating dish and use it to cover the beaker.

2. Gently place the paper cylinder into the beaker and cover the top with the plastic wrap. Remember that the spots must be above the liquid level for the experiment to work. Watch the eluent creep up the paper until it is about 2 cm from the top. It will take about 45-60 minutes for the solvent front to reach the finish line.
3. When the solvent front reaches the finish line, remove the paper from the beaker, being careful to touch only the top. Let excess eluent drip into the beaker. Gently remove the staple and lay the chromatogram on a piece of paper towel. Use a hair dryer to dry the chromatogram completely. Pour the eluting solution in the organic waste container under the fume hood.
4. Working in the fume hood, spray the chromatogram lightly with the ninhydrin solution. Dry the sprayed chromatogram with a hair dryer, distinct colored spots will appear as a result of the ninhydrin reacting with the amino acids.
5. CAUTION: Use the ninhydrin inside the fume hood. Do not breathe the fumes or get spray on your skin.

Part C: Interpretation of Chromatogram

1. Circle around each color spot.
2. Use a ruler and draw a plus sign in the center of each spot. Measure the distance from the starting line to each plus sign. Record this distance for each spot on your lab report. These are the DD values, in cm.
3. Measure the distance between the starting line and the finish line or, the farthest up that the solvent front reached. Record this distance. This is the FF value, in cm.
4. Calculate the retention factor (R_f) for each spot and record the values in your lab report.
5. You and your lab partner will hand in your lab reports at the same time, with the paper chromatogram stapled to one of the lab reports.

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14.2: Paper Chromatography of Amino Acids Pre-lab Assignment

Questions

1. What is the mobile phase and stationary phase used in this experiment?
2. Why is it important to determine the retention factor of each amino acid?
3. Give two safety precautions that must be following while carrying out this experiment.

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14.R: Paper Chromatography of Amino Acids (Report)

Experimental Data

Part C: Interpretation of Chromatogram

FF value (the distance traveled by the eluting solution): _____ cm

Record DD values for each eluted spot to the nearest 0.1 cm.

Amino Acid	Color	Distance(s) Traveled by Eluted Spot(s), D	Rf Value for Eluted Spot
Tyrosine			
Valine			
Glutamine			
Cysteine			
Arginine			
Glutamic acid			
Unknown			

Show calculations for Rf below for the unknown sample:

Identification of unknown #: _____

Data Analysis

1. This lab employs chromatography to identify amino acids. What other applications can we use chromatography for?
2. Are any of the Rf values in the table above the same (or similar, within 0.1)? What would the same or similar Rf values indicate?
3. What are the units for Rf values?
4. Which amino acid travelled farthest with the mobile phase? What does this indicate about that component's attraction to the mobile phase? To the stationary phase?
5. Which of the amino acids tested if any had an Rf value of zero? What does this indicate about that component's attraction to the mobile phase? To the stationary phase?
6. Why is it necessary to use a pencil to mark the lines and X's on the paper?

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15: Problem-Solving and Discussion Worksheets

Measurements Matter

1. Express the following values in scientific notation.
 - a. 150,000,000
 - b. 0.000043
 - c. 332000
 - d. 0.0293
 - e. 932
 - f. 0.1873
 - g. 78,000
 - h. 0.0001
 - i. 4500
 - j. 0.00290
 - k. 6281
 - l. 0.00700
2. Express the following values in decimal notation.
 - a. 3.8×10^{-3}
 - b. 9.21×10^5
 - c. 7.91×10^{-2}
 - d. 2.5×10^6
 - e. 3.42×10^{-8}
 - f. 5.4×10^5
 - g. 3×10^{-3}
 - h. 7.34×10^2
 - i. 9.8×10^{-4}
 - j. 6×10^7
 - k. 4.20×10^{-6}
 - l. 4.20×10^6
3. What SI base unit would be appropriate for each measurement?
 - a. the length of a room
 - b. the amount of carbon in a diamond
 - c. the mass of NaCl in a bottle
4. List the meaning of each abbreviation of the base units.
 - a. m
 - b. K
 - c. kg
 - d. s
 - e. mol
5. Give the meaning and name of each metric prefix abbreviation.
 - a. M
 - b. m
 - c. n
 - d. d
6. Convert each temperature to the missing one between Celsius and Fahrenheit.
 - a. 77°F
 - b. 212°F

- c. 37°C
- d. 22°C
- e. 95°F
- f. 15°C
- g. 0°F
- h. 0°C
- i. -10°C
- j. -10°F

Significant Figures

1. Explain the similarities and differences between accuracy and precision.
2. Determine the number of significant figures in each of the following values.
 - a. 406
 - b. 3.00
 - c. 3.20
 - d. 0.25
 - e. 0.0689
3. Write each value with three significant figures, use scientific notation if necessary.
 - a. 34500
 - b. 24
 - c. 0.0345
 - d. 0.012
 - e. 612.8
4. Find the result of each of the following calculations and report the value with the correct number of significant figures.
 - a. $0.23 + 12.2 =$
 - b. $13 - 1.03 =$
 - c. $0.839 + 0.28925 =$
 - d. $28 + 34.4 =$
 - e. $0.8 + 2.3 =$
 - f. $34.9 - 0.583 =$
5. Find the result of each of the following calculations and report the value with the correct number of significant figures.
 - a. $34 \times 0.12 =$
 - b. $68.2 / 0.78 =$
 - c. $3.29 \times 104 \times 16.2 =$
 - d. $0.8449 \times 29.7 =$
 - e. $5.92 \times 103 / 0.628 =$
 - f. $3.00 \times 2.6 =$

