# Chemistry Matters
## Second Edition
## Lab Manual

Oak Meadow

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## Unit 2: Molecular Structure and Properties



## Unit 3: Phase Changes and Behavior of Gases

## Unit 4: Stoichiometry, Solution Chemistry, and Acids and Bases



## Unit 5: Energy, Thermodynamics, and Oxidation-Reduction

# Introduction

Chemistry is the study of matter, and the best way to learn chemistry is through active explorations. The activities and experiments included in this lab manual are a required part of Oak Meadow's Chemistry Matters high school course. They can also be used as a supplement to any chemistry course.

A complete materials list for the lab manual activities is found in the appendix. In addition, you will need a box of disposable nitrile gloves to use during all inquiry and laboratory experiments.

Some of the activities in this lab manual use the textbook *Living by Chemistry* (W. H. Freeman, 2018).

Except for those in lesson 18, all the experiments are from Hands-On Labs (HOL) and use the materials in the following lab kits:

**HOL Chem 1 Kit (HOL SP-3005-CK-02)**

**HOL Chem 2 Kit (HOL SP-3006-CK-02)**

Below is a breakdown of the labs in each kit and the corresponding lesson. Some labs are optional, as noted in the table of contents.

### Chem 1 Kit (HOL SP-3005-CK-02)

| Lab Name | Corresponding Lesson |
|---|---|
| Laboratory Techniques and Measurements | 1 and 15 |
| Molecular Modeling and Lewis Structures | 7 |
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### Chem 2 Kit (HOL SP-3006-CK-02)

| Lab Name | Corresponding Lesson |
|---|---|
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Note: Working with chemicals can be dangerous. Please carefully follow all necessary safety protocols.

## Review of Laboratory Safety

Throughout the course, you will conduct home experiments that involve the use of potentially harmful chemicals. These labs are designed to give you a hands-on learning experience, but they need to be done in a safe manner. As such, it is important that we begin this course with a review of safe laboratory techniques.

1. Wear closed-toe shoes when working with chemicals.
2. Keep all nonessential lab items (such as bags, papers, food, cosmetics, lotions, etc.) out of the work space.
3. Never eat, drink, or chew gum when working on labs.
4. Before every lab, read all the directions carefully. Make sure you understand the overall goal of the lab.
5. Check that all equipment and supplies are clean and in working order before beginning.
6. Gather all the equipment needed for the lab. Keep all other lab materials packaged and out of the work space.
7. Always wear safety glasses and gloves. (The glasses are provided in your lab kit.)
8. Tie back long hair and loose clothing to keep them away from chemicals and flames.
9. Remove dangling jewelry.
10. Never touch, taste, or smell any chemical. To note odor, gently wave your hand over the opening of the container to direct the fumes toward your nose and smell carefully (wafting).

11. Never conduct your own experiments. Follow the directions provided, and use materials only as intended.
12. Hot glassware does not appear hot. Carefully check before touching.
13. Dispose of any unused or spilled chemicals by soaking them up with a paper towel and placing it in a trash can. **Never dispose of chemicals down the sink or toilet.**
14. Clean up your work space and all equipment after each experiment. Dispose of materials as noted above, or place them back in your lab kit for future use. Since you are conducting these experiments at home (and possibly in your kitchen), it is critically important that you clean up your work space before anyone else uses the area or food is prepared.
15. Wash your hands thoroughly after each experiment!

If at any point during this course you have questions about the laboratory directions or need assistance, stop working and contact your teacher immediately.

Lesson 1

# Introduction to Chemistry and Matter

## Activities

Complete Activity A. Activity B is optional and can be done for extra credit.

- Activity A: Measuring and Graphing
- Activity B: Understanding Material Safety Data Sheets

### Activity A: Measuring and Graphing

The purpose of this activity is to review and reinforce basic mathematical skills relevant to chemistry: making accurate measurements, using scientific notation, and using graphs for making inferences about a data set.

When performing calculations, always show your work. This applies for every activity and experiment throughout this course.

#### Materials

- scale (included in the HOL lab kit)
- scientific calculator
- graph paper (included in the appendix)
- ruler
- pencil
- cardboard or stock paper
- scissors

#### Before You Begin

1. Write one paragraph explaining the following. Properly cite any references you use.
    - Describe how to create a meaningful line graph. What titles, labels, and notations would you include?
    - What is the general equation for a line? Define all variables.

- What is the equation for determining the slope of a line?
- Optional extra credit: Define and explain interpolation and extrapolation as it relates to the slope of a line.

### Procedure, Part One: Data Collection

1. Using cardboard or thick paper, cut out four squares. Each square must be a different size. Label the squares A, B, C, and D.
2. Using the same cardboard or thick paper, cut out one irregular shape (such as a star, heart, squiggle, etc.). Label this shape E.
3. Measure the length and width of the cardboard squares (A, B, C, and D) to the nearest 0.1 cm. Do not attempt to measure the length and width of the irregular shape.
4. Convert each measurement to scientific notation (SN).
5. Calculate the area of each square using regular notation (RN) and scientific notation. Try to recall the formula for finding the area of a square. Look it up if you can't remember.
6. Using the electronic balance (scale), determine the mass of each square and the irregular shape E to the nearest 0.01 grams.

**Data Table: Measurement of Shapes**

| Shape | Length (RN) (cm) | Length (SN) (cm) | Width (RN) (cm) | Width (SN) (cm) | Area (RN) ($cm^2$) | Area (SN) ($cm^2$) | Mass (g) |
|---|---|---|---|---|---|---|---|
| A | | | | | | | |
| B | | | | | | | |
| C | | | | | | | |
| D | | | | | | | |
| E | | | | | | | |

### Procedure, Part Two: Graphing Data to Produce a Line of Best Fit

Create your graph, either by hand using graph paper or digitally using Excel, Google Sheets, or another online graphing tool.

1. Graph the relationship between mass (g) and area ($cm^2$) for squares A, B, C, and D.
2. Place mass on the *y*-axis and area on the *x*-axis. Remember to include the appropriate units when labeling your axes. Be sure to give your graph a descriptive title.
3. Draw a line of best fit. Your line of best fit must, by definition, come as close to as many points as possible and originate at the coordinate (0,0).

#### Graphing the Line of Best Fit

Need help with the line of best fit? Check out these videos:

"Line of Best Fit Equation"

www.youtube.com/watch?v=DmGLQkUm-4g

"Excel Basics—Linear Regression—Finding Slope & Y Intercept"

www.youtube.com/watch?v=KwQsV77bYDY

"Linear Regression T184 (Line of Best Fit)"

www.youtube.com/watch?v=0as2Jh_eDwg

Note: The last video shows how to use a graphing calculator. A graphing calculator is not required for this course, but you may have one from your previous math courses and already be familiar with using it. The last video is included for those students who prefer working by hand and like using a graphing calculator to support their results.

### Analysis

1. Using your graph, calculate the slope of your line of best fit. Show your work.

2. Using the constructed graph, determine the area of the irregular shape E. Include the correct units in your answer.
    a. Area of shape E = ________________
    b. Was this task completed through interpolation or extrapolation? Explain.

3. How was this graph useful when obtaining information? Was it more accurate than trying to make direct measurements? Explain.

## Activity B: Understanding Material Safety Data Sheets

(This activity is optional and can be done for extra credit.)

Material Safety Data Sheets (MSDS) are required by OSHA (Occupational Safety and Health Administration) to be maintained by any facility storing or using chemicals of any nature or quantity. Chances are you will, at some point, work at a facility that stores hazardous materials. You should know how to find information on chemicals and understand an MSDS.

Review OSHA's MSDS requirements. They have standardized the information that must be available on MSDSs and made that information publicly available here:

> "Hazard Communication Standard: Safety Data Sheets"
>
> www.osha.gov/Publications/OSHA3514.html

Next, review the MSDS linked below for the concentrated version of vinegar, a common household item, and use this to answer the following questions.

> "Flinn Scientific Safety Data Sheet (SDS): Acetic Acid Solution 4.3 M–6 M"
>
> www.flinnsci.com/sds_5.2-acetic-acid-solution-4.3-m---6-m/sds_5.2/

1. What does MSDS stand for? ________________________________
2. What is the name of your chemical? ________________________________
3. What is its formula? ________________________________
4. Identify three important types of information found on an MSDS and provide that information for this compound. ________________________________

    ________________________________

# Experiments

Complete the following lab experiment.

## Laboratory Techniques and Measurements

The materials for this experiment are found in the HOL Chem 1 Kit.

Note: You will be completing Exercise 1: Length, Temperature, and Mass and Exercise 2: Volume and Density in this lesson. Exercise 3: Concentration, Solution, and Dilution will be covered in lesson 15. Include photos of your lab setup and results. Complete all the required lab questions. You will generate data tables to be included with your lab.

# Laboratory Techniques and Measurements

## Learning Objectives

- Perform measurements with a graduated cylinder, volumetric flask, graduated pipet, ruler, digital scale, and thermometer.
- Perform the water displacement method for measuring the volume of an irregularly shaped object.
- Calculate experimental error.
- Practice basic math and graphing skills.

