INTRODUCTORY CHEMISTRY AT HOME LABS



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Introductory Chemistry At Home Labs

Georgia Fritz

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This text was compiled on 11/03/2025

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Glossary

Detailed Licensing

Detailed Licensing

Licensing

*A detailed breakdown of this resource's licensing can be found in **Back Matter/Detailed Licensing**.*

Disclaimer

The experiments in this laboratory manual are designed to be conducted safely at home. However, this assumes that the user of the manual begins at Experiment 1 and learns about laboratory safety. Additionally, the user agrees to follow the safety protocols given with the experiments. Ultimately, the user must maintain a safe work environment. The author cannot control the each user's environment. By conducting the experiments, the user assumes all risks.

CHAPTER OVERVIEW

1: Safety First

Laboratory Safety At Home

This lab procedure is designed to help prepare you to conduct chemistry experiments at home. You will be watching 6 videos from the American Chemical Society about how to conduct lab experiments at home. Please note, that if at any time the links to the videos may not be working on the Procedure page, a full URL can be found in on the reference page for this chapter.

After viewing the videos, there are several laboratory questions for you to answer.

1.1: Background

1.2: Laboratory Safety Videos

1.3: Lab Questions

1.4: References

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1.1: Background

Whenever you work with chemicals, it is important to consider the risks and hazards associated with the project. This is true whether you are in a laboratory or working with chemicals at home. As this is a fully online course, you will be conducting experiments in your home environment. This will involve additional consideration as compared to being in a lab. Do you have small children or pets in your home environment? How will you clean your workspace before and after conducting an experiment? This course will use mild chemicals that you may already use around your house. But the principles apply anytime you are using chemicals at home – whether cleaning, cooking, gardening, or doing home repairs.

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1.2: Laboratory Safety Videos

Task

To better understand how to use chemicals safely, watch the following videos from the American Chemical Society (ACS). Take notes as you watch to help you answer the laboratory questions.

Laboratory Safety Overview

The ACS College Safety Video 1 provides a basic overview of what is meant by laboratory safety. It also introduces the acronym RAMP. Be sure to take notes on what this acronym means. Additionally, think about what it means to have a culture of safety, as you will be creating that culture in your home environment.



Assessing Chemical Risk

The ACS College Safety Video 2 provides information about assessing chemical risk. Take note of the places you can find information about chemicals, as well as what the hazards mean. Whenever you work with a new chemical, the best practice is to look up safety information to make sure you understand the hazards associated with that chemical. This is true whether you are in a laboratory or working around the house. Common household items such as bleach can be deadly if used inappropriately.



Assessing Procedural Risk

The ACS College Safety Video 3 provides information about assessing risk while conducting an experiment. Be sure to take note of the difference between a hazard and a risk.



Preparing for Lab

The ACS College Safety Video 4 provided information about to prepare for conducting experiments. Take note of the ways chemicals can enter your body. As you go about doing experiments at home, be sure to keep in mind that you do not want to eat the experiments!



Minimizing Risks

The ACS College Safety Video 5 provides information about minimizing risk in the laboratory. At home, you will not be able to share chemicals with other people. The waste from this course can be disposed of either down a drain (liquids), in a trash can (solids), or composted (plant materials).



Preparing for Emergencies

The ACS College Safety Video 6 provides information about preparing for emergencies when conducting an experiment. At home, you will not have an eyewash station or emergency shower. Eyes can be flushed in the sink if needed. If you have a shower, this can be used to rinse spills off your skin. Also pay attention to ways to put out fires. The fire triangle applies to all fires, including those that happen in your kitchen. How might the information in this video apply in the home environment?



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1.3: Lab Questions

Please complete the following questions in complete sentences:

1. What is meant by the acronym RAMP?
2. What type of clothing is needed to be safe when conducting laboratory experiments? In other words, what should you be wearing ***prior*** to putting on personal protective equipment?
3. How should long hair be worn when conducting experiments? Why?
4. What personal protective equipment is included in your laboratory kit?
5. What is one example of laboratory safety equipment?
6. When should you read the laboratory experiment?
7. What are the three parts of the fire triangle?
8. What are three sources of information about chemical hazards?
9. As this class will not meet in person, what are three things you need to do in your own space to prepare for conducting experiments?

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1.4: References

(1) American Chemical Society. From Chemical Safety Rules to Risk Management | ACS College Safety Video #1. *YouTube*, 2022. https://www.youtube.com/watch?v=DWSymRPCDN4. (Accessed 2024-02-06)

(2) American Chemical Society. Chemical Safety Information Resources | ACS College Safety Video #2. *YouTube*, 2022. https://www.youtube.com/watch?v=iQDbrR04tbU. (Accessed 2024-02-06)

(3) American Chemical Society. Assessing Risks in the Chemistry Laboratory | ACS College Safety Video #3. *YouTube*, 2022. https://www.youtube.com/watch?v=PhwyPQ0_5K0. (Accessed 2024-02-06)

(4) American Chemical Society. Minimizing Risks in the Chemistry Laboratory | ACS College Safety Video #4. *YouTube*, 2022. https://www.youtube.com/watch?v=DUm4hFUXEvw. (Accessed 2024-02-06)

(5) American Chemical Society. Minimizing Risks in the Chemistry Laboratory: Techniques | ACS College Safety Video #5. *YouTube*, 2022. https://www.youtube.com/watch?v=NHBRlHzifMU. (Accessed 2024-02-06)

(6) American Chemical Society. Emergencies in the Chemistry Laboratory | ACS College Safety Video #6. *YouTube*, 2022. https://www.youtube.com/watch?v=rPNQH3usdzI. (Accessed 2024-02-06)

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CHAPTER OVERVIEW

2: Measurements Lab

This experiment will introduce you to using a few laboratory tools and reporting data appropriately.

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2.1: Required Supplies

You will need the following supplies to complete this experiment:

- Balance
- Ruler, marked to millimeter markings
- 10-mL graduated cylinder
- 50-mL graduated cylinder
- Beaker OR Cup
- Water
- Nickel (American)
- 10 Pennies (American), minted after 1983

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2.2: Background - Significant Figures in Measurements

Discrete Objects vs Measured Numbers

When we count objects, we can count exact numbers with precision. For example, in this experiment, you will use 10 pennies. You will be able to count out 10 pennies. This is an exact number. In contrast, you will also make measurements using a graduated cylinder, a ruler, and an electronic balance. In all these measurements, an estimate will be made for the last digit.

Significant Figures in Measurement

Measuring Volume

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

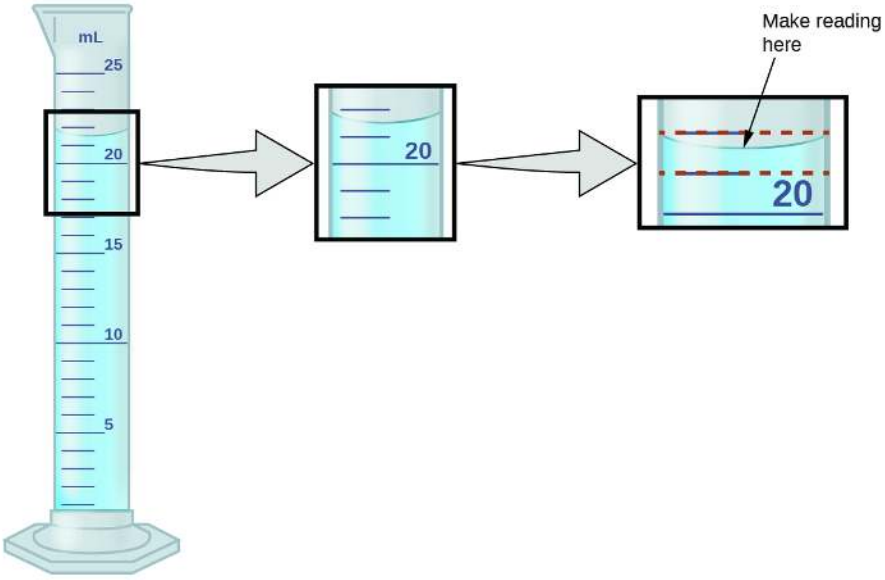


Figure **2.2.1**: To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus.

Refer to the illustration in Figure **2.2.1**. The bottom of the meniscus in this case clearly lies between the 21 and 22 markings, meaning the liquid volume is *certainly* greater than 21 mL but less than 22 mL. The meniscus appears to be a bit closer to the 22-mL mark than to the 21-mL mark, and so a reasonable estimate of the liquid's volume would be 21.6 mL. In the number 21.6, then, the digits 2 and 1 are certain, but the 6 is an estimate. Some people might estimate the meniscus position to be equally distant from each of the markings and estimate the tenth-place digit as 5, while others may think it to be even closer to the 22-mL mark and estimate this digit to be 7. Note that it would be pointless to attempt to estimate a digit for the hundredths place, given that the tenths-place digit is uncertain. In general, numerical scales such as the one on this graduated cylinder will permit measurements to one-tenth of the smallest scale division. The scale in this case has 1-mL divisions, and so volumes may be measured to the nearest 0.1 mL.

Measuring Mass

This concept holds true for all measurements, even if you do not actively make an estimate. If you place a quarter on a standard electronic balance, you may obtain a reading of 6.72 g. The digits 6 and 7 are certain, and the 2 indicates that the mass of the quarter is likely between 6.71 and 6.73 grams. The quarter weighs *about* 6.72 grams, with a nominal uncertainty in the measurement of ± 0.01 gram. If the coin is weighed on a more sensitive balance, the mass might be 6.723 g. This means its mass lies between 6.722 and 6.724 grams, an uncertainty of 0.001 gram. Every measurement has some **uncertainty**, which depends on the device used (and the user's ability). All of the digits in a measurement, including the uncertain last digit, are called **significant**

figures or **significant digits**. Note that zero may be a measured value; for example, if you stand on a scale that shows weight to the nearest pound and it shows “120,” then the 1 (hundreds), 2 (tens) and 0 (ones) are all significant (measured) values.

Measuring Width

Consider using a ruler to measure the width of an object, as shown in Figure **2.2.2**. The object is definitely more than 1 cm long, so we know that the first digit in our measurement is 1. We see by counting the tick marks on the ruler that the object is at least three ticks after the 1. If each tick represents 0.1 cm, then we know the object is at least 1.3 cm wide. But our ruler does not have any more ticks between the 0.3 and the 0.4 marks, so we can't know exactly how much the next decimal place is. But with a practiced eye we can estimate it. Let us estimate it as about six-tenths of the way between the third and fourth tick marks, which estimates our hundredths place as 6, so we identify a measurement of 1.36 cm for the width of the object.

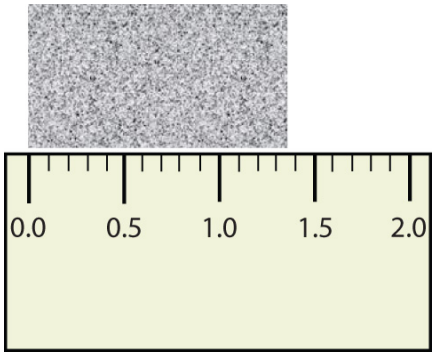


Figure **2.2.2**: Expressing Width. What is the proper way to express the width of this object?

Does it make any sense to try to report a thousandths place for the measurement? No, it doesn't; we are not exactly sure of the hundredths place (after all, it was an estimate only), so it would be fruitless to estimate a thousandths place. Our best measurement, then, stops at the hundredths place, and we report 1.36 cm as proper measurement.

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- **2.4: Significant Figures** by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/beginning-chemistry.

2.3: Background - Counting Significant Figures and Using Scientific Notation

Determine How Many Significant Figures

When we have measurements, we need to decide if the number is exact or involves an estimate. With exact numbers, such as counting an exact number of pennies, the number is considered infinitely significant. In contrast, measured values have the last digit estimated due to the limitation of the tool used. Therefore, we can count both how many decimal places and how many digits each number has. The volume of water estimated above was 22.6 mL. This number has three digits that are significant and one decimal place. In contrast, the number 6.723 g has four significant digits and three decimal places. All non-zero numbers are significant. Zeros are tricky because they are sometime significant, sometimes not and sometimes it is hard to determine. We can use the following rules to help us determine when a zero will be significant.

Rules to Determine if a Number is Significant

- All nonzero numbers are significant.
- A zero that leads a small number is between 0 and 1 in value is **not** significant.
  - 0.00345 has zeros that are not significant. They are just place holders. This number has 3 significant digits.
- A zero after decimal point after a nonzero digit **is** significant.
  - 0.00345**0** has 4 significant digits. The zero to the right of the 5 is significant.
- A zero between two nonzero numbers **is** significant.
  - 4**0**6**0**5 has 5 significant digits. The zeros are sandwiched between nonzero numbers.
- A zero at the end of a number without a decimal point is ambiguous.
  - 140 could either have 2 or 3 significant figures. We don't know.
  - Use scientific notation!

Scientific Notation

To avoid ambiguity, scientific notation is used to communicate how many significant digits are in a number. The format for scientific notation is always a nonzero digit, a decimal place if needed and digits to the right of the decimal place if needed multiplied by 10 to a power.

For example, we saw above in the electronic balance section that a weight of 120 pounds was measured, with the ones place being the last measured number. However, 120 pounds has an ambiguous number of significant figures. In scientific notation, 120 is written $1.20*10^2$. Now it is easier to see that the zero is significant. It appears to the right of the decimal point after a nonzero number. We do not count the 10^2 when determining significant figures as it is an exact number.

We could also have a measurement that estimated 120 to the tens place. Thus, only the 1 and the 2 are significant. We can communicate this by writing $1.2*10^2$. Now the number clearly has only 2 significant figures.

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2.4: Background - Significant Figures in Calculations

Once we have significant figures from measurements, we need to keep track of them when those numbers are used in further calculations. There are many rules for different situations, but for this lab, we will focus on multiplication/division and addition/subtraction.

