CHEM1111-GENERAL INORGANIC CHEMISTRY I LAB MANUAL-ONLINE SECTION



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Licensing

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1.1: Preface

Chemistry is an empirical science. Hence, experiments play an essential role in teaching chemistry. Other methods of teaching cannot substitute experimentation. The concepts learned in lectures can be reaffirmed through lab experiments. Students entering this course should be able to improve basic laboratory skills like technical comprehension, planning, execution, observation, analysis and scientific reporting. All the lab experiments included in this manual are tied to the materials covered in the lectures. Ideally, the lecture should precede the related lab experiment. However, the instructor may choose a different order depending on other factors. The questions at the end of the lab sheets are designed to improve the problem solving skills by applying the learning that is gained through each experiment. CHEM 1111: General Inorganic Chemistry-I course online section will not meet at the campus laboratory for doing lab experiments. The lab experiments for this course need to be done remotely, at student's home. There are many lab procedures that are borrowed from standard lab manuals meant for real laboratory settings, but are remixed and modified to make them more greener and affordable so that the online students are not required to purchase hazardous and expensive supplies. Chemical reactions have always been fun if performed safely. Read the laboratory safety protocols before working on any experiment.

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1.2: Laboratory Safety Protocols-Home Version

Note

When you are in a chemistry lab (wherever you are conducting the experiment) you need to adhere to the following practices for the sake of your own safety as well as for the safety of others around you.

- **Do not eat or drink in the lab. If you have food items with you keep that outside the lab.** Do not leave any containers of food/drinks inside the lab.
- **Wear closed-toe footwear, long pants, and long sleeved shirts/tops to the lab.** *This is to protect you from serious burns or allergies in case your body comes in close contact with chemicals, especially in the case of a chemical spill.*
- **Do not wear contact lenses.** *Even minute amount of corrosive chemicals can damage your eye sight by the time you remove your contact lenses and wash your eyes. Chemicals can get into your eyes due to evaporation or a chemical spill.*
- **Tie your hair.** *For example, if you are using an open flame long hair might catch fire if left untied.*
- **No dangling ear rings or jewelry.** *For the same reason as noted above.*
- **Wear safety glasses or goggles and a lab coat or apron as advised**. Safety glasses are available in the lab and you do not need to purchase them. Once you finish your experiment return them back to their original cabinet.
- **Wear appropriate gloves before handling chemicals.** Your instructor will advise you when to wear gloves.
- **Do not fool around in the lab.** *It may be fun to be around your classmates, and you may have a lot to share especially when you are waiting on an experiment to record the observation. However, it can lead to serious accidents if you do not pay close attention to the experiment that you are conducting because chemicals can cause serious accidents.*
- **Clear the clutter on the work bench/table**. Leave only the lab manual, pen/pencil, textbook, and calculator on the workbench. *Even traces of chemicals on the workbench can ruin your belongings.*
- **Never lean on the work bench.** *The same reason as above.*
- **Never taste, touch, or smell chemicals unless you are instructed to do so.**
- **Do not perform any unauthorized experiments.** Do not mix chemicals unless instructed.
- **Read chemical labels on the bottles carefully before you handle them.** Read Safety Data Sheet (SDS) provided with each chemical.
- **Unused chemicals should never be returned back into the chemical stock bottles.**
- **Use caution while handling hot objects and glass wares.** Use tongs or gloves as necessary.
- **Broken glass wares should be disposed properly. Do not touch them with your bare hands.**
- **Do not drain any chemicals or solids into the sink unless specified.**
- **Never leave any hot plate or burners unattended.**
- **Never reach across an open burner or hot glassware or a glassware that contain chemicals.**
- **Never pour water into acid to avoid corrosive acid splashes.**
- **Seek help immediately if a chemical spill or accident occurs.**
- **Read the experiment procedure a couple of times carefully before you start the experiment.**
- **Before you dispose any chemical waste as instructed. If you are unsure about how to dispose any unused chemicals check with your instructor.**
- **Clean the glassware and return them back to their proper place.**

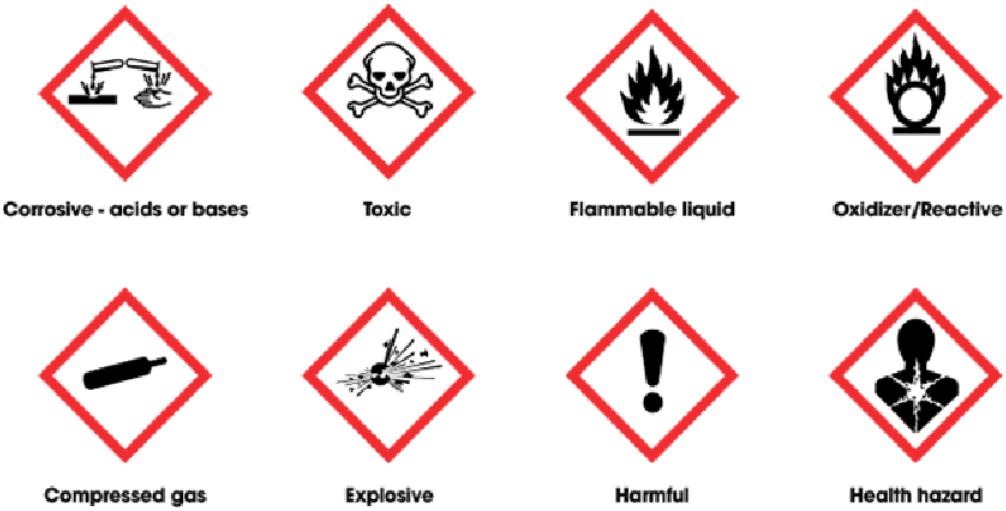
- **Clean your work bench after the experiment.**
- **Check the location of first aid measures and wash stations like fire blanket, fire extinguisher, eye wash stations, shower, first aid box, and SDS**
- **Wash your hands thoroughly with soap after you handle any chemical, and also before you leave the lab.**

Hazard Identification GHS Pictograms

Note

Watch for the following hazard identification pictograms while handling, storing, and disposing chemicals including household ones.

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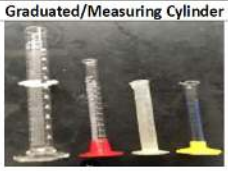
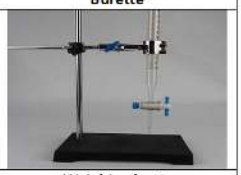
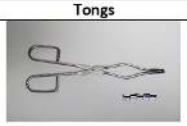


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1.3: Common General Chemistry Labwares and Equipments

Note

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Beaker	Erlenmeyer's or Conical flask	Graduated/Measuring Cylinder
		
Test tubes	Pipette	Burette
		
Electronic scale/balance	Stir/glass rod in a beaker	Weighing boats
		
Spatula	Tongs	Bunsen burner
		
Hot plate	Pasteur pipette	Funnel and filter paper
		

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1.4: Materials Required for the Lab Experiments

	Materials Required		
1	Beaker, 100-ml, 400-mL	1	Aluminum, one small solid piece from the lid of a soda can (careful about the sharp edges) rolled or a bottle cap
2	Electronic balance, pocket size	2	A wider clear glass bowl to hold the filter paper for elution in no bigger beaker is available
3	Erlenmeyer flask, 250 ml	3	Cardboard or Styrofoam lids or rectangular pieces to cover the Styrofoam cups,
4	Funnel	4	Chromatography paper (or filter or absorbent kitchen paper cut into a rectangle)
5	Graduated cylinders, 10-mL and 100-mL	5	Coins of the same type, 15-20, to make a solid cylindrical column
6	PH strips	6	Five (16 Fl. Oz) empty water or cola bottles with narrow mouth (any type of bottles of the same size and shape with narrow mouth).
7	Pipettes, small (and long graduated ones to use instead of a burette) with bulb	7	Graph papers
8	Thermometer	8	Medium sized balloons, 5-6
9	Goggles	9	Metric ruler
10	Gloves	10	Shape sheet (find a rectangle and circle shaped materials available at home)
1	Baking soda (sodium bicarbonate)-16 Oz Not Baking Powder	11	Small pieces of the following metals that you can find at home; copper (copper wire or a penny), silver (wire or a piece of broken jewelry), aluminum (aluminum foil), and iron (nail or steel wool)
2	Cola (with phosphoric acid ingredient), minimum 16 fl. Oz	12	Spoons, plastic
3	Epsom salt, 10 grams	13	Styrofoam cups, 5-10
4	Sugar, 5g	14	Washable-marker pens (Red, Blue, Green, Yellow, Brown)
5	Table salt, 10 g		
6	Vinegar (5% acidity-Acetic acid), 1 gallon		

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1.5: Lab Practices and Charts

Some Good Lab Practices

Measurement using any instrument

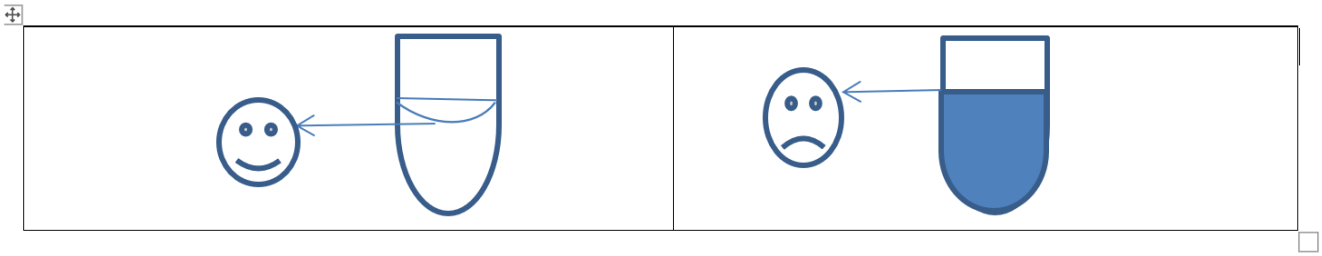
Make sure you are comparing the measurement value to the required unit. For example to measure the length of a book using a ruler in “mm”; with a ruler that has both “mm” and “cm” units on either side, make sure you are measuring using the right side or in the right unit (“mm”) that you need. If it is an electronic device check the “mode, and change it to the appropriate unit mode you need.

Using an electronic balance (scale)

An electronic balance is used to measure the mass of an object. Sometimes you are required to measure the mass of both the container and the object, and sometime you are required to measure the mass of the object only that is placed in a container. The first step as mentioned earlier is to check that the unit shown in the instrument is the one that you want. Otherwise press the mode button and bring it to the desired unit (it is usually the SI units). Then place the container and the object together to find the reading. Write down the measurement followed by the unit. If you do not want the mass of the container, keep the container on the scale, press the “zero” or “tare” button, so that the mass of the container is zeroed out. Now you can place the object in the container to find the mass of the object only.

Measuring volumes using graduated tubes or columns (volumetric measurements)

Volumes can be measured using either an instrument with a base like graduated cylinder or a tube like buret. In both the cases hold or place them so that the grading or marking on the tube and your eyes are on the same straight line (leveled). It is easier to keep the graduated cylinder on the work bench and kneel down to read the measurement. For a buret it can be clamped on a stand to read the measurement. For a colorless liquid the reading matching with the lower meniscus is taken, and for an opaque or colored liquid upper meniscus level is measured.



Using a pipet

Always use a mechanical or electric suction device to draw liquid into a pipet. Draw a little more than you need, and then slowly release the suction with great care to the nearest mark, pull it out from the stock container and then slowly release the liquid to the receiving container.

Using a buret

Open the knob (keep the knob vertical) and place the tip towards the drain, and rinse with distilled water using a wash (squirt) bottle. Rinse the buret two or three times. Then hold the buret slanted over a waste bottle or collection beaker and slowly rinse the inside of the tube with the liquid that you want to fill it with. Rinse the buret with 3-5ml twice. Then close the knob (horizontal), and fill with the liquid you

want. Fill it a little more than the top most reading. Hold the closed buret over the collection beaker again, and slowly open it (drip) to level with the top most reading (usually zero). This way you are removing the trapped air bubbles near the knob that could give rise to errors in your measurement.

Rinsing a glassware

Before you use any glassware make sure it is clean. Wash with a wire brush and soap until the glass is clean. After washing rinse it with 5-10ml of distilled water.

Washing used test tubes

Dispose the chemicals in the test tube into appropriate waste containers as instructed. Rinse with water and transfer the first two rinsing into the waste container itself. Then bring it to the sink and wash with wire brush, soap, and cold water. Then you can wash with hot water and soap if necessary. Rinse with distilled water. Make sure you wear gloves during the whole process. After removing gloves wash your hands thoroughly with soap and water.

Handling hot glassware

Never touch any hot glassware with your bare hands. Wear gloves, or use tongs or test tube holders as necessary. Never leave them on the table without a heat resistance mat or platform to cool off.

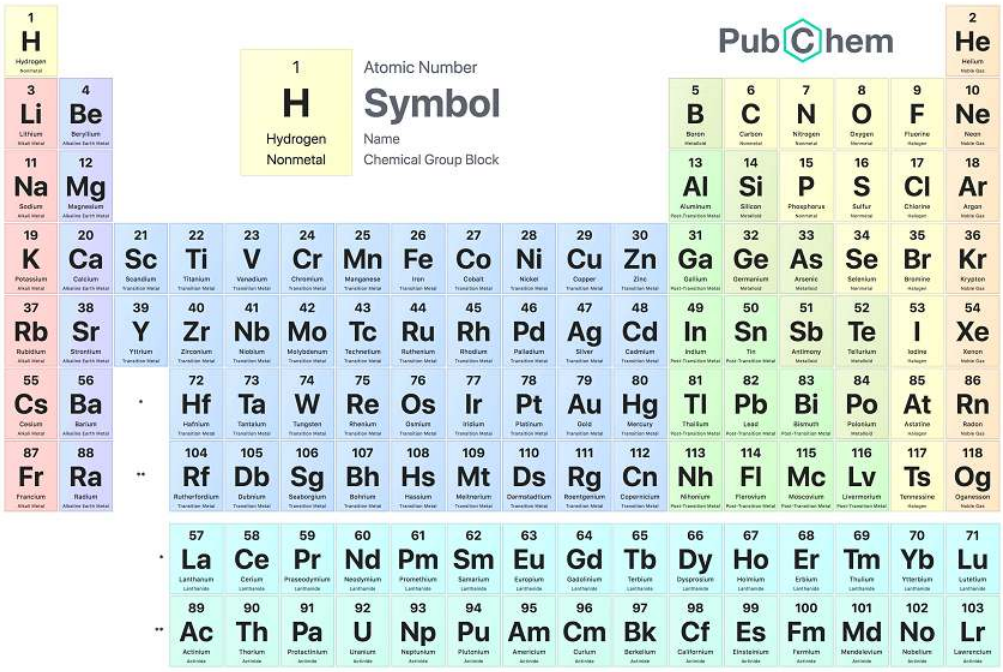
Diluting acids

Mixing acid and water will release heat. If you want to dilute an acid, take the required amount of water in bigger glassware in the fume hood and add the acid drop-by-drop into the water with vigorous shaking or mixing. It is also advisable to keep the water in an ice bath or cold water bath while you drip the acid into it. Wear corrosion resistant gloves, goggles, and lab coat.

Filtering a heterogeneous liquid-solid mixture

Find the appropriate filter paper with right dimension and mesh size. If it is a conical funnel, fold the filter paper into four or eight fold. Keep it in the funnel with the pointed side towards the bottom. Wet the paper with a very small amount of distilled water so that the paper sticks to the funnel. Slowly pour the top liquid in portions (never add more than the $3/4^{th}$ of the funnel height. Let the first portion drain, and then add the second. When the filtration is done rinse the remaining solid with fresh liquid, mix, and then repeat the filtration. Never touch the filter paper with sharp tips, like spatula or glass road.

PERIODIC TABLE OF ELEMENTS



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1.6: Acknowledgements

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This project reused multiple contents and materials available on the LibreTexts platform. All the good souls behind the email address info@libretexts.org are the life savers of this project by providing timely help for all the technical issues. Manjusha would like to extend warm and heart-felt gratitude to all of them including the Libretexts platform. All the contributors and attributions are duly mentioned on each page of the manual. If anybody finds any lapse or omissions in the attributions please email Manjusha. It is not intentional.

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2.1: Measurements in the Laboratory (Experiment)-Home Version

Objectives

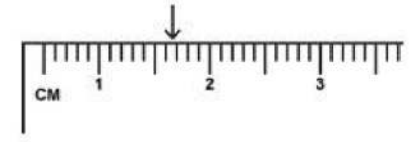
- To use standard laboratory measurement devices to measure length, volume and mass amounts.
- To use these measurements to determine the areas of shapes and volumes
- To determine the density of water.
- To determine the density of a solid and use this to determine further quantities.
- To determine the density of aluminum (applying the technique of water displacement) and use that value to determine the thickness of a piece of aluminum foil.

Chemistry is the study of matter. Our understanding of chemical processes thus depends on our ability to acquire accurate information about matter. Often, this information is quantitative, in the form of *measurements*. In this lab, you will be introduced to some common measuring devices, and learn how to use them to obtain correct measurements, each with correct precision. A metric ruler will be used to measure length in centimeters (cm).

All measuring devices are subject to error, making it impossible to obtain exact measurements. Students will record all the digits of the measurement using the markings that we know exactly and one further digit that we estimate and call uncertain. The *uncertain digit* is our best estimate using the smallest unit of measurement given and estimating between two of these values. These digits are collectively referred to as significant figures. Note, the electronic balance is designed to register these values and the student should only record the value displayed.

When making measurements, it is important to be as accurate and precise as possible. Accuracy is a measure of how close an experimental measurement is to the true, accepted value. Precision refers to how close repeated measurements (using the same device) are to each other.

Example 2.1.1 : Measuring length



Here the “ruler” markings are every 0.1-centimeter. The correct reading is 1.67 cm. The first 2 digits **1.6**7 are known exactly. The last digit 1.6**7** is uncertain. You may have instead estimated it as 1.68 cm.

The measuring devices used in this lab may have different scale graduations than the ones shown Precision is basically how many significant figures you have in your measurement. To find the precision, you basically take the smallest unit on your measuring device, and add a decimal place (the uncertain digit).

Note

In general, the more decimal places provided by a device, the more precise the measurement will be.

Measurements obtained in lab will often be used in subsequent calculations to obtain other values of interest. Thus, it is important to consider the number of significant figures that should be recorded for such calculated values. If multiplying or dividing measured values, the result should be reported with the lowest number of *significant figures* used in the calculation. If adding or subtracting measured values, the result should be reported with the lowest number of *decimal places* used in the calculation.

Example 2.1.2 : Significant Figures in Calculated Values

(a) A student runs 18.752 meters in 54.2 seconds. Calculate his velocity (or speed).

$$velocity = \frac{distance}{time} \tag{2.1.1}$$

$$= \frac{18.752m}{54.2s} \tag{2.1.2}$$

$$= 0.345978m/s \text{ from calculator} \tag{2.1.3}$$

$$= 0.346m/s \text{ to 3 significant figures} \tag{2.1.4}$$

(b) The mass of a glass is measured to be 12.456 grams. If 10.33 grams of water are added to this glass, what is the total combined mass?

$$\text{total mass} = 12.456g + 10.33g \tag{2.1.5}$$

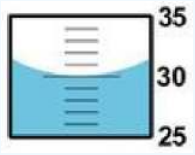
$$= 22.786g \text{ from calculator} \tag{2.1.6}$$

$$= 22.79g \text{ to 2 decimal places} \tag{2.1.7}$$

In this lab, students will also determine the density of water as well as aluminum. Volume is the amount of space occupied by matter. An extensive property is one that is dependent on the amount of matter present. Volume is an extensive property.

The *volume* of a liquid can be directly measured with specialized glassware, typically in units of milliliters (mL) or liters (L). In this lab, a beaker, two graduated cylinders and a burette will be used to measure liquid volumes, and their precision will be compared. Note that when measuring liquid volumes, it is important to read the graduated scale from the lowest point of the curved surface of the liquid, known as the liquid *meniscus*.