## Exploration

### Explore International System of Units (SI)

Chemistry is the science of matter and its changes. In order to study matter and its changes, scientists make qualitative and quantitative observations. Quantitative observations, or measurements, always consist of a numerical value and a unit of measurement. Scientists use the International System of Units (SI), which is derived from the metric system, as the standard system of measurement. Table 1 shows the basic SI units to measure the five fundamental properties of length, mass, time, temperature, and amount of substance. The sixth property shown, volume, is considered a derived unit, as it is the cubic (3D) version of length. As shown in Table 1, the SI system and the metric system are closely related, differing only in the size of the fundamental unit. Table 1 also includes the units more likely to be used in the laboratory, where very small quantities are used.

| Measurement | International System (SI) | Metric System | Common Laboratory Units |
|---|---|---|---|
| **Length** | meter (m) | meter (m) | centimeter (cm) |
| **Mass** | kilogram (kg) | gram (g) | gram (g) |
| **Time** | second (s) | second (s) | second (s) |
| **Temperature** | Kelvin (K) | Celsius (°C) | Celsius (°C) |
| **Amount of Substance** | mole (mol) | mole (mol) | mmole (mmol) |
| **Volume** | cubic meter ($m^3$) | liter (L) | milliliter (mL) |

**Table 1. Comparison of units used to measure fundamental properties**

US customary units require conversion to metric/SI units for scientific reporting. Table 2 illustrates the relationships (conversion factors) between US customary units and metric/SI units.

| Property | Factor | Factor |
|---|---|---|
| **Length** | 1 in = 2.54 cm | 1 mi = 1.609 km |
| **Mass** | 1 lb = 454 g | 1 kg = 2.206 lb |
| **Volume** | 1 qt = 0.946 L | 1 L = 1.06 qt |
| **Temperature** | °F = (°C × 1.8) + 32 | K = °C + 273.15 |

**Table 2. Metric-US conversions**

Table 3 lists the meanings of the prefixes used in the metric and SI systems. Each prefix is a multiplication factor for the base unit. For example, the prefix *kilo* means 1,000, so 1 kilogram is equivalent to 1,000 grams.

| Prefix | Symbol | Meaning | Exponential Notation |
|---|---|---|---|
| **mega** | M | 1,000,000 | $10^6$ |
| **kilo** | k | 1,000 | $10^3$ |
| **hecto** | h | 100 | $10^2$ |
| **deka** | da | 10 | $10^1$ |
| **deci** | d | 0.1 | $10^{-1}$ |
| **centi** | c | 0.01 | $10^{-2}$ |
| **mili** | m | 0.001 | $10^{-3}$ |
| **micro** | µ | 0.000001 | $10^{-6}$ |

**Table 3. Prefixes used in the SI and metric systems**

## Explore Length

Length is defined as the distance (amount of space) of an object from end to end. The SI system unit of length is the meter (m) and was originally intended to represent one ten-millionth of the distance between the North Pole and the Equator. However, over time the definition of the meter changed. The current definition, which has been in place since 1983, is that a meter equals the distance that light travels in a vacuum in $\frac{1}{299{,}792{,}458}$ seconds. There are many tools to measure length, including a caliper, a calibrated ruler, a tape measure, and even a laser. See Figure 1. Each of these different measuring devices measures length to a different degree of accuracy and precision.

**Figure 1. Common equipment used to measure length. A. Vernier caliper B. Tape measure C. Calibrated ruler. (Image A copyright Paul Paladin, 2013. Image B copyright Jiri Hera, 2013. Image C copyright Quang Ho, 2013. All images used under license from Shutterstock.com.)**

### Did You Know . . . ?

A ruler is a tool used to measure length. However, a ruler (rule) is actually defined as an instrument used to rule (create) straight lines. A calibrated ruler is a ruler that contains measurements to measure length along a straight line.

## Explore Temperature and Time

Temperature is defined as a measure of the average kinetic energy of a system. The SI system unit of temperature is the kelvin (K), and the standard value (0 K) is defined as $\frac{1}{273.16}$ of the temperature when water exists as a solid, liquid, and gas at 1 atmosphere of pressure. This point in which water co-exists as a solid, a liquid, and a gas is called the triple point. While the SI unit for temperature is the kelvin, the majority of thermometers are calibrated to degrees Celsius (°C or C) and/or degrees Fahrenheit

(°F or F). Today, scientific measurements are commonly taken in the Celsius scale. Converting between Fahrenheit, Celsius, and Kelvin scales is common and can be performed with one of the following conversion formulas:

$$T_K = T_C + 273.15$$

$$T_F = (1.8)T_C + 32$$

$$T_C = \frac{T_F - 32}{1.8}$$

A comparison of common temperatures between the three different scales (K, °C, °F) is shown in Figure 2.

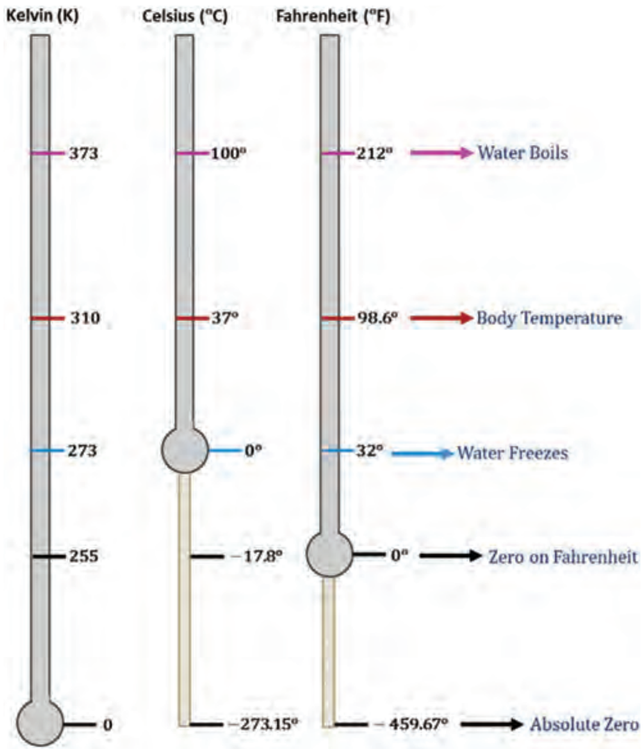



**Figure 2. Comparison of common temperatures on the three different scales**

The second (s) is the basic SI unit of time and is defined as the duration of 9,192,631,770 cycles of radiation in an energy level change of the cesium atom. While smaller quantities of time than the second are described using the standard SI prefixes (Table 1), quantities of time larger than a second (minute, hour, day, week) are not SI units. Time is measured using watches (clocks), which are calibrated to atomic clocks. Atomic clocks are extremely precise clocks that are regulated by the vibrations (resonance frequencies) of cesium atoms.

## Explore Volume

Volume is defined as the amount of space occupied by a three-dimensional object or area of space. The SI unit for volume is the cubic meter ($m^3$), which is equal to the volume of a cube measuring 1 meter on each side (1 m × 1 m × 1 m). See Figure 3. The units for volume most commonly used in science laboratories include the liter (L) and the milliliter (mL). A liter is equal to the volume of a cube, with sides of 0.1 m (0.1 m × 0.1 m × 0.1 m = 0.001 $m^3$ = 1 L). A milliliter is $\frac{1}{1,000}$ of a liter, and it is also equal to a cubic centimeter (1 mL = 1 $cm^3$).

$$1 \text{ L} = 1,000 \text{ mL} = 1,000 \text{ cm}^3$$

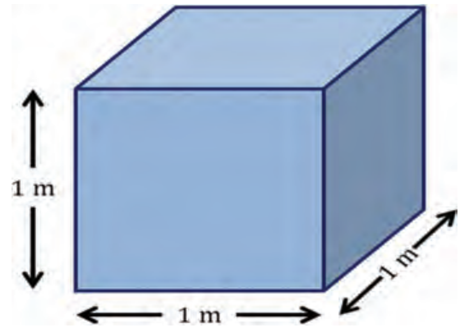



**Figure 3. Volume of a cube. A cube that is 1 meter in length on all sides has a volume of 1 $m^3$ or 1,000 L.**

In addition to the milliliter, the microliter (µL), which is one millionth of a liter, is also a common unit of volume used by scientists. When measuring the volume of a liquid, common laboratory equipment options include a graduated cylinder, volumetric flask, and graduated pipet.

## Explore Graduated Cylinder, Meniscus, Volumetric Flask, and Graduated Pipet


**Figure 4. Graduated cylinder. A 50-mL graduated cylinder made of glass. Note the markings along the cylinder, which represent milliliters.**

A graduated cylinder is a slender container that is calibrated by specific volumetric amounts, such as milliliters or liters, and can measure a range of volumes depending on its capacity. See Figure 4.

When reading a graduated cylinder made of glass, or any measuring device for volume that is made of glass, it is important to read the volume, at eye level, from the bottom of the meniscus. A meniscus is the curve that forms between the liquid and the surface of the container as the result of surface tension, cohesion, and adhesion. See Figure 5.

While graduated cylinders are designed to measure a variety of volumes, a volumetric flask is calibrated to measure only one volume, and is often used to prepare a specific volume of solution. A volumetric flask has a bulb-shaped bottom and a very long slender neck. Each flask is calibrated with a mark on that slender neck, allowing for very careful and accurate measurements. See Figure 6.

**Figure 6. Volumetric flask, 25 mL. The line on the flask, as noted by the black arrow, marks 25 mL calibrated measurement.**

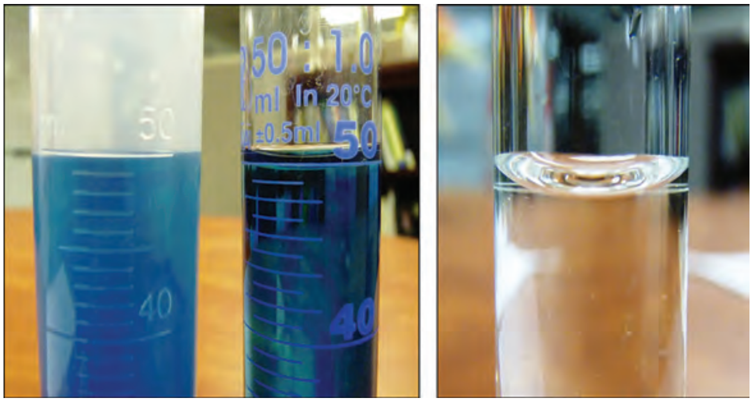



**Figure 5. Meniscuses. A. Colored water in a graduated cylinder made of plastic (left). Note that the liquid forms a straight, non-curved, line at the 50 mL mark. Colored water in a graduated cylinder made of glass (right). Note that the liquid forms a curved line. B. When reading the volume from a meniscus, the volume is read from the bottom of the curve.**

For smaller volumes, scientists use graduated pipets (also referred to as serological pipets) to measure liquids. Just as for graduated cylinders, graduated pipets can be used to measure a range of volumes depending on the capacity of the pipet. Similar to volumetric flasks, the calibrated measurement area of graduated pipets is very narrow, allowing for very precise measurements. Graduated pipets use a suction mechanism to fill and release liquid. See Figure 7. A pipet is designed to dispense or deliver a specific volume of solution to a container.

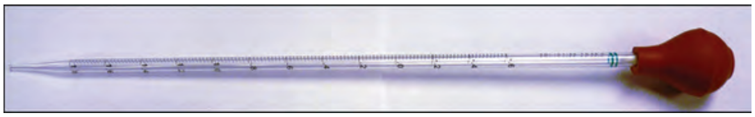


**Figure 7. Graduated pipet. Shown is a 2 mL graduated pipet, calibrated to 0.1 mL. The red bulb on the right end of the pipet is used to fill and release the liquid from the pipet.**

The order of markings on a graduated pipet is opposite from the order of markings on a graduated cylinder. When a 50 mL graduated cylinder contains 50 mL of liquid, the liquid is at the 50 mL marking. However, when using a 2 mL graduated pipet, when liquid reaches the marking labeled “0,” it means that the pipet will deliver 2 mL of liquid. There are many different types of graduated pipets, and each has their own mechanism for dispensing liquid. For the graduated pipet that will be used in this experiment, to dispense the full 2 mL from the pipet the liquid is completely released from the pipet. Likewise, to dispense only 1 mL from the pipet the liquid is released from the “0” marking to the “1.0” marking. See Figure 8.