- When multiplying and dividing, report the least number of significant figures.
  - Consider the equation: $1.42 * 2.0 = 2.84$
    - 1.42 has three significant figures. 2.0 has two significant figures. We need to round 2.84 to 2.8.
  - Thus, $1.42 * 2.0 = 2.8$ when reporting the correct number of significant figures.
- When adding and subtracting, count decimal places and report to the least number of decimal places.
  - For example: $1002.345 + 1.2796 = 1003.6246$
    - 1002.345 has three decimal places. 1.2796 has four decimal places. We need to round to three decimal places.
  - Thus, $1002.345 + 1.2796 = 1003.625$ when accounting for significant figures.
- When rounding, look at the number to the right of the last significant digit or decimal place:
  - if 0 to 4, just drop the digits not needed.
    - 2.84 was rounded to 2.8.
  - If 5 to 9, round the last significant digit up by one.
    - 1003.6246 was rounded to 1003.625.

Keeping track of significant figures in calculations communicates how certain we can be about a number based on the measured data used.

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2.5: Background - Percent Error, Reporting Data, and Density

Percent Error

Percent error is a measurement of how accurate or how close to a known value a given measurement it. In this experiment, you will be measuring both the diameter and the mass of both a penny and a nickel. The known values for these coins are published by the U.S. Mint, so we can determine how accurate our measurements are. The equation used to determine percent error is:

$$\text{Percent Error} = \left|\frac{\text{actual} - \text{experimental}}{\text{actual}}\right| \times 100 \tag{2.5.1}$$

Reporting Data

When data is used in calculations, it is important to include the units. Units can act as a guidepost to help determine if a calculation has been done correctly. For example, in this experiment, density is calculated. Density has units of g/mL. If the calculation done results in a different set of units, go back and check your work!

When reporting data in a data table, the goal is to easily see the data. Thus, units are included in the column heading on a data table, rather than in each cell. Tables are the one place units will not be right next to the numbers.

Density

Density is a measure of how much matter (mass) is in a given volume. Density is calculated:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \tag{2.5.2}$$

There are many possible units for reporting density. In this experiment, as the mass will be in grams and the volume is measured in milliliters, density will be reported in units of g/mL. You will be asked to think about density more in your lab questions after you have collected your data.

The volume of some objects are easy to measure. Liquids or finely ground solids can be poured into a measuring device. Cubes, spheres, and cylinders can have their density determined by uses geometric formulas. For example, the volume of a cylinder can be determined the following equation:

$$\text{Volume} = \pi r^2 h \tag{2.5.3}$$

$$r = \text{radius of circle} = \frac{1}{2} \times \text{diameter} \tag{2.5.4}$$

$$h = \text{height of the cylinder} \tag{2.5.5}$$

The volume of other objects is harder to determine. Solids with irregular shapes do not have easy formulas to determine their volume. One way that the density of irregular solids can be measured is to see how much water the object displaces. The volume of water displaced is calculated using the following equation:

$$\begin{gathered}\text{Volume of water displaced} \\ = (\text{volume of water and object}) - (\text{initial volume of water})\end{gathered} \tag{2.5.6}$$

In this experiment, you will determine the volume of coins using two methods. First, you can approximate the shape of a coin as a cylinder. However, this approximation does not factor in the fact that the top and bottom of the coin are not flat. You will also determine the volume of the coin by seeing how much water is displaced by the coin.

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2.6: Procedure

Table 1 Data Collection – Mass of Coins

1. Begin by turning on and taring your balance so it reads 0.00 g.
2. Measure the mass of the penny. Record this mass in data Table 2.7.1.
3. Measure the mass of the nickel. Record this mass in data Table 2.7.1.
4. With your ruler, measure the diameter of the penny in centimeters. Record this in data table 2.7.1 with appropriate significant figures.
5. With your ruler, measure the diameter of the nickel in centimeters. Record this in data table 2.7.1 with appropriate significant figures.
6. Calculate the Percent Errors for all measurements in data table 2.7.1. Show your calculations in the calculation section.

Table 2 Data Collection – Volume of a Penny and Nickel, Determined Two Ways

1. Measure the height of the penny in centimeters. Record this data in Table 2.7.2.
2. Measure the height of the nickel in centimeters. Record this data in Table 2.
3. Calculate the percent error in the coin height for both the penny and the nickel. Record these numbers in Table 2.
4. Calculate the volume for each coin by assuming each is a cylinder. Use the equation for a cylinder given in the background information. Remember the diameter of the coins is recorded in Table 1. Show your calculations in the calculations section.
5. Fill your 50 mL graduated cylinder with about 10 mL of water. Record the volume in Table 2 in the “Penny” row.
6. Add the penny to the graduated cylinder. Record the volume of the penny with the water in Table 2.
7. Empty the graduated cylinder to remove the penny.
8. Refill with about 10 mL of water. Record the volume in the graduated cylinder in Table 2 in the “Nickel” row.
9. Add the nickel to the graduated cylinder. Record the volume of the water plus the nickel in Table 2.
10. Calculate the volume of each coin. Show your calculations in the calculations section.

Table 3 Data Collection – Mass, Volume, and Density of Water

1. Begin by turning on and taring your balance so it reads 0.00 g.
2. Put some water in your beaker or cup (about 20 mL).
3. Measure and record the mass of the 10-mL graduated cylinder.
4. Add 6 mL of water to your 10-mL graduated cylinder. Record the exact volume in Table 3.
5. Measure the mass of the 10-mL graduated cylinder with the water. Record this mass in Table 3.
6. Calculate the mass of the water by subtracting the mass of the empty graduated cylinder from the mass of the graduated cylinder with water. Show your calculation in the calculations section and record the mass of water in Table 3.
7. Calculate the density of the water based on the data from the 10-mL graduated cylinder. Show your calculation in the calculations section.
8. Measure and record the mass of the 50-mL graduated cylinder.
9. Add 6 mL of water to your 50-mL graduated cylinder. Record the exact volume in Table 3.
10. Measure the mass of the 50-mL graduated cylinder with the water. Record this mass in Table 3.
11. Calculate the mass of the water by subtracting the mass of the empty graduated cylinder from the mass of the graduated cylinder with water. Show your calculation in the calculations section and record the mass of water in Table 3.
12. Calculate the density of the water based on the data from the 50-mL graduated cylinder. Show your calculation in the calculations section.
13. Determine the percent error in the measurement of the density of water.

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2.7: Data

Table 2.7.1 – Mass and Diameter of Coins

Table for reporting mass and diameter of coins

Coin	Measured Mass (g)	Known Mass (g)	Mass – Percent Error (%)	Measured Diameter (cm)	Known Diameter (cm)	Diameter – Percent Error (%)
Penny		**2.500**			**1.905**	
Nickel		**5.000**			**2.121**	

Table 2.7.2 – Height and Volume of a Penny and Nickel

Table for reporting coin height and volume data

Coin	Measured Height (cm)	Known Height (cm)	Height – Percent Error (%)	Calculated Volume (cm^3)	Initial Water Volume (mL)	Water Volume with Coin (mL)	Coin Volume (mL)
Penny		**0.152**					
Nickel		**0.195**					

Table 2.7.3 – Mass, Volume, and Density of Water

Table for reporting mass, volume, and density of water

Graduated Cylinder Volume (mL)	Mass of Empty Cylinder (g)	Mass of Cylinder with Water (g)	Mass of Water (g)	Volume of Water (mL)	Measured Density of Water (g/mL)	Known Density of Water (g/mL)	Density – Percent Error (%)
10						**0.99802**	
50						**0.99802**	

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2.8: Calculations

Show your calculations (including units!) for the following:

- Percent Error of Penny Mass
- Percent Error of Penny Diameter
- Percent Error of Nickel Mass
- Percent Error of Nickel Diameter
- Percent Error Penny Height
- Percent Error Nickel Height
- Volume of Penny Cylinder Approximation
- Volume of Nickel Cylinder Approximation
- Volume of Water Displaced by Penny
- Volume of Water Displaced by Nickel
- Mass of Water in 10-mL Graduated Cylinder
- Density of Water in 10-mL Graduated Cylinder
- Percent Error in Density of Water in 10-mL Graduated Cylinder
- Mass of Water in 50-mL Graduated Cylinder
- Density of Water in 50-mL Graduated Cylinder
- Percent Error in Density of Water in 50-mL Graduated Cylinder

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2.9: Lab Questions

1. Compare the percent error in the mass of the penny to the percent error to the mass of the nickel. Which coin has a higher percent error?
2. Compare the percent error in the diameter of the two coins. Which coin has a higher percent error?
3. Are there any observations you can make about percent error? Is it easier to measure a larger item or a smaller item with more accuracy?
4. Which graduated cylinder gave the more precise measurement of water? Why?
5. Compare the density of water determined in this experiment. Are they the same or different? Why do you think you obtained the results you did?
6. The density of liquids depends on the temperature of the liquid. The density of water reported as the known value was reported at 70°F. What is this temperature in Celsius? Be sure to show your calculation in your work. Use the following equation: $°C = \frac{5}{9}(°F - 32)$. Report your answer to two significant figures.

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2.10: References

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CHAPTER OVERVIEW

3: Chemical Nomenclature Practice

3.1: Background - Naming Ionic Compounds

3.2: Background - Naming Bimolecular Covalent Compounds

3.3: Activity 1 – Naming and Structure Determination Based on Compound Type

3.4: Activity 2 - Expanded Polyatomic Ions

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3.1: Background - Naming Ionic Compounds

Ionic compounds contain cation and anions. At the simplest level, to name an ionic compound is simply to name its cation and to name its anion, in that order. Cations have a positive charge. Anions have negative charge. The charges of ionic compounds balance to give a neural charge. We can determine the cation-to-anion ratio simply by balancing the charges to determine the structure. Typically, if a compound is a mixture of a metal and a nonmetal, we will use the naming convention for ionic compounds.

Naming Cations

For single atom cations in groups 1A, 2A, and 3A, the name is the same as the atomic name. The charge is the same as the group number. Examples are shown in Table 1.

Table **3.1.1** – Examples of Single Atom Cations with Consistent Charge

Element	Cation	Ion Name
Sodium	Na^{+}	Sodium
Magnesium	Mg^{2+}	Magnesium
Aluminum	Al^{3+}	Aluminum

For single atom cations that are transition metals, the charge can vary. Thus, the name of the cation also needs to include the charge. This is shown as a parathesis after the name with a Roman numeral that indicates the charge. These ions also have had common names in the past. These are listed in the table as they are still in use in some places, but you do not need to memorize them! They are there for your reference.

Table **3.1.2** – Examples of Single Atom Cations with Variable Charge

Element	Cation	Ion Name	Common Name
Iron	Fe^{2+}	Iron(II)	Ferrous
Iron	Fe^{3+}	Iron(III)	Ferric
Copper	Cu^{+}	Copper(I)	Cuprous
Copper	Cu^{2+}	Copper(II)	Cupric
Manganese	Mn^{2+}	Manganese(II)	N/A
Manganese	Mn^{4+}	Manganese(IV)	N/A
Manganese	Mn^{7+}	Manganese(VII)	N/A

A polyatomic ion contains a group of atoms that are covalently bonded together. The charge is for the overall group of atoms. We discuss one polyatomic cation, ammonium. The ammonium ion is NH_4^{+}. This is the one exception we will see to ionic compounds being a mixture of nonmetals and metals.

Naming Anions

We will discuss two types of anions: single atom ions and polyatomic ions. Single atom ions are typically nonmetals. The charge of these single atom ions is determined by the number of electrons an atom needs to gain to become a stable anion. For hydrogen, this is one electron. For group 5A elements, this is 3 electrons. For group 6A elements, this is 2 electrons. For group 7A elements, this is one electron. To name a single atom anion, first list the prefix and “-ide.” The prefix is determined by dropping letters from the end of the chemical name. Table 3 demonstrates how to name single atom anions.

Table **3.1.3** – Single Atom Anions

Element	Anion	Ion Name
Hydrogen	H^-	Hydride
Nitrogen	N^{3-}	Nitride
Phosphorus	P^{3-}	Phosphide
Oxygen	O^{2-}	Oxide
Sulfur	S^{2-}	Sulfide
Fluorine	F^-	Fluoride
Chlorine	Cl^-	Chloride
Bromine	Br^-	Bromide
Iodine	I^-	Iodide

Naming Polyatomic Ions

Polyatomic ions are groups of atoms that are covalently bonded together and have an overall charge on the structure. The names for polyatomic ions are more variable, but a few trends exist. For oxygen containing polyatomic ions, for ions with the same central atom (non-oxygen atom), the ion with more oxygen atoms in the structure will end in "-ate" while the one with fewer will end in "-ite." If even more oxygen atoms can be added, it will have a prefix of "per-" added to the "-ate" ion. If even fewer oxygen atoms are present than in the "-ite" ion, the prefix of "hypo-" is added. Hypochlorite (ClO^-), chlorite (ClO_2^-), chlorate (ClO_3^-), and perchlorate (ClO_4^-), are examples of this style of naming polyatomic ions.

Another common change to polyatomic ions is to add a hydrogen cation and the charge becomes more positive by one. This change can be indicated by adding the word "hydrogen" in front of the polyatomic ion name or by adding the prefix "bi-" to the polyatomic ion name. For example, carbonate has a formula of CO_3^{2-} while bicarbonate has a formula of HCO_3^-. Bicarbonate is also called hydrogen carbonate.

Be very careful with spelling! For example, nitri**d**e (N^{3-}) is chemically very different from nitri**t**e (NO_2^-), although the spelling is similar.