Example 2.1.3 : Measuring the Volume of a liquid



Here, the graduated cylinder markings are every 1-milliliter. When read from the lowest point of the meniscus, the correct volume 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.

The volume of a solid must be measured indirectly based on its shape. For regularly shaped solids, such as a cube, sphere, cylinder, or cone, the volume can be calculated from its measured dimensions (length, width, height, diameter) by using an appropriate equation.

Formulas for Calculating Volumes of Regularly Shaped Solids:

$$\text{Volume of a cube} = l \times w \times h \tag{2.1.8}$$

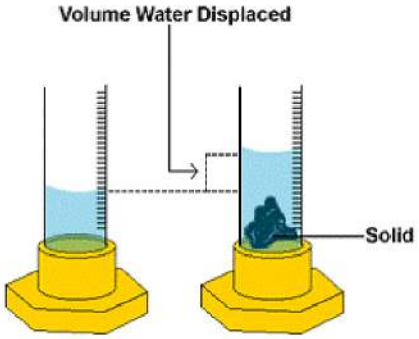
$$\text{Volume of a sphere} = \frac{4}{3}\pi r^3 \tag{2.1.9}$$

(where r = radius = 1⁄2 the diameter)

$$\text{Volume of a cylinder} = \pi r^2 h \tag{2.1.10}$$

For irregularly shaped solids, the volume can be indirectly determined via the volume of water (or any other liquid) that the solid displaces when it is immersed in the water (*Archimedes Principle*). The units for solid volumes are typically cubic centimeters (cm^3) or cubic meters (m^3). Note that 1 mL = 1 cm3.

Measuring the Volume of an Irregularly Shaped Solid



The volume water displaced is equal to the difference between the final volume and the initial volume , or:

$$V = V_f - V_i \tag{2.1.11}$$

where the volume water displaced is equal to the volume of solid.

Density is defined as the mass per unit volume of a substance. Density is a physical property of matter. Physical properties 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 also an intensive property. An intensive property is one that is independent of the amount 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. Density is determined by dividing the mass of a substance by its volume:

$$density = \frac{mass}{volume} \tag{2.1.12}$$

Density is commonly expressed in units of g/cm^3 for solids, g/mL for liquids, and g/L for gases.

Procedure

Materials and Equipment

Metric ruler, shape sheet (find a rectangle and circle available at home, say a notebook or circular filter paper), 250-mL Erlenmeyer flask, 100-mL beaker, sugar, 400-mL beaker, spoon, burette (instead of burette, use a long graduated pipette with a bulb), 10-mL and 100-mL graduated cylinders, aluminum pellets/bar, aluminum foil, electronic balance, water (you do not need distilled water, tap water is just fine).

Safety

Be careful when adding the aluminum to your graduated cylinder, as the glass could break. *Personal protective equipment (PPE)* needed: lab coat, safety goggles, closed-toe shoes

Part A: Measuring the Dimensions of Regular Geometric Shapes

1. Find a ruler and “shape sheet” (use any rectangular or circular shaped flat objects like notebook or filter paper) . Measure the dimensions of the two geometric shapes: length and width of the rectangle, and the diameter of the circle. Record these values on your lab report
2. Use your measurements to calculate the area of each shape:

- Area of a rectangle: $A = l \times w$
- Area of a circle: $A = \pi r^2$ (r = radius = 1⁄2 the diameter)

Part B: Volumes of Liquids and Solids

Volumes of Liquids

1. Find a burette (use a long graduated pipette with a bulb instead), 10-mL graduated cylinder, 100-mL graduated cylinder and 100-mL beaker, each filled with a certain quantity of water. Measure the volume of water in each. Remember to read the

volume **at the bottom of the meniscus.** It is useful to hold a piece of white paper behind the burette/cylinder/beaker to make it clearer.

Volume of a Regularly Shaped Solid

2. Find a wooden block or cylinder and ruler from your home.
3. Measure the dimensions of the block. If it is a cube or a rectangular box, measure its length, width and height. If it is a cylinder or cone, measure its height and the diameter of its circular base.

Part C: The Density of Water

1. Using the electronic balance determine the mass of a clean, dry, 100-mL graduated cylinder.
2. Pour 40-50 mL of **distilled water** into the graduated cylinder and weigh. Make sure that the outside of the graduated cylinder is dry before placing it on the electronic balance.
3. Measure the liquid volume in the cylinder
4. Use the mass and volume to calculate the density of water.

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

Density of Aluminum

1. Using the electronic balance to determine the mass of a clean, dry, small beaker.
2. Obtain an aluminum pellet/bar. Transfer pellet/bar to the beaker weighed in the previous step, and measure the mass of the beaker and pellet/bar together.
3. Pour 30-35 mL of water into your 100-mL graduated cylinder. Precisely measure this volume.
4. Carefully add all the aluminum pellet/bar to the water, making sure not to lose any water to splashing. Also make sure that the pellet/bar are all completely immersed in the water. Measure the new volume of the water plus the pellet/bar.
5. When finished, retrieve and dry the aluminum pellet/bar.
6. *Analysis:* Use your measured mass and volume (obtained via water displacement) of the aluminum pellet/bar to calculate the density of aluminum.

The Thickness of Aluminum Foil

7. Take a rectangular piece of aluminum foil and ruler. Use the ruler to measure the length and width of the piece of foil.
8. Fold the foil up into a small square and measure its mass using the electronic balance
9. *Analysis:* Use these measurements along with the density of aluminum to calculate the thickness of the foil.

Lab Report: Measurements in the Laboratory

Part A: Measuring the Dimensions of Regular Geometric Shapes

Experimental Data

Shape	Dimensions	Precision	Measurement	# Significant Figures
Rectangle	Length			
	Width			
Circle	Diameter			

Data Analysis

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

- Convert the measured rectangle length to hm.
- Convert the measured circle diameter to nm.

2. 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

Part B: The Volumes of Liquids and Solids

Table 1: The Volume of Liquid Water

Measuring Device	Precision	Volume Measurement	# Significant Figures
Burette			
Beaker			
100-mL Graduated Cylinder			
10-mL Graduated Cylinder			

Table 2: The Volume of a Regular Solid, shaped as a

Dimensions Measured	Measurement	# Significant Figures

Data Analysis

Use your measured block dimensions (in Table 2) to calculate the block volume, in cm^3. Show your work, and report your answer to the correct number of significant figures.

Part C: The Density of Water

Table 1: The Density of Water

Mass of Empty, Dry Graduated Cylinder	
Mass of Graduated Cylinder + Water	
Mass of Water	
Volume of the Water in Graduated Cylinder	

Calculate the density of water, in g/mL. Show your work, and report your answer to the correct number of significant figures.

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

Experimental Data

Table 1: The Density of Aluminum

Mass of empty beaker	
Mass of beaker and pellet/bar	
Mass of pellet/bar	
Initial volume of water in cylinder	
Final volume of water and pellet/bar	
Volume of pellet	

Table 2: The Thickness of Aluminum Foil

Mass of foil	
Length of foil	
Width of foil	

Data Analysis

1. Use your measured mass and volume of the aluminum pellet (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. Use your measurements for the aluminum foil (in Table 2) along with the true density of aluminum (Double subscripts: use braces to clarify = 2.70 g/cm^3) to calculate the foil thickness, in cm. Consider the foil to be a very flat rectangular box, where Volume of foil = V = length × width× height (thickness). Show your work, and report your answer in scientific notation.

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2.2: The Density of Liquids and Solids (Experiment)-Home Version

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;
- To measure the mass and volume (via measured dimensions) of several cylinders of an unknown material, and to determine the density of this material via graphical analysis of the collected data.

Density

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:

$$Density = \frac{Mass}{Volume} \tag{2.2.1}$$

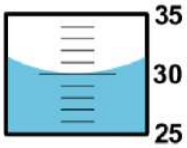
The units of density are commonly expressed as g/cm^3 for solids, g/mL for liquids, and g/L for gases.

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 mater 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.

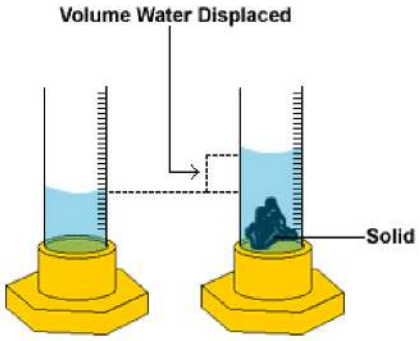
Measuring the Volume of a Liquid



The graduated cylinder markings are every 1-milliliter. When read from the lowest point of the meniscus, 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



$$\text{Volume water displaced} = \text{Final volume} - \text{Initial volume} \tag{2.2.2}$$

$$\text{Volume water displaced} = \text{Volume of solid} \tag{2.2.3}$$

Note that 1 mL = 1 cm^3.

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

$$\text{Volume of foil} = length \times width \times thickness \tag{2.2.4}$$

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.

Density and Graphical Analysis

Laboratory investigations involve collecting data, which is often numeric. One common method of interpreting data is graphical analysis.

In Part C of this lab, the mass and volume of several cylindrical pieces of an unknown solid material will be measured. Once again mass will be obtained using an electronic balance, in grams (g). But since the cylinders are regularly-shaped solids, their volumes (in cubic centimeters, cm^3) will be calculated from their measured dimensions by using the appropriate volume formula:

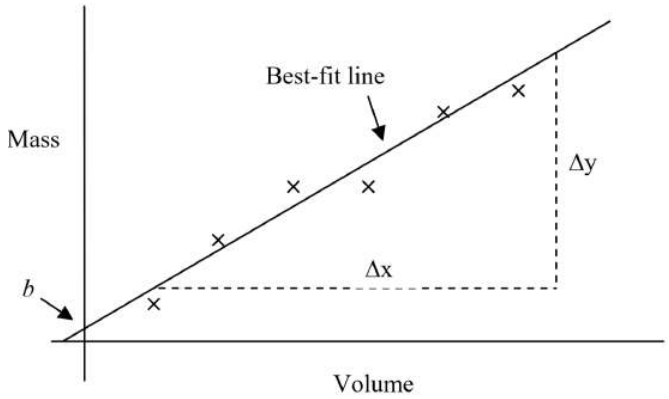
$$\text{Volume of a cylinder} = \pi r^2 h \tag{2.2.5}$$

$$h = \text{cylinder height or length} \tag{2.2.6}$$

$$r = \text{cylinder radius } = \frac{1}{2} \text{ the diameter} \tag{2.2.7}$$

Each pair of mass and volume values will then be plotted on graph paper as a scatter plot, with mass plotted on the y-axis and volume plotted on the *x*-axis. Since the plotted data generate (or at least approximate) a straight line, a "best-fit line" can be added to the graph. A best-fit line is a single line that comes as close as possible to all the plotted points.

The equation of this best-fit line will have the familiar form $y = mx + b$, where m represents the slope of the line, and b represents the *y*-intercept. This is illustrated in the figure below.



Best-fit line equation:

$$y = mx + b \tag{2.2.8}$$

where

- b is the y-intercept and
- m is the slope

The *y*-intercept (b) is the point on the y-axis where the line crosses the axis. In this experiment, the value of b should be equal to zero. This is because if there is no mass, the volume must also be zero. However, note that your best-fit line might not pass exactly through the origin (0,0) due to experimental error – but it should be quite close.

The slope of the line (m) is the change in the y-axis values divided by the change in x-axis values (or, rise over run):

$$\begin{aligned} m &= \frac{\Delta y}{\Delta x} \\ &= \frac{y_2 - y_1}{x_2 - x_1} \end{aligned}$$

Since Δy is really the change in mass (Δmass), and Δx is really the change in volume (Δvolume), this means that the slope of the best-fit line yields the density of the unknown material:

$$m = \frac{\Delta y}{\Delta x} = \frac{\Delta\text{mass}}{\Delta\text{volume}} = \text{density} \tag{2.2.9}$$

Once the density is determined in this manner, it will be used to identify the unknown material analyzed.

Procedure

Materials and Equipment

100-mL graduated cylinder, metric ruler*, aluminum pellets, small beaker, aluminum foil, thermometer, electronic balance, distilled water, tube of unknown solid cylinders* and graph paper.

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

1. Using the electronic balance, obtain the mass of your 100-mL graduated cylinder. Make sure it is dry before you weigh it.
2. Add 20-25 mL of distilled water to the graduated cylinder. Precisely measure this volume of water. Then measure the combined mass using the electronic balance.
3. Add another 20-25 mL of distilled water to the graduated cylinder. Again, precisely measure this volume of water, and then measure the combined mass using the electronic balance.
4. Repeat Step 3 to obtain a third set of mass and volume measurements.
5. Use your thermometer to record the temperature of the water in your graduated cylinder.
6. *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, 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. Transfer it to the beaker, and measure the mass of the beaker and aluminum pieces.
3. Pour 30-35 mL of water into your 100-mL graduated cylinder. Precisely measure this volume.

4. Carefully add all the aluminum piece(pieces) to the water, making sure not to lose any water to splashing. Also make sure that the pieces are all completely immersed in the water. Measure the new volume of the water plus the aluminum.
5. When finished, retrieve and dry the aluminum pieces.
6. *Analysis:* Use your measured mass and volume (obtained via water displacement) of the aluminum piece 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

7. Now obtain a rectangular piece of aluminum foil. Use the ruler to measure the length and width of the piece of foil.
8. Measure the mass of the foil using the electronic balance.
9. *Analysis:* Use these measurements along with the density of aluminum to calculate the thickness of the foil.

Part C: Graphical Analysis of Mass and Volume Data of an Unknown Solid (Any One Type of Coin)

1. Obtain 15-20 coins of the same type (say dimes or pennies). Make cylindrical columns of coins of various sizes by placing coins on top of each other. Say 4 coins in the smallest column, 7 in the second, 10 in the third, and 15 in the extra large column. You can make one at a time and do the measurements if you do not have enough coins.
2. Using the ruler, measure the dimensions (diameter and height) of each cylindrical coin column. Start with the smallest column (cylinder) first and progress in order of increasing object size.
3. Measure the mass of each cylindrical columns using an electronic balance. Again, start with the smallest column first and progress in order of increasing column size.
4. *Analysis:* Use the measured dimensions to calculate the volume of each solid coin column. The equation for calculating the volume of a cylinder is give at the beginning of this page. Then, on a graph paper, plot the mass (Y) versus the volume (X) of each measured coin column. Add a best- fit line to this plot. Calculate the slope of this line, which is the density of the unknown solid. Then use this density to compare with the true density of the coin listed in the table below. Calculate the percentage error using your calculated density value and the true density value given in the table below.

Substance	Density (g/cm^3)
Penny	5.77
Nickel	7.26
Dime	6.68
Quarter	7.01

Pre-laboratory Assignment: The Density of Solids and Liquids

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 and Metal sample	124.400 g
Initial volume of water in cylinder	12.7 mL
Final volume of water and Metal sample	21.6 mL

5. In Part C of this lab, you will measure the mass, height and diameter of four cylinders composed of some unknown material.

- Calculate the volume (in cm^3) of a cylinder with a measured height of 11.76 cm and a diameter of 7.22 cm. Show your work clearly.
- Each pair of mass and volume values (for each cylinder) will be plotted on a scatter plot, with mass on the y- axis and volume on the x-axis. A best-fit line will then be applied to the plotted data.
  - How will you calculate the value of the slope of this best-fit line?
- How will the value of the slope help you identify the unknown material that the cylinders are made of?

Lab Report: The Density of Liquids and Solids

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 only			
Volume of Water			
Density of Water			
Average Density of Water			

Temperature of Water: _____

Data Analysis

1. Look up the true density of water at the temperature recorded:_____

- Obtain this value from jupiter.plymouth.edu/~jsdunca...ityOfWater-vs- Temp.pdf. 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

Experimental Data

Table 1 – The Density of Aluminum

Mass of Empty Beaker	
Mass of Beaker and Al pellets	
Mass of 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: _____

- Use this to calculate the percent error in your experimentally determined density value. Show your work.

3. 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: $\textbf{Volume of foil} = \textit{length} \times \textit{width} \times \textit{thickness}$

Part C: Graphical Analysis of Mass and Volume Data of an Unknown Solid (Any One Type of Coin)

Experimental Data

ID Code of Unknown Solid:

	Small Cylinder (Coin Column)	Medium Cylinder (Coin Column)	Large Cylinder (Coin Column)	EX Large Cylinder (Coin Column)
Mass				
Length				
Diameter				
Calculated Volume				

Show a sample calculation for volume using your measured dimensions for the small cylinder(coin column) below.

Data Analysis

1. On the graph paper supplied, plot “Mass (Y) versus Volume (X)” for all four cylinders measured. Staple your graph to this report form.

Instructions for Graphing Data

- Use a sharpened pencil.
- Use a ruler to draw your axes.
- Choose axis scales that make use of the entire sheet of graph paper.
- Clearly number and label your axes.
- Use “X” symbols for each plotted point.
- Draw a best-fit straight line through your data points using a ruler.
- Give your graph an appropriate title.

2. Choose two points on your best-fit line separated far from each other. The points chosen cannot be any of your plotted data points. Circle the two points selected on your graph, and complete the table below.

	X Value	Y Value
Point 1		
Point 2		

Now calculate the slope (m) of your best-fit line using the equation: $m = \frac{y_2 - y_1}{x_2 - x_1}$. Show your work, and report your result to the correct number of significant figures.

3. The value of the slope obtained above in #2 yields the density of your unknown solid (coin), in units of g/cm^3. Using this value, compare your percentage error using the calculated density and true density (given in the table) of the coin (see Procedure, Part C, #5).:

- Coin Type: Penny/Nickel/Dime/Quarter
- True Density:

4. You are supplied with another cylinder made of the same material. If the cylinder length is 1.83 feet, calculate the mass of this cylinder, in grams. Use the true density of the solid in this calculation, and assume that the cylinder diameter is the same as all the other cylinders you measured. Show your work.

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2.3: Paper Chromatography of Washable-marker Ink Pens (Experiment)-Home Version

Objectives

- To use paper chromatography to identify whether certain colored inks are pure substances or mixtures.
- To obtain a paper chromatogram of various washable-marker inks
- To identify components of inks by calculating R_f values

Chromatography is a method of physically separating *mixtures* into its individual components. It is a common laboratory technique used to identify unknown components in mixtures.

There are several types of chromatography; all types employ a *mobile phase* or *eluent* (it can be liquid or gas), which is forced through a *stationary phase* (a solid or semi-solid). Mixtures are separated because some components will be more attracted to the stationary phase (and stick to it) while some components will be more attracted to the mobile phase (and travel with it).

By eye, we cannot know if each ink color is a mixture or pure substance. Using chromatography, the components in a sample will migrate along the filter paper at different rates such that they become spread out and separated from each other. The mobile phase takes advantage of differing solubility or polarity of the components in order to separate them. This component traveling process is called *elution*. Chromatography can be used to separate inks, dyes, pigments in plants, or used to determine the chemical composition of many substances.