Scientific Dimensional Analysis

1. Complete each of the following conversions.
 - a. 34 cm to m
 - b. 3.7 ft to in
 - c. 345 mg to Mg
 - d. 5.3 km to mm
 - e. 4.0 L to mL
 - f. 3.45×10^3 mm to km
 - g. 78 cm³ to mL
 - h. 0.85 kg to dg

- i. 13 pints to gallon
- j. 0.35 L to cm³

Percentages

1. Solve each of the following.
 - a. What percent of 35 is 8.2?
 - b. What percent of 56 is 12?
 - c. What percent of 923 is 38?
 - d. What percent of 342 is 118?
2. Solve each of the following?
 - a. What is 42% of 94?
 - b. What is 83% of 239?
 - c. What is 16% of 45?
 - d. What is 38% of 872?
3. A patient originally weighs 182 pounds and loses 15.0% of their body weight. What is their final weight?
4. A patient's original weight was 135 pounds and they lose 12 pounds. What percent of their body weight did they lose?
5. A patient needs to increase their calcium supplement by 25% a week. If they are currently taking a 300. mg supplement, how much more will they need to take?
6. A patient needs a 20.0% decrease in their medication dosage from 125 mg. What will his dosage be after the decrease?

Elements and Atomic Structure

1. Using the periodic table, identify the atomic symbols and names of the elements that correspond to the following atomic numbers:
 - a. 8
 - b. 27
 - c. 33
 - d. 59
 - e. 90
2. How many electrons do each of the following elements contain? Hint: These are not ions, but neutral elements.
 - a. Fluorine
 - b. Magnesium
 - c. Iodine
 - d. Copper
 - e. Silicon
3. Which elements contain the following numbers of protons and electrons?
 - a. 15 protons and 15 electrons
 - b. 77 protons and 77 electrons
 - c. 42 protons and 42 electrons
 - d. 4 protons and 4 electrons
 - e. 99 protons and 99 electrons
4. What is an isotope? How does one isotope differ from another within the same element?

Periodic Table of Elements

1. One of the main groups on the periodic tables is the Metals. How many different subgroups of metals are there on the periodic table? Name each of them. Where in the periodic table can each subgroup be located? Be specific.
2. Determine if the following descriptions are examples of a physical or chemical property.
 - a. Iron nails rust when left out in the rain
 - b. Glass is transparent and rigid

- c. Table salt creates an electrical current when dissolved in water
 - d. Diamond is virtually unbreakable
 - e. Metallic sodium explodes when exposed to water
3. Of the following elements, which of the three categories, metal, non-metal and metalloid, do they each belong to?
- a. Manganese
 - b. Bromine
 - c. Silicon
 - d. Uranium
 - e. Phosphorus
4. For periods 1-6, how many electrons are in each energy level? How many valence electrons do the following elements from levels 2 and 3 have?
- a. Fluorine
 - b. Aluminum
 - c. Nitrogen
 - d. Magnesium
 - e. Silicon

Half-Life

1. The half-life for Tc-99m is 6 hours. A 100-gram sample of technetium-99 is stored for 30 hours. How much of the radioisotope is left after that time period? How many half-lives have elapsed after the 30 hours?
2. The half-life for Kr-79 is 34.5 hours. How long would it take for 50 grams of Krypton-79 to decay to 6.25 grams? What percentage of the initial 50 grams does the 6.25 grams represent? How many half-lives did it take to decay down to 6.25 grams?
3. If 64 grams of fluorine-18 is administered to a patient, how much will be left after 4 half-lives? If the doctor expects the patient to have less than 5% of it left in their system, how many half-lives will that take?

Compounds Containing Polyatomic Ions

1. Translate the formula of each polyatomic ion into its name.
 - a. O_2^{2-}
 - b. SO_4^{2-}
 - c. CO_3^{2-}
 - d. OH^- , ClO^-
2. Give the names of each polyatomic ion in the following groups.
 - a. HCO_3^- , CH_3CO_2^-
 - b. SO_3^{2-} , SO_4^{2-}
 - c. CN^- , OH^- , NO_2^- , NO_3^-
 - d. PO_4^{3-} , HPO_4^{2-} , H_2PO_4^-
3. Use the names of the given polyatomic ion names to write their formulas. Include the charge of each ion.
 - a. ammonium ion
 - b. bicarbonate ion
 - c. hypochlorite ion
 - d. bisulfate ion
4. For each set of polyatomic ions named here, give the formulas for each.
 - a. Hypochlorite and chloride
 - b. Phosphate and biphosphate
 - c. Carbonate and bicarbonate
 - d. Sulfate and bisulfate
 - e. Hydroxide and hydride
5. Provide the formula for each ionic compound named below.

- a. Sodium hypochlorite
- b. Potassium hydroxide
- c. Ammonium nitrate
- d. Aluminum sulfate
- e. Iron(III) cyanide

Writing Formula Units and Naming Compounds

1. Provide the name of the individual ions that make up each of the following ionic compounds and indicate how many of each ion there are.
 - a. $\text{Al}(\text{OH})_3$
 - b. Li_3PO_4
 - c. KCl
 - d. ZnF_2
 - e. Na_2O
2. Combine the following pairs of elements (or ions) to express them as an ionic compound.
 - a. Sodium and Sulfur
 - b. Barium and Iodine
 - c. Iron(III) and Oxygen
 - d. Aluminum and Bromine
 - e. Lithium and Nitrogen
3. Provide the name of each ionic compound below.
 - a. Ag_2O
 - b. ZnCl_2
 - c. $\text{Au}(\text{NO}_3)_3$
 - d. KOH
 - e. BaBr_2
4. Name each ionic compound AND give the names of the individual ions that make up the formula.
 - a. Hg_2S
 - b. Cr_2O_3
 - c. K_2SO_4
 - d. $\text{Fe}_2(\text{CO}_3)_3$
 - e. $\text{Zn}(\text{ClO})_2$
5. Determine whether each compound is ionic or covalent
 - a. P_2O_5
 - b. MgCl_2
 - c. $\text{NaC}_2\text{H}_3\text{O}_2$
 - d. NH_3
 - e. CCl_4
6. Transition metals tend to have varying charges. Determine the charges on each transition metal in the following ionic compounds.
 - a. AuN
 - b. SnO_2
 - c. Cu_2S
 - d. PbBr_4
 - e. FeI_2