Table **3.1.4** – Common Polyatomic Ions

Polyatomic Ion	Ion Name
NH_4^+	Ammonium
ClO^-	Hypochlorite
ClO_2^-	Chlorite
ClO_3^-	Chlorate
ClO_4^-	Perchlorate
NO_2^-	Nitrite
NO_3^-	Nitrate
CO_3^{2-}	Carbonate
HCO_3^-	Bicarbonate (Hydrogen Carbonate)
PO_4^{3-}	Phosphate
OH^-	Hydroxide
O_2^{2-}	Peroxide
CH_3COO^-, $C_2H_3O_2^-$	Acetate
SO_4^{2-}	Sulfate
CN^-	Cyanide

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3.2: Background - Naming Bimolecular Covalent Compounds

Bimolecular covalent compounds contain two different nonmetals. As the electrons are shared, the atoms do not have a charge. Thus, the names must indicate the number of each atom type. The less electronegative element is typically listed first in the name, followed by the more electronegative element. Fluorine is the most electronegative element. The closer on the periodic table an element is to fluorine (excluding the noble gases, family 8A), the more electronegative the element is.

The same elements can bond in a variety of combinations. For example, NO_2 and N_2O are both bimolecular compounds that contain nitrogen and oxygen. NO_2 is called nitrogen dioxide. N_2O is named dinitrogen monoxide. As can be seen from this example, the first element retains its name. If only one atom of the first type is present, it does not have a prefix. If more than one of the first element is present, it does get a prefix (see Table 5). The second element is named with a prefix plus the monoatomic anion name.

Many molecules exist that contain just hydrogen and carbon atoms. The names of these molecules do not follow the rules for bimolecular covalent compounds, but instead have a separate set of rules. For example, CH_4 is called methane. As you can see, methane is a very different name than carbon tetrahydride!

Table **3.2.1** – Prefixes for Bimolecular Covalent Compounds

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

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3.3: Activity 1 – Naming and Structure Determination Based on Compound Type

Instructions

Using the background information for both ionic compounds and covalent compounds, complete the missing space on the following table. You should not need to look up names elsewhere online. Also, names online are not always systematic. Some molecules are known by common names that are not systematic names.

Table **3.3.1** – Determining the Formula, Compound Type, and Systematic Name

Chemical Formula	Compound Type	Systematic Name
MgO	Ionic	
		Sodium Hydroxide
PCl_5	Covalent	
		Potassium Perchlorate
CuI		
NH_4NO_3		
		Manganese(II) Acetate
$NaHCO_3$		
		Iron(III) Bromide
		Potassium Phosphate
		Silicon Disulfide
N_2O_5		

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3.4: Activity 2 - Expanded Polyatomic Ions

Activity 2 - Expanded Polyatomic Ions

Instructions

You have been given the names and formulas for several common polyatomic ions in the background information. Based on this information, you can determine either the name or the formula of the following polyatomic ions.

Table **3.4.1** – Determining Polyatomic Names and Structures

Polyatomic Ion	Name
BrO^-	
	bromite
BrO_3^-	
	perbromate
	iodate
IO_2^-	
	hypoiodite
IO_4^-	
PO_3^{3-}	
	dihydrogen phosphate

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CHAPTER OVERVIEW

4: Limiting Reactant

4.1: Safety
4.2: Lab Supplies
4.3: Background
4.4: Procedure
4.5: Data and Results
4.6: Calculations
4.7: Laboratory Questions

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4.1: Safety

This reaction generates a gas. Wear safety goggles to protect your eyes. A lab apron will protect your clothes from splashes. Use at least a 9-inch balloon when conducting this experiment. Do not increase the amount of vinegar beyond 30 mL.

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4.2: Lab Supplies

- Baking soda (Sodium bicarbonate)
- Vinegar (5% Acetic Acid)
- 3x 9-inch Balloons (12-inch would also be acceptable)
- 3x Plastic bottles (16 to 20 oz) with a narrow neck that a balloon can fit on.
- Funnel/Wax paper squares

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4.3: Background

What is a limiting reactant?

Have you ever played the game "Musical Chairs?" In the game, there are chairs set up in a circle, with one fewer chairs than people. Music is played and everyone walks in a circle around the chairs. When the music stops, you sit down in a chair. The person without a chair is out of the game. In this game, the chairs limit how many people can sit down.

In chemical reactions, one chemical may limit how far the reaction can occur. If the chemical is used up, then the reaction cannot keep occurring. In this experiment, you will react baking soda (sodium bicarbonate) with acetic acid (vinegar), as shown in equation **4.3.1**. When bicarbonate reacts with an acid, it forms carbonic acid, which breaks down into water and carbon dioxide. Carbon dioxide is a gas. It will be caught in a balloon so you can determine how much gas was generated in the reaction.

$$\mathrm{NaHCO_3}(s) + \mathrm{CH_3COOH}(aq) \rightarrow \mathrm{NaCH_3COO}(aq) + \mathrm{H_2O}(l) + \mathrm{CO_2}(g) \tag{4.3.1}$$

How to calculate the volume of a sphere

To determine the amount of gas generated, you will estimate the volume of balloon to be a sphere. You will measure the circumference of the balloon at its widest point. Recall that the circumference is the distance around the surface of the balloon. From this, you can determine the radius and the volume of the balloon.

The radius (r) of the balloon is related to the circumference (C) by the following equation:

$$C = 2\pi r \tag{4.3.2}$$

Once you know the radius (r), you can solve for the volume (V) with the following equation:

$$V = \frac{4}{3}\pi r^3 \tag{4.3.3}$$

If your calculator has a π button, you may use this. If not, estimate π as 3.14159.

Molarity – How to Find Moles in a Solution

We have seen in this class that the mass of a chemical can be related to the moles of the chemical using the molar mass of the compound. But what if we have a solution, which has a chemical dissolved in another chemical? In this lab, we will use vinegar, which is a solution of acetic acid dissolved in water. We cannot easily obtain the mass of acetic acid. Instead, we measure the concentration of the solution – how many moles of a chemical, in this case, acetic acid, are present in one liter of solution. Thus, the units for molarity are mol/L. Molarity is also represented by a capital "M" as the unit. Regular white vinegar is about 0.87 M. This means that there are 0.87 moles of acetic acid in 1 L of solution.

Remember when we use a molar mass, it has a unit of g/mol. This compound unit means it works as a conversion factor. Likewise, molarity with its units of mol/L is also a conversion factor. However, as solution concentration is a number which is frequently used in chemistry, the units for molarity are typically shortened to M. Remember:

$$M = \frac{\text{moles of solute (chemical dissolved)}}{\text{L of solution}} \tag{4.3.4}$$

Therefore, we could, for example, determine how many moles of acetic acid are present in 50.0 mL sample of vinegar.

$$50.0\ \text{mL vinegar} \times \frac{0.87\ \text{moles acetic acid}}{1\ \text{L vinegar}} \times \frac{1\ \text{L}}{1000\ \text{mL}} = 0.0435\ \text{mol acetic acid} \tag{4.3.5}$$

Rounding to the correct number of significant figures, **0.044 mol acetic acid are present in 50.0 mL of vinegar**. 0.87 M is not an exact number, so it has the least number of significant digits of the non-exact numbers in the equation. When multiplying and dividing, we need to count digits to determine significant figures.

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4.4: Procedure

1. Stretch three balloons a little bit so they are not stiff.
2. Label the balloons A, B, and C.
3. Tare Balloon A on the balance. Measure 1.7 g of sodium bicarbonate (baking soda) into the balloon. Record the exact mass in the data table.
4. Tare Balloon B on the balance. Measure 2.2 g of sodium bicarbonate (baking soda) into the balloon. Record the exact mass in the data table.
5. Tare Balloon C on the balance. Measure 2.7 g of sodium bicarbonate (baking soda) into the balloon. Record the exact mass in the data table.
6. Label the three bottles A, B, and C. Use the balloon with the bottle that has the same letter.
7. Measure 30 mL of vinegar into each bottle. Record the exact volume for each bottle into the data table.
8. Calculate the moles of sodium bicarbonate in each balloon. Show your calculations with units in the calculations section. Record the moles in the data table.
9. Calculate the moles of acetic acid in each bottle using 0.87 mol/L as the molarity of the vinegar solution. Report this in the data table. Show your calculations with units in the calculations section.
10. Carefully put the opening of balloon A onto the opening of bottle A. Be sure to keep the bottom of the balloon tipped down so the baking soda does not spill.
11. Once the balloon is securely on the bottle, lift the top of the balloon up and add the baking soda into the bottle. Hold the top of the bottle to keep the balloon on the bottle. If you hear a lot of gas escaping from the balloon, restart your trial.
12. Once the bubbling has stopped, swirl the bottle carefully while still holding onto the neck of the bottle where the balloon is attached. Allow the bottle to rest. Repeat until bubbling no longer happens when the bottle is swirled.
13. Pinch the bottom of the balloon and tie off the bottom of the balloon.
14. Measure the circumference of the balloon at its widest point.
15. Calculate the radius of the balloon. Be sure to show your calculation, with units, in the calculations section. Record the radius in the data table.
16. Calculate the volume of the balloon. Be sure to show your calculation, with units, in the calculations section. Record the radius in the data table.
17. Calculate the moles of gas in the balloon using the ideal gas relationship that 1 mol gas = 24.0 L of gas at 20°C. You also need to remember that 1 mL = 1 cm^3. Be sure to show your units on the calculations when you record them in the calculations section.
18. Repeat steps 8 to 15 for balloons B and C.

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4.5: Data and Results

Data Table **4.5.1 - Amounts of Reactants and Balloon Size**

Trial	Mass of Baking Soda (g)	Moles of Baking Soda (mol)	Volume Vinegar (mL)	Moles of Vinegar (mol)	Balloon Circumference (cm)	Balloon Radius (cm)	Balloon Volume (cm^3)	Moles of Gas in Balloon (mol)
A								
B								
C								

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4.6: Calculations

Show calculations for each of the following. Be sure to show units on all numbers with units in your math equations.

Trial A

- Moles of Baking Soda
- Moles of Vinegar
- Balloon Radius
- Balloon Volume
- Moles of Gas in Balloon

Trial B

- Moles of Baking Soda
- Moles of Vinegar
- Balloon Radius
- Balloon Volume
- Moles of Gas in Balloon

Trial C

- Moles of Baking Soda
- Moles of Vinegar
- Balloon Radius
- Balloon Volume
- Moles of Gas in Balloon

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4.7: Laboratory Questions

1. What is the maximum number of moles of carbon dioxide that can be generated in trial A? Show your calculations.
2. What was the limiting reactant in trial A? Explain how you know.
3. What is the maximum number of moles of carbon dioxide that can be generated in trial B? Show your calculations.
4. What was the limiting reactant in trial B? Explain your answer.
5. What is the maximum number of moles of carbon dioxide that can be generated in trial C? Show your calculations.
6. What was the limiting reactant in trial C? Explain how you know.
7. Were any of your balloons similar in volume? Why or why not?
8. For each trial, A, B, and C, do the moles of carbon dioxide that could have been generated match the moles of gas in each balloon. Why or why not? Give three reasons to help explain the result.

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CHAPTER OVERVIEW

5: Impact of Temperature on Balloon Volume

5.1: Safety
5.2: Lab Supplies
5.3: Background
5.4: Procedure
5.5: Data and Results
5.6: Calculations
5.7: Laboratory Questions

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5.1: Safety

Only inflate your balloon about 80% full when blowing it up.

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5.2: Lab Supplies

- Ruler
- Thermometer
- Balloon
- String/Measuring Tape
- Large Container/Pot
- Ice
- Cold Water
- Warm Water
- Towel

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5.3: Background

How Gases Behave

When you blow up a balloon, the size increases because you are adding more gas to the balloon. If you do not tie off the balloon, the gas will quickly try to fill the room around it, causing the balloon to deflate. However, if the balloon is tied off, the amount of gas remains relatively constant, at least initially. Over time, most balloons are not perfectly sealed, and some gas will leak out, causing the balloon to deflate.

When a balloon is filled and tied off, the pressure in the balloon will be equal to the pressure around the balloon. At sea level, this pressure will be approximately 1 atm. **Thus, the pressure inside the balloon is constant in this experiment.**

This experiment will investigate the impact temperature has on the volume of a balloon. Remember that this experiment is at constant pressure. We will also assume that on the time frame of the experiment, all the gas molecules we put into the balloon remain in the balloon. **Therefore, the moles of gas are also constant in this experiment.**

When we discuss pressure and volumes, we have a sense that these properties should not have a negative value. A balloon with a negative volume does not make sense! Thus, we will also need temperatures to remain positive as well. We need to work on an absolute scale, where 0 means absolute zero. The scale often used is the Kelvin scale. Temperature can be converted between Kelvin (K) and Celsius ($^\circ C$) using the following equation:

$$K = {}^\circ C + 273.15 \tag{5.3.1}$$

How to calculate the volume of a sphere

To estimate the volume of balloon, assume the balloon to be a sphere. You will measure the circumference of the balloon at its widest point. From this, you can determine the radius and the volume of the balloon.

The radius (r) of the balloon is related to the circumference (C) by the following equation:

$$C = 2\pi r \tag{5.3.2}$$

Once you know the radius (r), you can solve for the volume (V) with the following equation:

$$V = \frac{4}{3}\pi r^3 \tag{5.3.3}$$

If your calculator has a π button, you may use this. If not, estimate π as 3.14159.

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5.4: Procedure

1. Stretch a balloon. Blow it up until it is between 60% and 80% full.
2. Measure the circumference of the balloon. Record the circumference in the data table.
3. Measure the temperature of the room in Celsius. Record the data in the data table.
4. Fill the container with ice and cold water. Add the balloon and cover for 10 minutes.
5. While waiting for 10 minutes, calculate the radius and volume of the balloon at room temperature. Show your work in the calculations section. Report the data in the data table.
6. Convert room temperature from Celsius to Kelvin. Show your work in the calculations section. Report in the data table.
7. After 10 minutes, measure the temperature of the ice water in Celsius. Record the data in the data table.
8. Remove the balloon from the ice water. Measure the circumference of the balloon while it is cold. Record the data in the data table.
9. Empty the cold water from the container and add hot tap water. Add the balloon and cover for 10 minutes.
10. While waiting for 10 minutes, calculate the radius and volume of the balloon at the cold temperature. Show your work in the calculations section. Report the data in the data table.
11. Convert the cold temperature from Celsius to Kelvin. Show your work in the calculations section. Report in the data table.
12. After 10 minutes, measure the temperature of the hot water in Celsius. Record the data in the data table.
13. Remove the balloon from the hot water. Measure the circumference of the balloon while it is hot. Record the data in the data table.
14. Calculate the radius and volume of the balloon at the warm temperature. Show your calculations. Report the data in the data table.
15. Convert the warm temperature from Celsius to Kelvin. Show your calculations. Report in the data table.