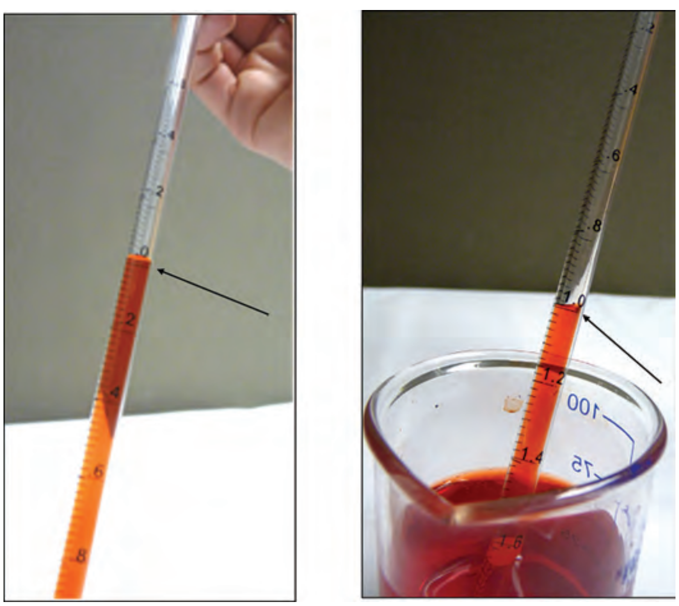


**Figure 8. Measuring with a graduated pipet. Notice the “0” and “1.0” markings, both noted with black arrows.**

Just as with graduated cylinders, graduated pipets are available in a variety of sizes, allowing for measurements as small as 0.1 µL.

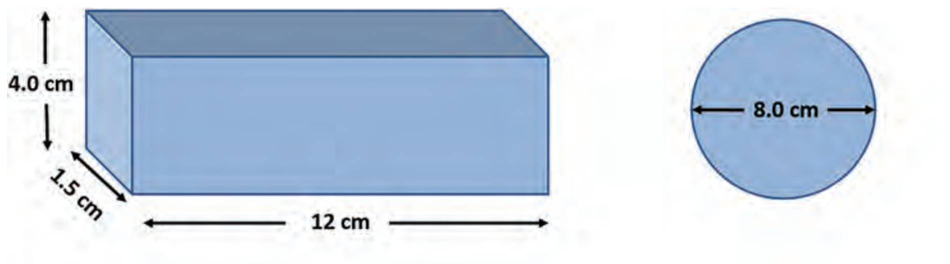
 

### Did You Know . . . ?

The majority of liquids in glass will form a concave-shaped meniscus. However, some liquids, such as mercury, will form a convex meniscus. In the case of a convex meniscus, the volume is read from the top of the curve rather than the bottom.

## Explore Measuring the Volume of a Solid

There are multiple techniques to measure the volume of a solid, three of which are discussed here. For a solid with defined edges, such as a cube, a box, or a sphere, the length, width, and height or the radius of the object can be measured and used to calculate volume. See Figure 9.




Volume = 4.0 cm × 12 cm × 1.5 cm
Volume = 72 $cm^3$ or 72 mL

Volume = $\left(\frac{4}{3}\right)\pi r^3 = \left(\frac{4}{3}\right)\pi(4.0\ cm^3)$
Volume = 270 $cm^3$ or 270 mL

**Figure 9. Calculating the volume of solid objects with defined edges**

When measuring the volume of an irregularly shaped object, there are two common methods: the water displacement method and Archimedes's method. In the water displacement method, the irregularly shaped object is placed into a known amount of water in a graduated cylinder. The increase in water volume, as measured by the graduated cylinder, is equal to the volume of the irregularly shaped object. See Figure 10.


**Figure 10. A. Irregularly shaped object. B. Graduated cylinder with known amount of water. C. Graduated cylinder with irregularly shaped object. The difference (increase) in volume from 30.0 mL to 32.5 mL is equal to the volume of the irregularly shaped object.**

Archimedes's method incorporates buoyancy into the water displacement method. Buoyancy is the upward force that a fluid places onto an object, which is equal to the weight of the displaced fluid. Put simply, when an object is submerged and suspended in water (1 mL of water = 1 gram of water), the change in mass as the result of the object being submerged is equal to the volume of the object. See Figure 11.

**Figure 11. Archimedes's method. A. Water in a beaker on a tared scale. B. Irregularly shaped object is submerged in the water, creating a buoyant force equal to 2.5 grams. As 1 mL of water has a mass of 1 gram, the volume of displaced water is 2.5 mL, which is equal to the volume of the irregularly shaped object.**

### Did You Know . . . ?

Archimedes, a Greek mathematician and scientist, was asked by King Hiero II to determine if a crown made for him was composed of pure gold or contained a mixture of gold and silver. As Archimedes could not damage the crown, he had to find a way to determine its composition. It is from this challenge that he developed what is now known as Archimedes's Principle. Knowing that density is a physical property that does not change, Archimedes could determine if the density of the crown was equal to the density of pure gold. As the densities of the crown and pure gold did not match, he concluded that the crown was not made from pure gold.

## Explore Mass and Density

Mass is defined as the measure of the amount of matter contained in a physical body. The SI unit of mass is the kilogram (kg) and is standardized (equal) to the mass of the International Prototype Kilogram (IPK). The IPK is composed of a platinum-iridium alloy and is stored at the International Bureau of Weights and Measurements in Sevres, France. See Figure 12.

It is important to note that the terms *mass* and *weight* are often used interchangeably but are not the same thing. Mass is a quantifiable measure of matter, while weight refers to the gravitational force of attraction exerted on an object. The SI system uses mass measurements, yet the verb *weigh* is the common verb used to describe or obtain the mass of an object. In the laboratory, scientists usually work with the gram (g), which represents one thousandth of a kilogram, and the milligram (mg), which equals one thousandth of a gram. Mass is measured using an instrument called a balance, which is commonly referred to as a scale.

**Figure 12. International Prototype Kilogram. The IPK is kept under two bell jars. (Image courtesy of the U.S. Federal Government)**

Density is defined as mass per unit of measure, which is most often volume. It is a way to describe how heavy something is for its size. Like volume, density is not a fundamental SI unit of measure, but is derived from the SI units for mass and volume. The density of a liquid is usually reported as grams per milliliter (g/mL), while the density of a solid is usually reported as grams per cubic centimeter (g/cm$^3$). The density of water is 1 g/mL. Substances with a density greater than 1 g/mL will sink when placed in water, while objects with a density less than 1 g/mL will float when placed in water. The density of an object is determined with the following equation:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

## Explore Significant Figures and Percent Error

Significant figures are a combination of the certain and first uncertain digit of a measurement. The number of significant figures in a measurement is dependent on the precision of the instrument. Measurements may be estimated to one additional place beyond the markings on the device. This would be considered the one allowed, uncertain digit. For example, if a ruler is calibrated to the millimeter length, a measurement may be taken to a tenth of a millimeter. Likewise, if a ruler is calibrated to the centimeter length, a measurement may only be taken to the tenth of a centimeter. See Figure 13.

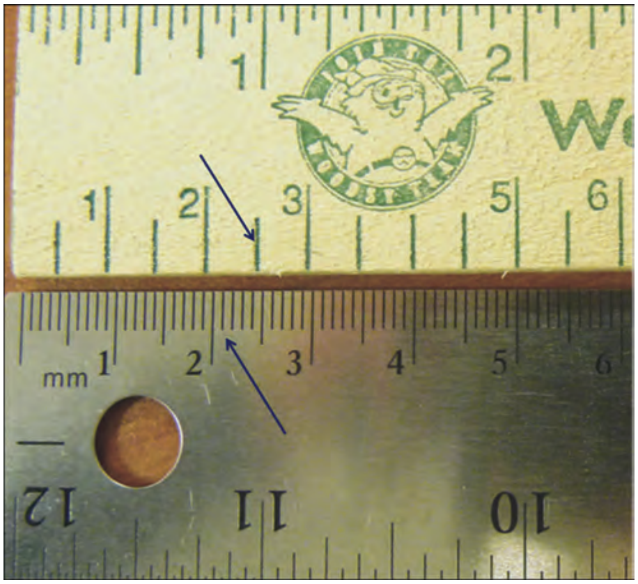



**Figure 13. Significant figures in measurements. The upper ruler is calibrated to 0.5 cm, whereas the bottom ruler is calibrated to 0.1 cm.**

The two rulers depicted in Figure 13 can be used to conduct measurements with different significant figures. The top ruler is calibrated to 0.5 cm. If a piece of string was measured, and its length was between 2 cm and 2.5 cm, the total length might be estimated as 2.3 cm. In this case, the ones place (2) is one certain digit, and the tenths place (.3) is the estimated, uncertain digit. The bottom ruler is calibrated to 0.1 cm (or 1 mm). The length of the same string could be reported as 2.25 cm. In this case, there are two certain digits: one in the ones place and one in the tenths place (2.2). The estimated, uncertain digit is in the hundredths place (.05). In general, measurements are reported one decimal place beyond the instrument's calibration.

Each time a particular measuring device is used, the scientist records the measurement to the same decimal place. That is, if a scientist is using a centimeter ruler similar to the bottom one in Figure 13, every measurement will end with a value in the hundredths column. The length of an item that is exactly 5 cm long must be recorded as 5.00 cm using that ruler. The zeros in the tenths and hundredths places indicate the precision of the measuring device. Remember, the last digit recorded is the uncertain one. If a student records a length of 5 cm, they are telling the instructor that the 5 is uncertain, and the actual length is somewhere between 4 and 6 centimeters. When a student records a length of 5.0 cm, the 0 is uncertain or estimated and the true length is somewhere between 4.9 and 5.1 centimeters. A measurement of 5.00 centimeters indicates that the 0 in the hundredths column is uncertain and the true length is between 4.99 and 5.01 centimeters.

Calculations should be consistent with the precision of measurements. When multiplying and dividing measurements, the measurement with the fewest number of significant figures determines the number of digits in the calculated result. For example:

$$2.36 \text{ cm} \times 3.1 \text{ cm} = 7.316 \text{ cm}^2$$

$$\text{Recorded as } 7.3 \text{ cm}^2$$

The product should be recorded as $7.3 \text{ cm}^2$ since there are only two significant figures in the least precise factor, 3.1 cm.