Molecular Mass, Formula Mass, and Molar Mass

1. What value is held by Avogadro's number?
2. What is the mass of each of the following quantities:
 - a. 1.0 mole of Hydrogen gas
 - b. 2.0 moles of water
 - c. 0.1 moles of lead
 - d. 0.50 mole of hydrochloric acid (HCl)
 - e. 3.0 moles of Helium
3. What is the atomic mass of CH_3OH , ethanol, in AMU? What is the molar mass?
4. For each of these different compounds, using the given masses, calculate how many moles each sample contains.
 - a. In 3.1 grams of sodium chloride
 - b. In 4.8 grams of hydrogen peroxide
 - c. In 8 grams of lithium hydroxide
 - d. In 10. grams of oxygen gas
 - e. In 22 grams of ammonia
5. Contrasting with the previous problem, this time find the mass of each sample by using the given number of moles.
 - a. 6.0 moles of phosphorus
 - b. 4.0 moles of iron
 - c. 5.20 moles of fluorine gas
 - d. 1.3 moles of hydrochloric acid (HCl)
 - e. 7.7 moles of neon
6. Combine the following ions into an ionic compound and calculate the molar mass.
 - a. Ammonium ion and sulfide ion
 - b. Strontium ion and nitride ion
 - c. Aluminum ion and fluoride ion
 - d. Calcium ion and iodide ion
 - e. Silver ion and oxide ion

Mole Questions

1. What is the molar mass of sodium bicarbonate (baking soda)?
2. How many moles are in a 42g sample of sodium bicarbonate?
3. What is the mass in grams of 0.42 moles of sodium bicarbonate?
4. How many formula units are in a 25g sample of sodium bicarbonate? Hint: Ionic compounds do NOT form molecules, so we use the term formula unit.
5. How many sodium ions are in a 3.2g sample of sodium bicarbonate?
6. What is the molar mass of glycerine ($\text{C}_3\text{H}_8\text{O}_3$)?
7. How many moles are in a 42g sample of glycerine?
8. What is the mass in grams of 0.42 moles of glycerine?
9. How many molecules are in a 25g sample of glycerine?
10. How many carbon atoms are in a 3.2 g sample of glycerine?
11. Answer the following questions for both caffeine ($\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$) and sodium phosphate.
 - a. Calculate the molar mass of each.
 - b. Is each compound covalently or ionically bonded? Explain your reasoning.
 - c. We have a pure sample of each compound that weighs 2.3 g. How many moles are present in each sample?
12. There is a 50 g sample of each of the following substances: $\text{Ca}(\text{NO}_3)_2$, NaHCO_3 , CuCl_2
 - a. Which sample contains the largest number of formula units?
 - b. Which sample contains the largest number of ions?

Writing and Balancing a Chemical Equation

- Balance the following chemical equations.
 - $\text{Al(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Al(OH)}_3\text{(s)} + \text{H}_2\text{(g)}$
 - $\text{N}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow \text{NO}_2\text{(g)}$
 - $\text{H}_2\text{CO}_3\text{(aq)} \rightarrow \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$
 - $\text{P}_2\text{O}_5\text{(s)} \rightarrow \text{P}_4\text{(s)} + \text{O}_2\text{(g)}$
 - $\text{N}_2\text{(g)} + \text{H}_2\text{(g)} + \text{Cl}_2\text{(g)} \rightarrow \text{NH}_4\text{Cl(s)}$
- Balance each of the given combustion reactions.
 - $\text{C}_6\text{H}_6\text{(l)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(g)}$
 - $\text{C}_2\text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(g)}$
 - $\text{C}_8\text{H}_{18}\text{(l)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(g)}$
 - $\text{C}_4\text{H}_9\text{OH(l)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(g)}$
 - $\text{CH}_3(\text{CH}_2)_4\text{OCH}_2\text{CH}_3\text{(l)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(g)}$
- The next set of chemical equations are NOT properly balanced. Alter the coefficients as needed so that the equations are properly balanced.
 - $\text{N}_2\text{(g)} + \text{H}_2\text{(g)} \rightarrow 2\text{NH}_3\text{(l)}$
 - $\text{Ba(OH)}_2\text{(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{BaSO}_4\text{(s)} + \text{H}_2\text{O(l)}$
 - $2\text{CH}_4\text{(g)} + \text{N}_2\text{(g)} \rightarrow \text{HCN(g)} + 4\text{H}_2\text{(g)}$
 - $8\text{Cl}_2\text{(g)} + 8\text{H}_2\text{O(l)} \rightarrow 2\text{HClO}_4\text{(aq)} + 14\text{HCl(aq)}$
 - $2\text{PH}_3\text{(aq)} + 2\text{O}_2\text{(g)} \rightarrow 2\text{H}_3\text{PO}_4\text{(aq)}$