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5.5: Data and Results

Data Table **5.5.1** – Balloon Measurements at Various Temperature

Trial	Circumference (cm)	Radius (cm)	Volume (cm^3)	Temp (°C)	Temp (K)
Room Temp					
Cold					
Warm					

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5.6: Calculations

Show your work for the following calculations. Be sure to show units on all numbers that have units.

Room Temperature Calculations

- Radius Calculation
- Volume Calculation
- Temperature Calculation

Cold Temperature Calculations

- Radius Calculation
- Volume Calculation
- Temperature Calculation

Warm Temperature Calculations

- Radius Calculation
- Volume Calculation
- Temperature Calculation

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5.7: Laboratory Questions

1. What was the difference in volume between the room temperature balloon and the cold balloon? Show your calculation.
2. What was the difference in volume between the room temperature balloon and the warm balloon? Show your calculation.
3. Based on your experimental data, what can you conclude about the relationship between the temperature and volume of a gas that is at constant pressure and amount of gas?
4. Why do you think you saw the relationship between temperature and volume that you did? What is happening with the molecules of gas?
5. A thermos (which has constant volume) was filled halfway with hot oatmeal, closed tightly and then left for 24 hours. The oatmeal had cooled to room temperature. The cap of the thermos was very difficult to remove. Why do you think this was? Use gas laws to help you explain.
6. Include a picture of the measuring tape OR string and ruler, as well as the thermometer used. Please be sure to show the tick marks on your tools.

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CHAPTER OVERVIEW

6: Periodic Trends

This laboratory assignment is adapted from the Periodic Trends Lab developed by Ryan Shirley at Clovis Community College.

6.1: Background

6.2: Task

6.3: Laboratory Questions

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6.1: Background

When Mendeleev developed the periodic table, one key feature in his work is that he predicted properties of elements yet to be discovered based on the elements around them. The arrangement of elements on the periodic table allows for trends in reactivity and atomic size to be noticed. In this experiment, you will use Microsoft Excel to analyze data about the size, electronegativity, first ionization energy and electron affinity of the first 36 elements on the periodic table.

Definitions

Atomic radius: one half the distance between the nuclei of two atoms that are bonded together. Often measured in units of picometers (pm) or angstroms (Å)

Electronegativity: the ability of an atom to attract electrons in a bond. Measured on a relative scale between 0 and 4. This scale does not have units.

Ionization Energy: the energy required to remove an electron from an atom. The first ionization energy is the energy required to remove the first electron. This is often measured in units of kJ/mol.

Electron Affinity: the energy released when an electron is added to an atom. Due to this definition, when a negative electron affinity is reported, this means energy is required to add an electron. The units often used to measure this are kJ/mol.

Effective nuclear charge (Z_{eff}): the nuclear charge experienced by the valence electrons. This is calculated by determining the number of protons in an atom's nucleus and subtracting the number of inner (non-valence) electrons: $Z_{eff} = \text{protons} - \text{inner electrons}$

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6.2: Task

1. Open the Excel file.
2. Each sheet has a different set of data. The tabs at the bottom of the sheet are labeled with the data set.
3. Begin on the atomic radius tab. Highlight the data set from cells A3 to E39. Format the data as a table that has headers.
4. Sort the data by atomic number from smallest to largest.
5. Highlight cells D3 to E39. Insert a chart (graph) of this data using a scatter plot that has the lines connected. Be sure to title your graph, the x-axis and the y-axis. Move the chart to cover the box in the Excel file that is provided for the graph.
6. Repeat steps 3 to 5 for the electronegativity sheet, ionization energy sheet, and the electron affinity sheet.
7. Now that you have created tables and graphs, you can begin to analyze the data to help you answer the lab questions. Return to the atomic radius sheet. Using the periodic table as well as your graph, answer lab question 1.
8. Next, filter the data by group number. The group number matches the column on the periodic table where the element is listed. Look at the data for the following groups one at a time: group 1, group 2, group 14, and group 18. You may need to expand the graph or you can look at the numbers in the data table. Use this analysis to help you answer lab question 2.
9. Repeat steps 7 and 8 for the electronegativity sheet and answer lab questions 3 and 4.
10. Repeat steps 7 and 8 for the ionization energy sheet and answer lab questions 5 and 6.
11. Repeat steps 7 and 8 for the electron affinity sheet and answer lab questions 7 and 8.

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6.3: Laboratory Questions

1. What trend in atomic radius do you observe as you move across a row from left to right on the periodic table?
2. What trend in atomic radius do you observe as you move down a column from top to bottom on the periodic table?
3. What trend in electronegativity do you observe as you move across a row from left to right on the periodic table?
4. What trend in electronegativity do you observe as you move down a column from top to bottom on the periodic table?
5. What trend in ionization energy do you observe as you move across a row from left to right on the periodic table?
6. What trend in ionization energy do you observe as you move down a column from top to bottom on the periodic table?
7. What trend in electron affinity do you observe as you move across a row from left to right on the periodic table?
8. What trend in electron affinity do you observe as you down a column from top to bottom on the periodic table?
9. Which periodic trends were easy to observe?
10. Which periodic trends are less clear?
11. How does the concept of effective nuclear charge explain the trends in atomic radius?
12. How does the concept of effective nuclear charge explain the trends in electronegativity?
13. Which elements do not have an electronegativity value? Filter the electronegativity data to show only “(blanks)” to help you answer this question.
14. Think about the definition of electronegativity. What does a blank value for electronegativity tell you about an element’s ability to form bonds?
15. Which group of elements had the highest ionization energies in their rows?
16. Which elements had a negative electron affinity? To answer this question, filter the electron affinity data. Use the “Number Filter” option. Select “Less Than” and enter 0 into the “Less than” field.
17. Energy is a measure of stability. Based on ionization energy and electron affinity data, which group of elements is the most stable? Explain.

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CHAPTER OVERVIEW

7: Molecular Modeling

7.1: Supplies
7.2: Background - Benefits and Limitations of Models
7.3: Background - Lewis Structures
7.4: Background- Molecular Shape
7.5: Activity
7.6: Data Table
7.7: Laboratory Questions

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7.1: Supplies

- Toothpicks
- Gum drops, multicolor mini marshmallows, modeling clay, or colored dough
- Periodic Table

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7.2: Background - Benefits and Limitations of Models

Molecular modeling is a way to help visualize molecules. There are many ways to model molecules, such as using toothpicks and gumdrops, plastic model kits, 3-D printed models or computer programs. It is important to note that these are models. All models will have some limitations. The advantage of using models is that they help us visualize atoms and molecules which are difficult or expensive to see in other ways. Models are useful when they allow us to make accurate predictions about how molecules behave.

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7.3: Background - Lewis Structures

A Lewis structure is a model that depicts the electrons around atoms. We can also put atoms together to make molecules.

Atomic Lewis Structures

To draw an atomic Lewis structure, use the following steps. The carbon atom will be shown as a demonstration.

1. Write the chemical symbol for the element from the periodic table. For carbon:



2. Determine the number of valence electrons based on the periodic table. For carbon, as it is group 4A, this number is 4. If looking at the 1 through 18 column numbers, for groups 1 and 2, the valence electrons match the group number. For groups 13-18, the number of valence electrons is the group number minus 10. For carbon in group 14, 14-10=4 valence electrons.
3. Place the electrons around the atom, one to each of the four sides before pairing. For carbon:



Atomic Lewis structures are an easy way to predict how many bonds an element will often make. The number of single electrons shown on an atomic Lewis structure shows the number of bonds the element is likely to form in neutral molecules. This is a general guideline but many exceptions exist. The bonding patterns in polyatomic ions will often vary from the prediction made based on the neutral atomic structure.

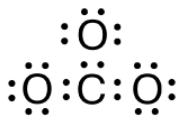
Lewis Structures for Molecules and Polyatomic Ions

To draw a Lewis structure, we will follow a similar process. We will use carbonate (CO_3^{2-}) as our example.

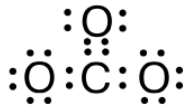
1. Determine the central atom. This is typically the less electronegative atom or the one listed first in the formula (but not always). Hydrogen is never the central atom! For CO_3^{2-}, carbon is less electronegative and listed first, so we will assume it is the central atom.
2. Write out the atoms, with the central atom in the middle and the other atoms arranged around it in space. For carbonate, we can show a carbon atom with one oxygen to the right of the carbon, one oxygen to the left of the carbon and one oxygen above the carbon:



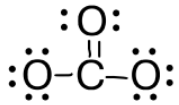
3. Determine the total number of electrons in the molecule/ion.
    a. Add all the electrons in the valence shells of the atoms. For carbonate, 4 electrons come from the carbon and 6 electrons come from each oxygen. Thus, the total number of valence electrons in carbonate is 4 electrons + 3 * 6 electrons = 22 electrons.
    b. Add or subtract electrons from the neutral atoms based on the charge of the ion. For neutral molecules, omit this step. For carbonate, the 2- charge indicates that two more electrons must be added. (Recall that electrons have a negative charge.) The total number of electrons in the carbonate Lewis structure will be 22 electrons + 2 electrons = 24 electrons.
4. Distribute the electrons, ensuring that all outside atoms have 8 electrons either shared or as lone pairs around it. The exceptions to 8 electrons are H, which has only 2 electrons; Be, which usually only has 4 electrons; and B, which often has 6 electrons. For carbonate, each oxygen atom has four pairs of electrons, with each oxygen sharing one pair of electrons with the carbon as shown:



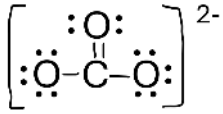
5. See if the central atom has 8 electrons (an octet) around it. The exceptions for Be and B still apply. When we look at carbonate, we see that the C only has 6 electrons around it, due to three oxygens each sharing one pair of electrons.
6. Share more electrons to fill the octet for the central atom. Remember to remove a lone pair of electrons from the atom that is being shown to share more electrons. If more than one atom could share electrons, pick one to share. For carbonate, any of the oxygens could share more electrons (top, left, or right). It does not matter which oxygen is picked to share more electrons, so we will pick the top one.



7. Count the electrons again to make sure that the correct number are shown. In the structure above in step 6, we have shown 24 electrons.
8. To simplify the structure, replace the bonding electrons with one line for each pair. The lone pairs remain as dots. For carbonate, there will four lines between carbon and the oxygens. One oxygen will have 2 lines between it and carbon, the other two will have one line:



9. For ions, draw brackets around the Lewis structure and indicate the charge. For carbonate, we will use the structure shown in step 8, add brackets, and place the 2- charge in the upper righthand corner:



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7.4: Background- Molecular Shape

The three-dimensional shape of a molecule has an impact on its physical and chemical properties. In this lab activity, you will investigate why molecules and ions have the shape they do. This is the first block in understanding molecular shapes. As you advance in chemistry courses, you will learn how molecular shape can impact how effective medicines are as well as why enzymes in your body work.

For this discussion, we will begin with simple binary molecules and ions that have one central atom. The central atom is the atom to which all the other atoms are bonded. This will be the atom whose geometry we are investigating.

Why do these molecules or ions adopt the shapes they do? We can answer this by noticing that there is Valence Shell Electron Pair Repulsion (VSEPR). In other words, we notice that the electrons in the outermost shell (valence electrons) of an atom determine the geometry of that atom. We defined a covalent bond as sharing pairs of electrons. We will also assume that electrons that are nonbonding also exist in pairs (this assumption is true for many molecules, but not all). This is where the “pair” comes from. Finally, as we are discussing electrons, we know they have a negative charge. Like charges repel or push away from each other.

Each bonding region and each nonbonded pair of electrons counts as a region of electron density that can repel other areas of electron density. What is meant by a bonding region is a “side” of the central atom as represented in the Lewis structure. Thus, a bonding region may contain a single bond, double bond, or triple bond. Because of this, a bonding region may contain 2, 4, or 6 electrons and still be one bonding region. For this discussion, a nonbonding region will be a lone pair of electrons.

When the structure of molecules is determined by x-ray diffraction, the nucleus of atoms can be seen, but the electrons cannot. Thus, when we determine the shape of a molecule, we are describing the location of the atoms in space. The effect of the nonbonding electrons helps determine the shape, but we do not expressly name the location of all the electron groups.

Linear – Two atoms attached to the central atom, no nonbonding electrons on central atom

If a central atom has two regions of electron density, the farthest apart these regions can be from each other is 180°. This arrangement will cause all the atoms to be in a line. Thus, the name of this shape is linear. The bond angle will be 180°. This can be shown using the letter A for the central atom and R for the attached group, as shown in Figure **7.4.1**.

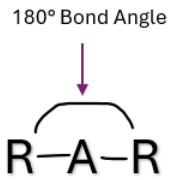


Figure **7.4.1**

Trigonal Planar – Three atoms attached to the central atom, no nonbonding electrons on central atom

If a central atom has three regions of electron density, the farthest apart these regions can be from each other is 120°. This arrangement will cause all the atoms to be in a triangle, with the central atom in the center of the triangle and the surrounding atoms at the corners of the triangle. Triagonal means three and planar means that it is all in one plane. Hence the name trigonal planar is used to describe this shape. The bond angle will be 120°, as shown in Figure **7.4.2**.