When adding or subtracting, the result of the calculation cannot be more precise, have more decimal places, than the least precise measurement:

$$22.55 \text{ g} + 75.3 \text{ g} + 13 \text{ g} = 110.85 \text{ g}$$

$$\text{Recorded as } 111 \text{ g}$$

The sum should be recorded as 111 g since the least precise number was precise to the full gram, 13 g.

Often measured results will differ from known values. Percent error (the difference between the accepted/true value and the value that was measured) may be calculated using the following formula:

$$\text{Percent Error (\%)} = \frac{|\text{Measured Value} - \text{Accepted Value}|}{\text{Accepted Value} \times 100}$$

## Exercise 1: Length, Temperature, and Mass

In this experiment, you will make measurements using the SI system units for length, mass, and temperature.

Read the procedure for each experiment in its entirety before you begin.

### Student-Supplied Materials

- aluminum pie pan
- matches or lighter
- CD or DVD
- 4 dimes
- fork
- key
- pen or pencil
- 5 pennies
- plastic or glass cup
- 3 quarters
- metric ruler
- ice cubes
- water
- spoon
- gloves

### Lab Kit Materials

The materials for this experiment are found in the HOL Chem 1 Lab Kit.

- aluminum cup, 2 oz
- burner fuel
- burner stand
- digital scale
- glass beaker, 100 mL
- safety glasses
- thermometer

### Procedure, Part 1: Length Measurements

1. Gather the metric ruler, CD or DVD, key, spoon, and fork.
2. Look at the calibration marks on your ruler to determine the degree of uncertainty and number of significant figures that can be made when measuring objects with the ruler.

   Note: Record every measurement you make with this ruler to the same decimal place. Remember to do this any time you use this ruler throughout the experiment.
3. Measure the length of each of the following objects (CD or DVD, key, spoon, fork) with the ruler in centimeters (cm) to the correct level of precision. Record each measurement in Data Table 1: Length Measurements.
4. Convert the measurements for each of the objects from centimeters to millimeters and record in Data Table 1.

5. Convert the measurements for each of the objects to meters and record in Data Table 1.

**Data Table 1: Length Measurements**

| | Length (cm) | Length (mm) | Length (m) |
|---|---|---|---|
| CD or DVD | | | |
| Key | | | |
| Spoon | | | |
| Fork | | | |

## Procedure, Part 2: Temperature Measurements

1. Gather the 100 mL glass beaker, cup (plastic or glass), matches or lighter, burner stand, burner fuel, thermometer, 2 oz aluminum cup, and aluminum pie pan.

   Note: The thermometer is shipped in a protective cardboard tube labeled “thermometer.”

2. Look at the calibration marks on the thermometer to determine the degree of uncertainty and number of significant figures that can be made when measuring temperature.

   Note: Record every measurement you make with this thermometer to the same decimal place. Remember to do this any time you use this measuring device throughout the experiment.

3. Turn on the tap water to hot. Let the water run as hot as possible for approximately 15 seconds.
4. Fill the 100 mL glass beaker with approximately 75 mL of hot tap water.
5. Measure the temperature of the hot tap water with the thermometer in degrees Celsius (°C) to the correct level of precision. Record the measurement in Data Table 2: Temperature Measurements.

   Note: When measuring the temperature, place the thermometer into the water so that the silver bulb is fully submerged without touching the bottom or any sides of the glass beaker. The measurement is complete when the thermometer remains the same temperature without changing.

6. Put on safety glasses and gloves.
7. Assemble the burner setup and light the fuel, using the following directions. See Figure 14.
   a. Place an aluminum pie pan on a solid work surface away from flammable objects.
   b. Set the burner stand toward the back of the pie pan.
   c. Place the beaker on the center of the stand.
   d. Uncap the burner fuel and set cap aside. Place the burner fuel on the pie pan just in front of the stand.
   e. Use matches or a lighter to ignite the fuel. **BE CAREFUL**: the flame may be nearly invisible.
   f. Gently slide the fuel under the stand without disturbing the beaker.

g. The small, 2 oz aluminum cup will be placed over the fuel to extinguish the flame. Set the aluminum cup next to the burner setup so you are ready to extinguish the flame at any point.

**Figure 14. Burner fuel setup**

8. Allow the water to heat until it comes to a full boil. As soon as the water is boiling, measure the temperature with the thermometer in degrees Celsius (°C) to the correct level of precision. Record the measurement in Data Table 2.
9. Allow the water to continue boiling for approximately 5 minutes. After 5 minutes, measure the temperature with the thermometer in degrees Celsius (°C) to the correct level of precision. Record the measurement in Data Table 2.
10. Use the 2 oz aluminum cup to extinguish the burner fuel flame, using the following directions. See Figure 15.

    a. Do not touch the metal stand or the beaker; they may be hot.

    b. Carefully slide the burner fuel canister out from underneath the burner stand. The sides of the burner fuel canister will be warm but not hot.

    c. Place the aluminum cup directly over the flame to smother it. The cup should rest on top of the fuel canister, with little or no smoke escaping. Do not disturb the burner stand and beaker; allow everything to cool completely.

**Figure 15. Using the aluminum cup to extinguish the flame**

    d. Once all equipment is completely cool, remove the aluminum cup and place the plastic cap back on the fuel. Ensure that the plastic cap snaps into place to prevent fuel leakage and evaporation. The aluminum cup, fuel, and all other materials may be used in future experiments.

11. Allow the 100 mL beaker to cool before touching it.
12. Turn on the tap water to cold. Let the water run as cold as possible for approximately 15 seconds.
13. Fill the cup (plastic or glass) approximately half full with cold tap water.
14. Measure the temperature of the cold tap water with the thermometer in degrees Celsius (°C) to the correct level of precision. Record the measurement in Data Table 2.
15. Add a handful of ice cubes to the cup of cold tap water and allow them to sit in the cold water for approximately 1 minute.
16. After 1 minute, stir the ice water with the thermometer.

17. Measure the temperature of the ice water after 1 minute with the thermometer in degrees Celsius (°C) to the correct level of precision. Record the measurement in Data Table 2.
18. Allow the ice to remain in the water for an additional 4 minutes.
19. After 4 minutes, stir the ice water with the thermometer.
20. Measure the temperature of the ice water with the thermometer in degrees Celsius (°C) to the correct level of precision. Record the measurement in Data Table 2.
21. Convert the temperature measurements for each of the 6 water samples from °C to °F and K. Record the converted temperatures in Data Table 2.

**Data Table 2: Temperature Measurements**

| | Temperature (°C) | Temperature (°F) | Temperature (K) |
|---|---|---|---|
| Hot from Tap | | | |
| Boiling | | | |
| Boiling for 5 minutes | | | |
| Cold from Tap | | | |
| Ice Water, 1 minute | | | |
| Ice Water, 5 minutes | | | |

## Using the Digital Scale

1. Remove the scale from the box and discard any excess packaging, including the strip inside the cover, next to the batteries.
2. Remove the lid from the scale and turn it on. Make sure that the scale is set to read in grams.
3. Place a small sheet of clean paper on the scale.
4. Tare the scale by pressing the Φ/T button so that the scale reads 0.0 g.
5. Add the item to the scale and read the mass.
6. Record the mass and continue with your experiment.

## Procedure, Part 3: Mass Measurements

1. Gather the pen or pencil, 5 pennies, 3 quarters, 4 dimes, and the key.
2. Read the instructions on how to use the digital scale. The lid of the scale must be opened to expose its weighing surface and make mass measurements.

   Note: There may be a cardboard protector between the scale base and top. If so, remove the cardboard from the scale.

3. Turn the scale on by pressing the Φ/T button.
4. Make sure the scale is reading in grams by looking for the letter *g* in the upper right corner of the scale. If the *g* is not showing, then press the M button until the scale is reading in grams.
5. Review the different objects listed in Data Table 3: Mass Measurements.
6. Estimate the masses for each of the objects in grams and record in Data Table 3. To help you with this process, a penny has a mass of approximately 2.5 grams.
7. Tare the scale by pressing the Φ/T button so that the scale reads 0.0 g.
8. Place the pen or pencil on the scale to measure the mass of the object. Record the mass in Data Table 3 under Actual Mass (g).
9. Repeat steps 7 and 8 for the remaining objects in Data Table 3.
10. For each item or group of objects, convert the actual mass (in grams) to kilograms (kg). Record in Data Table 3.

**Data Table 3: Mass Measurements**

| | Estimated Mass (g) | Actual Mass (g) | Actual Mass (kg) |
|---|---|---|---|
| Pen or Pencil | | | |
| 3 Pennies | | | |
| 1 Quarter | | | |
| 2 Quarters, 3 Dimes | | | |
| 4 Dimes, 5 Pennies | | | |
| 3 Quarters, 1 Dime, 5 Pennies | | | |
| Key | | | |
| Key, 1 Quarter, 4 Pennies | | | |

## Exercise 1 Questions

1. Water boils at 100°C at sea level. If the water in this experiment did not boil at 100°C, what could be the reason?

2. While heating two different samples of water at sea level, one boils at 102°C and one boils at 99.2°C. Calculate the percent error for each sample from the theoretical 100.0°C.

## Exercise 2: Volume and Density

### Student-Supplied Materials

- distilled water
- isopropyl (rubbing) alcohol ($C_3H_8O$)
- metric ruler
- gloves

### Lab Kit Materials

The materials for this experiment are found in the HOL Chem 1 Lab Kit.

- digital scale
- glass beaker, 100 mL
- graduated cylinder, 25 mL
- magnet bar
- metal bolt
- short-stem pipet
- string, 1 m
- safety glasses

### Procedure, Part 1: Volume and Density Measurements (Liquid)

1. Gather the graduated cylinder, distilled water, short-stem pipet, and isopropyl alcohol. Put on gloves and safety glasses.
2. Place the clean, dry 25 mL graduated cylinder on the tared scale. Record the mass of the graduated cylinder (g) in Data Table 4: Liquid Measurements under the Mass A column for water.
3. Fill the graduated cylinder with 5.0 mL of distilled water; use the short-stem pipet to measure exactly 5.0 mL of water. Record the volume in Data Table 4.
4. Place the 25 mL graduated cylinder with 5.0 mL distilled water on the tared scale. Record the mass of the graduated cylinder + liquid (g) in Data Table 4 under Mass B.
5. Calculate the mass of the water by subtracting Mass A from Mass B. Record the mass of the water in Data Table 4.
6. Pour the water down the drain and fully dry the graduated cylinder.