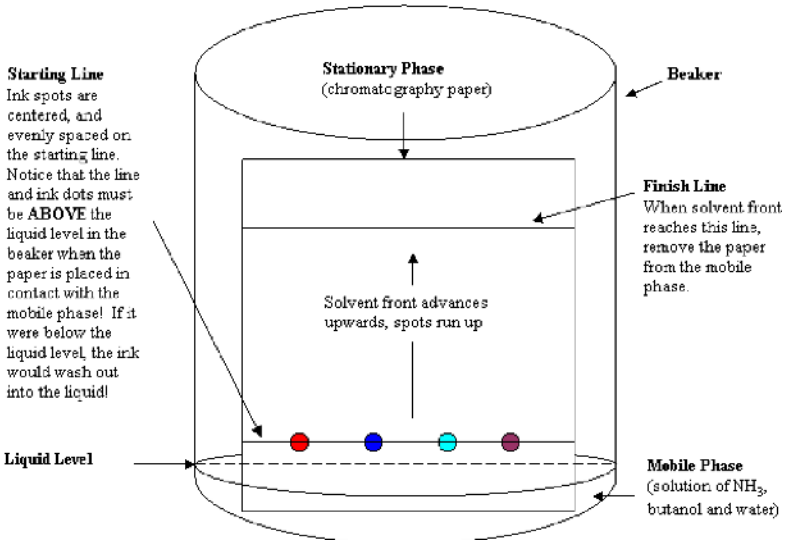


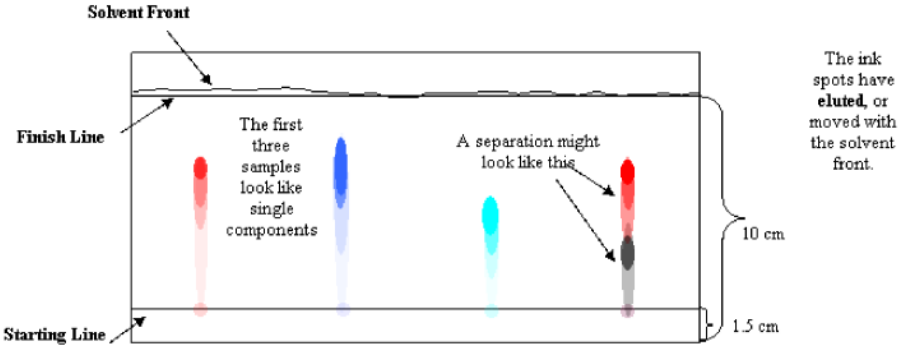
Figure 1 shows a beaker containing mobile phase and a prepared paper stationary phase. A line drawn at the bottom edge of the paper is the *starting line*. The starting line and ink dots *must be above* the level of the mobile phase when the paper is placed inside the beaker. If the starting line is below the liquid level, the inks 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.

Figure 2 shows a typical paper chromatogram. There are a few difficulties commonly encountered in the elution process. One problem is that spots tend the 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.



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, $\boldsymbol{D}$, traveled by the spot divided by the distance traveled by the eluting solution, or *Solvent Front*, $\boldsymbol{F}$.

$$R_f = \frac{D}{F} \tag{2.3.1}$$

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

Procedure

Materials and Equipment

Materials: chromatography paper (filter paper cut into a rectangle), washable-marker pens (Red, Blue, Green, Yellow, Brown) and eluting solution (0.1g table salt dissolved in 10ml water), a wider mouth clear container or a larger beaker to keep the filter paper for elution

Equipment: 500 ml beaker or a similar size flat bottom container, pencil, ruler, tape, and paper towels.

Safety

The eluting solution is safe to handle. Dispose it into the sink after the experiment.

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 your instructor. 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**. Draw another line about 10 cm above the bottom edge. This is the **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 seven more, 1.5 cm apart.
4. In the center of each X, make a small spot of ink color in this order:

brown, red, yellow,green, blue, and any other color if there is enough space

When you have finished, you should have something that looks like Figure 3.

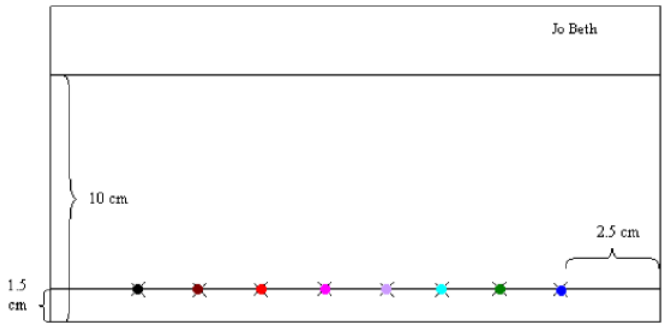


Figure 3: Prepared Chromatography Paper

5. Go back over each ink spot a second time to ensure there is enough ink in the spot.
6. Obtain a small piece of tape and gently curl the paper into a cylinder, with the spots on the outside. Tape 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.

Part B: Acquisition of Chromatogram

1. Take a 500-mL beaker and pour about 25-mL of eluting solution into the beaker.
2. Gently place the paper cylinder into the beaker. Remember that the spots must be above the liquid level for the experiment to work. Watch the eluent creep up the paper until it begins to move some of the ink. It will take about 45-90 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 tape and lay the chromatogram on a piece of paper towel in the hood. Leave the paper in the fume hood, where it will dry completely. If needed, use a heat lamp (in the fume hood) to dry the chromatogram; if using the heat lamp, allow 5-10 minutes to dry.

Part C: Interpretation of Chromatogram

1. Write the names of the original ink colors beneath each “X” mark (from the order you added the colors). Each ink sample should no longer be on the “X” mark, and have travelled up the paper, becoming one or more separate color spots between the starting and finish lines. 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 $\boldsymbol{D}$ 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 $\boldsymbol{F}$ 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 the lab reports.

Lab Report: Paper Chromatography of Gel Ink Pens

Part C: Interpretation of Chromatogram

Experimental Data

$\boldsymbol{F}$ value (the distance traveled by the eluting solution):

Record $\boldsymbol{D}$ values for each eluted spot to the nearest 0.1 cm. Draw an X through any unused boxes.

Ink Color	Distance(s) Traveled by Eluted Spot(s), *D*		

Calculate and record the R_f value for each eluted spot, using Equation 2.3.1:

Ink Color	R_f Value for Eluted Spot		

Show calculations for R_f below for the Purple marker:

Data Analysis

1. This lab employs chromatography to separate the components in ink. What other applications can we use chromatography for?
2. Are any of the R_f values in the table above the same (or similar, within 0.1)? What would the same or similar R_f values indicate?
3. What are the units for R_f values?
4. Record ink colors in the appropriate column; for mixtures, list the colors of the components.

Mixtures/Component Colors	Single-Component Colors

5. Which color travelled farthest with the mobile phase? What does this indicate about that component's attraction to the mobile phase? To the stationary phase?
6. Which of the materials tested (gel ink pen and colored markers) had an R_f value of zero? What does this indicate about that component's attraction to the mobile phase? To the stationary phase?
7. Why is it necessary to use a pencil to mark the lines and X's on the paper?
8. The air we breathe is a mixture of different components. What is the composition of dry air, by percentage? See Fig 1.2 in Chapter 1 of the text.

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2.4: Chemical Nomenclature (Experiment)

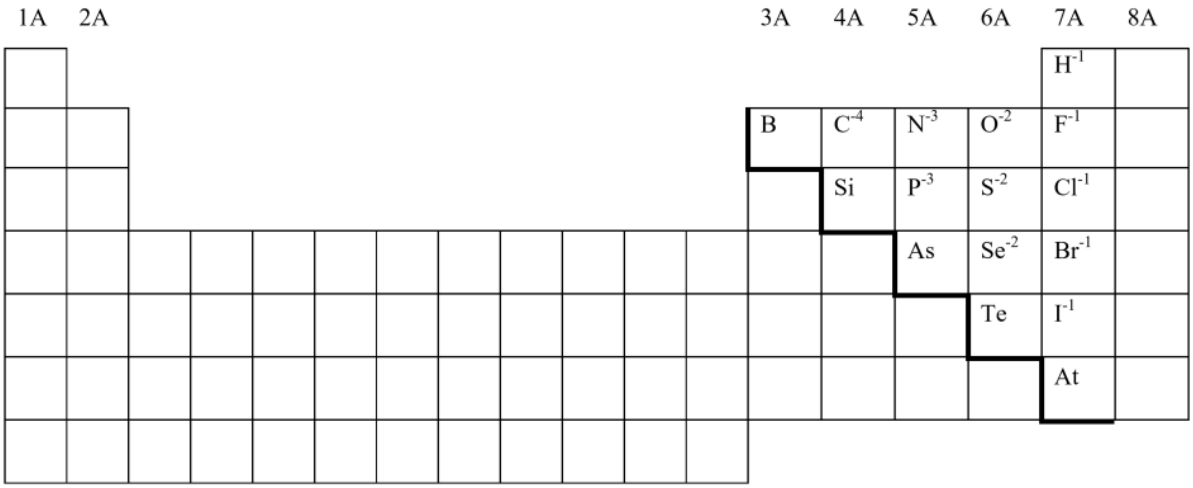
Nomenclature of Ionic Compounds

Ionic compounds are composed of ions. An ion is an atom or molecule with an electrical charge. Monatomic ions are formed from single atoms that have gained or lost electrons. Polyatomic ions are formed from molecules (groups of atoms bonded together) that have gained or lost electrons.

Negative ions are called anions, and are formed when an atom or molecule gains electrons. All non- metals form negatively charged ions. Positive ions are called cations, and are formed when an atom or molecule loses electrons. All metals form positively charged cations. Ions with opposite charges (positive metal cations and negative non-metal anions) will experience a strong electrostatic attraction and form an ionic bond, which leads to the formation of the ionic compound.

Non-metal Anions

Non-metals will form anions with only one possible negative charge. The following Periodic Table shows the charges for non-metal anions commonly found in ionic compounds:



Note that

- The magnitude of the negative charges on these anions is equal to 8 minus their Group Number.
- The names of these anions are based on the element names, but the endings are all changed to –ide.

Example 3.1 :

Sulfur in Group 6A forms anions with a $[8–6] = 2$, or, **–2 charge**.

Example 3.2 :

The Cl^{-1} anion is called the chloride ion.

Metal Cations

Most (but not all) main group metals will form cations with only one possible charge. Most (but not all) transition metals will form cations with more than one possible charge. The following Periodic Table shows the charges for metal cations commonly found in ionic compounds:

1A	2A	Transition Elements (B)									3A	4A	5A
H^{-1}													
Li^{+1}	Be^{+2}												
Na^{+1}	Mg^{+2}										Al^{+3}		
K^{-1}	Ca^{+2}		Ti^{+2}, Ti^{+4}		Cr^{+2} Cr^{+3} Cr^{+6}	Mn^{+2} Mn^{+3} Mn^{+4}	Fe^{+2} Fe^{+3}	Co^{+2} Co^{+3}	Ni^{+2} Ni^{+3}	Cu^{+1} Cu^{+2}	Zn^{+2}	Ga^{+3}	Ge	
Rb^{+1}	Sr^{+2}									Ag^{+1}	Cd^{+2}	In^{+1} In^{+3}	Sn^{+2} Sn^{+4}	Sb
Cs^{+1}	Ba^{+2}									Au^{+1} Au^{+3}	Hg_2^{+2} Hg^{+2}		Pb^{+2} Pb^{+4}	Bi^{+3} Bi^{-5}

Note that

- The magnitude of the positive charge on the main group metal cations is generally equal to their Group Number.
- The names of metal cations with only one possible charge are the same as the names of the metals themselves.
- For metal cations with more than one possible charge, the ion charge must be indicated in the ion name. In the IUPAC system, the ion charge is indicated in the name as Roman numerals in brackets.

Example 3.3 :

Magnesium in Group 2A forms cations with a **+2 charge**.

Example 3.4 :

Al^{+3} is called the **aluminum** cation.

Ag^{+1} is called the **silver** cation.

Example 3.5 :

Pb^{+2} is called the **lead(II)** cation.

Pb^{+4} is called the **lead(IV)** cation.

Polyatomic Ions

Polyatomic ions are formed from molecules (groups of atoms bonded together) that have gained or lost electrons. The table below includes a list of common polyatomic ions that must be memorized.

OH^{-1}	Hydroxide	O_2^{-2}	Peroxide
CN^{-1}	Cyanide	CO_3^{-2}	Carbonate
SCN^{-1}	Thiocyanate	SO_3^{-2}	Sulfite
HCO_3^{-1}	Bicarbonate (Hydrogen Carbonate)	SO_4^{-2}	Sulfate
HSO_3^{-1}	Bisulfite (Hydrogen Sulfite)	$S_2O_3^{-2}$	Thiosulfate
HSO_4^{-1}	Bisulfate (Hydrogen Sulfate)	$C_2O_4^{-2}$	Oxalate
$C_2H_3O_2^{-1}$	Acetate	CrO_4^{-2}	Chromate
NO_2^{-1}	Nitrite	$Cr_2O_7^{-2}$	Dichromate
NO_3^{-1}	Nitrate		
MnO_4^{-1}	Permanganate	PO_3^{-3}	Phosphite
ClO^{-1}	Hypochlorite	PO_4^{-3}	Phosphate
ClO_2^{-1}	Chlorite		
ClO_3^{-1}	Chlorate	NH_4^{+1}	Ammonium
ClO_4^{-1}	Perchlorate	Hg_2^{+2}	Mercury (I)

Note that

- Almost all the polyatomic ions are negatively charged anions.
- Most of the names of polyatomic anions end in either –ate or –ite. The –ate's always have one more oxygen than the -ite's.

Example 3.6 :

Sulfate, SO_4^{-2}, has one more oxygen atom present than sulfite, SO_3^{-2}.

Formulas and Names of Ionic Compounds

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

Writing Formulas

1. Determine the formulas and charges on the cation and anion involved in the compound.
2. Combine the ions in a ratio that results in the formation of a neutral ionic compound. In other words, the total charge of all the positive cations must equal the total charge of all the negative anions in the compound. The numbers of each element present in the compound are shown as subscripts after the element symbol.

Writing Names

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

Example 3.7 :

Write the formula and name for the compound formed between calcium and fluorine

Ca (metal) forms a +2 cation Ca^{+2}, the **calcium** cation.

F (non-metal) forms a -1 anion F^{-1}, the **fluoride** anion.

To obtain a neutral compound, 1 Ca^{+2} is needed for every 2 F^{-1}

The formula of the compound is CaF_2

The name of the compound is **Calcium Fluoride**

Example 3.8 :

Write the formula for iron(III) chloride.

First identify the cation and the anion in this compound.

Cation = iron(III) = Fe^{+3} (transition metal cation)

Anion = chloride = Cl^{-1} (non-metal anion)

To obtain a neutral compound, **1** Fe^{+3} is needed for every **3** Cl^{-1}

The formula of the compound is $FeCl_3$

Example 3.9 :

Write the formula for magnesium phosphate.

First identify the cation and anion in this compound.

Cation = magnesium = Mg^{+2} (metal cation)

Anion = phosphate = PO_4^{-3} (polyatomic anion)

To obtain a neutral compound, **3** Mg^{+2} are needed for every **2** PO_4^{-3}

The formula of the compound is $Mg_3(PO_4)_2$

Note in the above Example 9 that parentheses are placed around the polyatomic portion of compound, to indicate that it must be treated as a complete and whole unit.

Example 3.10 :

Name the ionic compound $Al(NO_3)_3$.

First identify the cation and anion in this compound.

The name of this compound is **Aluminum Nitrate**

Example 3.11 :

Name the compound TiO_2.

First identify the cation and anion in this compound.

The name of the compound is **Titanium(IV) Oxide**

Nomenclature of Simple Covalent Compounds

Covalent compounds are compounds formed between non-metals only. Simple binary covalent compounds contain just two different types of non-metal elements. When non-metals combine they can form several different covalent compounds. These

compounds must therefore be identified with unique names and formulas.

Example 3.12 :

Carbon and oxygen combine to form two common covalent compounds CO_2 and CO.

Formulas and Names of Simple Covalent Compounds

1. Always write/name the element with more metallic character first. Metallic character increases going from right to left, and top to bottom on the Periodic Table.
2. Then write/name the second (less metallic) element, changing the ending of its name to -ide.
3. Since nonmetals often combine in different proportions to form a number of different compounds, prefixes must be included in the names to indicate the numbers of each kind of atom present.Prefixes for 1-10 atoms are given in the following table.

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

Example 3.13 :

A compound contains 3 atoms of sulfur and 4 atoms of phosphorus. Write its name and formula.

Since **phosphorus is the more metallic element** (left of sulfur), it must be named first.

Sulfur being the less metallic element is named second with an -ide ending = **sulfide**.

The prefix for three is **tri**, and the prefix for four is **tetra**.

The name of this compound is **tetraphosphorus trisulfide**

The formula of this compound is P_4S_3

Example 3.14 :

Write the name of the compound N_2O.

The **two N atoms** require the prefix **di** in the name.

The **one O atom** requires the prefix **mono** in the name. The ending of oxygen must be changed to –ide = **oxide**.

The name of this compound is **dinitrogen monoxide**

There are two important exceptions to the naming rules outlined so far:

- Never use the prefix “mono” for the first element, even if just one atom is present.
- Never use any prefixes at all for simple covalent compounds containing Hydrogen.

Example 3.15 :

Write the name of the compound BCl_3.

Although just one B atom is present, **the prefix mono is not used** since it is the first element in this compound.

The **three Cl atoms** require the prefix **tri** in the name. The ending of chlorine must be changed to –ide = **chloride**.

The name of this compound is **boron trichloride**

Example 3.16 :

HF is **hydrogen fluoride** not hydrogen monofluoride or monohydrogen monofluoride.

Please note that many simple covalent compounds have common, rather than systematic names. Please memorize the common names of the following three compounds:

- $\mathbf{H_2O}$ water
- $\mathbf{NH_3}$ ammonia
- $\mathbf{CH_4}$ methane

Covalent compounds containing more than two non-metal elements become increasingly more difficult to name, and common names for these compounds are more extensively used. You will not have to learn these yet.

Nomenclature of Acids

Acids are compounds that release hydrogen cations ($\mathbf{H^{+1}}$) when dissolved in water. They are all found in the aqueous state (aq).

$$\mathbf{HCl\,(aq) \longrightarrow H^{+1}\,(aq) + Cl^{-1}\,(aq)} \tag{2.4.1}$$

$$\mathbf{\text{hydrochloric acid} \longrightarrow \text{hydrogen ions} + \text{chloride ions}} \tag{2.4.2}$$

$$\mathbf{HNO_3\,(aq) \longrightarrow H^{+1}\,(aq) + NO_3{}^{-1}\,(aq)} \tag{2.4.3}$$

$$\mathbf{\text{nitric acid} \longrightarrow \text{hydrogen ions} + \text{nitrate ions}} \tag{2.4.4}$$

$$\mathbf{H_2SO_3\,(aq) \longleftrightarrow 2\,H^{+1}\,(aq) + SO_3{}^{-2}\,(aq)} \tag{2.4.5}$$

$$\mathbf{\text{sulfurous acid} \longleftrightarrow \text{hydrogen ions} + \text{sulfite ions}} \tag{2.4.6}$$

In acids the element hydrogen actually behaves like a Group 1A metal cation. Since it behaves like a +1 cation, hydrogen is always written first in the formulas of all acids. The anion in the acid can be either monatomic or polyatomic, and affects how the acid is named.

Acids containing Non-Metal Anions

These acids contain the $\mathbf{H^{+1}}$ cation and a monatomic non-metal anion.

Example 3.17 :

Hydrochloric acid, **HCl** (*aq*), contains $\mathbf{H^{+1}}$ and the monatomic anion, **chloride**, $\mathbf{Cl^{-1}}$.

Acids containing monatomic anions are named using the prefix hydro + the name of the anion with the suffix -ic + the word acid.

Example 3.18 :

Name the acid **HBr** (*aq*).

This acid contains $\mathbf{H^{+1}}$ and the monatomic anion, **bromide**, $\mathbf{Br^{-1}}$.

The name of this acid is **hydro** + brom-**ic** + **acid** = **hydrobromic acid**

The formulas of these acids are obtained in an identical fashion to regular ionic compounds. The $\mathbf{H^{+1}}$ cation and the monatomic anion are combined in a ratio to yield a neutral compound.

Example 3.19 :

Write the formula and name for the acid containing the phosphide anion.

Just combine them together like any other cation and anion: $3\ H^{+1} : 1\ P^{-3}$

aq)

H_3P (*aq*) is named **hydro** + phosphor-**ic** + **acid** = **hydrophosphoric acid**

Acids containing Polyatomic Anions

These acids contain the H^{+1} cation and a polyatomic anion.

Example 3.20 :

Nitric acid, HNO_3 (*aq*), contains H^{+1} and the polyatomic anion, **nitrate**, NO_3^{-1}.

These acids have names that are based on the name of the polyatomic ion in the acid. If the polyatomic ion has the ending -ate, in the acid the ending is changed to -ic + acid. If the polyatomic ion has the ending -ite, in the acid the ending is changed to -ous + acid.

Example 3.21 :

Name the acid $HClO_3$ (*aq*).

This acid contains H^{+1} and the polyatomic anion ClO_3^{-1} = **chlorate**.