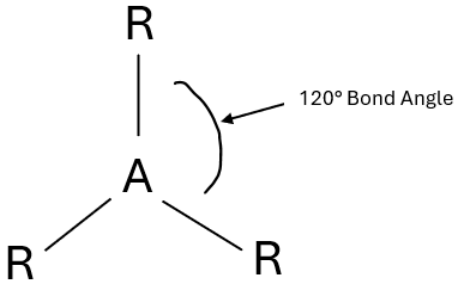


Figure **7.4.2**

Bent – Two atoms attached to the central atom, one pair of nonbonding electrons on central atom

As noted with the trigonal planar shape, the furthest apart three regions of electron density can be is 120°. However, the nonbonding electrons fill more space than the bonding electrons. Thus, the bond angle between the atoms is slightly less than 120°. Because an x-ray of the molecule only shows the atoms, we describe this shape as bent, as shown in Figure **7.4.3**.

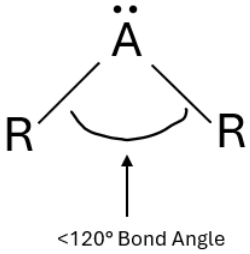
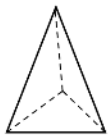


Figure **7.4.3**

Tetrahedral – Four atoms attached to the central atom, no nonbonding electrons on central atom

The first three shapes we have described can be depicted in one plane and are thus easy to visualize on paper. When four atoms are attached to a central atom, it is easy to think that they would be spread out in a square. This error does not account for the electron density spreading out in three dimensions, not two.

The four regions of electron density will spread out to have the central atom in the middle and the four groups attached at the vertices of a tetrahedron. The tetrahedron shape is shown below in Figure **7.4.4**. It is a solid that has four vertices and four sides. Each side looks like a triangle.



tetrahedron

Figure **7.4.4**

The bond angle between the groups is 109.5°, as shown in Figure **7.4.5**. To show this shape in two dimensions, a wedge is used to depict the group that is coming out of the page and a dash is used to depict the group going back behind the page.

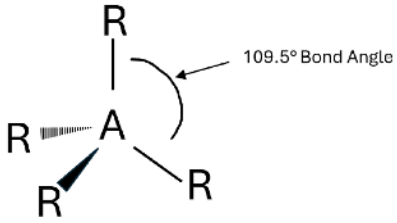


Figure **7.4.5**

Trigonal Pyramidal – Three atoms attached to the central atom, one pair of nonbonding electrons on central atom

As seen with the tetrahedral shape, four regions of electron density can spread out 109.5° from each other in space. Again, a nonbonding pair of electrons will take up more space than bonding electrons and push the groups around the central atom just a little bit closer together. Thus, the bond angle in the trigonal pyramidal shape will be less than 109.5°. This shape can be depicted with one group on a dash, one on a wedge and one in the plane to help emphasize that the atoms are not all in one plane, as shown in Figure **7.4.6**.

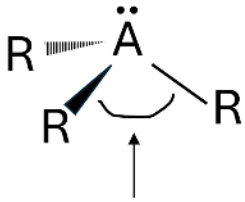


Figure **7.4.6**

Bent – Two atoms attached to the central atom, two pairs of nonbonding electrons on central atom

This shape, shown in Figure **7.4.7** is similar to the bent structure with only one nonbonding pair of electrons. The difference is that the bond angle is less than 109.5°. This is due to there being four regions of electron density instead of three as there are two sets of lone pairs on the central atom.

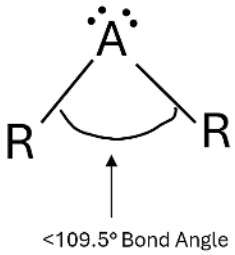


Figure **7.4.7**

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7.5: Activity

For this lab activity, you will be modeling molecules with toothpicks and gumdrops (or marshmallows). You will also practice drawing Lewis structures and assigning the shape of the molecule.

Instructions

Complete the following table with the missing data. For the picture of the model, please take a picture of the molecule and insert the picture into the table. BeH_2 was completed as an example. Please keep in mind that Lewis structures for non-linear molecules cannot be easily typed on a computer. Rather, use a drawing program or hand draw the Lewis structures and include photos of these as well.

In your model, use toothpicks on the central atom to represent pairs of electrons, whether they are bonding or nonbonding. However, when determining the shape of the molecule, only consider where the other "atoms" (gumdrops) are in your model.

Worked Example – BeH_2

1. Create a Lewis structure for BeH_2.
    a. Determine valence electrons. Be has 2 valence electrons. Each H has 1 valence electron.
        i. $\textbf{Valence electrons} = 2 \text{ electrons} + 2 * (1 \text{ electron}) = 4 \text{ electrons}$
2. Arrange the atoms in space, with Be in the center.
    a. H Be H
3. Share a pair of electrons between each of the atoms.
    a. H:Be:H
4. Count the electrons. 4 are distributed and there were only 4 valance electrons in this molecule. Be is an exception to the octet rule.
    a. If desired, redraw electron pairs as lines.
        i. H–Be–H
5. Check the charge. This is a neutral molecule, so no charge needs to be shown.
6. Build a model.
    a. Choose one color gumdrop for Be and another for H.
    b. Arrange the toothpicks on opposite sides of the Be "atom" or gumdrop.
    c. Add the other two gumdrops.
    d. Take a picture of the model.
    e. Insert the picture into the table. **Please be sure to complete this step yourself!**
7. Assign the shape.
    a. Use both the information in the background section and your model to help you assign the shape of the molecule.
    b. BeH_2 is linear in shape as there are only two regions of electron density on the central atom.

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7.6: Data Table

Table of Lewis Structure, Model, and Shape based on Molecular Formula

Molecular Formula	Lewis Structure	Picture of Model	Shape
BeH_2	H–Be–H		Linear
CO_2			
BH_3			
CH_4			
NH_3			
NH_4^+			
H_2O			

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7.7: Laboratory Questions

1. What are two benefits of using a Lewis structure?
2. What are two limitations of a Lewis structure?
3. What are two benefits of using a three-dimensional model?
4. What are two limitations of a three-dimensional model?
5. Compare BeH_2 and CO_2. When considering shape, are double bonds counted as one region or two regions of electron density?
6. Using the Lewis structure in the background information, what shape would you predict the carbonate ion to be?
7. Dichloromethane, which is also known as methylene chloride, is a chemical solvent with many uses, which include decaffeinating coffee and removing paint. However, due to health risks, the EPA has banned on consumer uses of dichloromethane. The formula of dichloromethane is CH_2Cl_2. What is its Lewis structure?
8. What are the current allowed uses of dichloromethane under the EPA restrictions? Be sure to include a citation of your source.

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CHAPTER OVERVIEW

8: Intermolecular Forces and Density

8.1: Safety
8.2: Supplies
8.3: Background - Density and Intermolecular Forces
8.4: Procedure
8.5: Data and Results
8.6: Calculations
8.7: Laboratory Questions
8.8: References

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8.1: Safety

Do not drink mineral oil. A laboratory apron will help keep your clothes from being stained by food coloring.

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8.2: Supplies

- 50-mL graduated cylinder
- Pocket balance
- 2 clear plastic cups (clear drinking glasses or 1 cup mason jars also work)
- Water
- Vegetable or Mineral Oil
- Food Coloring
- Toothpick, spoon, or chopstick
- Paper towel

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8.3: Background - Density and Intermolecular Forces

Density

Density is a measure of how much matter (mass) is in a given volume. Density is calculated:

$$Density = \frac{mass}{volume} \tag{8.3.1}$$

There are many possible units for reporting density. In this experiment, as the mass will be in grams and the volume is measured in milliliters, density will be reported in units of g/mL. You will be asked to think about density more in your lab questions after you have collected your data.

This section is adapted from OpenStax Chemistry 2e.

Intermolecular Forces

Within a covalent molecule, the atoms are attached via shared electrons. Intermolecular forces are the interactions between molecules. They are much weaker than chemical bonds. Vaporizing a liquid to the gas phase involves breaking the intermolecular forces as the molecules move apart from each other to the point that they no longer interact. The molecules themselves still exist in the gas phase – the chemical bonds have not been broken. Thus, boiling point is a way for us to compare the relative strength of intermolecular forces.

Types of Intermolecular Forces

There are four types of intermolecular forces:

1. Dispersion Force (also called London Dispersion Force)
2. Dipole-dipole interaction
3. Hydrogen bonding (much weaker than ionic or covalent bonds!)
4. Ion-dipole interaction

The first three can exist within a substance or within mixtures of substances. The fourth exists only within mixtures of substances.

Dispersion Force

Dispersion forces are caused by temporary dipoles within molecules. Electron density within a molecule shifts, momentarily causing one part of a molecule to have a partial negative charge (δ-). The region with less electron density then momentarily has a partial positive charge (δ+). The partially positive region then attracts the electron density of a neighboring molecule. As the electrons are constantly shifting, the dipoles change within the molecules. All molecules have electrons so all molecules can interact via dispersion forces.

The strength of dispersion forces depends on how easily the electron density shifts within a molecule. The technical name for this process is the polarizability of the molecule. The more electrons and the large the molecule, the more polarizable a molecule becomes. Thus, as the number of atoms or the size of the atoms of a molecule increases, the more dispersion forces are present in a molecule.

Dipole-Dipole Interaction

Polar molecules have permanent regions of partial positive (δ+) and partial negative (δ-) charge due to the electronegativity difference between the atoms. The dipoles of the molecules can then align so the partially positive region in one molecule interacts with the partially negative region in another molecule. This is shown in Figure 8.3.1 .

Due to the dipoles in polar molecules being permanent instead of temporary, this intermolecular force is stronger. This can be seen by comparing the boiling point of SiH_4 (silane) to HCl (hydrogen chloride). Both have a similar molar mass, 32.117 g/mol for silane and 36.46 g/mol for hydrogen chloride, so the dispersion forces are expected to be similar. However, the boiling point at 1 atm for silane is -185°C as compared to hydrogen chloride boiling point at 1 atm of -85°C. Thus, more energy is required to separate molecules of the polar hydrogen chloride as compared with the nonpolar silane molecules.

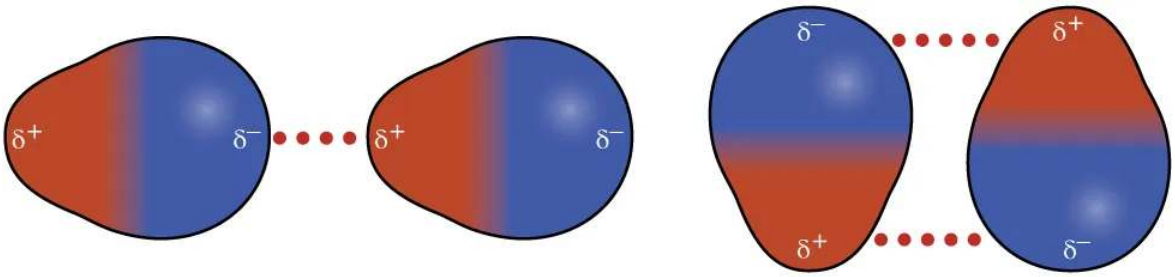


Figure **8.3.1** This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another. Image courtesy of OpenStax, licensed under CC-BY 4.0. Access this image for free at openstax.org.

Hydrogen bonding

First and foremost, **hydrogen bonding is not as strong as an ionic or covalent bond**! The name is misleading. Hydrogen bonding (H-bonding) interactions are extra strong dipole-dipole interactions. They occur only in systems that contain an O-H, N-H, or H-F bond. The extreme dipole is causes by the large electronegativity difference between oxygen, nitrogen, and fluorine relative to hydrogen. The second factor that explains this extreme dipole is that the hydrogen atom has a very small valence shell, so when the electron density is shifted away from the hydrogen atom, the positively charged proton in the hydrogen atom's nucleus is more exposed and is not shielded by inner electrons as happens in larger atoms.

The strength of the hydrogen bonding interaction can be seen in comparing the boiling points at 1 atm of hydrogen chloride and hydrogen fluoride. Hydrogen fluoride has a much lower molar mass than hydrogen chloride, so based on dispersion forces, the prediction would be that it would also have a lower boiling point. However, the boiling point of hydrogen fluoride is 20°C, which is much higher than -85°C for hydrogen chloride.

Ion-Dipole Interactions

Ion-dipole interactions occur when ions (which are charged) interact with polar molecules in solution. For example, when hydrogen chloride is added to water, it becomes hydrochloric acid, which has H_3O^+ and Cl^- ions surrounded by water molecules. The water molecules are polar and have partially negative regions and partially positive regions. The partially negative regions of water interact with the positively charged H_3O^+ (hydronium) ions. The partially positive regions of water interact with the negatively charged Cl^- (chloride) ions.

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8.4: Procedure

Part 1 A: Water and Food Coloring

1. Record the mass of the 50-mL graduated cylinder in the first data table.
2. Add 25 mL of water into the graduated cylinder. Record the mass and the volume with appropriate significant figures in the first data table.
3. Calculate the mass of the water by subtracting the mass from step 1 from the mass in step 2. Be sure to show your calculations in the calculations section.
4. Calculate the density of the water. Be sure to show your calculations in the calculations section.
5. Pour the water into a 10 oz clear plastic cup.
6. Add one drop of blue food coloring. Observe what happens. Record your observations in the second data table.
7. Stir the mixture with a clean toothpick, chopstick, or spoon. Observe what happens. Record your observations in the second data table.
8. Dry the graduated cylinder. A paper towel rubber banded to a chopstick works for this task.