7. Repeat steps 2–6 for the isopropyl alcohol.
8. Calculate the densities of both the water and the isopropyl alcohol and record in Data Table 4.
9. The accepted value for the density of water is 1.00 g/mL and the accepted density for isopropyl alcohol is 0.786 g/mL. Determine the percent error between your calculated densities and the accepted values for both water and isopropyl alcohol. Record the percent error in Data Table 4.

**Data Table 4: Liquid Measurements**

| | Water | Isopropyl Alcohol |
|---|---|---|
| Mass A: Graduated Cylinder (g) | | |
| Volume (mL) | | |
| Mass B: Graduated Cylinder with Liquid (g) | | |
| Mass B – A: Liquid (g) | | |
| Density (g/mL) | | |
| Percent Error (%) | | |

## Procedure, Part 2: Volume and Density Measurements (Solid)

### Direct Measurement Method

1. Gather the metal bolt, string, magnet, graduated cylinder, beaker, metric ruler, and scale.
2. Tare the scale by pressing the Φ/T button so that the scale reads 0.0 g.
3. Place the magnet on the scale to measure the mass of the object. Record the mass in Data Table 5: Magnet—Direct Measurement Method.
4. Use the ruler to measure the length, width, and height of the magnet in centimeters to the correct level of precision. Record the measurements in Data Table 5.
5. Calculate the volume of the magnet by multiplying the length × width × height, and record in Data Table 5.
6. Calculate the density of the magnet by dividing the mass by the volume and record in Data Table 5.

**Data Table 5: Magnet—Direct Measurement Method**

| | Magnet |
|---|---|
| Mass (g) | |
| Length (cm) | |
| Width (cm) | |
| Height (cm) | |
| Volume ($cm^3$) | |
| Density ($g/cm^3$) | |

### Water Displacement Method

1. Tare the scale by pressing the Φ/T button so that the scale reads 0.0 g.
2. Place the magnet on the scale to measure the mass of the object. Record the mass in Data Table 6: Water Displacement Method.
3. Fill the graduated cylinder with 6–8 mL of distilled water. Record the volume, to the correct decimal place, in Data Table 6.
4. Carefully slide the magnet into the graduated cylinder so that the water doesn't splash. Read the volume of the graduated cylinder. Record the volume in Data Table 6, next to Final Volume of Graduated Cylinder.
5. Determine the volume of the object by calculating the difference in water displacement volumes (final – initial). Record in Data Table 6.
6. Calculate the density of the magnet and record in Data Table 6.
7. Carefully pour the water from the cylinder down the drain and collect the magnet.
8. Repeat these steps for the metal bolt.

**Data Table 6: Water Displacement Method**

| | Magnet | Metal Bolt |
|---|---|---|
| Mass (g) | | |
| Initial Volume of Graduated Cylinder (mL) | | |
| Final Volume of Graduated Cylinder (mL) | | |
| Object Volume (mL) | | |
| Density (g/mL) | | |

## Cleanup

1. Clean all glassware and lab equipment with soap and water.
2. Rinse the equipment again with distilled water.
3. Dry all items with paper towels and return them to the lab kit for future use.

## Exercise 2 Questions

1. An unknown, rectangular substance measures 3.60 cm high, 4.21 cm long, and 1.17 cm wide. If the mass is 21.3 g, what is this substance's density (in g/mL)? Remember to always show your work.

2. A sample of gold (Au) has a mass of 26.15 g. Given that the theoretical density is 19.30 g/mL, what is the volume of the gold sample?

3. A student was given an unknown metal. The student determined that the mass of the metal was 30.2 g. The student placed the metal in a graduated cylinder filled with 20.0 mL of water. The metal increased the volume of water to 22.9 mL. Calculate the density of the metal and determine the identity of the metal using the table below.

**Table 1: Density of Metals**

| Metal | Density (g/mL) |
|---|---|
| Lead | 11.3 |
| Silver | 10.5 |
| Nickel | 9.90 |
| Zinc | 7.14 |

Lesson 2

# Basic Building Materials

## Activities

Complete the following activity.

### Activity: Color Coding the Periodic Table

Having your periodic table for use in each lesson is key to your success in this course. In this activity, you'll make a simplified version to help visualize how it is arranged.

#### Materials

- blank periodic table (included in the appendix)
- colored pencils

#### Procedure

1. Using the blank periodic table, create a color-coded periodic table according to the directions below. Provide a key. Include a photo of your work.
   a. Color the boron group **gray**
   b. Color the nitrogen group **light brown**
   c. Color the transition metals **yellow**
   d. Color the alkali metals **green**
   e. Color the alkaline earth metals **pink**
   f. Color the halogens **purple**
   g. Color the noble gases **red**
   h. Add **blue** stripes to all the nonmetals (don't forget hydrogen!)
   i. Color the lanthanide series **orange**
   j. Color the actinide series **blue**
   k. Add **black** dots to all the metalloids
   l. Add the labels 1A–8A and 1B–8B to your periodic table (see textbook pages 42–43)

2. Fill in the blanks in the following sentences.

   a. Rows of elements in the periodic table are called ______________________.

   b. Columns of elements in the periodic table are called ______________________ or ______________________.

   c. Using page 44 of the textbook, identify the name of the following groups:

   - Group 1 (1A) ______________________
   - Group 2 (2A) ______________________
   - Groups 3–12 (1B–8B) ______________________
   - Group 17 (7A) ______________________
   - Group 18 (8A) ______________________

Lesson 6

# Speaking of Molecules

## Activities

Complete both of the following activities:

- Activity A: Single, Double, and Triple Bond Experiment for Kids
- Activity B: Connect the Dots: Lewis Dot Structure Tetris

### Activity A: Single, Double, and Triple Bond Experiment for Kids

Design a short inquiry-based experiment using everyday household items to explain to an elementary school student the difference between single, double, and triple bonds as well as their relative strength, and what they look like. Your goal is to make something that could be sent to fifth-grade students, so make it fun! Provide a one-page lab handout for the students, answers for your experiment questions, and photos of the setup and experiment design. Be sure to teach your students something about single, double, and triple bonds in a way they can understand and relate to!

### Activity B: Connect the Dots: Lewis Dot Structure Tetris

#### Materials

- scissors
- glue or tape
- white paper
- Lewis dot puzzle pieces (included in the appendix)

#### Procedure

1. Cut out the Lewis dot puzzle pieces.
2. Complete the activity on page 160 of your textbook, answering the questions.
3. Photograph your completed structures to share with your teacher.

Lesson

17

# Toxic Cleanup

## Experiment

Complete the following lab experiment.

### Stoichiometry of a Precipitation Reaction

The materials for this experiment are found in the HOL Chem 1 Kit.

Answer all questions. Provide data tables. Provide photos of your setup and results.

Experiment

# Stoichiometry of a Precipitation Reaction

## Learning Objectives

- Calculate the theoretical maximum amount of product produced in a precipitation reaction using stoichiometry.
- Perform a precipitation reaction and measure the precipitate to calculate percent yield.
- Explain differences between theoretical and actual yield in a controlled experiment.

# Exploration

## Explore Chemical Equations

A chemical equation is an illustration of the reaction that occurs between two or more specific chemical compounds. Chemical equations use letters and numbers to represent the chemical elements and the amounts or ratios of those elements present in the compounds that are either participating in the reaction or are a product of the reaction. For example, one methane molecule contains one carbon atom and four hydrogen atoms and is denoted as $CH_4$. The chemical compounds that are present before a reaction occurs are called reactants, and the compounds produced from the reaction are called products. In addition to identifying the products and reactants in a balanced chemical reaction, a chemical equation will also quantitatively identify the proportion of reactants to products. This quantitative proportion is known as stoichiometry, and it can be used to determine how much of each reactant is needed to produce a specific amount of each product. See Figure 1.

$$\underbrace{CuSO_4(aq) + 2NaOH(aq)}_{\text{(Reactants)}} \rightarrow \underbrace{Cu(OH)_2(s) + Na_2SO_4(aq)}_{\text{(Products)}}$$

**Figure 1. A balanced chemical equation. The chemical equation shows the chemical reaction between copper(II) sulfate and sodium hydroxide. The equation shows that when 1 formula unit of copper(II) sulfate reacts with 2 formula units of sodium hydroxide, 1 formula unit of copper(II) hydroxide and 1 formula unit of sodium sulfate are produced.**

As shown in Figure 1, chemical equations often denote the physical states of the reactants and products. The reaction in Figure 1 is a precipitation reaction where two solutions are mixed and an insoluble substance (precipitate) forms, which is then able to be separated or removed from the solution. The (s) after $Cu(OH)_2(s)$ denotes that a solid was formed as a product from the two aqueous (aq) reactants. The stoichiometry of a balanced chemical equation can be used to calculate the mass and number of moles of each reactant and each product in a chemical reaction.

## Explore Moles and the Periodic Table

A mole (mol) is a unit of measure describing the amount of a chemical substance that contains as many atoms, formula units, or molecules as there are in exactly 12 grams of pure carbon ($^{12}C$). One mole of a substance has $6.022 \times 10^{23}$ atoms (for an element, this represents Avogadro's number), molecules (for a compound), or formula units (for an ionic compound) and is equal to its molecular weight (molecular mass or formula mass). For example, the element sodium has a molecular mass of 22.99 grams; thus, 1 mole of sodium is equal to 22.99 grams. Likewise, the compound $H_2O$ has a molecular mass of H + H + O (1.008 + 1.008 + 16.00); thus, 1 mole of $H_2O$ is equal to 18.016 grams. The molar mass of each element is found in the periodic table.

## Explore Stoichiometric Quantities and Calculations

In addition to determining the amount of product formed in a reaction, stoichiometry can be used to determine how much of each reactant is required for all reactants to be used up at the same time. The quantities of reactants that are needed to fully react with one another at the same time are known as stoichiometric quantities. Stoichiometric quantities can be used to maximize the amount of product produced from the chemical reaction. For example, if you were performing the reaction in Figure 1 and had 5.70 grams of $CuSO_4$, you can use the balanced chemical equation and stoichiometry to determine how many grams of NaOH you would need to create the maximum amount of $Cu(OH)_2$. More specifically, to quantitatively calculate the maximum amount of product expected through a chemical reaction, you need only a balanced chemical equation, the molar mass of each substance, and the quantity of substance available for only one of the reactants.

A step-by-step example of this process, using the balanced equation from Figure 1, is shown below:

$$CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$$

Assuming there are only 5.70 grams of $CuSO_4$ available, how many grams of NaOH are necessary to reach stoichiometric quantities? How many grams of solid $Cu(OH)_2$ are expected to be produced?

**Step 1.** Check that the equation is balanced. To do this, ensure that there is the same number of atoms from each element on both sides of the equation.