Energy and Chemical Reactions

- Follow the directions in each of the following sets of conversions. Report your answers to two (2) significant figures.
 - Convert 6.0 grams of fat into kilocalories, and the kilocalories into joules
 - Convert a sample containing 9.0 grams of carbohydrates into calories, then into kilocalories, then into joules
 - Convert a 7.0 gram serving of protein into kilocalories, then into joules
 - Convert a meal consisting of 14. grams of protein and 12. grams of carbohydrates into kilocalories, then joules
 - Convert a full course serving of 8.0 grams of fat, 8.0 grams of protein and 6.0 grams of carbohydrates into kilocalories, then joules

Pressure

- Name at least three different units of measurement for pressure and relate them using conversion factors.
- Perform the following conversions of pressure from one unit to another.
 - 120 psi into Torr
 - 6.0 atmospheres into pascals
 - 2800 mmHg into atmospheres

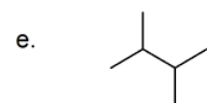
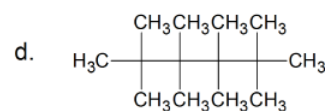
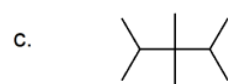
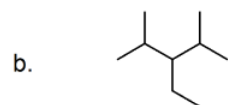
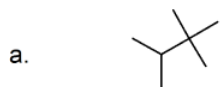
Gas

- A helium tank has a volume of 4.0 Liters. The pressure on the tank is 5.0 atmospheres. To save room the helium is transferred directly into a tank with a volume of 1.5 liters. What will the new pressure be in the new tank?
- If an athlete is training in a high-altitude area where the atmospheric pressure is 0.90 atmospheres and their healthy lung capacity is 3.2 liters, what would the pressure be if they developed a respiratory disease that decreased their lung capacity to 2.5 liters?
- Answer the following True or False questions.
 - As pressure increases, volume increases.
 - As pressure decreases, volume increases
 - As volume decreases, pressure increases

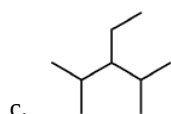
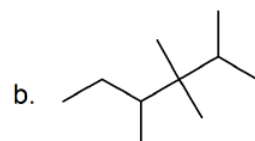
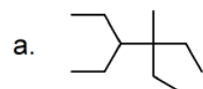
4. A mixture of the halogen gases has four partial pressures that compose the total pressure. If chlorine gas has a partial pressure of 2.00 atmospheres, bromine gas has a partial pressure of 13.0 psi, fluorine gas has a partial pressure of 800. mmHg and iodine gas has a partial pressure of 900. Torr, what is the total pressure in mmHg? What is the percent composition for each of the gases in the mixture?

Writing Structures

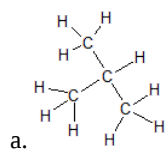
1. Translate the following structures into their condensed formulas.

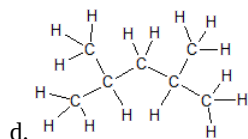
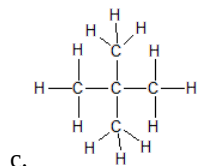
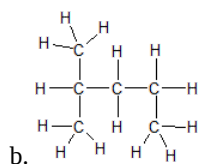


2. Give the IUPAC name for each of the following compounds.



3. Translate these expanded structures into their skeletal equivalents.





Solution Concentrations

- Many different biological molecules must be monitored to maintain a healthy body. Perform the following operations to convert one unit of measurement to another.
 - Blood sugar (glucose) level of 60 mg/dL to g/L
 - Blood cholesterol level of 185 mg/dL to mcg/mL
 - Iron level of 142 mcg/dL to mg/L
 - Blood alcohol level of 10 mg/dL to g/L
 - Sodium level of 150 mg/dL to mcg/mL
- What mass in grams of glucose will you need if you wish to make a solution of 5.00% (m/v) of 250 mL?
- You have 23.0 g of potassium phosphate with which to make a 20.0% (m/v) solution. What volume of solvent will you need to get this concentration?
- Convert the following quantities to the desired units.
 - 860 mmol/L to mol/L
 - 70 mol/dL to mol/L
 - 8 mol/L to mmol/mL
 - 900 mmol/mL to mol/L
 - 17 mol/L to mmol/dL
- How many moles of hydroxide are there in a 300. mL solution of 0.50 M aluminum hydroxide? How many moles of aluminum are there?
- A blood test reveals different concentrations of various essential minerals. A concentration of 0.65 mmol/L of copper (II), 0.09 mmol/L of manganese (IV), and 0.80 mmol/L of zinc. For each concentration, convert them to milliequivalents/L.
- A solution of calcium sulfate has a concentration of 0.20 M. How many equivalents/L of calcium ions are there? How many equivalents/L of sulfate ions?
- A doctor prescribes a dosage of 150,000 units of medication. 75 mg of the medication is equivalent to 1.0×10^5 units. The method of delivery provides only 50. mg of the active ingredient per mL. How many milliliters of the medication will be administered in each dose?
- A prescription calls for 20. g of medication for a patient. Two possible suspensions are available, one with a dosage of 7.0g/10. mL and 10.g/25 mL. How many milliliters of each suspension will be needed to meet the dosage the prescription calls for?
- The dosage is 4.0 mg and must be equally distributed throughout a ten-minute period. If the solution is only available in 3.0 mg/10 mL, what will the flow rate be in mL/min?
- A doctor calls for a dosage to be increased. Currently the patient is receiving a flow of 50. mL/hr, with the concentration of 30. mg/200. mL. The doctor calls for the dosage, in mg/min, to be increased by 50.%. What will the new flow rate, in mg/min, be?
- An infant weighing 20. lbs. is prescribed a dose of 100. mg three times a day. A report emerges that the safe dosage of that medication is between 25 and 45 mg/kg of body weight per day. Is the prescribed dosage within the safe limit for the infant?