Part 1 B: Oil and Food Coloring

1. After the word “oil” in the first data table, record the type of oil used for the experiment.
2. Record the mass of the dried 50-mL graduated cylinder in the first data table.
3. Add 25 mL of oil into the graduated cylinder. Record the mass and the volume with appropriate significant figures in the first data table.
4. Calculate the mass of the oil subtracting the mass from step 2 from the mass in step 3. Be sure to show your calculations in the calculations section.
5. Calculate the density of the oil. Be sure to show your calculations in the calculations section.
6. Pour the oil into a 10 oz clear plastic cup.
7. Add one drop of red food coloring. Observe what happens. Record your observations in the second data table.
8. Stir the mixture with a clean toothpick, chopstick, or spoon. Observe what happens. Record your observations in the second data table.

Part 2: Mixing Solutions

1. Pour the oil solution into the cup with the water. Observe what happens. Take a picture from the side of the cup. Record your observations.
2. Over a sink, pour the mixture back and forth between the two cups 30 times.
3. Take a picture of the mixture from the side of the cup. Record your observations.
4. Allow the mixture to stand for 1 hour.
5. Observe the mixture after 1 hour. Take a picture from the side of the cup. Record your observations.

Clean Up

1. Wash your graduated cylinder with soap and hot water. You will need to gently scrub with a bottle brush or a paper towel on a chopstick to remove the oil.
2. Separate the oil and water. The water can go down the drain. Oil should be disposed of in the trash or in a container for oil collection. Please see your local regulations about disposing of used cooking oil.
3. The cups can be washed with hot soapy water and air dried for use in future experiments.

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8.5: Data and Results

Part 1 Data

Table **8.5.1** - Mass, Volume, and Density of Liquids

Liquid	Mass of empty Graduated Cylinder (g)	Mass of empty Graduated Cylinder +Liquid (g)	Mass of Liquid (g)	Volume of Liquid (mL)	Density of Liquid (g/mL)
Water					
Oil					

Table **8.5.2** - Observations When Food Coloring is Added to Liquid

Liquid	Observation Upon Initial Addition of Food Coloring to Liquid	Observation of Food Coloring in Liquid After Stirring
Water		
Oil		

Part 2 Data: Mixing Solution

Record your observations of the the following:

- Initial Observations of Both Liquids Together in One Cup
- Photo of Both Liquids Together in One Cup
- Observations of After Mixing Liquids Together
- Photo of After Mixing Liquids Together
- Observations of Mixture After 1 Hour
- Photo of Mixture After 1 Hour

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8.6: Calculations

Be sure to show your calculations for the following:

- Mass of Water
- Density of Water
- Mass of Oil
- Density of Oil

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8.7: Laboratory Questions

1. Was the initial mass of graduated cylinder the same before you added the oil and before you added the water? Why or why not?
2. Which liquid had a higher density, water or oil?
3. Which liquid floated on top of the other? What conclusions can you draw about the relationship between the density of a substance and how it interacts with another substance?
4. Did the food coloring sink or float in the oil? How does the density of the food coloring compare to the oil?
5. Did the food coloring sink or float in the water? How does the density of the food coloring compare to the water?
6. Do the food coloring and the oil have similar intermolecular forces? What data did you use from your experiment to reach your conclusion?
7. Do the food coloring and the water have similar intermolecular forces? What data did you use from your experiment to reach your conclusion?
8. Water is H-O-H and bent in shape. This molecule is polar. What intermolecular forces are present in water?
9. Oil is made of molecules that have only C-H and C-C bonds. Is this molecule polar or nonpolar? Why?
10. What type of intermolecular forces are present in oil?
11. Consider the following statement:

    - Nonpolar molecules with only dispersion forces are still attracted to polar molecules with dipole-dipole or H-bonding intermolecular forces. However, the polar molecules are more attracted to each other and thus exclude the nonpolar molecules.

    Does your experimental data support this statement? Why or why not? Think about what happened when you mixed the oil and water.

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8.8: References

(1) Flowers, P.; Theopold, K.; Langley, R.; Robinson, W. R. *Intermolecular Forces - Chemistry 2e - OpenStax*. Openstax.org. CC-BY. https://openstax.org/books/chemistry-2e/pages/10-1-intermolecular-forces (accessed July 31, 2024).

(2) PubChem. *Hydrochloric Acid*. pubchem.ncbi.nlm.nih.gov. https://pubchem.ncbi.nlm.nih.gov/compound/Hydrochloric-acid (accessed July 31, 2024).

(3) PubChem. *Hydrofluoric Acid*. pubchem.ncbi.nlm.nih.gov. https://pubchem.ncbi.nlm.nih.gov/compound/14917 (accessed July 31, 2024).

(4) PubChem. *Silane*. pubchem.ncbi.nlm.nih.gov. https://pubchem.ncbi.nlm.nih.gov/compound/silane (accessed July 31, 2024).

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CHAPTER OVERVIEW

9: All About That Base and Acid

9.1: Safety
9.2: Supplies
9.3: Background
9.4: Procedure
9.5: Data and Results
9.6: Laboratory Questions
9.7: References

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9.1: Safety

- While all the substances used in this lab are found around your house, some of them should not get in your eyes. Thus, wear your goggles when testing the pH and making solutions.
- To avoid stains from the red cabbage juice, wear your lab apron.
- Be careful when cutting the cabbage to not have your fingers under the knife.
- Be careful with a hot stove to avoid burns.

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9.2: Supplies

- Test tubes & rack OR condiment cups
- Red/purple cabbage
- Water
- Stove
- Pot
- Strainer
- Knife and cutting board
- Lime OR lemon juice
- White vinegar OR white wine vinegar
- Baking soda
- Washing soda OR borax
- Sparkling water
- Clear dish soap
- Clear or white shampoo

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9.3: Background

Definitions of Acids and Bases

Acids and bases are compounds found in everyday life. Stomach acid, as the name implies, is an acid that helps break down food in the stomach. Acids are also useful to safely preserve canned food. Both acids and bases are used in cooking. Bases are also used in soap making.

But what is an acid or a base? There are several definitions, but for this lab we will use the Brønsted-Lowery definitions of acids and bases. By this definition, an acid is a compound that generates H^+ ions in solution. A base accepts H^+ ions in solution. A hydrogen atom that has lost an electron is merely a proton. Thus, an acid is also defined as a proton donor in solution. A base is a proton acceptor in solution.

Autoionization of Water and Neutral Solutions

For the scope of this discussion, all solutions will be aqueous solutions. This means water is the solvent, or major component, of the solution. Water, H_2O, can itself be either a base or an acid. Another way to say this is that water is amphoteric. In the following reaction, the left $\mathbf{H_2O}$ is accepting a proton from the right H_2O. Thus, the left $\mathbf{H_2O}$ is the base and the right H_2O is the acid.

$$\mathbf{H_2O(l)} + H_2O\ (l) \rightarrow H_3O^+(aq) + OH^-(aq) \qquad (9.3.1)$$
$$K_w = 1.0 * 10^{-14}$$

In water, H^+ ions do not exist by themselves. Instead, they exist as hydronium ions (H_3O^+). The concentration of hydronium ions is a measure of how acidic or basic a reaction is. We saw in equation 9.3.1 that water can autoionize to form hydronium ions and hydroxide ions. When the concentration of hydronium ions and hydroxide ions is equal, the solution is neutral. K_w is the equilibrium constant for the autoionization of water.

$$K_w = [H_3O^+][OH^-] \qquad (9.3.2)$$

The brackets around the ions in equation 9.3.2 mean that the concentration of the ions is considered. When $[H_3O^+] = [OH^-]$, as it does in a neutral solution, then $K_w = [H_3O^+] * [H_3O^+]$. As we saw in equation 9.3.1, $K_w = 1.0 * 10^{-14}$. Thus, in a neutral solution, $1.0 * 10^{-14} = [H_3O^+]^2$. Solving for $[H_3O]^+$ in a neutral solution, $[H_3O^+] = 1.0 * 10^{-7}$.

pH Scale

The pH scale, or power of hydrogen scale, is a way to discuss hydronium ion concentration without using extremely small numbers.

$$pH = -\log[H_3O^+] \qquad (9.3.3)$$

Thus, for a neutral solution, $pH = -\log(1.0 * 10^{-7})$. Solving, $pH = 7$ in a neutral solution.

If $[H_3O^+] > [OH^-]$, then the solution is considered acidic. This means that $[H_3O^+] > 1.0 * 10^{-7}$. The exponent on the 10 will become larger than -7. Recall that when a number becomes less negative, it gets closer to 0. Thus, for an acidic solution, $pH < 7$. For example, gastric secretions, also know as stomach acid, have a pH of about 3 after eating.[1]

If $[H_3O^+] < [OH^-]$, then the solution is considered basic or alkaline. This means that $[H_3O^+] < 1.0 * 10^{-7}$. The exponent on the 10 will become less than -7. Recall that when a number becomes more negative, it is farther from zero and its magnitude is larger. Thus, for an alkaline (basic) solution, $pH < 7$. For example, the pH of alkaline bottled water typically has a pH of 9.5.

In summary:

- The pH scale ranges from 0 to 14.
- A pH of 7 is neutral.
- A pH below 7 is acidic.
- A pH above 7 is basic/alkaline.

Figure 9.3.1. pH Scale. Created by Christinelmiller. CC-BY.

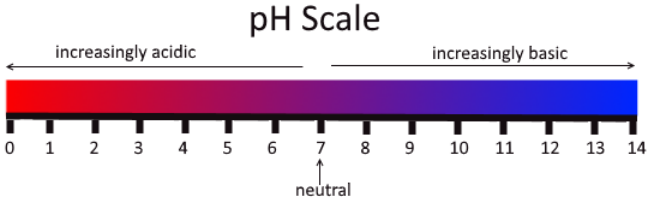


Figure 9.3.1. pH Scale. Created by Christinelmiller. CC-BY.

Measuring pH

Three main methods exist to determine the pH of a solution. They are:

- pH paper/test strips (see figure 9.3.2)
- pH indicators
- pH meters (see figure 9.3.3)



Figure 9.3.2. pH Test Strips. Created by ScienceStockPhotos. CC-BY



Figure 9.3.3. pH meter. Cropped from image by Datamax. Public Domain.

The first two methods are similar in that both the test strips and the pH indicators contain chemicals that change color when they react with acids or bases. Some indicators can undergo several different reactions and change several different colors. Other indicators can only undergo one reaction and just change color once at a specific pH.

The main difference between a test strip and an indicator is how these tools are used. A drop of solution can be added to a test strip and the color checked. In contrast, the pH indicator is added directly to a sample of the solution that is large enough to easily see.

Instead of using color, a pH meter determines the electrical current in the solution. For the current to be correctly related to the pH, the instrument must be calibrated with solutions of a known pH.

This experiment will use red cabbage juice as an indicator to help determine the pH of a variety of household solutions. Red cabbage juice can range in color from red to pink to blue to green to yellow, depending on the pH of the solution. The color

depends on how the anthocyanins in the juice react with the solution being tested. The anthocyanins are amphoteric just as water is, so they can react either as a base or an acid.

[1]Gastric Secretion. In *Encyclopidea Britannica*. https://www.britannica.com/science/human-digestive-system/Contractions-and-motility (accessed 18 March 2024).

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9.4: Procedure

9.5: Data and Results

Experiment Overview

For this experiment, you will first need to make red cabbage juice. Then you will test a variety of household items to determine the pH of the substance. As you complete the data tables, be sure to record the item and its source/brand. For example, if you test lemon juice that is freshly squeezed, you can record that it is fresh lemon juice. In contrast, if you used bottled lemon juice, record the brand.

Some of the items to be tested are solids. However, pH is a measure of the concentration of hydronium ions in solution. Thus, it will be necessary to make a solution. Be sure to record the mass of the solid and the volume of water used to make your solution.

Be sure to take a picture of all your solutions before you clean up!

Experimental Procedure – Part 1: Preparing Red Cabbage Juice

To prepare the red cabbage juice:

1. Chop one quarter of a head of red cabbage into small pieces.
2. Place in a medium pot. Add enough water to cover the cabbage.
3. Heat the water to a boil, then simmer for 20 minutes or until the solution is dark purple.
4. Remove from heat and cool.
5. Pour the liquid into a bowl through a strainer to remove the cabbage.

Experimental Procedure – Part 2: Make Solutions from Solids

To prepare solutions of solids:

1. Label a small cup with the name of the solid.
2. Tare the labeled cup.
3. Measure out about 1 gram of the solid. Record the mass in the first data table.
4. Add 5 mL of water to the cup. Record the exact volume of water added in the first data table.
5. Stir until the solid is dissolved.

Experimental Procedure – Part 3: Test pH

1. Number 8 test tubes or condiment cups from 1 to 8.
2. In test tube/cup 1, add either lime juice or lemon juice. Fill the test tube about halfway. Record the juice type and source in the second data table.
3. In test tube/cup 2, add either white vinegar or white wine vinegar. Fill the test tube about halfway. Record the vinegar type and source in the second data table
4. In test tube/cup 3, add the solution of baking soda made in part 2. Fill the test tube about halfway. Record the baking soda brand in the second data table.
5. In test tube/cup 4, add the solution of either washing soda or borax made in part 2. Fill the test tube about halfway. Record the chemical and brand in data the second data table.
6. In test tube/cup 5, add sparkling water. Fill the test tube about halfway. Record the sparkling water brand and flavor in the second data table.
7. In test tube/cup 6, add clear dish soap. Fill the test tube about halfway. Record the brand and scent in the second data table.
8. In test tube/cup 7, add clear or white shampoo. Fill the test tube about halfway. Record the brand and scent in the second data table.
9. In test tube/cup 8, add tap or drinking water. Fill the test tube about halfway. Record the source in the second data table.
10. Predict the pH of each solution. This is only a prediction, so give it your best guess.

11. **STOP! Before you do this step, did you make your prediction in step 10?** Using a dropper, add three drops of red cabbage juice to each test tube. Swirl the test tube to mix. Record the color in each test tube in the data table 2. If you do not observe a color change, add three more drops of indicator.
12. Using the information in red cabbage juice infographic, shown in Figure **9.4.1**, estimate the pH of each solution based on its color to the nearest integer. Record the number in the second data table.
13. Take a picture of your **labeled** test tubes/cups on a white background. Be sure the labels are clearly visible in the picture. Submit the picture with your lab report.