**Step 2.** Convert the 5.70 grams of $CuSO_4$ to moles of $CuSO_4$.

$$5.70 \text{ g } CuSO_4 \times \frac{1 \text{ mol } CuSO_4}{159.62 \text{ g } CuSO_4} = 0.0357 \text{ mol } CuSO_4$$

**Step 3.** Evaluate the molar ratio between $CuSO_4$ and NaOH.

$$0.0357 \text{ mol } CuSO_4 \times \frac{2 \text{ mol NaOH}}{1 \text{ mol } CuSO_4} = 0.0714 \text{ mol NaOH}$$

The chemical equation states that for 1 mole of $CuSO_4$, 2 moles of NaOH are needed for stoichiometric quantities. Using the information calculated in step 2, if there are 0.0357 moles of $CuSO_4$, then 0.0714 moles of NaOH are required for a complete reaction.

**Step 4.** Convert moles of NaOH to grams of NaOH.

$$0.0714 \text{ mol NaOH} \times \frac{40.0 \text{ g NaOH}}{1 \text{ mol NaOH}} = 2.86 \text{ g NaOH}$$

This shows that 2.86 grams of NaOH are required to completely react with the 5.70 grams of $CuSO_4$.

**Step 5.** Determine the amount (moles) of $Cu(OH)_2$ expected from the reaction.

$$0.0357 \text{ mol } CuSO_4 \times \frac{1 \text{ mol } Cu(OH)_2}{1 \text{ mol } CuSO_4} = 0.0357 \text{ mol } Cu(OH)_2$$

The chemical equation states that for every 1 mole of $CuSO_4$ used, 1 mole of $Cu(OH)_2$ is expected. This means that the 0.0357 moles of $CuSO_4$ should produce 0.0357 moles of $Cu(OH)_2$.

**Step 6.** Calculate the theoretical yield by converting moles of $Cu(OH)_2$ to grams of $Cu(OH)_2$.

$$0.0357 \text{ mol } Cu(OH)_2 \times \frac{97.56 \text{ g } Cu(OH)_2}{1 \text{ mol } Cu(OH)_2} = 3.48 \text{ g } Cu(OH)_2$$

To double-check the results of the calculations, the law of the conservation of mass can be applied. According to this law, the mass of the products in a chemical reaction must equal the mass of the reactants.

**Step 7.** To double-check the conservation of mass, first calculate the mass of $Na_2SO_4$ that is expected from the reaction. With a 1:1 ratio, 0.0357 moles of $Na_2SO_4$ are expected.

$$0.0357 \text{ mol } CuSO_4 \times \frac{1 \text{ mol } Na_2SO_4}{1 \text{ mol } CuSO_4} = 0.0357 \text{ mol } Na_2SO_4$$

$$0.0357 \text{ mol } Na_2SO_4 \times \frac{142.04 \text{ g } Na_2SO_4}{1 \text{ mol } Na_2SO_4} = 5.07 \text{ g } Na_2SO_4$$

**Step 8.** To verify the results of the calculations and have confidence in the series of stoichiometric calculations, calculate the sum of the reactants and the sum of the products.

$$5.70 \text{ g } CuSO_4(aq) + 2.86 \text{ g } 2NaOH(aq) = 3.48 \text{ g } Cu(OH)_2(s) + 5.07 \text{ g } Na_2SO_4(aq)$$

However, in practical experimentation, a system is seldom completely closed. As a result, one should realistically expect a slightly smaller amount of product, as the theoretical yield is rarely obtained. This deviation, from theoretical yield to actual yield, is called the percent yield and can be calculated.

Note: Always watch significant figures during calculations, or the theoretical yield of the products and reactants may differ slightly.

**Step 9.** Determine the percent yield.

Assume that the actual yield was 3.26 g $Cu(OH)_2$.

$$\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% = \text{Percent Yield}$$

Using the yields both given and calculated:

$$\frac{3.26 \text{ g } Cu(OH)_2}{3.48 \text{ g } Cu(OH)_2} \times 100\% = 93.7\%$$

## Explore Hydrates

In this experiment, you will use stoichiometry to determine the quantities necessary for a complete precipitation reaction between sodium carbonate ($Na_2CO_3$) and calcium chloride dihydrate ($CaCl_2 \cdot 2H_2O$). A hydrate is a solid compound that contains water molecules. Hydrates are named by adding a prefix plus the word hydrate to the end of the standard name of the compound. These prefixes (*mono-*, *di-*, *tri-*, *tetra-*, etc.) indicate the number of moles of water molecules present for each mole of the compound. For example:

$CuSO_4 \cdot 5H_2O$ = Copper(II) sulfate pentahydrate

$MgSO_4 \cdot 7H_2O$ = Magnesium sulfate heptahydrate

The equations above state that 1 mole of $CuSO_4 \cdot 5H_2O$ contains 1 mol $CuSO_4$ and 5 mol $H_2O$; and 1 mol $MgSO_4 \cdot 7H_2O$ contains 1 mol $MgSO_4$ and 7 mol $H_2O$. Water molecules in a hydrate are easily removed by heating the solid hydrate or by dissolving it in water. Thus, while the molecular weight of the $CaCl_2 \cdot 2H_2O$ compound includes the 2 water molecules, only the $CaCl_2$ portion of the compound is available to react with the sodium carbonate when placed into solution.

Assume that there were 5.0 g of $CaCl_2 \cdot 2H_2O$, and you needed to determine the moles of $CaCl_2$ available in that 5.0 g to react in an aqueous solution with $Na_2CO_3$. Convert the 5.0 g of $CaCl_2 \cdot 2H_2O$ to mol $CaCl_2 \cdot 2H_2O$.

$$5.0 \text{ g } CaCl_2 \cdot 2H_2O \times \frac{1 \text{ mol } CaCl_2 \cdot 2H_2O}{147.01 \text{ g } CaCl_2 \cdot 2H_2O} = 0.034 \text{ mol } CaCl_2 \cdot 2H_2O$$

Thus, in 5.0 g $CaCl_2 \cdot 2H_2O$ there are 0.034 mol $CaCl_2$ available to react in an aqueous solution with $Na_2CO_3$. If the stoichiometry of the reaction was 1:1, then 0.034 mol $Na_2CO_3$ would be required to reach stoichiometric quantities and fully react with all the $CaCl_2$ in solution.

### Did You Know . . . ?

Stoichiometry is used in everyday life. Converting standard food recipes to produce larger or smaller amounts is one example. If 4 tablespoons of butter and 1 egg are used to produce 12 cookies, then 8 tablespoons of butter and 2 eggs would be needed to yield 24 cookies. That's stoichiometry!

## Exercise 1: Stoichiometry of a Precipitation Reaction

In this exercise, you will use stoichiometry to determine the amount of reactant needed to create the maximum amount of product in a precipitation reaction. After performing the reaction, you will calculate the percent yield of product. Share photos of your setup and results with your teacher.

### Student-Supplied Materials

- distilled water
- dish soap
- scissors
- paper towels
- white paper
- tap water

### Lab Kit Materials

The materials for this experiment are found in the HOL Chem 1 Lab Kit.

- digital scale
- funnel, 70 mm
- 2 glass beakers, 100 mL
- graduated cylinder, 25 mL
- gloves
- safety glasses

- metal spatula
- glass stirring rod
- filter paper, 12.5 cm
- Experiment Bag: Chemistry 1
- calcium chloride dihydrate ($CaCl_2 \cdot 2H_2O$) – 2.5 g
- sodium carbonate ($Na_2CO_3$), 2.0 g

**Procedure**

1. Review the following reaction, where sodium carbonate and calcium chloride dihydrate react in an aqueous solution to create calcium carbonate (solid precipitate formed in the reaction), a salt (sodium chloride), and water.

$$Na_2CO_3(aq) + CaCl_2 \cdot 2H_2O(aq) \rightarrow CaCO_3(s) + 2NaCl(aq) + 2H_2O(aq)$$

2. Put on your safety glasses and gloves.
3. Use the graduated cylinder to measure 25 mL of distilled water. Add 25 mL of distilled water to each of the two 100 mL glass beakers.
4. Cut a small square of white paper and place it on the scale. Tare the scale by pressing the Φ/T button so that the scale reads 0.0 g. See Figure 3.
5. Use the metal spatula to measure 1.50 grams of the $CaCl_2 \cdot 2H_2O$.
6. Carefully add the $CaCl_2 \cdot 2H_2O$ to one of the beakers containing 25 mL of distilled water. Use the metal spatula to transfer any residue remaining on the piece of paper. Swirl the mixture until the $CaCl_2 \cdot 2H_2O$ is fully dissolved into the water.
7. Rinse the metal spatula with distilled water and fully dry it with paper towels.


**Figure 3. Tared scale with the small square of paper**

8. Use the information and examples provided in the Exploration section to calculate how many moles of $CaCl_2 \cdot 2H_2O$ are present in 1.50 grams of $CaCl_2 \cdot 2H_2O$, and then calculate how many moles of pure $CaCl_2$ are present in the 1.50 grams of $CaCl_2 \cdot 2H_2O$. Record the answers in Data Table 1: Stoichiometry Values.
9. Use the information and examples provided in the Exploration section and the values input into Data Table 1 (from step 8) to determine how many moles of $Na_2CO_3$ are necessary to reach stoichiometric quantities. From that calculation, determine how many grams of $Na_2CO_3$ are necessary to reach stoichiometric quantities. Record both values in Data Table 1.
10. Cut a new small square of white paper and place it on the scale. Tare the scale by pressing the Φ/T button so that the scale reads 0.0 g.

11. Use the metal spatula to measure the calculated amount of $Na_2CO_3$, and carefully add it to the 25 mL of distilled water in the second 100 mL glass beaker. Use the metal spatula to transfer any residue remaining on the piece of paper.

    Note: As the scale weighs to the nearest tenth of a gram, round up when weighing the $Na_2CO_3$. This will ensure that there is sufficient $Na_2CO_3$ to react with the $CaCl_2$.

12. Use the stirring rod to stir the $Na_2CO_3$ and break apart any clumps until it is fully dissolved in the water.

13. Pour the $Na_2CO_3$ solution from the 100 mL glass beaker into the beaker containing the $CaCl_2 \cdot 2H_2O$ solution. Rinse the beaker containing $Na_2CO_3$ with 2–3 mL of distilled water and transfer the rinse to the beaker containing the $CaCl_2 \cdot 2H_2O$. Swirl the contents of the beaker to fully mix the two solutions and the precipitate of calcium carbonate will form instantly.

14. Use the information and examples provided in the Exploration section to determine the maximum (theoretical) amount of $CaCO_3$, in grams, that can be produced from the precipitation reaction. Record this value in Data Table 1.