Converting Between Different Concentration Units

1. Convert a 0.65 % (m/v) calcium chloride concentration to molarity. How many Eq/L of calcium ions are present? How many Eq/L of chloride ions are present?

Colloids and Suspensions

1. Differentiate between the following pairs, which is the colloid and which the solution.
 - a. Chocolate milk and rice pudding
 - b. Tuna salad and vinegar-based dressing
 - c. Light syrup and pureed fruit
 - d. Rocky road ice cream and fruit juice
 - e. Mashed potatoes and vegetable oil

Membranes, Osmosis and Dialysis

1. Given the following situations, wherein two tanks of different solutions are separated by a semipermeable membrane, determine the direction of the flow of solvent (water).
 - a. Solution A contains a 0.40 M concentration of CaCl_2 , while Solution B contains a 0.45 M concentration of KI
 - b. Solution A contains a 1.00 M concentration of NH_4Cl , while Solution B contains a 1.00 M concentration of CH_2O
 - c. Solution A contains 0.20 M NaCl and Solution B contains 0.30 M KCl

Conjugate Acids and Bases

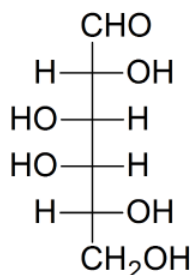
1. For each of the following acids, determine the form of their conjugate bases.
 - a. HSO_4^-
 - b. H_2CO_3
 - c. HF
 - d. HPO_4^{2-}
 - e. HCOOH
2. For each of the following bases, determine the form of their conjugate acids.
 - a. Br^-
 - b. NO_3^-
 - c. S^{2-}
 - d. NH_3
 - e. HCO_3^-

pH

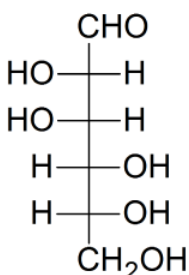
1. Examine the following pH measures and tell whether they are in the range to be considered acidic, basic or neutral.
 - a. 6.8
 - b. 13.5
 - c. 1.9
 - d. 4.2
 - e. 9.7
2. Using the formula for converting concentration into pH, determine whether each of the following molarities are acidic, basic or neutral. Remember, $\text{pH} = -\log[\text{H}^+]$
 - a. $1.0 \times 10^{-7} \text{ M H}^+$
 - b. $4.7 \times 10^{-12} \text{ M H}^+$
 - c. $6.4 \times 10^{-4} \text{ M H}^+$
 - d. $7.3 \times 10^{-10} \text{ M H}^+$
 - e. $9.8 \times 10^{-2} \text{ M H}^+$

Carbohydrates

1. Study the following Fischer projections to answer the questions below.



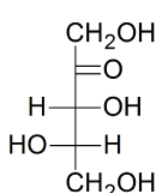
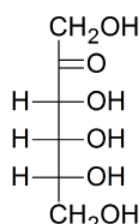
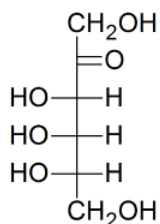
galactose



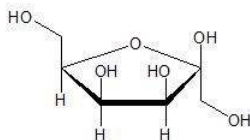
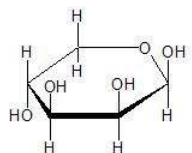
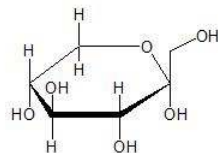
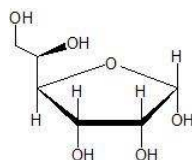
mannose

- Is galactose a D-sugar or an L-sugar?
- Is mannose a D-sugar or an L-sugar?
- Choose either one, and sketch it as it would appear if it was an L-sugar.
- Are these two carbohydrates enantiomers? If not, in how many places do they differ?
- What is the term to describe the relationship between galactose and mannose?

1. Are any of the following carbohydrates L-sugars? What makes something an L-sugar as opposed to a D-sugar? Rewrite each structure into an L-sugar if they are not already in that conformation.

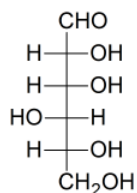


2. Study these Haworth projections to answer the following for each of them.

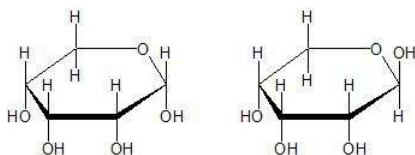


- Is it a furanose or a pyranose?
- Is it an alpha or beta anomer?
- Circle the anomeric carbon.

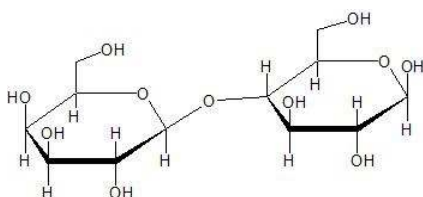
3. Examine the structure of this Fischer projection of D-gulose.



- Number each carbon from 1 to 6. What number is the anomeric carbon?
 - Circle the –OH group that determines whether it is a D- or L-sugar.
 - Sketch the structure of L-gulose for comparison. Are the two versions of gulose enantiomers of one another or diastereomers?
 - Now sketch the structure of D-glucose for comparison. Are D-gulose and Dglucose enantiomers or diastereomers?
4. Examine the following Haworth projections to answer the questions below.



- Circle each anomeric carbon
 - Which of the two is the alpha anomer and which is the beta anomer?
 - Are these structures considered enantiomers or diastereomers?
 - Are these monosaccharides reducing sugars? Explain.
 - Can you convert the alpha anomer to a beta anomer? Explain.
5. Some people cannot digest the disaccharide lactose. The term for this is known as lactose intolerance. Lactose is shown below. Answer the following questions.