To name this acid, the ending **-ate** is switched to **-ic**.

The acid $HClO_3$ (*aq*) is thus called **chlor<u>ic</u> acid**

Example 3.22 :

Name the acid H_2SO_3 (*aq*).

This acid contains H^{+1} and the polyatomic anion SO_3^{-2} = **sulfite**.

To name this acid, the ending **-ite** is switched to **-ous**.

The acid H_2SO_3 (*aq*) is thus named **sulfur<u>ous</u> acid**

Again, the formulas of these acids are obtained in an identical fashion to regular ionic compounds. The H^{+1} cation and the polyatomic anion are combined in a ratio to yield a neutral compound.

Example 3.23 :

Write the formula for oxalic acid.

Oxalic acid must contain (by reverse logic) the polyatomic anion **oxalate** = $C_2O_4^{-2}$, as well as H^{+1} (the cation in all acids).

Once the ions have been identified, just combine them together like any other cation and anion: $2\ H^{+1} : 1\ C_2O_4^{-2}$

The formula of oxalic acid is $H_2C_2O_4$ (*aq*)

Nomenclature of Hydrates

A hydrate is typically an ionic compound with a certain number of water molecules loosely bound to it.

The general formula of a hydrate is $MX \cdot n\,H_2O$ (*s*), where M is the cation in the ionic compound, X is the anion in the ionic compound and $n\,H_2O$ are the water molecules loosely bound to the ionic compound.

Hydrates are named by writing the name of the ionic compound first, followed by the word “hydrate”. To indicate the number of water molecules present, prefixes must be used.

Example 3.24 :

Name the hydrate $\mathbf{MgSO_4 \cdot 7\,H_2O}$.

$\mathbf{MgSO_4}$ is the ionic compound **magnesium sulfate**.

Since there are **seven water molecules** present, the correct prefix to use is **hepta**.

The name of this hydrate is **magnesium sulfate heptahydrate**

Example 3.25 :

Write the formula for copper(II) chloride dihydrate.

Copper(II) chloride has the formula $\mathbf{CuCl_2}$

The prefix **di** indicates that there are **two** water molecules present.

The formula of this hydrate is $\mathbf{CuCl_2 \cdot 2\,H_2O}$

The water molecules in a hydrate can be removed with relative ease by heating the hydrate. The ionic compound that remains after heating is called an anhydrous salt.

$$\mathbf{MX \cdot \mathit{n}\,H_2O\,(s) \longrightarrow MX\,(s) + \mathit{n}\,H_2O} \tag{2.4.7}$$

$$\mathbf{\text{Hydrate} \longrightarrow \text{Anhydrous salt} + \text{Free Water}} \tag{2.4.8}$$

Often the anhydrous salt has a completely different color and texture from the hydrate.

Example 3.26 :

Copper(II) sulfate pentahydrate is blue and crystalline, whereas anhydrous copper(II) sulfate is white and powdery

Lab Report: Chemical Nomenclature

Write the names and formulas for the following inorganic compounds in the spaces provided.

Part 1: Ions and Ionic Compounds

Write formulas/charges or names as appropriate for each of the following monatomic ions.

1. Calcium ion
2. Phosphide ion
3. Iodide ion
4. Gallium ion
5. Titanium(IV) ion
6. $\mathbf{C^{-4}}$
7. $\mathbf{Rb^{+1}}$
8. $\mathbf{Pb^{+4}}$
9. $\mathbf{S^{-2}}$
10. $\mathbf{Cr^{+2}}$

Write formulas or names as appropriate for each of the following ionic compounds.

1. Magnesium nitride
2. Lithium oxide
3. Aluminum sulfite

4. Copper(II) bicarbonate
5. Sodium nitrate
6. $\mathrm{SrI_2}$
7. $\mathrm{Ba_3(PO_4)_2}$
8. $\mathrm{(NH_4)_2O}$
9. $\mathrm{Fe(ClO)_3}$
10. $\mathrm{ZnCrO_4}$

Part 2: Covalent Compounds

Write formulas or names as appropriate for each of the following covalent compounds.

1. Dichlorine monoxide
2. Disulfur dichloride
3. Carbon tetrafluoride
4. Phosphorus pentachloride
5. Nitrogen tribromide
6. $\mathrm{AsI_3}$
7. $\mathrm{P_4O_{10}}$
8. $\mathrm{Cl_2O_7}$
9. $\mathrm{SeCl_6}$
10. NO

Part 3: Acids

Write formulas or names as appropriate for each of the following acids.

1. Hydroiodic acid
2. Carbonic acid
3. Chlorous acid
4. Sulfuric acid
5. Phosphorous acid
6. HCN (*aq*)
7. $\mathrm{H_2C_2O_4}$ (*aq*)
8. $\mathrm{HNO_2}$ (*aq*)
9. $\mathrm{H_2Cr_2O_7}$ (*aq*)
10. $\mathrm{HMnO_4}$ (*aq*)

Part 4: Hydrates (optional – check with your instructor to see if you are responsible for this section)

Write formulas or names as appropriate for each of the following hydrates.

1. Magnesium sulfate heptahydrate
2. Copper(I) sulfate pentahydrate
3. Potassium phosphate decahydrate
4. Calcium chloride hexahydrate
5. Iron(III) nitrate nonahydrate
6. $\mathrm{CoSO_4 \cdot H_2O}$
7. $\mathrm{Na_2CrO_4 \cdot 4\,H_2O}$
8. $\mathrm{CuF_2 \cdot 2\,H_2O}$
9. $\mathrm{Sr(NO_3)_2 \cdot 6\,H_2O}$
10. $\mathrm{ZnSO_4 \cdot 7\,H_2O}$

Part 5: Nomenclature of Ionic Compounds, Covalent Compounds and Acids

	Classification	Name or Formula
C_3O_2		
IF_7		
Rb_2CO_3		
SnS_2		
$Au(CN)_3$		
H_2CrO_4 (*aq*)		
H_3P (*aq*)		
Li_3PO_4		
Mg_3N_2		
$Ti(C_2H_3O_2)_4$		
Fe_2O_3		
NaH		
Br_3O_8		
MnS_2O_3		
NH_4NO_2		
$Cd(ClO_2)_2$		
$Ba(HSO_3)_2$		
Cu_2O		
$NiBr_3$		
$Sr(OH)_2$		
Perchloric acid		
Potassium permanganate		
Calcium hydride		
Vanadium(II) bicarbonate		
Bismuth(V) nitrate		
Rubidium peroxide		

Strontium hydrogen phosphite		
Hydrofluoric acid		
Chromium(III) thiocyanate		
Acetic acid		
Molybdenum(IV) carbonate		

Tetraiodine nonaoxide		
Diphosphorus tetrafluoride		
Aluminum sulfate		
Ammonium hydroxide		
Sodium dichromate		
Carbon disulfide		
Nickel(II) oxalate		
Barium selenide		
Silver bisulfate		

Questions

1. How are the following types of compounds recognized from their *formulas*?

- Ionic
- Covalent
- Acid

2. When do *parentheses* appear in the formulas of ionic compounds?
3. Do *Roman Numerals* appear in the names of ionic or covalent compounds? Explain why they are used.
4. Do *Greek Prefixes* appear in the names of ionic or covalent compounds? Explain why they are used.
5. What is the relationship between the *number of hydrogens* in an acid and the *charge on the anion* that they are combined with?

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2.5: Explore and Identify the Types of Chemical Reactions around You-Home

Learning Objectives

- To classify chemical reactions into the five basic types
- To balance chemical equations
- To write ionic equations for aqueous reactions
- To conduct reactions using house hold items and analyze the result based on the above mentioned objectives

Theory

Watch the following video to review the general types of reactions mentioned in there including the types of aqueous displacement reactions based on the visible products. *You are not required to purchase any study materials mentioned in the video.*



All chemical reactions are atom-economical, all atoms found in the reagents at the beginning should be present in the products at the end of the reaction, even though the bonding and arrangements would be different. This is in accordance with the law of conservation of mass. *Matter is neither created nor destroyed in a chemical reaction*.The total mass of reagents before the reaction, and the total mass of products after the reaction should be the same. In order to check the mass balance, the number of atoms on each side of reaction are tallied for each type of element present. If the numbers are balanced so also the mass. Watch the following video to review and practice how to balancing chemical equations.



Double displacement reactions can be further classified based on the visible products (evidence). The visible products are no more in the aqueous phase. Aqueous reactions are classified as precipitation (if a solid **(s)** is formed), gas evolution (if a gas, **(g)** is

formed), and acid-base (if **$H_2O(l)$** is formed, and drastic color change is observed) reactions based on the products being formed (chemicals on the right side of the equation arrow in a chemical equation).

Observation on mixing the reagents	**Inference**
Turbidity/opaqueness/solid	Precipitation (PPT) or (s)
Gas bubbles	Gas evolution reaction (GAS) or (g)
Color change to the pH indicator or the pH strips	Acid-base reaction (AB) or (H_2O(l)

You need to check the product side of an equation to identify that particular product, and that is the "net product" of the reaction.

For example, **$Na_2SO_4(aq) + CaCl_2(aq)$ à $CaSO_4$ (s) + 2NaCl(aq)** This is a precipitation reaction, because one of the products is a solid (s). This is the "net product".

The rest of the compounds labeled as aqueous are soluble in water and they exist as solvated ions floating around in water, not as a neutral compounds. Converting the molecular equation into the ionic form is called ionic equation. All ionic compounds soluble in water (aq) are split as ions except the net product which exist in a different phase than aqueous phase ((s), (g), (l)).

$2Na^+(aq) + (SO_4)^{2-}(aq) + Ca^{2+}(aq) + 2Cl^-(aq)$ à $CaSO_4$ (s) + 2Na(aq) + $2Cl^-$(aq)

Don't forget to insert the ionic charge for the ions. All equations should be balanced as well. The following video will help you review the steps involved in writing the ionic and net-ionic equations.



Procedure

Materials Required

Vinegar (acetic acid) 5%, Baking Soda (sodium bicarbonate), paper towels, table salt, cola (with phosphoric acid in the ingredients), Epsom salt, graduated cylinder, pipet, electronic scale, spoon, beakers or cups to hold the solutions, pH strips, apple, small pieces of the following metals that you can find at home; copper (copper wire or a penny), silver(optional) (wire or a piece of broken jewelry), aluminum (aluminum foil), and iron (nail or other iron metal pieces)

Part A: What Reactions do the Metals around You Undergo?

In the first video on the types of reactions you might have seen the type of reaction that iron metal undergo if left unprotected. Have you ever observed some other common metals around you and see how they react if left unprotected in the nature. In order to speed up the process, a little bit of vinegar is applied on the metal surface.

- Try to see whether you can find a pure copper wire. Otherwise, take a penny. Write down the appearance of the metal. Apply some vinegar/water on it to speed-up the reaction. Leave it on a moist paper towel for 4-5 days, and see whether you see any changes in its appearance.
- Optional-Take a silver piece, or a piece of silver jewelry and repeat the process. Record the change.

- Repeat the above experiment with a piece of iron, say iron nail or steel wool.
- Repeat the process with a piece of Aluminum foil.
- Have you ever observed any changes happening to gold ornaments above 18-carat?
- Cut an apple and leave a small piece outside for 4-5 hours. Observe the change in its appearance. Apples are rich in iron content.
- Write balanced equations for all the reactions where you observe a change. Identify the reaction types for all those reactions.

Part B: Compare and Contrast the Types of Reactions

You are going to watch one reaction in the following video and perform two others on your own. At the end compare and contrast the three reactions and explain the chemistry behind the reactions.

Reaction 1

- The work bench should be cleared out before this reaction. Not even your lab book or laptop should be on the work bench. Measure out 10.0ml of 5% vinegar (acetic acid) in a graduated cylinder. Check the pH of the acid using a pH strip. Note the color of the strip. Weigh out 0.7 gram of baking soda (sodium bicarbonate) and add into the measured vinegar in the graduated cylinder. Write down your observation. Check the pH of the left over solution after the reaction. Note the color of the pH strip.
- Take 20 ml of any type of fresh (from an un-opened bottle) soda or cola in a graduated cylinder, add a couple of scoops of table salt into this. Record your observation.
- You might have seen the elephant toothpaste experiment before. Here is a video of that experiment. You do not have to do this experiment. Just watch the video carefully, record the reagents used. Write down the observation.



- Compare all the three reactions to see the similarity and difference between the reaction types. Write an equation for each chemical reaction. Circle/highlight the visible product.
- Identify the net products in the first reaction involving acetic acid. Write the ionic and net-ionic equation for that reaction.

Part C: Reactions o Epsom Salt

- Weigh out 5 gram Epsom salt ($MgSO_4$) into a beaker or a cup. Measure out 10 ml water into this using a graduated cylinder. Stir this well with a spoon to dissolve the salt.
- Add a cola (with phosphoric acid ingredient) into this drop-by-drop using a pipet. Keep on adding until you see a visible product. Write down your observation .
- Assuming that the active ingredient in the cola as Phosphoric acid (H_3PO_4), write down a balanced molecular equation for the reaction between the acid and Epsom salt ($MgSO_4$). Identify the visible net product. Write the ionic and net-ionic equation for the reaction.

Report Sheet

Part A: What Reactions do the Metals around You Undergo?

Metal	Appearance before the reaction	Appearance after	Write down the type of reaction from the five types mentioned. Write a balanced equation if you have observed any change.
Cu			
Ag (Optional)			
Fe			
Al			
Apple			

Part B: Compare and Contrast the Types of Reactions

Reaction	Reaction Type	Observation	Balanced Equation for the Reaction
1-Acetic acid and sodium bicarbonate			
1. Continues-Complete this row for the above reaction		pH/color before pH/color after	Net products- Ionic Equation- Net-ionic equation
2. Cola and Table salt			
3. Elephat toothpaste reaction (video)			

Part C: Reaction of Epsom Salt

Reaction	Reaction Type	Observation	Balanced Equation for the Reaction
Epsom salt and Phosphoric acid (from the cola)			
Write down the following,			

Net product-
Ionic Equation-
Net-ionic equation

Questions

a. Based on what you have learned, how would you explain the color of the Statue of Liberty. What would have been a better choice of metal for making such statues?

b. Hope you have heard about stalagmites. If not search a little bit and explain what stalgmites are, how are they formed. Explain the type of reaction for the formation with a balanced equation.

c. Which out of the five types of reaction is used in water softners? Write a generic equation to show the reaction.

d. Complete and balance the following reaction between heptane (a hydrocarbon) and oxygen gas.

$O_2 + C_7H_{16}$ -->

Contributors and Attributions

Manjusha Saraswathiamma, Minnesota State Community and Technical College, Moorhead has developed this experiment to perform at a homeschool setting with less hazardous and cost effective lab supplies. The author would like to acknowledge the creators of all the YouTube videos embedded on this page.

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2.6: Mole Ratios and Limiting Reactant-Home

Learning Objectives

- To calculate the moles of the reagents used
- To perform stoichiometric calculations to identify the limiting and excess reactants
- Predict the theoretical yield in mols and grams

Theory

Learning Objectives

- Explain the concepts of theoretical yield and limiting reactants/reagents.
- Derive the theoretical yield for a reaction under specified conditions.
- Calculate the percent yield for a reaction.

The relative amounts of reactants and products represented in a balanced chemical equation are often referred to as *stoichiometric amounts*. All the exercises of the preceding module involved stoichiometric amounts of reactants. For example, when calculating the amount of product generated from a given amount of reactant, it was assumed that any other reactants required were available in stoichiometric amounts (or greater). In this module, more realistic situations are considered, in which reactants are not present in stoichiometric amounts.

Limiting Reactant

Consider another food analogy, making grilled cheese sandwiches (Figure **2.6.1**):

$$\textbf{1 slice of cheese} + \textbf{2 slices of bread} \rightarrow \textbf{1 sandwich} \tag{2.6.1}$$

Stoichiometric amounts of sandwich ingredients for this recipe are bread and cheese slices in a 2:1 ratio. Provided with 28 slices of bread and 11 slices of cheese, one may prepare 11 sandwiches per the provided recipe, using all the provided cheese and having six slices of bread left over. In this scenario, the number of sandwiches prepared has been *limited* by the number of cheese slices, and the bread slices have been provided in *excess*.

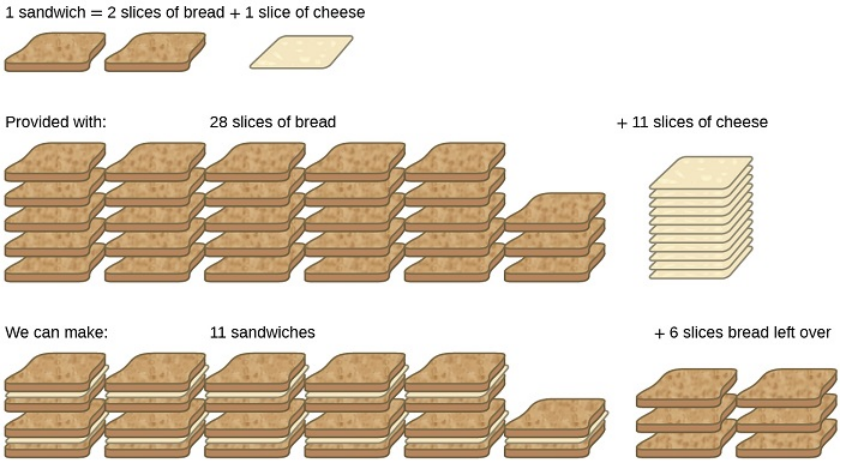


Figure **2.6.1**: Sandwich making can illustrate the concepts of limiting and excess reactants.

This figure has three rows showing the ingredients needed to make a sandwich. The first row reads, “1 sandwich = 2 slices of bread + 1 slice of cheese.” Two slices of bread and one slice of cheese are shown. The second row reads, “Provided with: 28 slices of bread + 11 slices of cheese.” There are 28 slices of bread and 11 slices of cheese shown. The third row reads, “We can make: 11 sandwiches + 6 slices of bread left over.” 11 sandwiches are shown with six extra slices of bread.

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$\mathrm{H_2}(g) + \mathrm{Cl_2}(g) \rightarrow 2\,\mathrm{HCl}(g)$$

The balanced equation shows the hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the limiting reactant, and the other substance is the excess reactant. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation.

For example, imagine combining 6 moles of H_2 and 4 moles of Cl_2. Identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant.

For the example in the previous paragraph, the complete reaction of the hydrogen would yield

$$\text{mol HCl produced} = 6 \text{ mol } H_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } H_2} = 12 \text{ mol HCl}$$

The complete reaction of the provided chlorine would produce

$$\text{mol HCl produced} = 4 \text{ mol } Cl_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } Cl_2} = 8 \text{ mol HCl}$$

The chlorine will be completely consumed once 8 moles of HCl have been produced. Since enough hydrogen was provided to yield 12 moles of HCl, there will be non-reacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (Figure **2.6.2**). To determine the amount of excess reactant that remains, the amount of hydrogen consumed in the reaction can be subtracted from the starting quantity of hydrogen.

The amount of hydrogen consumed is

$$\text{mol } H_2 \text{ produced} = 8 \text{ mol HCl} \times \frac{1 \text{ mol } H_2}{2 \text{ mol HCl}} = 4 \text{ mol } H_2$$

Subtract the hydrogen consumed from the starting quantity

$$\text{mole of excess } H_2 = 6 \text{ mol } H_2 \text{ starting} - 4 \text{ mol } H_2 \text{ consumed} = 2 \text{ mol } H_2 \text{ excess}$$

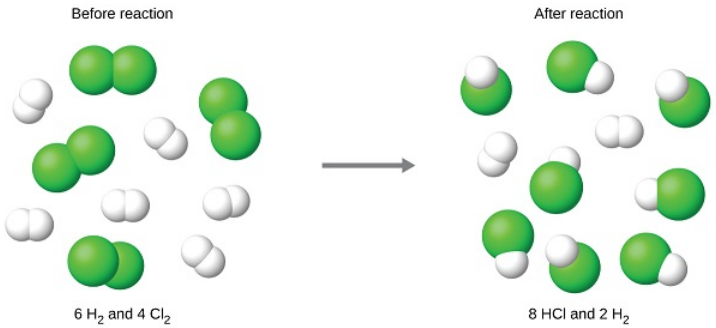


Figure **2.6.2**: When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant.