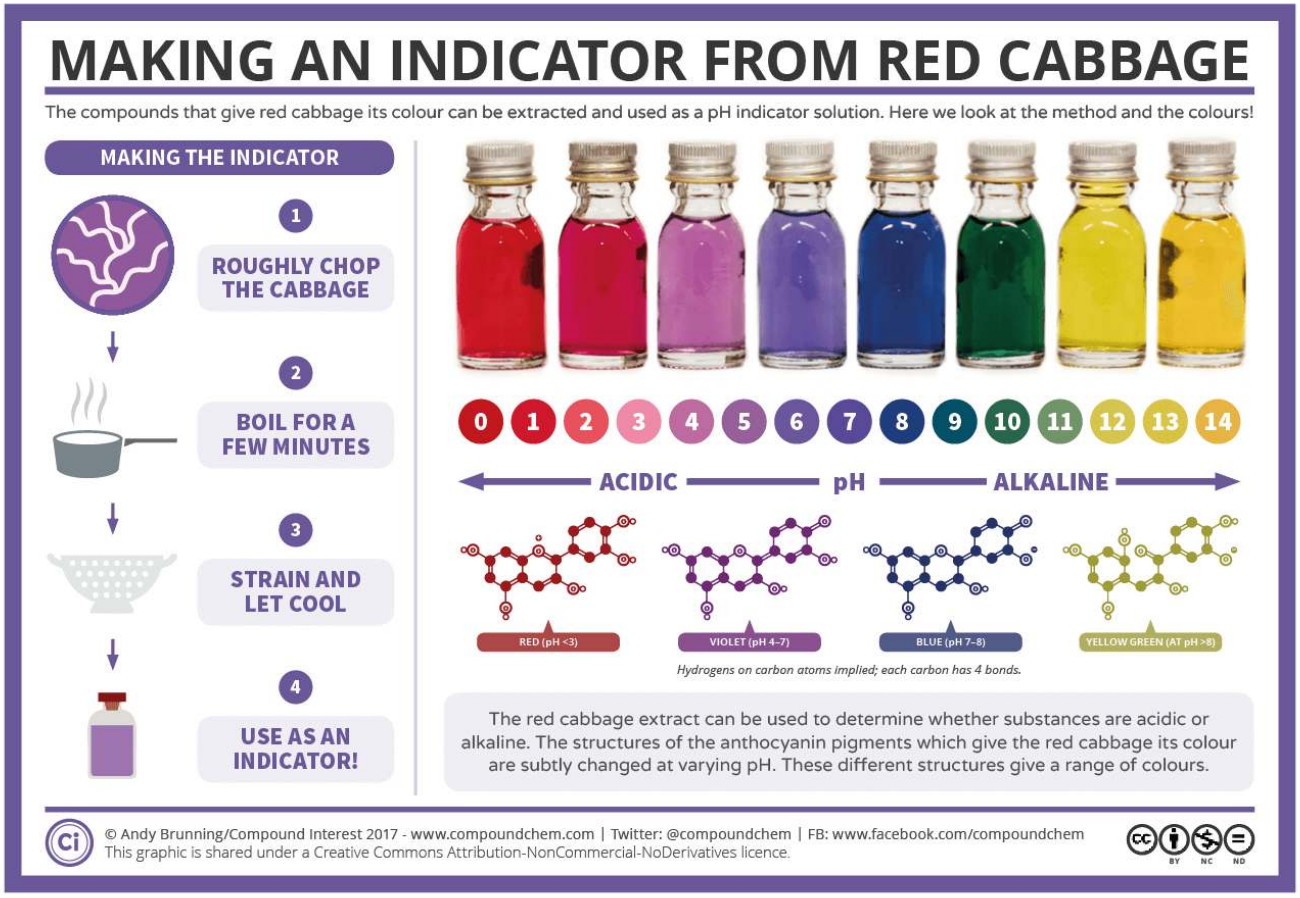


Figure **9.4.1** - Red Cabbage Indicator Scale By Andy Brunning CC-BY-NC-ND

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9.5: Data and Results

Data Table **9.5.1** – Solutions of Solids

Solid	Source	Mass (g)	Volume of Water (mL)

Data Table **9.5.2** – pH of Solutions Based on Color

Test Tube	Solution	Source	Predicted pH	Color	Actual pH
1					
2					
3					
4					
5					
6					
7					
8					

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9.6: Laboratory Questions

1. Which solution was the most acidic of all the solutions you tested? What experimental data led to your conclusion?
2. Which solution is the most basic/alkaline of all the solutions you tested? What experimental data led to your conclusion?
3. Did any of the solutions surprise you by their pH? Explain why or why not.
4. When testing juice, you tested either lime juice or lemon juice. Would you easily be able to test the pH of orange juice with red cabbage juice indicator? Why or why not?
5. Submit the picture of your solutions.

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9.7: References

(1) Gastric Secretion. In *Encyclopidea Britannica*. https://www.britannica.com/science/h...s-and-motility (accessed 18 March 2024).

(2) Christinelmiller. *PH_Scale*; 2018. https://en.wikipedia.org/wiki/File:P....png#filelinks (accessed 2024-03-18).

(3) sciencestockphotos.com. *PH Litmus Paper Chart and Strips*. https://sciencestockphotos.com/free/.../ph_paper.html (accessed 2024-03-18).

(4) Datamax. *PH Meter*; 2008. https://en.m.wikipedia.org/wiki/File:PH_Meter.jpg (accessed 2024-03-18).

(5) Brunning, A. *Making Indicator from Red Cabbage*. Compound Interest. https://i0.wp.com/www.compoundchem.c...ator.png?ssl=1 (accessed 2024-03-18)

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CHAPTER OVERVIEW

10: Hot and Cold Processes

10.1: Safety
10.2: Supplies
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10.1: Safety

- Wear a lab apron to keep your clothes dry. Wear goggles and gloves to protect you from the hydrogen peroxide solution.
- Look at the temperature from the side of the beaker or cup when running the experiment. Do NOT put your face directly over the beaker or cup.

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10.2: Supplies

- Baking Soda (Sodium Bicarbonate)
- Baker’s yeast (instant or active)
- 3% hydrogen peroxide solution in water (make sure it’s not expired)
- 250 mL beaker
- Thermometer
- 50 mL graduated cylinder
- Stopwatch

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10.3: Background

Did a Reaction Occur?[1]

When a process happens, chemists want to know if a chemical change occurred or if only a physical change occurred. Essentially – did a new substance get made? If a substance only changes its state of matter, then a chemical reaction did not occur. If ions mixed but no new liquid, solid, or gas was formed, no chemical reaction occurred.

What evidence do chemists look for to determine if a chemical reaction occurred? On a macroscopic level (can been seen with the human eye):

- Clear solution becomes cloudy/solid precipitates from solution
- Gas bubbles form in solution
- Color change occurs
- New liquid has formed (this can be more difficult to observe for most reactions, as they often occur in solution).

Chemists can also use other tools that allow for detection of chemical changes on a molecular level, which cannot be seen by human eyes. (Young, 2024)

Endothermic vs Exothermic

Some physical changes and some chemical changes involve the transfer of heat. If the system absorbs heat, this process is considered endothermic. The observation made is that the surroundings get colder due to the heat being taken from the surroundings. Conversely, if the system releases heat, the process is considered exothermic. The observation made is that the surroundings get warmer.

Dissolving Salts for Cold and Hot Packs

A common endothermic process you may have used is an instant cold pack. Often these cold packs contain water and ammonium nitrate in separate pouches. By breaking the pouches and mixing the water with the solid ammonium nitrate, the water dissolved the solid ammonium nitrate.

$$\mathrm{NH_4NO_3(s) + heat \xrightarrow{H_2O} NH_4NO_3(aq)} \tag{10.3.1}$$

This is not a chemical reaction, because the ammonium nitrate just changes phase from a solid to aqueous.

Instant hot packs work in a similar way to instant cold packs. However, instead of ammonium nitrate being in the pouch, magnesium sulfate is in the pouch instead. When the pouches break, the magnesium sulfate and water mix, dissolving the magnesium sulfate. This process releases heat to the surroundings, so it is exothermic.

$$\mathrm{MgSO_4(s) \xrightarrow{H_2O} MgSO_4(aq) + heat} \tag{10.3.2}$$

Burning Fuel

Another common exothermic process is the combustion reaction of methane or propane. This is used in many homes to cook food either on a stove or on a gas grill. The heat given off is what allows the food to be cooked. The reaction of methane with oxygen is:

$$\mathrm{CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) + heat} \tag{10.3.3}$$

This process is a chemical reaction because methane and oxygen are transformed into carbon dioxide and water.

Heat is sometimes written on the left side of the arrow for an endothermic process and the right side of the arrow for exothermic processes.

Dissolving Sodium Bicarbonate in Water

In this lab, you will examine two processes. In the first process, you will dissolve baking soda in water. Dissolving sodium bicarbonate can be represented:

$$\mathrm{NaHCO_3(s) \xrightarrow{H_2O} NaHCO_3(aq)} \tag{10.3.4}$$

Yeast in Hydrogen Peroxide

In the second process, you will put baker's yeast in a hydrogen peroxide solution. Yeast has an enzyme called catalase that breaks down hydrogen peroxide into water and oxygen.

$$\mathrm{H_2O_2(aq)} \xrightarrow{\text{catalase}} \mathrm{H_2O(l) + O_2(g)} \tag{10.3.5}$$

[1] This section adapted from work Paul R. Young. ChemistryOnline.com Chapter 5.1. Accessed May 3, 2024.

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10.4: Procedure

Baking Soda (Sodium Bicarbonate)

1. Measure 50 mL of water into a 250 mL beaker or cup. Record the volume water in the first data table.
2. Measure 10 g of baking soda into a weigh boat or a small piece of wax paper or clean muffin cup. Record the brand used and the mass in the first data table.
3. Place your thermometer into the water. Record the temperature of the water as timepoint 0 in the second data table.
4. Carefully pour the baking soda into the water and start the stopwatch.
5. **Gently** stir the baking soda and water with the thermometer.
6. Record the temperature every 10 seconds in the second data table.
7. Graph your time and temperature data in Excel. Include your graph in your lab report.
8. Dispose of the baking soda solution down the drain. Clean out the beaker/cup and dry it for the next part of the experiment

Yeast

1. Measure 50 mL of hydrogen peroxide into the 250 mL beaker or cup. Record the volume in the first data table.
2. Measure 7 g of yeast into a weigh boat or a small piece of wax paper or clean muffin cup. Record the yeast type and brand used along with the mass of yeast in the first data table.
3. Place your thermometer into the hydrogen peroxide solution. Record the temperature of the water as timepoint 0 in the second data table.
4. Carefully pour the yeast into the water and start the stopwatch.
5. **Gently** stir the mixture with the thermometer.
6. Record the temperature every 10 seconds in the second data table.
7. Graph your time and temperature data in Excel. Include your graph in your lab report.

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10.5: Data and Results

Table **10.5.1** – Substance Data

Substance	Brand/Type	Liquid Volume (mL)	Mass (g)
Baking Soda			
Yeast			

Table **10.5.2** – Temperature and Time Data

Time (s)	Baking Soda Solution Temperature (°C)	Hydrogen Peroxide and Yeast Solution Temperature (°C)
0		
10		
20		
30		
40		
50		
60		
70		
80		
90		

Graph of Baking Soda Process

Use Excel or another similar graphing program to make a graph of temperature over time about the baking soda process. Use your data from table **10.5.2**.

Graph of Yeast with Hydrogen Peroxide

Use Excel or another similar graphing program to make a graph of temperature over time about the yeast process. Use your data from table **10.5.2**.

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10.6: Laboratory Questions

1. Which if any of the processes were chemical reactions? How do you know?
2. Which if any of the processes were physical changes? How do you know?
3. Was either process exothermic? Explain.
4. Was either process endothermic? Explain.
5. Did you have any challenges with this experiment? Explain.

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10.7: Reference

Paul R. Young. ChemistryOnline.com Chapter 5.1. Accessed May 3, 2024.

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CHAPTER OVERVIEW

11: Copper and Aluminum Electrochemistry

11.1: Safety
11.2: Supplies
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11.6: Lab Questions
11.7: References

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11.1: Safety

Vinegar and salt can both irritate your eyes. Goggles will minimize the risk of these substances entering your eyes. Do not eat or drink the experiment. Be careful to dispose of any solids in the trash and flush liquids down the drain with water.

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11.2: Supplies

- Pocket balance
- Clear Plastic cups
- 50 mL graduated cylinder
- Goggles and apron
- Tweezers
- Weigh boat
- Vinegar (150 mL)
- Table salt (21 g)
- 12 pennies with no deep scratches, made in 1983 or late
- Paper towels
- Aluminum foil

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11.3: Background

History of Copper and Aluminum

Copper metal has many uses, including in plumbing pipes, electrical wires, and statues. Over time, when exposed to air, copper will oxidize to copper(II) oxide, which is black. In the presence of water, copper oxide can further react to generate a green patina called verdigris. Verdigris contains several different copper containing compounds, depending on the environment.[1]

Aluminum is a common component of the earth's crust. However, it does not exist in nature as aluminum metal, but rather in minerals. Early methods of isolating aluminum were expensive – Napoleon used aluminum utensils for special guests. The less favored guest has gold cutlery instead.[2] In the later 1800s, it was discovered that aluminum metal could be generated by running electricity through molten alumina. This vastly reduced the cost of aluminum metal.

Electrochemistry

Metals give up their electrons with varying amounts of ease. The reactivity of different metals has been studied and the metals have been put in an activity series. The lower a compound is on the activity series, the less reactive it is.