- What is the name of the enzyme that would be required for someone to be able to digest lactose?
- Classify lactose as a mono-, di-, oligo-, or polysaccharide.
- Label anomeric carbons by circling them. Is lactose a reducing sugar?
- Draw an arrow pointing to the glycosidic bond. Is the glycosidic bond connected to both anomeric carbons?
- Classify the glycosidic bonds using the alpha or beta-(#,#) format.
- If the glycosidic bond is hydrolyzed, what are the names of the monosaccharides produced. Remember to include the alpha or beta classification for the anomeric carbon.

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12.1: Synthesis of Aspirin Lab Procedure

water displacement

2.1: Density of Liquids and Solids Lab Procedure

weight change

[7.R: Osmosis and Dialysis \(Report\)](#)

Z

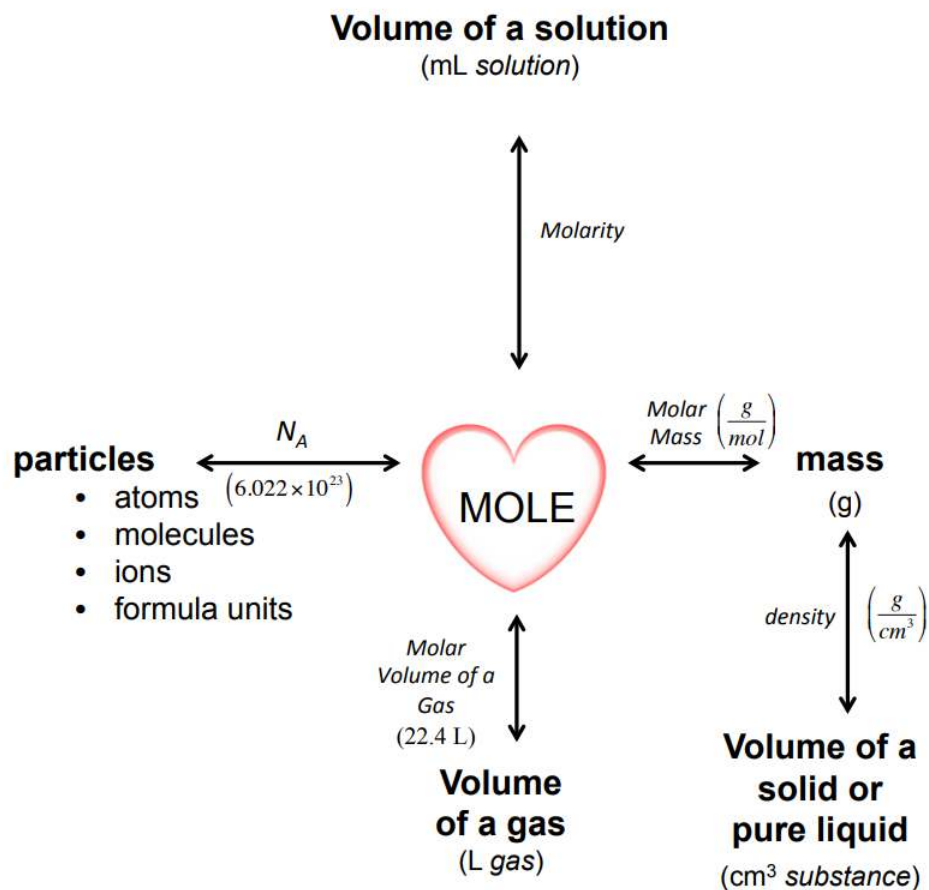
zinc metal

[5.1: Single and Double Replacement Reactions Lab Procedure](#)

Glossary

Sample Word 1 | Sample Definition 1

The Mole is the Heart of Chemistry



Metric Prefixes

Prefix	Symbol	Multiple	Multiple
Exa	E	10^{18}	1,000,000,000,000,000,000
Peta	P	10^{15}	1,000,000,000,000,000
Tera	T	10^{12}	1,000,000,000,000
Giga	G	10^9	1,000,000,000
Mega	M	10^6	1,000,000
kilo	k	10^3	1,000
hecto	h	10^2	100
deka	dk	10^1	10
<base unit>		10^0	1
deci	d	10^{-1}	0.1
centi	c	10^{-2}	0.01
milli	m	10^{-3}	0.001
micro	μ	10^{-6}	0.000 001
nano	n	10^{-9}	0.000 000 001
pico	p	10^{-12}	0.000 000 000 001
femto	f	10^{-15}	0.000 000 000 000 001
atto	a	10^{-18}	0.000 000 000 000 000 001

Polyatomic Ions

The following is a list of selected, common polyatomic ion names & formulas.

Ion Name	Ion Formula
ammonium	NH_4^+
cyanide	CN^-
hydroxide	OH^-
nitrate	NO_3^-
nitrite	NO_2^-
sulfate	SO_4^{2-}
sulfite	SO_3^{2-}
hydrogen sulfate (bisulfate)	HSO_4^-
carbonate	CO_3^{2-}
hydrogen carbonate (bicarbonate)	HCO_3^-
phosphate	PO_4^{3-}
hydrogen phosphate	HPO_4^{2-}
dihydrogen phosphate	H_2PO_4^-
permanganate	MnO_4^-
perchlorate	ClO_4^-
chlorate	ClO_3^-
chlorite	ClO_2^-
hypochlorite	ClO^-

Gas Laws Summary

Boyle's Law – Pressure vs. Volume

- Pressure and Volume are inversely proportional – as one increases, the other decreases by the same factor. $P \propto 1/V$
- $PV = k$ (a constant) (when T and n are held constant)
- $P_1V_1 = P_2V_2$