15. Wash the now empty 100 mL glass beaker (that contained the $Na_2CO_3$ solution) with soap and water. Rinse the beaker with distilled water and thoroughly dry with a paper towel.

16. Fold the round filter paper into a cone shape, as shown in Figure 4.

17. Place the folded filter paper onto the tared scale and record the mass of the filter paper in Data Table 1.

18. Place the folded filter paper into the funnel and dampen it with a small amount of distilled water. Swirl the contents of the beaker to dislodge any precipitate from the sides and while holding the filter paper open, slowly pour the contents of the beaker into the filter-paper lined funnel.

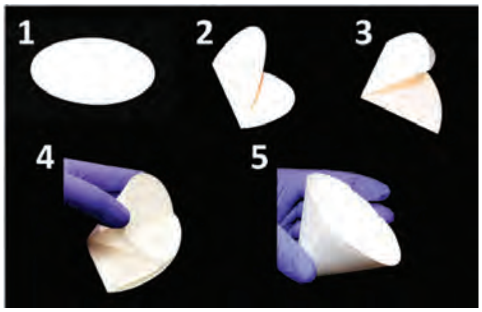



**Figure 4. Folding of filter paper**

    Note: Be careful not to overfill the funnel. It may be necessary to gently swirl the contents in the funnel to keep the precipitate from clogging the paper, but be careful not to touch the filter paper so that it does not tear.

19. Add 2–3 mL of distilled water to the beaker and swirl the water around to collect any precipitate stuck to the sides of the beaker. Pour into the filter-paper lined funnel.

20. Allow all the liquid to drain from the funnel into the beaker. This may take 10–15 minutes.

    Note: Lift the funnel periodically during this time to facilitate the water draining into the beaker. Again, do not touch the filter paper so that it does not tear.

21. After all liquid has drained from the funnel, carefully remove the filter paper from the funnel and place it on paper towels in a warm location, such as a windowsill that receives a lot of sunlight, where it will not be disturbed. See Figure 5.

22. Allow the filter paper to completely dry, which will require at least an overnight drying period.

    Note: Ensure the filter paper is placed in a location that does not have flowing air, such as a fan or air vent.

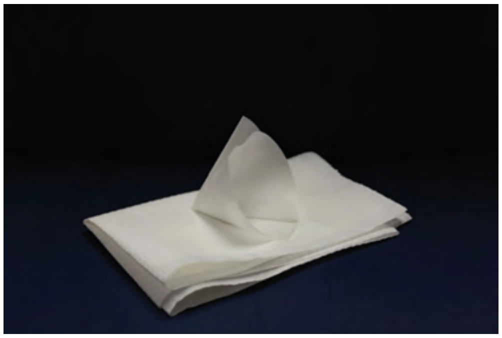


**Figure 5. Filter paper with precipitate set on paper towel to dry**

23. When the filter paper with precipitate is completely dry, tare the scale and place the paper on the scale to obtain the mass. Record the mass of the filter paper and precipitate in Data Table 1.

24. Calculate the actual mass of the precipitate and record in Data Table 1.

25. Calculate the percent yield of the precipitate and record in Data Table 1.

**Data Table 1: Stoichiometry Values**

| | Value |
|---|---|
| Initial: $CaCl_2 \cdot 2H_2O$ (g) | |
| Initial: $CaCl_2 \cdot 2H_2O$ (mol) | |
| Initial: $CaCl_2$ (mol) | |
| Initial: $Na_2CO_3$ (mol) | |
| Initial: $Na_2CO_3$ (g) | |
| Theoretical: $CaCO_3$ (g) | |
| Mass of Filter Paper (g) | |
| Mass of Filter Paper + $CaCO_3$ (g) | |
| Actual: $CaCO_3$ (g) | |
| % Yield: | |

## Cleanup

1. Dispose of chemicals properly.
2. Clean all equipment and thoroughly dry.
3. Return cleaned materials to the lab kit for future use.

**Exercise 1 Questions**

1. A perfect percent yield would be 100%. Based on your results, describe your success in recovery of the calcium carbonate and suggest possible sources of error.

2. What impact would adding twice as much $Na_2CO_3$ than required for stoichiometric quantities have on the quantity of product produced?

3. Determine the quantity (g) of pure $CaCl_2$ in 7.5 g of $CaCl_2 \cdot 9H_2O$. Show your work.

4. Determine the quantity (g) of pure $MgSO_4$ in 2.4 g of $MgSO_4 \cdot 7H_2O$. Show your work.

5. Conservation of mass was discussed in the Exploration section. Describe how conservation of mass (actual, not theoretical) could be checked in the experiment performed.

Lesson

# 23 Chemical Equilibrium

## Experiment

Complete the following experiment.

### Using Buffers

The materials for this experiment are found in the HOL Chem 2 Kit.

This lab builds on your understanding of pH from lesson 16. Review lesson 16 if you need a refresher before you begin.

# Experiment

# Using Buffers

## Learning Objectives

- Create an acetic acid/sodium acetate buffer solution.
- Test a buffer solution by the addition of acids and bases.
- Evaluate buffering capacity in response to additions of concentrated and dilute acids and bases.

# Exploration

## Explore Chemical Buffers

A chemical buffer is an aqueous solution that resists pH changes when small quantities of an acid or base are added to it. Buffer solutions are important to both industrial and physiological systems because many chemical reactions proceed only when pH remains within a narrow range. Pharmaceutical production, textile dyeing, electroplating, and glue manufacturing all require the use of buffers. Furthermore, living organisms utilize internal buffers to maintain cellular respiration, metabolism, and fluid levels. See Figure 1.

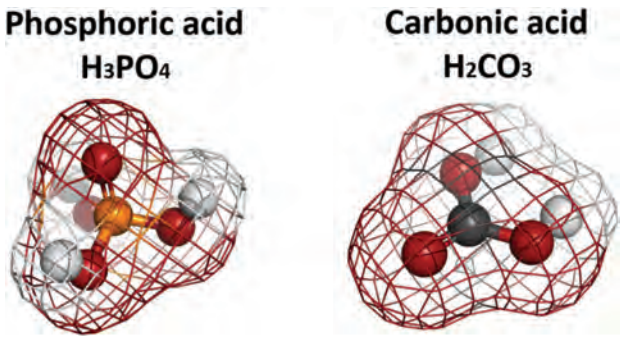



**Figure 1. Common buffers. A. Phosphoric acid, a buffer component used in industry. B. Carbonic acid, a component of the blood buffering system important for respiration and metabolism in living organisms. (Image copyright molekuul.be, 2013. Used under license from Shutterstock.com)**

The function of a buffer is not to keep a solution neutral but rather to minimize the change in pH when a base or acid is added to the solution. Every buffer that is made has a certain buffer capacity and buffer range. The buffer capacity is the amount of acid or base that can be added before the pH begins to change significantly. It can be also defined as the quantity of strong acid or base that must be added to change the pH of 1 liter of solution by 1 pH unit. The buffer range is the pH range where a buffer effectively neutralizes added acids and bases while maintaining a relatively constant pH. The pH range of the buffer will be close to the $pK_a$ of the weak acid in the buffer. See Table 1.

| Buffer | pK | Approximate pH Range |
|---|---|---|
| Phosphoric acid ($H_3PO_4$) | 2.16 | 1–3 |
| Formic acid (HCOOH) | 3.75 | 3–5 |
| Acetic acid ($CH_3COOH$) | 4.75 | 4–6 |
| Ammonium ($NH_4^+$) | 9.26 | 8–10 |
| Bicarbonate ($HCO_3^-$) | 10.25 | 9–11 |

**Table 1. Examples of chemical buffer ranges**

### Did You Know . . . ?

Human blood needs to maintain a neutral pH of 7.4 for the body to function. If the blood pH rises above 7.8 or below 6.8, the body will shut down. The carbonic acid/bicarbonate buffer system keeps the blood pH in the range of 7.35 to 7.45.

## Explore Buffer Composition

Buffers typically consist of a weak acid and its conjugate base or a weak base and its conjugate acid. Recall that acids ionize in water to produce a hydronium ion ($H_3O^+$) and a conjugate base ($A^-$). See the equation below.

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

The strength of an acid is dependent on its ability to fully ionize in water. A strong acid fully ionizes into hydronium ions ($H_3O^+$) and a conjugate base ($A^-$). See the example with HCl below. The stronger the tendency of an acid to ionize in water, the more hydrogen ions ($H^+$) that end up in the solution. Therefore, a strong acid will have a low pH.

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Weaker acids form solutions with higher pH values than stronger acids of the same concentration because weak acids exhibit only partial ionization in water. Acid and conjugate base strength are inversely related. For example, a stronger acid has a weaker conjugate base, with water having a higher affinity for the proton than the weaker conjugate base. On the other hand, a weaker acid has a stronger conjugate base, with the stronger conjugate base having a higher affinity for the proton than water. Therefore, less of the weaker acid ionizes in water, which results in less [$H^+$] in the solution and a higher pH. Buffers act most effectively when [$HA^+$] and [$A^-$] are equal. Weak acids make better buffers than strong acids since their conjugate bases have measurable strength.

Acetic acid ($CH_3COOH$) and its conjugate base, the acetate ion ($CH_3COO^-$), can form an acetic acid buffer. The relationship between acetic acid and acetate at equilibrium is shown below.

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

In this experiment, you will make an acetic acid buffer by adding sodium hydroxide (NaOH) to commercial vinegar. Commercial vinegar is usually composed of 5% acetic acid by volume. The sodium hydroxide will react with the acetic acid to form sodium acetate ($CH_3COONa$). See the equation below.

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$$

Sodium acetate completely dissociates in water to form acetate and sodium ions, forming the conjugate base for acetic acid in the buffer solution.

### Did You Know . . . ?

Acetic acid has many commercial uses. Aside from being the active ingredient in vinegar, acetic acid is used in the production of photographic film, wood glue, and synthetic fibers and fabrics. Additionally, acetic acid is used as a descaling agent in households and also as a food additive to regulate acidity. The global demand for acetic acid is around 6.5 million tons per year!

## Explore the Henderson-Hasselbalch Equation

The pH of a buffered solution can be calculated using the Henderson-Hasselbalch equation. This equation is also used to determine how to prepare the buffer and how to calculate the effect of adding additional acid and base to the existing buffer.