Another way to determine the activity of a compound is to measure change in voltage across a galvanic cell relative to a standard electrode. The basic set up of a galvanic cell is depicted in Figure **11.3.1**. The standard conditions required for measuring reduction potentials are as follows:

- Use the standard hydrogen electrode
- Pressure is 1 atm
- Temperature is 298 K
- Solutions are 1 M in concentration

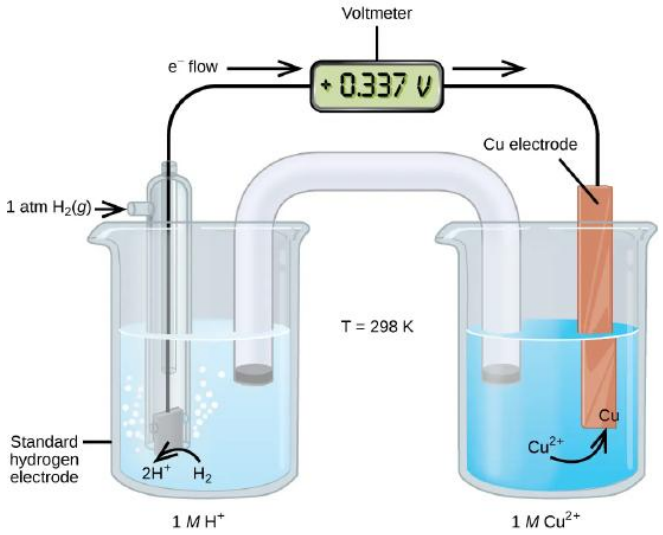


Figure **11.3.1** Galvanic Cell with Standard Hydrogen Electrode Created by: OpenStax. CC-BY 4.0. This image is available for free at openstax.com.

Drawing of a Galvanic Cell with hydrogen and copper electrode. Electrons flow from hydrogen to copper. Voltage reads +0.337 V

The voltage is always measured as the difference in potential between two electrodes. Thus, there is not an absolute value for any metal's reduction potential. Rather, the reduction of a chemical is always measured relative to another chemical, hydrogen. The standard reduction potential is also called the standard electrode potential. It is important to note that this data always assumes the reduction or gaining of electrons. If an oxidation occurs, the reverse reaction is occurring and the sign on the standard reduction potential must be reversed as well. For example, Figure **11.3.1** shows copper(II) ions from solution being reduced to copper metal on the electrode with a standard potential of +0.337 V. If the oxidation of copper metal to copper(II) ions occurs, the standard potential for this oxidation would be -0.337 V. We can depict these reactions as follows:

$$\mathrm{Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)} \quad E = +0.337V \quad (reduction) \tag{11.3.1}$$

$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^- \quad E = -0.337V \quad (oxidation) \tag{11.3.2}$$

Standard potential data can be used to predict the cell potential between two chemicals via the following reaction:

$$E_{cell} = E°reduced\ chemical - E°oxidized\ chemical \tag{11.3.3}$$

Because oxidation or reduction can happen multiple times at an electrode, **do not** use multiples of standard reduction potentials. Remember that the standard reduction potential is a relative number, not an absolute number. A positive E_{cell} value means the reaction will occur spontaneously. A negative E_{cell} value means the reaction will not be spontaneous.

Select Standard Reduction Potential Data[3]

Reduction Reaction	E° (V)
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	-1.662
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.337
$Cu(OH)_2(aq) + 2e^- \rightarrow Cu(s) + 2OH^-(aq)$	-0.222
$2H^+(aq) + 2e^- \rightarrow 2H_2(g)$	0.0000
$O_2(g) + 2H^+(aq) + 2e^- \rightarrow 2H_2O(l)$	+1.229

Acid-Base Chemistry

Along with electrochemistry occurring in this experiment, acid-base chemistry may also be occurring. Vinegar contains acetic acid (CH_3COOH). Acetic acid can react with a base to form acetate ions (CH_3COO^-). Acetate ions are very water soluble, regardless of the metal cation.

Experimental Design

When making observations about differences between different systems, it is important to change only one variable at a time. This makes it easier to determine which variables impact experiment. If multiple variables are changed, replicating the results can be challenging.

Sometimes, scientists are not even aware of all the variables. For example, iron(III) chloride was thought to catalyze reactions to carbon-nitrogen bonds and carbon-oxygen bonds. Later, chemists discovered that there were copper impurities in the iron. The copper was the actual catalyst, not the iron.[4] If iron that did not contain the copper impurity was used, the reaction did not occur.

As you conduct your experiment, consider the experimental design. Why are you doing the steps in the order you are? Where could you change the experiment?

[1] Fitzgerald, K. P.; Nairn, J.; Atrens, A. The Chemistry of Copper Patination. *Corrosion Science* **1998**, *40* (12), 2029–2050. https://doi.org/10.1016/S0010-938X(98)00093-6

[2] A Short History Of Aluminum, From Precious Metal To Beer Can : Short Wave : NPR by Scott Neuman. Dec. 10, 2019.

[3] Lide, David R. *Ed.* CRC Handbook of Chemistry and Physics. 77th ed. 1996.

[4] Buchwald, S. and Bolm, C. (2009), On the Role of Metal Contaminants in Catalyses with $FeCl_3$. Angewandte Chemie International Edition, 48: 5586-5587. https://doi.org/10.1002/anie.200902237

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11.4: Procedure

NOTE: There are two different procedures listed in steps 15 and 16. Please only conduct the procedure you have been assigned. The class will then compare data and see if the difference in procedure had any effect on the outcome of the experiment.

1. Measure 7 g of table salt, sodium chloride, into a plastic cup. Record the actual mass in the first data table.
2. Add 50 mL of white vinegar into the cup. Record the volume in the first data table.
3. Stir the salt and vinegar until the salt is dissolved.
4. Add 12 pennies (made after 1983) and stir for 5 minutes.
5. Place paper towel on two plastic weigh boats or a plastic lid. Carefully wet the paper towel with some of the vinegar/salt solution.
6. Use tweezers/forceps to remove the pennies from solution. Lay the pennies onto the paper towel. Make sure there is no overlap of the pennies.
7. Leave the pennies to sit for 24 to 48 hours. Record the time in the data section.
8. Take four clear plastic cups. It is easier to make observations if you have clear cups. Label one "water", one "water+salt", one "vinegar" and one "vinegar+salt."
9. Add 7 g of table salt to each cup that has salt in the label. Record the exact masses in the second data table.
10. Add 50 mL of water to each cup labeled water. Record the exact volumes of water in the second data table.
11. Add 50 mL of vinegar to each cup labeled vinegar. Record the exact volumes of vinegar in the second data table.
12. Stir the salt to dissolve in the cups that contain salt.
13. Cut aluminum foil into 2.0 cm by 2.0 cm squares. Place one square into each cup.
14. Record initial observations about the pennies, the aluminum foil, and the solutions into the appropriate data tables.
15. Procedure A: Add three pennies to each cup, making sure the aluminum foil and pennies **are** touching. Begin timing. Circle the word “are” in the first sentence in the data and results section.
16. Procedure B: Add three pennies to each cup, making sure the aluminum foil and pennies are **not** touching. Begin timing. Circle the words “are not” in the first sentence in the data and results section.
17. Record observations every 10 minutes for an hour for each cup, both the foil and the pennies.
18. Leave overnight (at least 16 hours). Record a final set of observations. **Be sure to indicate how many minutes the pennies have sat in solution.**

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11.5: Data and Results

Data and Results

In this experiment, the pennies **are/are not** touching the aluminum foil.

Data Table **11.5.1** : Initial Soak of 12 Pennies

Mass of salt (g)	Volume of vinegar (mL)	Time of soak (h)

Data Table **11.5.2** : Composition of 4 Different Solutions

Solution	Mass of Salt (g)	Volume of Liquid (mL)
Water	0	
Water + Salt		
Vinegar	0	
Vinegar + Salt		

Data Table **11.5.3** : Observations Of Changes in Pennies

Time (min)	Water	Water + Salt	Vinegar	Vinegar + Salt
0				
10				
20				
30				
40				
50				
60				
Long: _____				

Data Table **11.5.4** : Observations Of Changes in Aluminum Foil

Time (min)	Water	Water + Salt	Vinegar	Vinegar + Salt
0				
10				
20				
30				
40				
50				
60				
Long: _____				

Data Table **11.5.5** : Observations Of Changes in Solution

Time (min)	Water	Water + Salt	Vinegar	Vinegar + Salt
0				
10				
20				
30				
40				
50				
60				
Long: _____				

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11.6: Lab Questions

1. Copper metal has been used for pipes that transport water. Based on this use, would you predict copper metal to react much with water?
2. Which solution produced the largest change in the appearance of the pennies?
3. Which solution produced the largest change in the appearance of the aluminum foil?
4. Which solution had the largest change in the appearance of the liquid?
5. Did any of the results surprise you? Why or why not?
6. In this experiment, we will compare data among classmates. As each of you conducted these experiments at home, what are three variables that may be different in your experiment from your classmates even if you followed the same lab procedure?
7. What effect did solutions with salt have on the experiment? Why do you think this is?
8. What is one way you might change the experimental design?
9. Copper(II) oxide is a black solid that can react with acetic acid to form copper(II) acetate and water. Write out the chemical equation for this reaction. Did this reaction occur in any step of your experiment? Why or why not?
10. Copper(II) hydroxide is a blue solid that is insoluble in water. Did this form in any of your reactions? What evidence do you have?
11. Copper(II) acetate is a blue-green solid that dissolves in water. Did this form in any steps of your experiment? Why or why not?
12. Calculate the cell potential for the reaction of copper(II) ions with aluminum metal. Show your work. Is the process spontaneous? Why or why not?
13. Hydroxide ion is a strong base. Do you predict it to remain in solution with acetic acid? Why or why not?

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11.7: References

(1) Buchwald, S.; Bolm, C. On the Role of Metal Contaminants in Catalyses with $FeCl_3$. *Angewandte Chemie* **2009**, *48* (31), 5586–5587. https://doi.org/10.1002/anie.200902237.

(2) Fitzgerald, K. P.; Nairn, J.; Atrens, A. The Chemistry of Copper Patination. *Corrosion Science* **1998**, *40* (12), 2029–2050. https://doi.org/10.1016/S0010-938X(98)00093-6.

(3) Flowers, P.; Theopold, K.; Langley, R.; Robinson, W. R. *17.3 Electrode and Cell Potentials - Chemistry 2e | OpenStax*. openstax.org. https://openstax.org/books/chemistry-2e/pages/17-3-electrode-and-cell-potentials (accessed 2024-05-23).

(4) Lide, D. R. *CRC Handbook of Chemistry and Physics*, 77th ed.; CRC Press: Baco Raton, 1996.

(5) Neuman, S. *Aluminum's Strange Journey from Precious Metal to Beer Can : Short Wave*. NPR.org. https://www.npr.org/2019/12/05/785099705/aluminums-strange-journey-from-precious-metal-to-beer-can (accessed 2024-05-23).

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CHAPTER OVERVIEW

12: Edible Half-Life

12.1: Supplies
12.2: Background
12.3: Procedure
12.4: Data and Results
12.5: Laboratory Questions
12.6: References

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12.1: Supplies

- 100 M&Ms or Skittles (Pennies also work, but don’t eat!)
  - All 100 objects must be the same size
- Cup
- Rimmed baking sheet, tray or plate

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12.2: Background

The half-life a radioactive element is the amount of time required for one half of the original material to decay. The rate of decay is different for every radioactive isotope.[1] Some decay very quickly, such as radium-215, which has a half-life of 1.55 ms. Others take a very long time to decay, such as bismuth-209, which has a half-life of $1.90*10^{19}$ years!

In this experiment, you will investigate exponential decay. However, given that using radioactive material at home would be unsafe, you will use candy as a model.

[1] Information about half-lives of different isotopes can be found at Ptable.com - Isotopes - Count.

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12.3: Procedure

1. Count out 100 M&Ms or Skittles. (Any round candy with a letter on one side and blank on the other will work.)
2. Place them in the cup.
3. Gently shake the cup to mix up the candy.
4. Pour the candy out onto the tray. Remove any candy where you can see the letter and set it aside. (If using pennies, remove all the pennies with heads showing.)
5. Count the pieces of candy that are left on the tray. Record this in the data table for Round 1. Put this candy back in the cup.
6. Repeat sets 3 through 5 for 5 more rounds. Enter the data in the corresponding round in the data table.
7. After 6 rounds, repeat steps 1 through 6 for Trials 2 and 3.
8. In Excel, calculate the average of all three trial for each round using the "=AVERAGE" function. You will need to select the data cells to be averaged. Report this in the "Average Amount Remaining" column of the data table.
9. Using Excel, make a scatter plot without dots connected of the average data set with the "Round" on the x-axis and the "amount remaining" on the y-axis. Fit an exponential trend line to your data. Display the equation and R^2 value on your graph.

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12.4: Data and Results

Data Table **12.4.1** – Amount of Candy Remaining After Each Round

Round	Trial 1 Amount Remaining	Trial 2 Amount Remaining	Trial 3 Amount Remaining	Average Amount Remaining
0	100	100	100	100
1				
2				
3				
4				
5				
6				

Graph – Amount of Candy Remaining After Each Round

Make a graph of the average amount of candy versus the round and include it in your lab report.

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12.5: Laboratory Questions

1. For each trial, what percent of the initial material was remaining after 6 rounds of decay? What was the average percentage of initial material that remained after 6 round of decay?
2. For medical imaging, there can be two issues with radiotracers. The half-life could be too short or too long. Why would a half-life that is too short be an issue for medical imaging?
3. What issues would arise from a long half-life for radiotracers in medical imaging?
4. Read the nuclear waste fact sheet from the U.S. Nuclear Regulatory Commission. What impact does half-life have on considering how to handle disposal of radioactive waste?

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12.6: References

Dayah, M. *Periodic Table - Ptable*. ptable.com. https://ptable.com/#Isotopes/Count (accessed 2024-05-24).

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