Charles's Law – Volume vs. Temperature

- Volume and Temperature (in Kelvin) are directly proportional – as one increases, the other increases by the same factor. $V \propto T$
- $\frac{V}{T} = \text{constant}$ (when P and n are held constant)
- $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Gay-Lussac's Law – Pressure vs. Temperature

- Pressure and Temperature are directly proportional. $P \propto T$
- $\frac{P}{T} = \text{constant}$ (when V and n are held constant)
- $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Combined Gas Law – Pressure, Temperature, and Volume

- This combines the other three gas laws into a single ratio.
- $\frac{PV}{T} = \text{constant}$ (when n is held constant)
- $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Avogadro's Law – Moles vs. Volumes

- The number of moles of any gas is directly proportional to the volume of the sample. $n \propto V$
- $\frac{n}{V} = \text{constant}$ (when P and T are held constant)
- $\frac{n_1}{V_1} = \frac{n_2}{V_2}$
- At STP, 1.00 mole of a gas occupies 22.4 Liters.

Ideal Gas Law

- The ideal gas law relates all of these different properties of a gas – pressure (P), volume (V), temperature (T in Kelvin), and number of moles (n) into a single equation that can be used to describe any ideal gas sample under any conditions.
- $PV = nRT$
- R is the ideal gas constant. There are different values for R, depending on what units the various properties are expressed in. Common R values are given at the bottom of the other side.

Standard Temperature and Pressure (STP)

- Standard Temperature = $0\text{ }^{\circ}\text{C} = 273.15\text{ K}$
- Standard Pressure = $1.000\text{ atm} = 760.0\text{ mmHg} = 101300\text{ Pa} = 101.3\text{ kPa}$

Values of the Ideal Gas Constant (R)

- $R = 0.08206 \frac{\text{L atm}}{\text{mol K}}$
- $R = 8314 \frac{\text{L Pa}}{\text{mol K}}$
- $R = 62.37 \frac{\text{L mmHg}}{\text{mol K}}$
- $R = 8.314 \frac{\text{L kPa}}{\text{mol K}} = 8.314 \frac{\text{J}}{\text{mol K}}$

Property	Conversion Factor
Mass	2.205 lb = 1 kg 1 lb = 16 oz
Volume	1.06 qt = 1 L 1 qt = 4 cups 1 fl oz = 29.6 mL 1 tsp = 5 mL
Length	1 mi = 1.6 km 1 mi = 5280 ft 1 in = 2.54 cm 1 ft = 12 in.
Energy	1 cal = 4.187 J 1 Calorie = 1000 calories

Radiation Activity Units

- 1 becquerel (Bq) = 1 disintegration per second
- 1 Curie (Ci) = 3.7×10^{10} disintegration per second
- 1 Ci = 1000 mCi
- 1 Ci = 1,000,000 μCi

Solutions

- Concentration = $\frac{\text{amount of solute}}{\text{amount of solution}}$
- Molarity, M = $\frac{\text{mole solute}}{\text{L solution}}$
- % concentration = $\frac{\text{parts of solute}}{100 \text{ parts of solution}}$
- Dilution Factor = $\frac{V_f}{V_i}$
- % (m/v) = $\frac{\text{g solute}}{\text{mL solution}} \times 100$
- % (m/m) = $\frac{\text{g solute}}{\text{g solution}} \times 100$

- $\% (v/v) = \frac{\text{mL solute}}{\text{mL solution}} \times 100$
- $C_{\text{initial}} \times V_{\text{initial}} = C_{\text{final}} \times V_{\text{final}}$
- $\text{ppm} = \frac{\text{g solute}}{\text{mL solution}} \times 1,000,000$
- $\text{ppb} = \frac{\text{g solute}}{\text{mL solution}} \times 1,000,000,000$

Calculating pH and pKa

- $\text{pH} = -\log[\text{H}_3\text{O}^+]$
- $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$
- $\text{pK}_a = -\log K_a$

Supplemental Notes: Naming

In this exercise, you will practice naming and writing chemical formulas for many inorganic compounds, both ionic and molecular. Before beginning the exercise, you should carefully read all the sections of your text (or notes) on the names and formulas of ionic compounds, simple covalent compounds, and acids. The following is a brief summary of the Nomenclature rules for each of these types of compounds.

Ionic Compounds

- Composed of metal cations and non-metal anions, or, of polyatomic ions.
- Names and formulas always start with the positively charged cation.
- Ions are combined in ratios so that the final ionic compound is neutral.
- Never use prefixes in the names of ionic compounds. The cation name is simply combined with the anion name only.
- If the cation is capable of having more than one possible charge, the cation charge is included in the name as a Roman numeral in brackets (Stock system).
- Several ion names, charges and formulas are provided in the following tables. They must be memorized as soon as possible.

Names of selected cations and anions (names are in alphabetical order)