The figure shows a space-filling molecular models reacting. There is a reaction arrow pointing to the right in the middle. To the left of the reaction arrow there are three molecules each consisting of two green spheres bonded together. There are also five molecules each consisting of two smaller, white spheres bonded together. Above these molecules is the label, "Before reaction," and below these molecules is the label, "6 H subscript 2 and 4 C l subscript 2." To the right of the reaction arrow, there are eight molecules each consisting of one green sphere bonded to a smaller white sphere. There are also two molecules each consisting of two white spheres bonded together. Above these molecules is the label, "After reaction," and below these molecules is the label, "8 H C l and 2 H subscript 2."

Example 2.6.1: Identifying the Limiting Reactant

Silicon nitride is a very hard, high-temperature-resistant ceramic used as a component of turbine blades in jet engines. It is prepared according to the following equation:

$$3\,Si(s) + 2\,N_2(g) \rightarrow Si_3N_4(s)$$

Which is the limiting reactant when 2.00 g of Si and 1.50 g of N_2 react?

Solution

Compute the provided molar amounts of reactants, and then compare these amounts to the balanced equation to identify the limiting reactant.

$$\text{mol Si} = 2.00\ \cancel{\text{g Si}} \times \frac{1\ \text{mol Si}}{28.09\ \cancel{\text{g Si}}} = 0.0712\ \text{mol Si}$$

$$\text{mol } N_2 = 1.50\ \cancel{\text{g } N_2} \times \frac{1\ \text{mol } N_2}{28.02\ \cancel{\text{g } N_2}} = 0.0535\ \text{mol } N_2$$

The provided Si:N_2 molar ratio is:

$$\frac{0.0712\ \text{mol Si}}{0.0535\ \text{mol } N_2} = \frac{1.33\ \text{mol Si}}{1\ \text{mol } N_2}$$

The stoichiometric Si:N_2 ratio is:

$$\frac{3\ \text{mol Si}}{2\ \text{mol } N_2} = \frac{1.5\ \text{mol Si}}{1\ \text{mol } N_2}$$

Comparing these ratios shows that Si is provided in a less-than-stoichiometric amount, and so is the limiting reactant.

Alternatively, compute the amount of product expected for complete reaction of each of the provided reactants. The 0.0712 moles of silicon would yield

$$\text{mol } Si_3N_4 \text{ produced} = 0.0712\ \text{mol Si} \times \frac{1\ \text{mol } Si_3N_4}{3\ \text{mol Si}} = 0.0237\ \text{mol } Si_3N_4$$

while the 0.0535 moles of nitrogen would produce

$$\text{mol } Si_3N_4 \text{ produced} = 0.0535\ \text{mol } N_2 \times \frac{1\ \text{mol } Si_3N_4}{2\ \text{mol } N_2} = 0.0268\ \text{mol } Si_3N_4$$

Since silicon yields the lesser amount of product, it is the limiting reactant.

Exercise 2.6.1

Which is the limiting reactant when 5.00 g of H_2 and 10.0 g of O_2 react and form water?

Answer

O_2

Percent Yield

The amount of product that *may be* produced by a reaction under specified conditions, as calculated per the stoichiometry of an appropriate balanced chemical equation, is called the theoretical yield of the reaction. In practice, the amount of product obtained is called the actual yield, and it is often less than the theoretical yield for a number of reasons. Some reactions are inherently inefficient, being accompanied by *side reactions* that generate other products. Others are, by nature, incomplete (consider the partial reactions of weak acids and bases discussed earlier in this text). Some products are difficult to collect without some loss, and so less than perfect recovery will reduce the actual yield. The extent to which a reaction's theoretical yield is achieved is commonly expressed as its percent yield:

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Actual and theoretical yields may be expressed as masses or molar amounts (or any other appropriate property; e.g., volume, if the product is a gas). As long as both yields are expressed using the same units, these units will cancel when percent yield is calculated.

Example 2.6.2: Calculation of Percent Yield

Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

$$\mathrm{CuSO_4}(aq) + \mathrm{Zn}(s) \rightarrow \mathrm{Cu}(s) + \mathrm{ZnSO_4}(aq)$$

What is the percent yield?

Solution

The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield is found by the approach illustrated in the previous module, as shown here:

Double subscripts: use braces to clarify

Using this theoretical yield and the provided value for actual yield, the percent yield is calculated to be

$$\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100$$

$$\text{percent yield} = \left(\frac{0.392\ \text{g Cu}}{0.5072\ \text{g Cu}}\right) \times 100$$
$$= 77.3\%$$

Exercise 2.6.2

What is the percent yield of a reaction that produces 12.5 g of the gas Freon $\mathrm{CF_2Cl_2}$ from 32.9 g of $\mathrm{CCl_4}$ and excess HF?

$$\mathrm{CCl_4} + 2\,\mathrm{HF} \rightarrow \mathrm{CF_2Cl_2} + 2\,\mathrm{HCl}$$

Answer

48.3%

Green Chemistry and Atom Economy

The purposeful design of chemical products and processes that minimize the use of environmentally hazardous substances and the generation of waste is known as *green chemistry*. Green chemistry is a philosophical approach that is being applied to many areas of science and technology, and its practice is summarized by guidelines known as the “Twelve Principles of Green Chemistry”. One of the 12 principles is aimed specifically at maximizing the efficiency of processes for synthesizing chemical products. The *atom economy* of a process is a measure of this efficiency, defined as the percentage by mass of the final product of a synthesis relative to the masses of *all* the reactants used:

$$\text{atom economy} = \frac{\text{mass of product}}{\text{mass of reactants}} \times 100\%$$

Though the definition of atom economy at first glance appears very similar to that for percent yield, be aware that this property represents a difference in the *theoretical* efficiencies of *different* chemical processes. The percent yield of a given chemical process, on the other hand, evaluates the efficiency of a process by comparing the yield of product actually obtained to the maximum yield predicted by stoichiometry.

The synthesis of the common nonprescription pain medication, ibuprofen, nicely illustrates the success of a green chemistry approach (Figure **2.6.3**). First marketed in the early 1960s, ibuprofen was produced using a six-step synthesis that required 514 g of reactants to generate each mole (206 g) of ibuprofen, an atom economy of 40%. In the 1990s, an alternative process was developed by the BHC Company (now BASF Corporation) that requires only three steps and has an atom economy of ~80%, nearly twice that of the original process. The BHC process generates significantly less chemical waste; uses less-hazardous and recyclable materials; and provides significant cost-savings to the manufacturer (and, subsequently, the consumer). In recognition of the positive environmental impact of the BHC process, the company received the Environmental Protection Agency's Greener Synthetic Pathways Award in 1997.

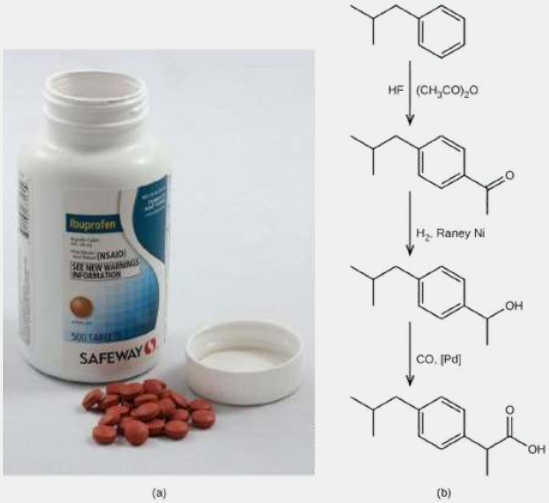


Figure **2.6.3**: (a) Ibuprofen is a popular nonprescription pain medication commonly sold as 200 mg tablets. (b) The BHC process for synthesizing ibuprofen requires only three steps and exhibits an impressive atom economy. (credit a: modification of work by Derrick Coetzee)

Summary

When reactions are carried out using less-than-stoichiometric quantities of reactants, the amount of product generated will be determined by the limiting reactant. The amount of product generated by a chemical reaction is its actual yield. This yield is often less than the amount of product predicted by the stoichiometry of the balanced chemical equation representing the reaction (its theoretical yield). The extent to which a reaction generates the theoretical amount of product is expressed as its percent yield.

Key Equations

- $\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100$

Glossary

actual yield
amount of product formed in a reaction

excess reactant
reactant present in an amount greater than required by the reaction stoichiometry

limiting reactant
reactant present in an amount lower than required by the reaction stoichiometry, thus limiting the amount of product generated

percent yield
measure of the efficiency of a reaction, expressed as a percentage of the theoretical yield

theoretical yield
amount of product that may be produced from a given amount of reactant(s) according to the reaction stoichiometry

Materials Required

Five (16 Fl. Oz) empty water or cola bottles with narrow mouth (any type of bottles of the same size and shape with narrow mouth). Five to six medium sized balloons. Baking soda (Sodium bicarbonate), 5% Vinegar (acetic acid), spoons, 5-cups, funnel (or a paper cone to bes used as a funnel to transfer the powder to the balloon, electronic scale, graduated cylinders, water

Watch the following Youtube video. We will se up a similar experiment.



Instead of the conical flask we will be using five identical water or cola bottles (16 Fl. Oz). Also, we will have a total of five sample instead of the four shown in the solution. The acetic acid we have is the 5% vinegar. 5% vinegar is approximately 0.83M not 1.0M as shown in the video. In order to get 0.19 mol of acetic acid in each bottle, you need to take 229 ml of the 5% (0.83M).

- Use the **M1V1=M2V2** equation to find out the new volume 1.0 M x 190 ml =0.83 M x X ml. On solving for X you will get 229 ml as the new volume.
  - Take five medium size water or cola bottles of the same size and shape (16 Fl. Oz). Label them as 3, 7, 12, 16, and 20 using a marker.
  - Measure out 229 ml of 0.83M acetic (5% vinegar) into each bottle.
  - Measure 3 g, 7 g, 12 g, 16 g, 20 g sodium bicarbonate ($NaHCO_3$) or baking soda separately in small containers.
  - Without spilling add the powder into five similar sized balloons using a funnel, and keep them aside.
  - Align each balloon by the side of each bottle. The balloons should be attached on the mouth of the bottles tightly while taking care not to spill the powder into the flask. Hold the balloons tight as shown in the video, as it might fly off as large amount of gas is produced.
  - Once all the balloons are attached tightly, all the balloons should be flipped to release the powder into the bottle as shown in the video.
  - Record the observation on the report sheet. Take a lab picture
  - Compare the volume of the balloons with the amount of the gas particles produced. Observe the trend in the volume and enter it in a separate sheet. On a scale of 1-4, 1 being the smallest volume. Use this for the gas law lab if you are yet to complete it.

- Following is the reaction between the acetic acid and sodium bicarbonate producing water, sodium acetate, and carbon dioxide gas.
- Keep this **observation and data sheet to reuse** in the lab for **Gas Law in case you are doing this lab first.**
- Calculate the number of moles of bicarbonate using the mole to gram conversion factor using the formula mass of the bicarbonate.
- Write the ratio between the moles of acid and bicarbonate as n1:n2. 1 is acid and 2 is bicarbonate.
- Since this is a 1:1 reaction, the smaller of the two will be the limiting reactant. Identify and label the limiting reactant of the two.
- Predict the theoretical yield of carbon dioxide. The reaction is also 1:1 reagents to products. The mols of product formed is limited by the mols of the limiting reactant. It should be the same as the mols of the limiting reactant as it is a 1:1 reaction.
- Convert the theoretical yield in mols of carbon dioxide to grams using the mole to gram conversion factor

Report Sheet: Wet Lab to Identify the Limiting and Excess Reactant and the Theoretical Yield

Constant Temperature= K (Room temperature) Constant Pressure = 1.0 Atm						
#	Mols of Acetic acid used, **n1**, mol	Grams of Sodium bicarbonate used	Moles, **n2**, of bicarbonate used	Mol ratios of Acetic acid to bicarbonate as (**n1:n2**). Find the smallest ratio by dividing through out with the smaller of the two numbers.	Identify the limiting reactant out of the two reagents. Write down the moles of the limiting reactant	Moles of Carbon dioxide produced (assume 1:1 mol ratio). This is equal the moles of the limiting reactant. Report in both mols and grams.
1		3.0 g	mol			= mol = grams
2		7.0 g	mol			
3		12.0 g	mol			
4		16.0 g	mol			
5		32.0 g	mol			

Data Analysis-Reaction between Acetic Acid and Sodium Bicarbonate

Questions

1. Identify the limiting reactant for entry number 2. Explain the reasoning.
2. Identify the limiting reactant for entry number 4. Explain the reasoning.
3. Identify the limiting reactant for entry number 5. Explain the reasoning.
4. Show the work. Convert 3 grams of Sodium bicarbonate to mols.
5. Show the work to convert 0.19 mols of Carbon dioxide to grams.

? Exercise 2.6.1

For the following reaction, identify the limiting reactant if 300.2 grams of Copper (Cu-Molar mass-63.55 g/mol) reacts with 567.8 grams of Sulfur (S8-Molar mass 256g/mol). Following is the balanced equation. Identify the limiting and excess reactant based on the mole ratios as shown in the examples given in the theory part at the beginning. Calculate the theoretical yield of copper sulfide (Molar mass 159.16g/mol) in mols and grams.

16 Cu + 1 S_8 à 8 Cu_2S

Answer

Cu is the limiting reactant.

Contributors and Attribution

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2.7: Gas Laws-Simulations and Wet Lab-Home

Learning Objectives

- To experiment with variables to reach gas law relationships
- To verify Avogadro's law
- To improve problem-solving skills by practicing gas-law problems

Theory

A gas is the state of matter that is characterized by having neither a fixed shape nor a fixed volume. Gases exert pressure, are compressible, have low densities and diffuse rapidly when mixed with other gases. On a microscopic level, the molecules (or atoms) in a gas are separated by large distances and are in constant, random motion.

Four measurable properties can be used to describe a gas quantitatively: pressure (P), volume (V), temperature (T) and mole quantity (n). The relationships among these properties are summarized by the Gas Laws, as shown in the table below.

Charles's Law	$V \propto T$	P and n are held constant As gas temperature increases, gas volume increases.	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Boyle's Law	$V \propto \frac{1}{P}$	T and n are held constant As gas pressure increases, gas volume decreases.	$P_1V_1 = P_2V_2$
Avogadro's Law	$V \propto n$	P and T are held constant As the number of moles of gas increase, gas volume increases.	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$
Combined Law	$V \propto \frac{T}{P}$	n is held constant Obtained by combining Boyle's Law and Charles's Law.	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

A closer look at the Combined Law reveals that the volume of a gas depends on both the pressure and temperature. Thus, if the volumes of two gases are to be compared, they must be under the same P and T. A commonly used set of P and T reference conditions is known as Standard Temperature and Pressure, or STP. Standard temperature is defined as exactly 0 °C (273 K) and standard pressure is defined as exactly 1 atm (760 mmHg).

The Ideal Gas Law is obtained by combining Boyle's Law, Charles's Law and Avogadro's Law together:

$$PV = nRT \tag{2.7.1}$$

Here, P represents as the gas pressure (in atmospheres); V is the gas volume (in Liters); n is the number of moles of gas in the sample; T is the gas temperature (in Kelvins). R is a proportionality constant called the Gas Constant, and has a theoretical value of 0.08206 $\frac{L \cdot atm}{mol \cdot K}$. Note that the units of R will allow the units of P, V, n and T in the Ideal Gas Law to cancel correctly.

If you observe a parameter (dependent variable), say volume by changing any one of the other parameters (independent variable) like pressure, (keep the rest at constant of fixed value) you may come across some kind of data trends regarding the dependence of volume on increasing or decreasing the pressure. If you follow the results of such experiments by systematic observation and analysis you can hypothesize the relationship between the parameters. You can repeat the same experiment many times. If the trend is generalizable it could be stated as a LAW. This is how many science laws are postulated, including the simple gas laws. There are three simple gas laws studying

the various parameters like temperature (T), pressure (P), and number of moles of gas particles (n) on the volume (V) of a gas. Following are the laws and the parameters they correlate

- Boyle's Law (relates Volume and Pressure: **V** and **P**)
- Charles Law (relates Volume and Temperature: **V** and **T**)
- Avogadro's Law (relates Volume and Number of particles: **V** and **n**)

In **PART A** of the experiment a free online simulation on all the four parameters (V, T, P, and n) will be accessed to play around with the parameters to see how they are related, and to predict the trend in the results.

In **PART B** of the experiment a gas evolution reaction will be followed to observe the effect of adding more particles (n) on the volume (V) of the gas to postulate the Avogadro's law. A graphical analysis of the data will also be done.

Procedure

Materials Required

A laptop to play the gas law simulation online, Five (16 Fl. Oz) empty water or cola bottles with narrow mouth (any type of bottles of the same size and shape with narrow mouth). Five to six medium sized balloons. Baking soda (Sodium bicarbonate), 5% Vinegar (acetic acid), spoons, 5-cups, funnel (or a paper cone to bes used as a funnel to transfer the powder to the balloon, electronic scale, graduated cylinders, water

Part A: Simulations to Verify Simple Gas Laws and Ideal Gas Law

The major difference between solid, liquid, and gas phases is how the particles in a phase are packed. Component particles in a gas are widely separated as if there is no interaction between them. Because of this reason, a gas can be compressed further and further down to smaller volumes (space). Most of the gases are colorless. Therefore, it is very difficult to observe them for experimental purposes. However, some parameters of gases are measurable like their volume, pressure, temperature, and number (moles) of particles present.

- Turn on your lap top. Access the internet and the go to the free online gas simulation website provided by your instructor
- Connect to the webpage, and play around with the simulation by changing volume, pressure, temperature and number of molecules by clicking and dragging the various bars after selecting one.

Boyle's law

- Click on the pressure button on the drag and drop bar on the right hand side. Keep the number of moles (n) and Temperature fixed at constant values. Record them. Now go ahead and drag the button on pressure bar and leave at any five different values and observe the volume reading. Record five different pressure values and the corresponding volumes on the observation sheet. This can be used to predict the relationship between the pressure and volume. This relationship is the concept behind the Boyle's law.

Charle's law

- Click on the volume button on the drag and drop bar on the right hand side. Keep the number of moles (n) and Pressure fixed at constant values. Write them down. Now go ahead and drag the button on temperature (T) bar and leave at any five different values and observe the volume

reading. Record all the five different Temperature values and the corresponding volumes on the observation sheet. This can be used to predict the relationship between the temperature and volume. This relationship is the concept behind the Charle's law.

Ideal gas law

- Click on the volume button on drag and drop bar on the right hand side. Play around with the temperature (T) and Pressure at various values. Now go ahead and . Record all the five different corresponding sets of Volume, Pressure, and Temperature values. Calculate the number of moles using the ideal gas equation. The value of Ideal gas constant, R should be taken as 0.08206 L.atm/mol.K.

Part B: Wet Lab to Test Avogadro's Law

Watch the following Youtube video. We will se up a similar experiment.



Instead of the conical flask we will be using five identical water or cola bottles (16 Fl. Oz). Also, we will have a total of five sample instead of the four shown in the solution. The acetic acid we have is the 5% vinegar. 5% vinegar is approximately 0.83M not 1.0M as shown in the video. In order to get 0.19 mol of acetic acid in each bottle, you need to take 229 ml of the 5% (0.83M).

- Use the **M1V1=M2V2** equation to find out the new volume 1.0 M x 190 ml =0.83 M x X ml. On solving for X you will get 229 ml as the new volume.
  - Take five medium size water or cola bottles of the same size and shape (16 Fl. Oz). Label them as 3, 7, 12, 16, and 20 using a marker.
  - Measure out 229 ml of 0.83M acetic (5% vinegar) into each bottle.
  - Measure 3 g, 7 g, 12 g, 16 g, 20 g sodium bicarbonate ($NaHCO_3$) or baking soda separately in small containers.
  - Without spilling add the powder into five similar sized balloons using a funnel, and keep them aside.
  - Align each balloon by the side of each bottle. The balloons should be attached on the mouth of the bottles tightly while taking care not to spill the powder into the flask. Hold the balloons tight as shown in the video, as it might fly off as large amount of gas is produced.

- Once all the balloons are attached tightly, all the balloons should be flipped to release the powder into the bottle as shown in the video.
- Record the observation on the report sheet. Take a lab picture
- Compare the volume of the balloons with the amount of the gas particles produced. Observe the trend in the volume and enter it in the observation sheet.
- Following is the reaction between the acetic acid and sodium bicarbonate producing water, sodium acetate, and carbon dioxide gas. Analyze the relationship between the volume and number of moles of gas particles produced and state your hypothesis.
- Keep this **observation and data sheet to reuse** in the lab for **Mole ratios and limiting reactant**

Report Sheet

Part A: Simulations to Verify Simple Gas Laws and Ideal Gas Law

Boyle's Law

Constant Temperature= K		Constant number of moles = mol
#	Pressure (Atm)	Volume (L)
1		
2		
3		
4		
5		

Charle' Law

Constant Pressure = Atm		Constant number of moles = mol
#	Temperature (K)	Volume (L)
1		
2		
3		
4		
5		

Ideal Gas Law

Calculate the number of mols, n for each set of data. Use the value of Ideal gas constant R as 0.08206 L.atm/mol.K				
#	P, Pressure (Atm)	V, Volume (L)	T, Temperature (K)	? mols (n)
1				
2				
3				
4				
5				

Report Sheet: Part B-Wet Lab to Test Avogadro's Law

Avogadro's Law

Constant Temperature= K (Room temperature) Constant Pressure = 1.0 Atm				
#	Grams of Sodium bicarbonate used	Moles, n, of bicarbonate used	Moles of Carbon dioxide produced (assume 1:1 mol ratio)	Rate the size of the inflated balloons on a scale of 1-4, 1 being the smallest
1	3.0 g	mol		
2	7.0 g	mol		
3	12.0 g	mol		
4	16.0 g	mol		
5 (Do not use this data for this lab. Reserve this for the Moles and Limiting Reactant Lab)	32.0 g	mol		

Data Analysis

1. On analyzing the data in Part A simulation for Boyle's law, how are the Volume and Pressure related? Write your statement. Does this match with the theoretical statement given at the beginning of the experiment?

2. On analyzing the data in Part A simulation for Charle's law, how are the Volume and Temperature related? Write your statement. Does this match with the theoretical statement given at the beginning of the experiment?

3. On analyzing the data in Part B for Avogadro's law, how are the Volume (size of the balloons) and the number of moles of the gas produced related? Write your statement. Does this match with the theoretical statement given at the beginning of the experiment?

4. On analyzing the data in Part A simulation for the Ideal gas law, how are the number of moles related? Are you getting similar values or very different values? Explain your findings

5. Plot a graph with your independent variable in Part A Boyle's law, pressure vs. volume. Give a proper title, and label the axes with the parameters and the units.

Practice Problems

? Exercise 2.7.1

If 3.7L of an ideal gas at 35.8 degree Celsius and 1.5atm is heated to 373.5 K, what would be the new pressure as the gas has expanded to 7.8 L?

Answer

0.59 atm

? Exercise 2.7.2

Pick one of the entries from Part A-Ideal gas law simulation. Write the number of moles calculated in one of those five data points. Convert the number of moles to the mass of the gas assuming that the gas in the simulation is Carbon monoxide or CO.

Answer

Clue: n mol=Mass (g)/Formula mass (g/mol). Therefore, Mass (g)= n mol X Formula mass (g/mol)

Contributions and Attributions

Manjusha Saraswathiamma, Minnesota State Community and Technical College, Moorhead has developed this experiment to conduct this at a homeschool setting using less hazardous and cost-effective materials. The author would like to acknowledge the creators (Ohio State University) of the Youtube video used in this experiment. A portion of the Theory part is a content reuse form Chem 10 Experiment **Experimental Determination of the Gas Constant (Experiment)** by Santa Monica College is licensed CC BY-NC 4.0.

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2.8: Heat of a Reaction and Coffee Cup Calorimeter-Home

Learning Objectives

- To observe and measure heat exchanged in a chemical reaction and to come up with appropriate inferences based on previous knowledge
- To demonstrate the concepts involved in calorimetric measurements
- To calculate the heat change involved in the reaction (q)
- To classify the reaction as endothermic or exothermic

Theory

One technique we can use to measure the amount of heat involved in a chemical or physical process is known as calorimetry. Calorimetry is used to measure amounts of heat transferred to or from a substance. To do so, the heat is exchanged with a calibrated object (calorimeter). The change in temperature of the measuring part of the calorimeter is converted into the amount of heat (since the previous calibration was used to establish its heat capacity). The measurement of heat transfer using this approach requires the definition of a system (the substance or substances undergoing the chemical or physical change) and its surroundings (the other components of the measurement apparatus that serve to either provide heat to the system or absorb heat from the system). Knowledge of the heat capacity of the surroundings, and careful measurements of the masses of the system and surroundings and their temperatures before and after the process allows one to calculate the heat transferred as described in this section.

A calorimeter is a device used to measure the amount of heat involved in a chemical or physical process. For example, when an exothermic reaction occurs in solution in a calorimeter, the heat produced by the reaction is absorbed by the solution, which increases its temperature. When an endothermic reaction occurs, the heat required is absorbed from the thermal energy of the solution, which decreases its temperature (Figure **2.8.1**). The temperature change, along with the specific heat and mass of the solution, can then be used to calculate the amount of heat involved in either case.

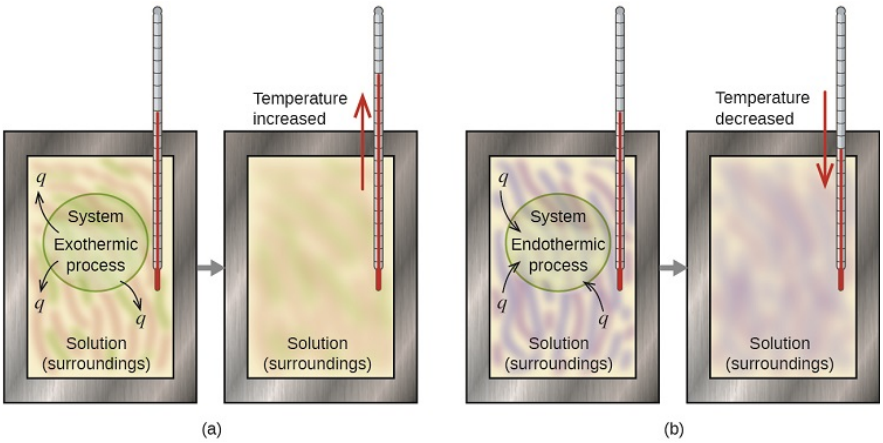


Figure **2.8.1**: In a calorimetric determination, either (a) an exothermic process occurs and heat, q, is negative, indicating that thermal energy is transferred from the system to its surroundings, or (b) an endothermic process occurs and heat, q, is positive, indicating that thermal energy is transferred from the surroundings to the system.

> By convention, q is given a negative (-) sign when the system releases heat to the surroundings (exothermic); q is given a positive (+) sign when the system absorbs heat from the surroundings (endothermic).

Scientists use well-insulated calorimeters that all but prevent the transfer of heat between the calorimeter and its environment. This enables the accurate determination of the heat involved in chemical processes, the energy content of foods, and so on. General chemistry students often use simple calorimeters constructed from polystyrene cups (Figure **2.8.2**). These easy-to-use “coffee cup” calorimeters allow more heat exchange with their surroundings, and therefore produce less accurate energy values.

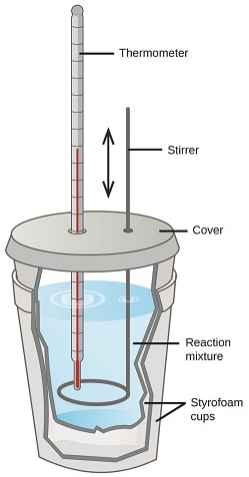


Figure 2.8.2: A simple calorimeter can be constructed from two polystyrene cups. A thermometer and stirrer extend through the cover into the reaction mixture.

Commercial solution calorimeters are also available. Relatively inexpensive calorimeters often consist of two thin-walled cups that are nested in a way that minimizes thermal contact during use, along with an insulated cover, handheld stirrer, and simple thermometer. More expensive calorimeters used for industry and research typically have a well-insulated, fully enclosed reaction vessel, motorized stirring mechanism, and a more accurate temperature sensor (Figure 2.8.3).

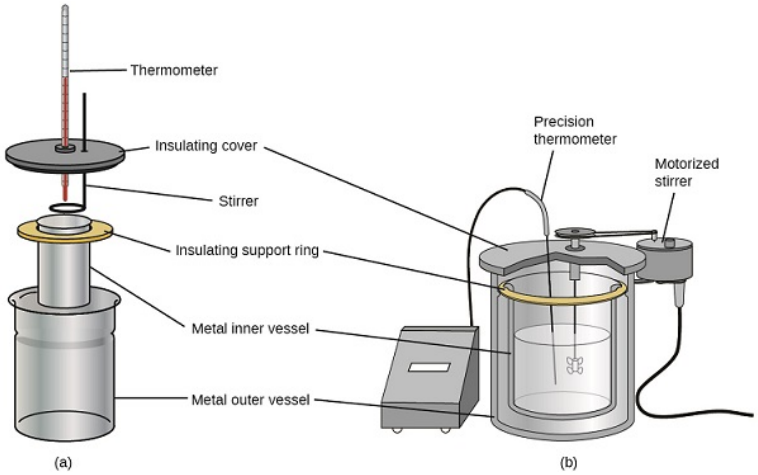


Figure 2.8.3: Commercial solution calorimeters range from (a) simple, inexpensive models for student use to (b) expensive, more accurate models for industry and research.

Before we practice calorimetry problems involving chemical reactions, consider a simple example that illustrates the core idea behind calorimetry. Suppose we initially have a high-temperature substance, such as a hot piece of metal (M), and a low-temperature substance, such as cool water (W). If we place the metal in the water, heat will flow from M to W. The temperature of M will decrease, and the temperature of W will increase, until the two substances have the same temperature—that is, when they reach thermal equilibrium (Figure 2.8.4). If this occurs in a calorimeter, ideally all of this heat transfer occurs between the two substances, with no heat gained or lost by either the calorimeter or the calorimeter's surroundings. Under these ideal circumstances, the net heat change is zero:

$$q_{\text{substance M}} + q_{\text{substance W}} = 0 \tag{2.8.1}$$

This relationship can be rearranged to show that the heat gained by substance M is equal to the heat lost by substance W:

$$q_{\text{substance M}} = -q_{\text{substance W}} \tag{2.8.2}$$

The magnitude of the heat (change) is therefore the same for both substances, and the negative sign merely shows that $q_{substance\ M}$ and $q_{substance\ W}$ are opposite in direction of heat flow (gain or loss) but does not indicate the arithmetic sign of either q value (that is determined by whether the matter in question gains or loses heat, per definition). In the specific situation described, $q_{substance\ M}$ is a negative value and $q_{substance\ W}$ is positive, since heat is transferred from M to W.

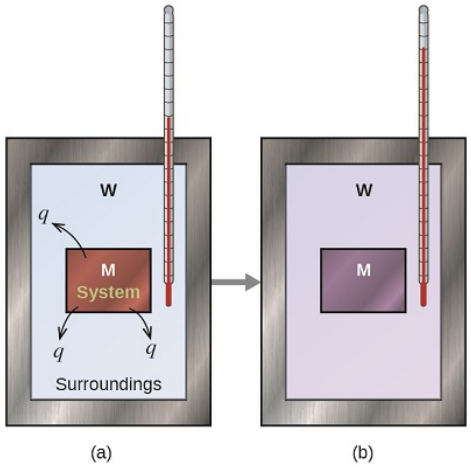


Figure **2.8.4**: In a simple calorimetry process, (a) heat, q, is transferred from the hot metal, M, to the cool water, W, until (b) both are at the same temperature.

✔ Example 2.8.1: Heat Transfer between Substances at Different Temperatures

A hot 360-g piece of rebar (a steel rod used for reinforcing concrete) is dropped into 425 mL of water at 24.0 °C. The final temperature of the water is measured as 42.7 °C. Calculate the initial temperature of the piece of rebar. Assume the specific heat of steel is approximately the same as that for iron (Table T4), and that all heat transfer occurs between the rebar and the water (there is no heat exchange with the surroundings).

Solution

The temperature of the water increases from 24.0 °C to 42.7 °C, so the water absorbs heat. That heat came from the piece of rebar, which initially was at a higher temperature. Assuming that all heat transfer was between the rebar and the water, with no heat "lost" to the surroundings, then heat given off by rebar = − heat taken in by water, or:

$$q_{\text{rebar}} = -q_{\text{water}}$$

Since we know how heat is related to other measurable quantities, we have:

$$(c \times m \times \Delta T)_{\text{rebar}} = -(c \times m \times \Delta T)_{\text{water}}$$

Letting f = final and i = initial, in expanded form, this becomes:

$$c_{\text{rebar}} \times m_{\text{rebar}} \times (T_{\text{f,rebar}} - T_{\text{i,rebar}}) = -c_{\text{water}} \times m_{\text{water}} \times (T_{\text{f,water}} - T_{\text{i,water}})$$

The density of water is 1.0 g/mL, so 425 mL of water = 425 g. Noting that the final temperature of both the rebar and water is 42.7 °C, substituting known values yields:

$$(0.449\ \text{J/g °C})(360\text{g})(42.7\text{°C} - T_{\text{i,rebar}}) = -(4.184\ \text{J/g °C})(425\ \text{g})(42.7\text{°C} - 24.0\text{°C})$$

$$T_{\text{i,rebar}} = \frac{(4.184\ \text{J/g °C})(425\ \text{g})(42.7\text{°C} - 24.0\text{°C})}{(0.449\ \text{J/g °C})(360\ \text{g})} + 42.7\text{°C}$$

Solving this gives Ti,rebar= 248 °C, so the initial temperature of the rebar was 248 °C.

? Pre-lab Question-1- 2.8.1A

A 248-g piece of copper is dropped into 390 mL of water at 22.6 °C. The final temperature of the water was measured as 39.9 °C. Calculate the initial temperature of the piece of copper. Assume that all heat transfer occurs between the copper and the water.

Answer

The initial temperature of the copper was 335.6 °C.

This method can also be used to determine other quantities, such as the specific heat of an unknown metal.

Example 2.8.2: Identifying a Metal by Measuring Specific Heat

A 59.7 g piece of metal that had been submerged in boiling water was quickly transferred into 60.0 mL of water initially at 22.0 °C. The final temperature is 28.5 °C. Use these data to determine the specific heat of the metal. Use this result to identify the metal.

Solution

Assuming perfect heat transfer, heat given off by metal = −heat taken in by water, or:

$$q_{\text{metal}} = -q_{\text{water}}$$

In expanded form, this is:

$$c_{\text{metal}} \times m_{\text{metal}} \times (T_{\text{f,metal}} - T_{\text{i,metal}}) = -c_{\text{water}} \times m_{\text{water}} \times (T_{\text{f,water}} - T_{\text{i,water}})$$

Noting that since the metal was submerged in boiling water, its initial temperature was 100.0 °C; and that for water, 60.0 mL = 60.0 g; we have:

$$(c_{\text{metal}})(59.7\ \text{g})(28.5°\text{C} - 100.0°\text{C}) = -(4.18\ \text{J/g °C})(60.0\ \text{g})(28.5°\text{C} - 22.0°\text{C})$$

Solving this:

$$c_{\text{metal}} = \frac{-(4.184\ \text{J/g °C})(60.0\ \text{g})(6.5°\text{C})}{(59.7\ \text{g})(-71.5°\text{C})} = 0.38\ \text{J/g °C}$$

Comparing this with values in Table T4, our experimental specific heat is closest to the value for copper (0.39 J/g °C), so we identify the metal as copper.

Pre-lab Question 2 2.8.2

A 92.9-g piece of a silver/gray metal is heated to 178.0 °C, and then quickly transferred into 75.0 mL of water initially at 24.0 °C. After 5 minutes, both the metal and the water have reached the same temperature: 29.7 °C. Determine the specific heat and the identity of the metal. (Note: You should find that the specific heat is close to that of two different metals. Explain how you can confidently determine the identity of the metal).

Answer

$c_{metal} = 0.13\ J/g\ °C$

This specific heat is close to that of either gold or lead. It would be difficult to determine which metal this was based solely on the numerical values. However, the observation that the metal is silver/gray in addition to the value for the specific heat indicates that the metal is lead.



Conservation of Energy: The Movement of Heat between Substances: https://youtu.be/pGEYy-pNHBg

When we use calorimetry to determine the heat involved in a chemical reaction, the same principles we have been discussing apply. The amount of heat absorbed by the calorimeter is often small enough that we can neglect it (though not for highly accurate measurements, as discussed later), and the calorimeter minimizes energy exchange with the surroundings. Because energy is neither created nor destroyed during a chemical

reaction, there is no overall energy change during the reaction. The heat produced or consumed in the reaction (the "system"), $q_{reaction}$, plus the heat absorbed or lost by the solution (the "surroundings"), $q_{solution}$, must add up to zero:

$$q_{reaction} + q_{solution} = 0 \tag{2.8.3}$$

This means that the amount of heat produced or consumed in the reaction equals the amount of heat absorbed or lost by the solution:

$$q_{reaction} = -q_{solution} \tag{2.8.4}$$

This concept lies at the heart of all calorimetry problems and calculations.

Example 2.8.3: Heat Produced by an Exothermic Reaction

When 50.0 mL of 0.10 M HCl(aq) and 50.0 mL of 1.00 M NaOH(aq), both at 22.0 °C, are added to a coffee cup calorimeter, the temperature of the mixture reaches a maximum of 28.9 °C. What is the approximate amount of heat produced by this reaction?

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

Solution

To visualize what is going on, imagine that you could combine the two solutions so quickly that no reaction took place while they mixed; then after mixing, the reaction took place. At the instant of mixing, you have 100.0 mL of a mixture of HCl and NaOH at 22.0 °C. The HCl and NaOH then react until the solution temperature reaches 28.9 °C.

The heat given off by the reaction is equal to that taken in by the solution. Therefore:

$$q_{reaction} = -q_{solution}$$

(It is important to remember that this relationship only holds if the calorimeter does not absorb any heat from the reaction, and there is no heat exchange between the calorimeter and its surroundings.)

Next, we know that the heat absorbed by the solution depends on its specific heat, mass, and temperature change:

$$q_{solution} = (c \times m \times \Delta T)_{solution}$$

To proceed with this calculation, we need to make a few more reasonable assumptions or approximations. Since the solution is aqueous, we can proceed as if it were water in terms of its specific heat and mass values. The density of water is approximately 1.0 g/mL, so 100.0 mL has a mass of about 1.0×10^2 g (two significant figures). The specific heat of water is approximately 4.18 J/g °C, so we use that for the specific heat of the solution. Substituting these values gives:

$$q_{solution} = (4.184\ \mathrm{J/g\ °C})(1.0 \times 10^2\ \mathrm{g})(28.9°\mathrm{C} - 22.0°\mathrm{C}) = 2.89 \times 10^3\ \mathrm{J}$$

Finally, since we are trying to find the heat of the reaction, we have:

$$q_{reaction} = -q_{solution} = -2.89 \times 10^3\ J$$

The negative sign indicates that the reaction is exothermic. It produces 2.89 kJ of heat.

Pre-lab Question 3 2.8.3

When 100 mL of 0.200 M NaCl(aq) and 100 mL of 0.200 M $AgNO_3$(aq), both at 21.9 °C, are mixed in a coffee cup calorimeter, the temperature increases to 23.5 °C as solid AgCl forms. How much heat is produced by this precipitation reaction? What assumptions did you make to determine your value?