Al^{3+}	Aluminum	Pb^{2+}	Lead(II); Plumbous
NH_4^+	Ammonium	Pb^{4+}	Lead(IV); Plumbic
As^{3+}	Arsenic(III)	Li^+	Lithium
Ba^{2+}	Barium	Mg^{2+}	Magnesium
Cd^{2+}	Cadmium	Mn^{2+}	Manganese(II)
Ca^{2+}	Calcium	Mn^{4+}	Manganese(IV)
Cr^{2+}	Chromium(II)	Hg_2^{2+}	Mercury(I); Mercurous
Cr^{3+}	Chromium(III)	Hg^{2+}	Mercury(II); Mercuric
Cr^{6+}	Chromium(VI)	Ni^{2+}	Nickel(II); Nickelous
Co^{2+}	Cobalt(II); Cobaltous	K^+	Potassium
Co^{3+}	Cobalt(III); Cobaltic	Ag^+	Silver
Cu^+	Copper(I); Cuprous	Na^+	Sodium
Cu^{2+}	Copper(II); Cupric	Rb^+	Rubidium
Au^{3+}	Gold(III); Auric	Sr^{2+}	Strontium
H^+	Hydrogen	Sn^{2+}	Tin(II); Stannous
Fe^{2+}	Iron(II); Ferrous	Sn^{4+}	Tin(IV); Stannic
Fe^{3+}	Iron(III); Ferric	Zn^{2+}	Zinc
$\text{C}_2\text{H}_3\text{O}_2^-$	Acetate	I^-	Iodide
AsO_4^{3-}	Arsenate	IO_4^-	Periodate
BO_3^{3-}	Borate	MoO_4^{2-}	Molybdate
$\text{B}_4\text{O}_7^{2-}$	Tetraborate	N^{3-}	Nitride
Br^-	Bromide	NO_2^-	Nitrite
BrO^-	Hypobromite	NO_3^-	Nitrate
BrO_3^-	Bromate	$\text{C}_2\text{O}_4^{2-}$	Oxalate
CO_3^{2-}	Carbonate	O^{2-}	Oxide
HCO_3^-	Bicarbonate; Hydrogen carbonate	O_2^{2-}	Peroxide
Cl^-	Chloride	MnO_4^-	Permanganate
ClO^-	Hypochlorite	P^{3-}	Phosphide
ClO_2^-	Chlorite	PO_4^{3-}	Phosphate
ClO_3^-	Chlorate	HPO_4^{2-}	Hydrogen phosphate
ClO_4^-	Perchlorate	$\text{H}_2\text{PO}_4^{2-}$	Dihydrogen phosphate
CrO_4^{2-}	Chromate	Se^{2-}	Selenide

$\text{Cr}_2\text{O}_7^{2-}$	Dichromate	S^{2-}	Sulfide
$\text{C}_6\text{H}_5\text{O}_7^{2-}$	Citrate	SO_3^{2-}	Sulfite
CN^-	Cyanide	SO_4^{2-}	Sulfate
F^-	Fluoride	HSO_3^-	Bisulfite; Hydrogen sulfite
H^-	Hydride	HSO_4^-	Bisulfate; Hydrogen sulfate
OH^-	Hydroxide	$\text{S}_2\text{O}_3^{2-}$	Thiosulfate
		SCN^-	Thiocyanate

Examples

- K_2S
 - 2 K^+ cations and 1 S^{2-} anion
 - potassium sulfide
- FeCl_3
 - 1 Fe^{3+} cation and 3 Cl^- anions
 - iron(III) chloride, or ferric chloride
- $\text{Mg}_3(\text{PO}_4)_2$
 - 3 Mg^{2+} cations and 2 PO_4^{3-} anions
 - magnesium phosphate

Simple Covalent (Molecular) Compounds

- Composed of non-metal atoms only.
- The more metallic non-metal is written first.
- Prefixes are used in the name to indicate the number of each atom present. A list of prefixes 1-10 (and 12) is provided below, which must be memorized.
- The prefix “mono” is dropped if there is only one of the first element.
- The name of the second element always ends in ide.

Prefixes for Covalent Compounds

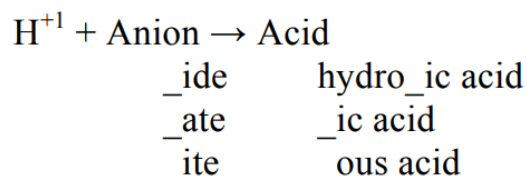
1	Mono
2	Di
3	Tri
4	Tetra
5	Penta
6	Hexa
7	Hepta
8	Octa
9	Nona
10	Deca
12	Dodeca

Examples

- P₄S₃
 - 4 P atoms and 3 S atoms
 - tetraphosphorus trisulfide
- N₂O
 - 2 N atoms and 1 O atom
 - dinitrogen monoxide
- BrCl₅
 - 1 Br atom and 5 Cl atoms
 - bromine pentachloride

Acids

- Composed of hydrogen cations and non-metal anions or polyatomic anions.
- H always leads the formula.
- Acids are in the aqueous state.
- Ions are combined in ratios so that the final acid is neutral.
- The acid name depends on the name of the anion involved:



Examples

- HBr (aq)
 - 1 H⁺¹ cation and 1 Br⁻¹ anion (bromide)
 - hydrobromic acid
- HNO₃ (aq)
 - 1 H⁺¹ cation and 1 NO₃⁻¹ anion (nitrate)
 - nitric acid
- H₂SO₃ (aq)
 - 2 H⁺¹ cations and 1 SO₃⁻² anion (sulfite)
 - sulfurous acid

Hydrates

- Hydrates are solid substances that also contain water molecules in their crystal structure.
- The number of water molecules for each formula unit is fixed.
- The formula is written the same as any other compound, but then we add a dot (•), (note that this is NOT a multiplication sign!) followed by the number of water molecules per ionic formula unit and the symbol H₂O.
- The name is the same as for any other compound, but we add the word “hydrate” at the end with a Latin prefix to indicate the number of water molecules per ionic formula unit.

Examples

- CuSO₄ • 5 H₂O
 - 1 Cu²⁺ ion, 1 SO₄⁻² ion, and 5 water molecules
 - Copper(II) sulfate pentahydrate
- Sodium carbonate decahydrate

- Na^+ ion(s), CO_3^{2-} ions(s), and 10 water molecules
- $\text{Na}_2\text{CO}_3 \cdot 10 \text{H}_2\text{O}$

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