Answer

$1.34 \times 10^3\ J$; assume no heat is absorbed by the calorimeter, no heat is exchanged between the calorimeter and its surroundings, and that the specific heat and mass of the solution are the same as those for water

Thermochemistry of Hand Warmers

When working or playing outdoors on a cold day, you might use a hand warmer to warm your hands (Figure 2.8.5). A common reusable hand warmer contains a supersaturated solution of $NaC_2H_3O_2$ (sodium acetate) and a metal disc. Bending the disk creates nucleation sites around which the metastable $NaC_2H_3O_2$ quickly crystallizes (a later chapter on solutions will investigate saturation and supersaturation in more detail).

The process $\mathrm{NaC_2H_3O_2}(aq) \longrightarrow \mathrm{NaC_2H_3O_2}(s)$ is exothermic, and the heat produced by this process is absorbed by your hands, thereby warming them (at least for a while). If the hand warmer is reheated, the $NaC_2H_3O_2$ redissolves and can be reused.



Figure 2.8.5: Chemical hand warmers produce heat that warms your hand on a cold day. In this one, you can see the metal disc that initiates the exothermic precipitation reaction. (credit: modification of work by Science Buddies TV/YouTube)

Another common hand warmer produces heat when it is ripped open, exposing iron and water in the hand warmer to oxygen in the air. One simplified version of this exothermic reaction is

$$2\,\mathrm{Fe(s)} + \tfrac{3}{2}\,\mathrm{O_2(g)} \longrightarrow \mathrm{Fe_2O_3(s)}.\ n$$

Salt in the hand warmer catalyzes the reaction, so it produces heat more rapidly; cellulose, vermiculite, and activated carbon help distribute the heat evenly. Other types of hand warmers use lighter fluid (a platinum catalyst helps lighter fluid oxidize exothermically), charcoal (charcoal oxidizes in a special case), or electrical units that produce heat by passing an electrical current from a battery through resistive wires.

Example 2.8.4: Heat Flow in an Instant Ice Pack

When solid ammonium nitrate dissolves in water, the solution becomes cold. This is the basis for an "instant ice pack" (Figure 2.8.5). When 3.21 g of solid NH_4NO_3 dissolves in 50.0 g of water at 24.9 °C in a calorimeter, the temperature decreases to 20.3 °C.

Calculate the value of *q* for this reaction and explain the meaning of its arithmetic sign. State any assumptions that you made.



Figure 2.8.5: An instant cold pack consists of a bag containing solid ammonium nitrate and a second bag of water. When the bag of water is broken, the pack becomes cold because the dissolution of ammonium nitrate is an endothermic process that removes thermal energy from the water. The cold pack then removes thermal energy from your body.

Solution

We assume that the calorimeter prevents heat transfer between the solution and its external environment (including the calorimeter itself), in which case:

$$q_{\mathrm{rxn}} = -q_{\mathrm{soln}}$$

with "rxn" and "soln" used as shorthand for "reaction" and "solution," respectively.

Assuming also that the specific heat of the solution is the same as that for water, we have:

$$\begin{aligned} q_{\mathrm{rxn}} &= -q_{\mathrm{soln}} = -(c \times m \times \Delta T)_{\mathrm{soln}} \\ &= -[(4.184\,\mathrm{J/g\,°C}) \times (53.2\,\mathrm{g}) \times (20.3\,\mathrm{°C} - 24.9\,\mathrm{°C})] \\ &= -[(4.184\,\mathrm{J/g\,°C}) \times (53.2\,\mathrm{g}) \times (-4.6\,\mathrm{°C})] \\ &+ 1.0 \times 10^3\,\mathrm{J} = +1.0\,\mathrm{kJ} \end{aligned}$$

The positive sign for *q* indicates that the dissolution is an endothermic process.

? Prelab Question-4 2.8.4

When a 3.00-g sample of KCl was added to 3.00×10^2 g of water in a coffee cup calorimeter, the temperature decreased by 1.05 °C. How much heat is involved in the dissolution of the KCl? What assumptions did you make?

Answer

1.33 kJ; assume that the calorimeter prevents heat transfer between the solution and its external environment (including the calorimeter itself) and that the specific heat of the solution is the same as that for water.

If the amount of heat absorbed by a calorimeter is too large to neglect or if we require more accurate results, then we must take into account the heat absorbed both by the solution and by the calorimeter.



Conservation of Energy: Coffee Cup Calorimetry: https://youtu.be/FwQcc17PN0k

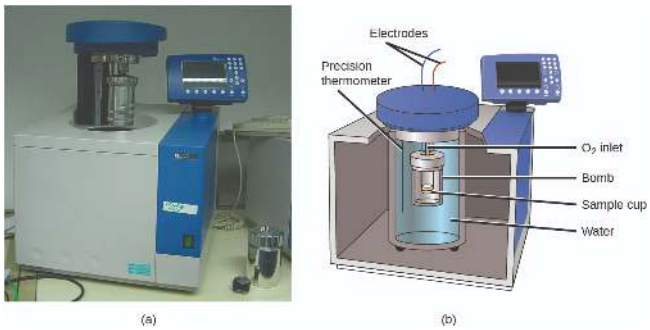
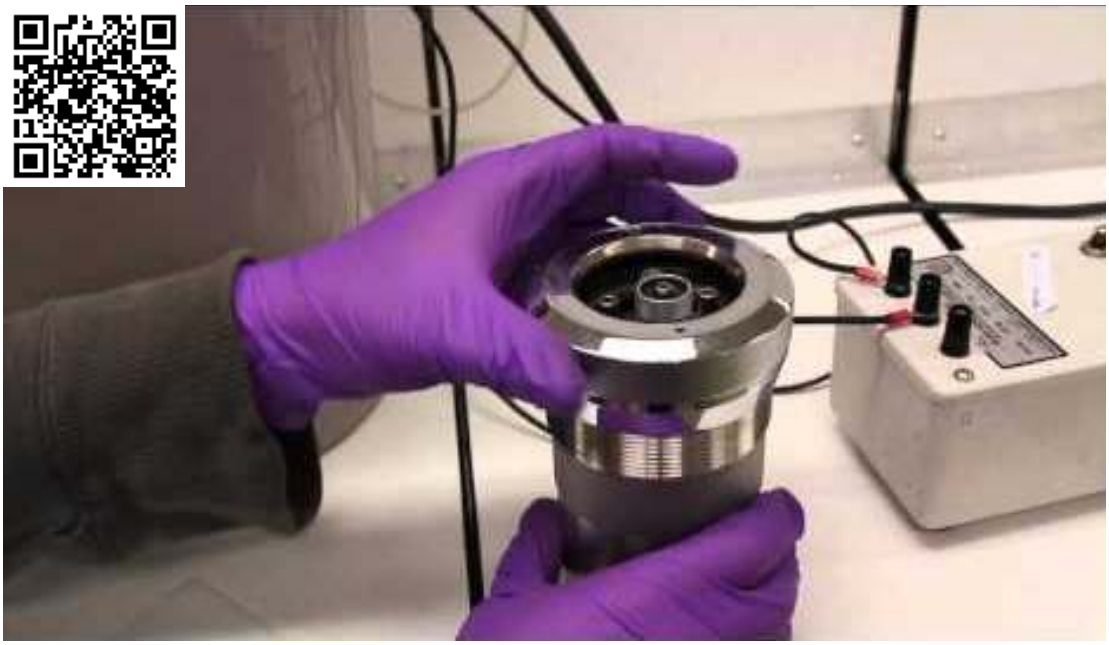


Figure **2.8.6**: (a) A bomb calorimeter is used to measure heat produced by reactions involving gaseous reactants or products, such as combustion. (b) The reactants are contained in the gas-tight "bomb," which is submerged in water and surrounded by insulating materials. (credit a: modification of work by "Harbor1"/Wikimedia commons)

The calorimeters described are designed to operate at constant (atmospheric) pressure and are convenient to measure heat flow accompanying processes that occur in solution. A different type of calorimeter that operates at constant volume, colloquially known as a bomb calorimeter, is used to measure the energy produced by reactions that yield large amounts of heat and gaseous products, such as combustion reactions. (The term "bomb" comes from the observation that these reactions can be vigorous enough to resemble explosions that would damage other calorimeters.) This type of calorimeter consists of a robust steel container (the "bomb") that contains the reactants and is itself submerged in water (Figure **2.8.6**). The sample is placed in the bomb, which is then filled with oxygen at high pressure. A small electrical spark is used to ignite the sample. The energy produced by the reaction is trapped in the steel bomb and the surrounding water. The temperature increase is measured and, along with the known heat capacity of the calorimeter, is used to calculate the energy produced by the reaction. Bomb calorimeters require calibration to determine the heat capacity of the calorimeter and ensure accurate results. The calibration is accomplished using a reaction with a known q, such as a measured quantity of benzoic acid ignited by a spark from a nickel fuse wire that is weighed before and after the reaction. The temperature change produced by the known reaction is used to determine the heat capacity of the calorimeter. The calibration is generally performed each time before the calorimeter is used to gather research data.



Video 2.8.1: Video of view how a bomb calorimeter is prepared for action.

Since the first one was constructed in 1899, 35 calorimeters have been built to measure the heat produced by a living person.[1] These whole-body calorimeters of various designs are large enough to hold an individual human being. More recently, whole-room calorimeters allow for relatively normal activities to be performed, and these calorimeters generate data that more closely reflect the real world. These calorimeters are used to measure the metabolism of individuals under different environmental conditions, different dietary regimes, and with different health conditions, such as diabetes. In humans, metabolism is typically measured in Calories per day. A nutritional calorie (Calorie) is the energy unit used to quantify the amount of energy derived from the metabolism of foods; one Calorie is equal to 1000 calories (1 kcal), the amount of energy needed to heat 1 kg of water by 1 °C.

Summary

Calorimetry is used to measure the amount of thermal energy transferred in a chemical or physical process. This requires careful measurement of the temperature change that occurs during the process and the masses of the system and surroundings. These measured quantities are then used to compute the amount of heat produced or consumed in the process using known mathematical relations. Calorimeters are designed to minimize energy exchange between the system being studied and its surroundings. They range from simple coffee cup calorimeters used by introductory chemistry students to sophisticated bomb calorimeters used to determine the energy content of food.

Footnotes

1. 1 Francis D. Reardon et al. “The Snellen human calorimeter revisited, re-engineered and upgraded: Design and performance characteristics.” *Medical and Biological Engineering and Computing* 8 (2006)721–28, http://link.springer.com/article/10....517-006-0086-5.

Glossary

bomb calorimeter

device designed to measure the energy change for processes occurring under conditions of constant volume; commonly used for reactions involving solid and gaseous reactants or products

calorimeter

device used to measure the amount of heat absorbed or released in a chemical or physical process

calorimetry

process of measuring the amount of heat involved in a chemical or physical process

nutritional calorie (Calorie)

unit used for quantifying energy provided by digestion of foods, defined as 1000 cal or 1 kcal

surroundings

all matter other than the system being studied

system

portion of matter undergoing a chemical or physical change being studied

Procedure

Materials and Equipments

Styrofoam cups-four, cardboard or Styrofoam lids or rectangular pieces to cover the Styrofoam cups, vinegar (acetic acid) (5%), baking soda (sodium bicarbonate), ice cubes, tap water, thermometer, spoon, some small containers to hold the reagents like beakers or clear cups

Part A-Heat Exchanged During Mixing Hot and Cold Water Assuming No Heat Exchange with the Calorimeter



- Stack two Styrofoam cups, one inside the other. Find an insulating lid like a rectangular Styrofoam or cardboard piece. Make a hole on the centre of the lid to insert the thermometer as shown in the figure. This is your calorimeter.
- Keep a container on the electronic scale, tare (zero) the weight of the container.
- Measure out exactly 50 ml of very hot tap water into the container on the scale and record the mass of the 50 ml hot water in grams. Keep it aside.
- Take another small container. Place it on the scale. Tare(zero) the weight of the container. Weigh out exactly 50 ml ice-cold water into the container.
- Add the cold water into the Styrofoam calorimeter that you have set-up in the first step. Cover with the lid. Stir with the thermometer placed through the center of the lid. Record the temperature. This is your initial Temperature, Ti.
- Add the the 50 ml hot water into the Styrofoam calorimeter with the cold water. Hold the lid tight and stir with the thermometer. If the temperature continue to rise on addition, stir continuously with the thermometer and wait until the maximum peak temperature is reached. Record the temperature. This is your final temperature, Tf.
- Add the mass of both cold and hot water together in grams. This is your mass.
- The heat capacity Cs of water is taken as 4.184J/g.^{0}C
- Calculate the heat exchanged in the mixing using the appropriate given in the theory part. The assumption is that the calorimeter is a perfectly insulated system with zero heat exchange and therefore heat gained by the cold water is equal to the heat lost by the hot water. Only the sign conventions are opposite.
- Heat exchanged in a reaction=q=m X Cs X ΔT.

Part B-Heat of the Reaction When Acetic Acid and Sodium Bicarbonate React

- Stack two new Styrofoam cups, one inside the other. Find an insulating lid like a rectangular Styrofoam or cardboard piece. Make a hole on the centre of the lid to insert the thermometer as shown in the figure. This is your calorimeter.
- Keep a container on the electronic scale, tare (zero) the weight of the container.
- Measure out exactly 20 ml of 5% vinegar (acetic acid) into the container on the scale and record the mass of the 20 ml vinegar in grams. Keep it aside.
- Take another small container. Place it on the scale. Tare(zero) the weight of the container. Weigh out exactly 5 grams of baking soda (sodium carbonate) into the container using a spoon. Keep it aside.
- Add the vinegar into the Styrofoam calorimeter that you have set-up in the first step. Record the temperature that is placed into the cup through the lid. This is your initial temperature, Ti.
- Add the weighed out bicarbonate into the Styrofoam calorimeter containing the acetic acid. Hold the lid tight and stir with the thermometer. If the temperature continue to rise on addition, stir continuously with the thermometer and wait until the maximum peak temperature is reached. If it is otherwise, the temperature continue to drop on addition, stir continuously with the thermometer and wait until the lowest temperature is

reached.Record the temperature. This is your final temperature, Tf. The final temperature may be higher or lower than the initial temperature depending on whether it is an exothermic or endothermic reaction.

- Add the mass of the acetic acid and bicarbonate in grams. This should be the mass, m for the calculation.
- The heat capacity Cs is taken for water (4.184J/g.^{0}C), because the reaction medium is water. Vinegar is acetic acid diluted with water.

Lab Report:Part A-Heat Exchanged During Mixing Hot and Cold Water

Mass of cold water, m1	g
Mass of hot water, m2	g
Total mass, m1+m2=m	g
Specific heat capacity of water, Cs	4.184 J/g^{0}C
Initial temperature of cold water, Ti	^{0}C
Final temperature, Tf	^{0}C
Temperature change,(Tf-Ti)= ΔT	
Heat gained by the cold water during mixing, q	J

Lab Report:Part B-Heat of the Reaction When Acetic Acid and Sodium Bicarbonate React

Mass of acetic acid, m1	g
Mass of baking soda, m2	g
Total mass, m1+m2=m	g
Specific heat capacity of water, Cs	4.184J/g.^{0}C
Initial temperature of acetic acid, Ti	^{0}C
Final temperature, Tf	^{0}C
Temperature change,(Tf-Ti)= ΔT	^{0}C
Heat exchanged to the surrounding, q-surrounding or q-solution	J
Heat of the reaction, q-reaction	J

Data Analysis

? Post lab-question 1 2.8.1

Based on the final calculated heat of reaction for Part B, identify the reaction as exothermic or endothermic.

Answer

Endothermic

Contributions and Attributions

The experiment is developed by Manjusha Saraswathiamma, Minnesota State Community College, Moorhead for doing experiments at a homeschool setting with less hazardous and cost-effective lab supplies. The theory part of the experiment is from section **5.2: Calorimetry** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

2.8: Heat of a Reaction and Coffee Cup Calorimeter-Home is shared under a CC BY-NC-SA 4.0 license and was authored, remixed, and/or curated by Manjusha T Saraswathiamma.

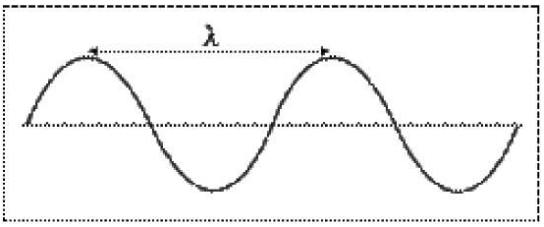
- **5.3: Calorimetry** by OpenStax is licensed CC BY 4.0. Original source: https://openstax.org/details/books/chemistry-2e.

2.9: Flame Tests and Atomic Spectra (Experiment)-Home Version

Objectives

- To perform flame tests of metal cations in order to observe their characteristic colors,
- To perform calculations to determine the frequency and energy of the emitted photons.
- To relate these results to the types of electronic transitions occurring in these elements.
- To observe and understand line emission spectra of atoms using gas-discharge tubes.
- To practice writing electron configurations for these (and other) elements.

Electromagnetic radiation is energy in the form of waves. Waves are characterized by their *wavelength* (λ) and *frequency* (ν). Wavelength is defined as the distance between successive crests (or troughs) on a wave, and is measured in meters. Frequency is defined as the number of waves that pass a given point every second, and is measured in 1/seconds (1/s or s^{-1}), or Hertz (Hz).



All electromagnetic waves travel at the *speed of light* (c), or $3.0 \times 10^8 m/s$. The relationship between the wavelength, frequency and speed of a wave is given by the equation:

$$c = \lambda \times \nu \tag{2.9.1}$$

Electromagnetic radiation also occurs as discreet "packets" called *photons*. The *energy* (E) of a photon (in Joules, J) is given by the equation:

$$E_{photon} = h \times \nu \tag{2.9.2}$$

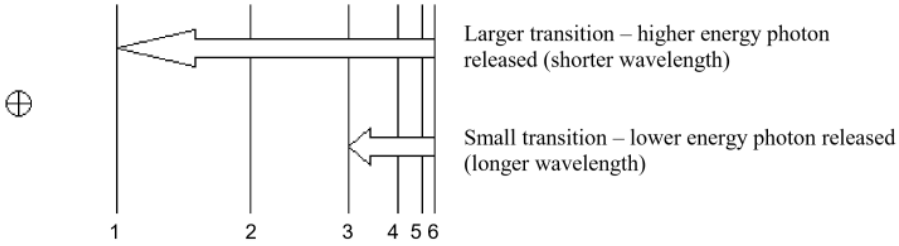
Here, h is *Planck's constant*, which has a value of $6.63 \times 10^{-34} J \cdot s$.

Visible light is the most familiar example of electromagnetic radiation. Differences in the wavelengths of visible light are manifested as different colors, shown in the color spectrum below (colors can be seen in the PDF document on-line). Other examples of electromagnetic radiation include X-rays, ultraviolet light, infrared light, microwaves and radio waves.

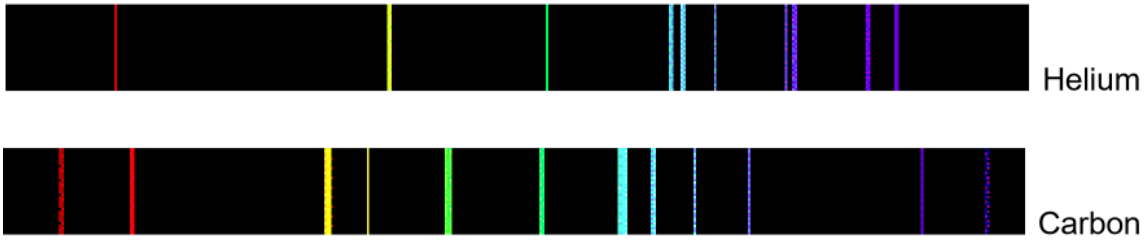


So, how does electromagnetic radiation relate to flame tests? Well, when an atom or ion absorbs energy, its electrons can make transitions from lower energy levels to higher energy levels. The energy absorbed could be in the form of heat (as in flame tests), or electrical energy, or electromagnetic radiation. However, when electrons subsequently return from higher energy levels to lower energy levels, energy is released predominantly in the form of *electromagnetic radiation*.

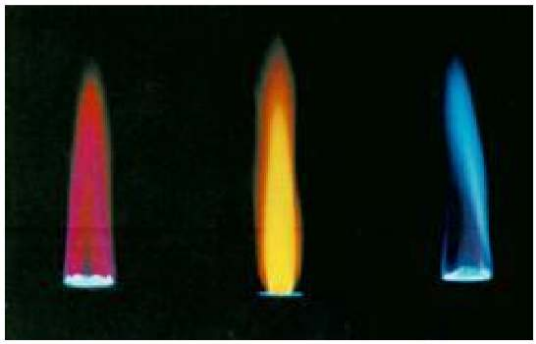
The spacing between energy levels in an atom determines the sizes of the transitions that occur, and thus the energy and wavelengths of the collection of photons emitted.



If emitted photons are in the visible region of the spectrum, they may be perceived as lines of different colors (note that photons outside the visible spectrum may also be emitted, but cannot be seen). The result is called a *line emission spectrum*, and can serve as a 'fingerprint' of the element to which the atoms belong. For example, the line spectra shown below for the elements helium and carbon are clearly quite different (colors can be seen in the PDF document on-line).



Unfortunately, techniques more sophisticated than those used in this lab are required to obtain such line spectra. To the naked eye, when an element is vaporized in a flame (or an electrical discharge)the emission spectrum will appear to be just *one color*. For example, helium gas when excited by an electrical discharge emits light that appears an orange- peach color. This one color results from a *combination of all lines* of the emission spectrum, in proportion to their intensities. As many elements will still produce distinctive colors under such conditions, simple flame tests can be used to identify these elements. In fact, flame tests were used to identify elements long before the invention of modern techniques, such as emission spectroscopy.



Procedure

Materials and Equipment

Usually aqueous solutions of the following metal salts are sparyed into a Bunsen burner : $\mathbf{LiCl}$ (*aq*), $\mathbf{NaCl}$ (*aq*), $\mathbf{KCl}$ (*aq*), $\mathbf{CuCl_2}$ (*aq*), $\mathbf{BaCl_2}$ (*aq*), $\mathbf{CaCl_2}$ (*aq*). Since this is a kitchen lab for online students no real wet experiment needs to be conducted, instead watch the following video to identify the flame color made by the above solutions and write the data in the following table.

safety

Since no real wet lab experiment is conducted, no safety measures are required in particular.

Experimental Procedure - *This will be shown as a pre-recorded video demonstration only.*

Part A: Flame Tests of Metal Cations

Observe the color of the flame produced with each metal compound shown in the video. Students will record the dominant flame color observed.

Analysis: For each metal cation flame test performed, determine the wavelength corresponding to the observed flame color from the table below.

*Wavelength values here are given for the mid- range of the color indicated.



Dominant Color	Approximate Wavelength (in nm)*
Red	701
Red-Orange	622
Orange	609
Orange-Yellow	597
Yellow	587
Yellow-Green	577
Green	535
Green-Blue	492
Blue	474
Blue-Violet	455
Violet	423

Part A: Flame Tests of Metal Cations Experimental Data and Observations

Solution	Dominant Flame Color	Wavelength (nm)	Flame Intensity (low, medium, high)	
$BaCl_2$				
$CaCl_2$				
$CuCl_2$				
KCl				
LiCl				
NaCl				

Part B: Line Emission Spectra-Research

Do some research to find the line emission spectra of the following elements and copy paste them to your lab sheet and label.

- Copy paste the spectral lines for the indicated elements. Indicate names colors below each line drawn.

Line Emission Spectrum for Helium

- colors:
- nm:

Line Emission Spectrum for Mercury

- colors:
- nm:

Line Emission Spectrum for Hydrogen

- colors:
- nm:

Data Analysis

Using the wavelengths recorded above, calculate the corresponding wavelengths, frequencies and photon energies for each compound tested. Record the values, in scientific notation, in the table below.

Solution	Wavelength (m)	Frequency (s^{-1})	Energy (J)
$BaCl_2$			
$CaCl_2$			
$CuCl_2$			
KCl			
LiCl			
NaCl			

Show a set of sample calculations for **NaCl** only below. Clearly show any equations you have used.

- Wavelength (in m):
- Frequency (in s^{-1}):
- Energy (in J):

Questions

1. Complete the following paragraph by circling the correct responses:

In this experiment, the metal cations in the solutions were initially in the (ground, excited) state. When placed in the flame, the metals then (absorbed, emitted) energy as (electricity, heat, EM radiation). When this occurred, electrons made transitions from (low, high) energy levels to (low, high) energy levels. The metals were then in the (ground, excited) state. The electrons in these metals then made transitions from (low, high) energy levels to (low, high) energy levels, resulting in the (absorption, emission) of energy as (electricity, heat, EM radiation).

2. What evidence is there that the colors observed in the flame tests are due to the metals, and not the nonmetals in the compounds tested?

3. Which metal cation was observed to emit radiation with the *longest wavelength*?

Compared to the other metals studied, did the radiation emitted by this metal cation have

• the highest or lowest frequency?
• the highest or lowest energy?

4. Why do the gas-discharge tubes need to be turned “on” before a color can be seen?

5. You observed the spectral lines for different elements. What is happening within an atom that causes it to emit light in specific lines in a spectrum?

6. Why did the different elements have different line emission spectra?

7. The radio station JACK FM broadcasts at a frequency of 93.5 kHz. Calculate its wavelength (\lamda) in nanometers. (Hz = 1/s)

8. For the same radio station, calculate the energy (E) in kilojoules, kJ.

Contributors and Attributions

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2.10: Lewis Structures and Molecular Shapes (Experiment)-Simulation-Home Version

Objectives

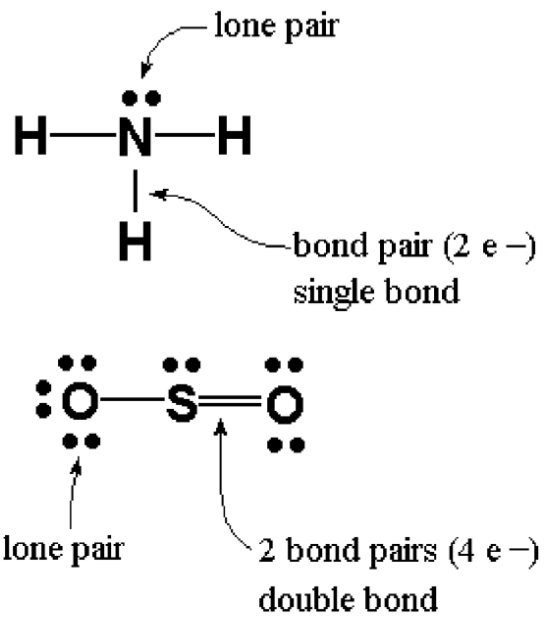
- To use PHET simulation (developed by the University of Colorado, Boulder) to construct these molecules/ions in order to explore their structure and shapes.
- To practice predicting molecular shapes (using VSEPR theory) and molecular polarity

Non-metal atoms bond covalently, resulting in the formation of either neutral molecules or polyatomic ions. A covalent bond is formed when non-metal atoms share their valence electrons, which they do in order to achieve filled valence orbitals like their nearest noble gas neighbor. This means that most bonded non-metal atoms will acquire a total of eight valence electrons via the sharing process – often referred to as the *octet rule*. A notable exception is hydrogen, which only needs to acquire two electrons to be like its nearest noble gas neighbor, helium.

Lewis Structures

A *Lewis Structure* is a representation of covalent molecules (or polyatomic ions) where all the valence electrons are shown distributed about the bonded atoms as either shared electron pairs (bond pairs) or unshared electron pairs (lone pairs). A shared pair of electrons is represented as a short line (a single bond). Sometimes atoms can share two pairs of electrons, represented by two short lines (a *double bond*). Atoms can even share three pairs of electrons, represented by three short lines (a *triple bond*). Pairs of dots are used to represent *lone pair electrons*.

Examples:



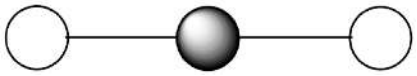
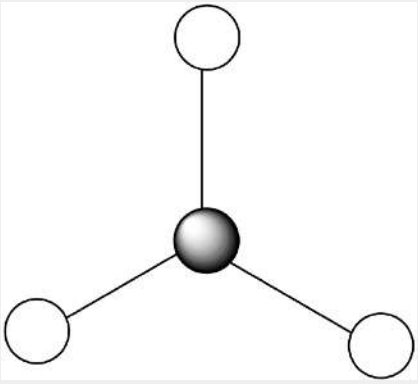
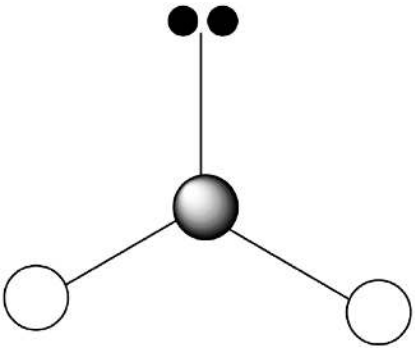
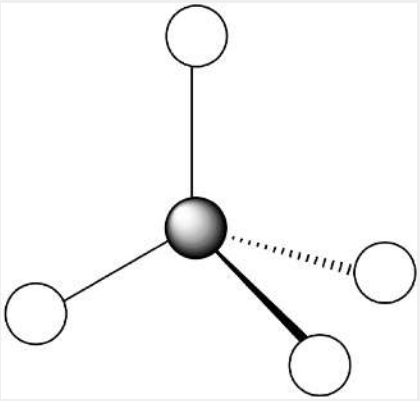
The rules for drawing Lewis structures can be found in the Procedure Section of this handout.

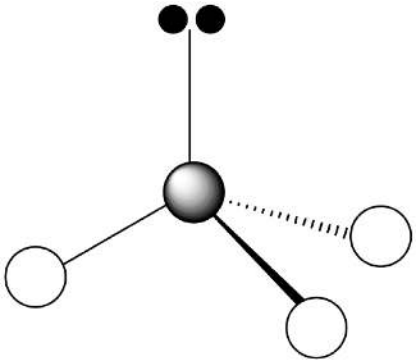
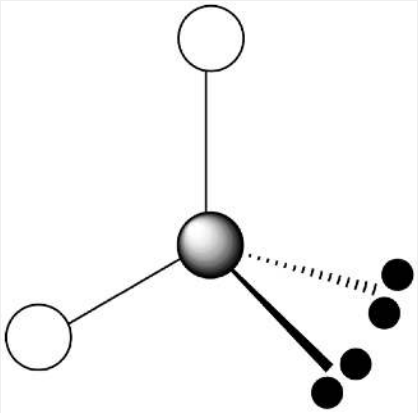
Molecular Shapes

The shape of a molecule depends on the distribution of atoms in space about the central atom, and their bond angles. Bond pair electrons and lone pair electrons repel one another, thus they will be arranged around a central atom as far apart as possible in order to minimize repulsions. This is known as:

Valence **S**hell **E**lectron **P**air **R**epulsion theory, or VSEPR theory.

The following VSEPR table supplies the names, sketches and descriptions of the most common types of molecular shapes that you will encounter. Note that several other molecular geometries do exist, however, they are beyond the scope of this course.

Shape Name	Sketch	Description
Linear		Only two outer atoms surround the central atom. There are no lone pairs on the central atom. Outer atoms are arranged opposite to each other. The bond angles are exactly 180°.
Trigonal Planar		Three outer atoms surround the central atom. There are no lone pairs on the central atom. The central and outer atoms all lie in the same plane (molecule is flat). Bond angles are exactly 120°.
Bent		Two outer atoms and one lone pair surround the central atom. Bond angles are slightly less than 120°.
Tetrahedral		Four outer atoms surround the central atom. There are no lone pairs on the central atom. The four outer atoms are evenly arranged in 3D around the central atom as if at the corners of a regular tetrahedron. The bond angles are exactly 109.5°.

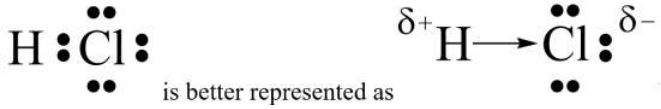
Shape Name	Sketch	Description
Trigonal Pyramidial		Three outer atoms and one lone pair surround the central atom. Here the central atom is located slightly above the three outer atoms, like a tripod. The bond angles are slightly less than 109.5°.
Bent		Two outer atoms and two lone pairs surround the central atom. Bond angles are slightly less than 109.5°.

Electronegativity and Bond Polarity

Some atoms in molecules have the ability to pull shared electrons closer to themselves than other atoms, an ability referred to as *electronegativity*. Electronegativity (\Chi) increases going across a period and decreases going down a group.

If two bonded atoms have different electronegativities, then the bond pair electrons will be *unequally shared*. The atom with the greater electronegativity will pull the bond electrons closer towards itself, causing it to obtain a very slight negative charge (δ^-). The atom with the lower electronegativity will have bond electrons pulled further away from it, causing it to obtain a very slight positive charge (δ^+). The result is a *polar covalent bond.* However, if two bonded atoms have the same electronegativity, then the bond pair electrons will be *equally shared.* The result is a *non-polar covalent bond.*

Example:



The bond electrons are pulled closer to Cl due to its greater electronegativity. Thus, HCl contains a polar covalent bond.

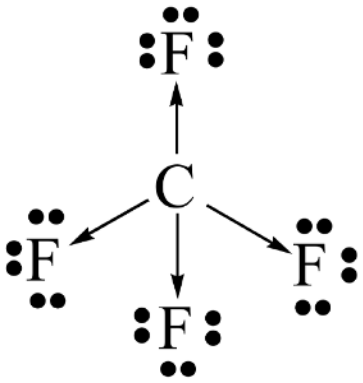
Molecular Polarity

Molecular polarity results when the *entire molecule* (not just a bond in the molecule) ends up with an unequal distribution of electrons. In general, a molecule will be *polar* if it contains polar bonds that are distributed in a *non-symmetrical* arrangement

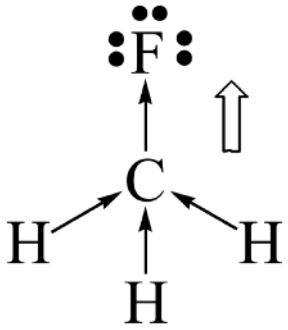
around the central atom. A polar molecule is said to have a net dipole moment. A non-symmetrical arrangement typically results when there are lone pairs on the central atom, or, when different outer atoms surround the central atom.

Not surprisingly, a molecule will be *non-polar* if it contains all non-polar bonds. It will also be non- polar if it contains polar bonds distributed in a *symmetrical* arrangement around the central atom. The symmetry causes the individual bond polarities to cancel out, resulting in a net non-polar molecule. A symmetrical arrangement typically results when there are no lone pairs on the central atom, and if all the outer atoms are identical.

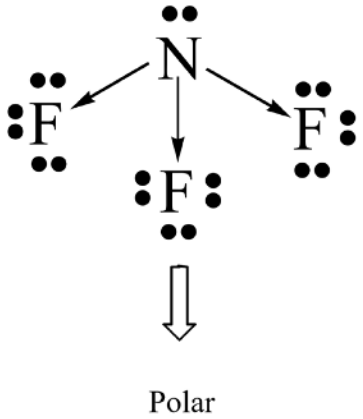
Examples:



Non-Polar



Polar



Polar

Procedure

Drawing Lewis Structures

Draw Lewis structures for each of the molecules and polyatomic ions given on your Report Form. Clearly show all bond pair electrons as lines and lone pair electrons as pairs of dots.

Rules for Drawing Lewis Structures

1. Total the number of valence electrons that each atom contributes to the molecule/polyatomic ion.

- For polyatomic anions, you must add electrons (equal to the negative charge) to the total number of valence electrons. For polyatomic cations, you must subtract electrons (equal to the positive charge) from the total number of valence electrons.

2. Draw a stick structure for the molecule.

- The least electronegative atom is the central atom. Hydrogen is the only exception to this, as it forms only 1 bond. The central atom will usually need to form multiple bonds.
- The other atoms are arranged around the central atom, and are attached to it via single bonds.

3. The octet rule must be obeyed for all elements except hydrogen (follows a “duet” rule).

- Starting with the outside atoms, place the remaining electrons around each atom until it has a total of 8 electrons (except hydrogen – it only requires 2 electrons).
- If there are not enough electrons available to obey the octet rule using single bonds, this indicates that double or triple bonds between two atoms are required in your structure. If short by two electrons, try a double bond, and if short by four electrons, try a triple bond or two double bonds.

Constructing Models, Determining Molecular Shapes and Molecular Polarity using Simulation

1. Click on the following link to access **molecular modelling simulation**
2. PhET: Molecule Shapes

13. Press the “play” arrow.

2. Click on **model**
3. Click on bond angle, molecular geometry, and electron geometry boxes
4. You can bring is as many bonds as you want, single, double, or triple
5. You can also add as lone pairs on the central atom by clicking the lone pairs
6. You can click the red cross on the right side to “remove” the added features
7. Just play around with the features before working with molecules on the lab sheet.
8. You can write down the bond angle, electron geometry, and molecular geometry from the simulation. Also, you can check and improve the molecular structure that you have already drawn on the lab sheet.

   The Colorado State University provides the link. It may be sensitive to browsers. It works best on explorer window. It is difficult open it on Mac. However, it works well on cell phones.

Rules for Constructing Molecules with the Simulation

- Use the *first one bond* for *single bonds.*
- Use *two bonds* for a *double bond* (second from the top right) and for a *triple bond.*
- *Use the balloon kind of lone pair of each pair of lone pairs on the central atom*
- *You can click on the red cross on them to remove the bonds or the lone pairs.*

2. Evaluate the polarity of the bonds in each molecule as well as its overall symmetry in order to determine whether it is polar or non-polar.

Lab Report: Lewis Structures and Molecular Shapes

1. CH_4

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

2. CO_2

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

3. NH_4

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

4. H_2O

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

5. N_2

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

6. SO_2

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

7. O_2

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

8. O_3 - **use yellow ball for central atom**

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

9. CO

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

10. CO_3^{-2}

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

11. **NO_3^- – use a blue ball for phosphorus when constructing this model**

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

12. CF_2Cl_2 (CFC = chlorofuorocarbon)

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

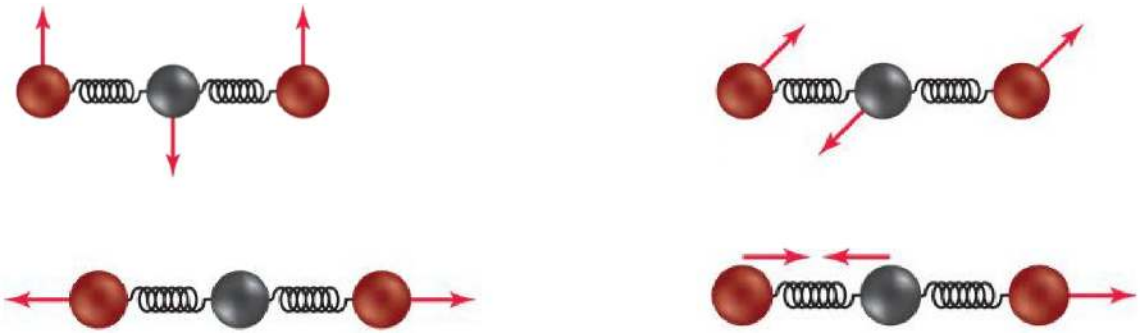
Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

Questions

1. When molecules and energy interact, there are different results. Atoms may dissociate, a molecule may rotate, or bonds may stretch and bend. Carbon dioxide has four possible vibrations. What is the effect of each vibration on the molecule: stretching, bending, or breaking?



2. What is the definition of a greenhouse gas?
3. Various atmospheric gases are listed below:

	Chemical Formula	VSEPR Shape (Sketch)	Greenhouse Gas (yes or no)
carbon dioxide			
methane			
nitrogen			
oxygen			
ozone			
sulfur dioxide			
water			

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2.3: Paper Chromatography of Washable-marker Ink Pens (Experiment)-Home Version

polyatomic ions

2.4: Chemical Nomenclature (Experiment)

Glossary

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