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

Where:

$[HA]$ = weak acid concentration

$[A^-]$ = conjugate base concentration

$pK_a$ = acid ionization constant

Buffers act most effectively when the ratio of $[HA]$ to $[A^-]$ is equal to 1:

$pH = pK_a + \log 1$

$pH = pK_a + 0$

$pH = pK_a$

It is expected that as $OH^-$ is added to the solution, the weak acid will give up its $H^+$, forming $H_2O$. As more and more $H^+$ ions are neutralized to form water, more conjugate base ions will be formed, and the pH will increase. (Fewer $H^+$ ions will be free in the solution and more acetate will be available.) As noted, when $pH = pK_a$, the concentration of weak acid equals the concentration of its conjugate base. In this balanced buffer solution, there are as many acid molecules available (to react with any base added to the solution) as there are conjugate base ions (to react with any acid added to the solution).

The following is an example of using the Henderson-Hasselbalch equation to determine the pH of a nitrous acid buffer solution consisting of 1.0 M $HNO_2$ and 0.225 M $NaNO_2$. The $pK_a$ for $HNO_2$ is 3.14.

$pH = pK_a + \log\frac{[A^-]}{[HA]}$

$pH = 3.14 + \log(0.225M/1.0M)$

$pH = 3.14 - 0.648 = 2.49$

Below is an example of using the Henderson-Hasselbach equation to determine what ratio of $[A^-]$ to $[HA]$ is required to create an acetic acid buffer of pH 5.0. The $pK_a$ of acetic acid is 4.74.

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$
$$5.0 = 4.74 + \log\left(\frac{[A^-]}{[HA]}\right)$$
$$0.26 = \log\left(\frac{[A^-]}{[HA]}\right)$$
$$100.26 = \left(\frac{[A^-]}{[HA]}\right)$$
$$1.8 = \left(\frac{[A^-]}{[HA]}\right)$$

## Exercise 1: Using Buffers

In this exercise, you will experiment with the weak acid and conjugate base pair of acetic acid and acetate ion.

### Student-Supplied Materials

- distilled water
- toothpicks
- dish soap
- scissors
- permanent marker
- paper towels
- tap water
- 120 mL white vinegar

### Lab Kit Materials

The materials for this experiment are found in the HOL Chem 2 Lab Kit.

- gloves
- safety glasses
- 2 short-stem pipets
- 2 small graduated pipets, 5 mL
- test tube cleaning brush
- well plate, 24-count
- Experiment Bag: Chemistry 2
    - hydrochloric acid (HCl), 6 M, 3 mL
    - sodium hydroxide (NaOH), 6 M, 3 mL
    - pH test strips, wide range

### Procedure, Part 1: Preparation of the Solutions and Materials

Using white vinegar, which is 5% acetic acid, you will prepare an acetic acid/acetate ion buffer as described below.

It is important that you:

- Keep the pH paper clean.
- Avoid air bubbles when transferring solutions by plastic pipets.

- Use clean toothpicks to stir the solutions after adding each drop of buffer.
- Avoid cross contamination of toothpicks.

1. Gather all materials from the materials list for this experiment.
2. Put on your safety glasses and gloves. Keep them on for the remainder of the experiment.
3. Use the permanent marker to label an empty graduated pipet "vinegar."
4. Use the labeled pipet to add 2.0 mL of white vinegar (5% $CH_3COOH$) into wells A1, A6, B1, and B6 of the 24-well plate.
5. Using scissors, carefully remove the tips of the HCl and NaOH pipets.

   Note: A good place to store the opened chemical pipets is bulb-side down in an empty well of the well plate. See Figure 2.

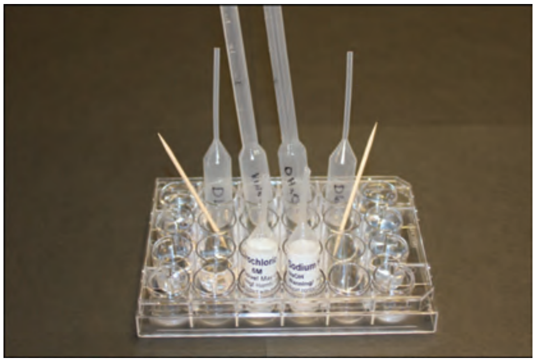


**Figure 2. Experimental setup**

6. Add 4 drops 6 M NaOH to wells A1, A6, B1, and B6 to create an acetic acid buffer solution.
7. Stir each of the 4 wells with a clean toothpick.
8. Use the permanent marker to label the second empty graduated pipet "distilled water."
9. Use this pipet to add 2.0 mL of distilled water to well D1.
10. Add 1 drop of 6 M HCl to the distilled water in well D1 to create a 0.1 M solution of HCl.
11. Stir the solution in well D1 with a clean toothpick.
12. Use the pipet labeled "distilled water" to add 2.0 mL of distilled water to well D6.
13. Add 1 drop of 6 M NaOH to the distilled water in well D6 to create a 0.1 M solution of NaOH.
14. Stir the solution in well D6 with a clean toothpick.
15. Remove 27 pH test strips from the wide-range pH booklets and cut them in half, lengthwise, creating 54 half-sized strips. The half-sized strips will be referred to as "strips" for the remainder of the experiment. See Figure 3.

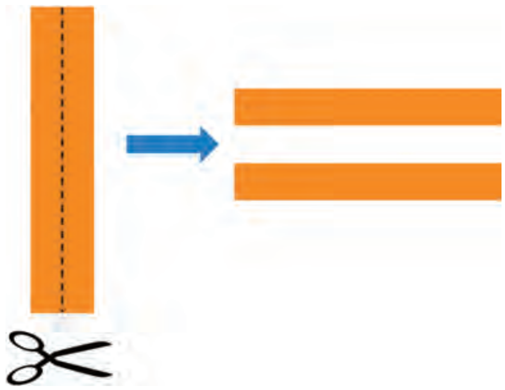


Figure 3. Cutting the pH strips in half, lengthwise

16. Place the pH color range guide next to the well plate so that it is easy to see while conducting the experiment.

    Note: The test solutions have been created. Wells A1, A6, B1, and B6 contain the same acetic acid/acetate buffer solution. Well D1 contains 0.1 M HCl, and well D6 contains 0.01 M NaOH. The pipets contain 6 M HCl and 6 M NaOH.

### Procedure, Part 2: Adding Different Concentrations of Acid and Base to a Buffer

1. Use the permanent marker to label one of the empty short-stem pipets "Well D1." Label the second empty, short-stem pipet "Well D6."
2. Measure the pH of the solution in well A1 by dipping the end of a strip of unused pH paper into the solution. See Figure 4 for proper technique.

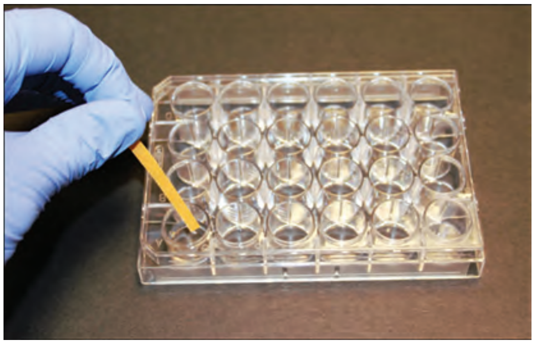


Figure 4. Using proper technique to measure pH

3. Use the color chart to determine the pH and record in Data Table 1: Adding 0.1 M HCl from D1 to A1.
4. Use the short-stem pipet labeled "Well D1" to add 2 drops of 0.1 M HCl from well D1 into the buffer in well A1.
5. Use a clean toothpick to stir the solution in the well.
6. Measure the pH of the solution in well A1 by dipping the end of a strip of unused pH paper into the solution.

7. Use the color chart to determine the pH and record in Data Table 1.
8. Repeat steps 4–7 seven more times, adding 0.1 M HCl to well A1. Complete Data Table 1.

   Note: Use a clean toothpick and a new strip of pH paper after each addition of 0.1 M HCl.
9. Measure the pH of the solution in well A6 and record in Data Table 2: Adding 0.1 M NaOH from D6 to A6.
10. Use the short-stem pipet labeled "Well D6" to add 2 drops of 0.1 M NaOH from well D6 into the buffer in well A6.
11. Use a clean toothpick to stir the solution in the well.
12. Measure the pH of the solution in well A6 and record in Data Table 2.
13. Repeat steps 10–12 seven more times, adding 0.1 M NaOH to well A6. Complete Data Table 2.

    Note: Use a clean toothpick and a new strip of pH paper after each addition of 0.1 M NaOH.
14. Measure the pH of the solution in well B1 and record in Data Table 3: Adding 6 M HCl from Pipet into B1.
15. Add 2 drops of 6 M HCl directly from the chemical pipet into the buffer in well B1.
16. Use a clean toothpick to stir the solution in the well.
17. Measure the pH of the solution in well B1 and record in Data Table 3.
18. Repeat steps 15–17 four more times, adding 6 M HCl to well B1. Complete Data Table 3.

    Note: Use a clean toothpick and a new strip of pH paper after each addition of 6 M HCl.
19. Measure the pH of the solution in well B6 and record in Data Table 4: Adding 6 M NaOH from Pipet into B6.
20. Add 2 drops of 6 M NaOH directly from the chemical pipet into the buffer in well B6.
21. Use a clean toothpick to stir the solution in the well.
22. Measure the pH of the solution in well B6 and record in Data Table 4.
23. Repeat steps 20–22 four more times, adding 6 M NaOH to well B6. Complete Data Table 4.

    Note: Use a clean toothpick and a new strip of pH paper after each addition of 6 M NaOH.

### Data Table 1: Adding 0.1 M HCL from D1 to A1

| Number of Drops | pH of Solution |
|---|---|
| 0 | |
| 2 | |
| 4 | |
| 6 | |
| 8 | |
| 10 | |
| 12 | |
| 14 | |
| 16 | |

### Data Table 2: Adding 0.1 M NaOH from D6 to A6

| Number of Drops | pH of Solution |
|---|---|
| 0 | |
| 2 | |
| 4 | |
| 6 | |
| 8 | |
| 10 | |
| 12 | |
| 14 | |
| 16 | |

### Data Table 3: Adding 6 M HCl from Pipet into B1

| Number of Drops | pH of Solution |
|---|---|
| 0 | |
| 2 | |
| 4 | |
| 6 | |
| 8 | |
| 10 | |

### Data Table 4: Adding 6 M NaOH from Pipet into B6

| Number of Drops | pH of Solution |
|---|---|
| 0 | |
| 2 | |
| 4 | |
| 6 | |
| 8 | |
| 10 | |

### Part 3: Adding Dilute Concentrations of Acid and Base to Distilled Water

1. Use the pipet labeled "distilled water" to add 2.0 mL of distilled water into wells C1 and C6.
2. Measure the pH of the distilled water in well C1 and record in Data Table 5: Adding 0.1 M HCl from D1 to C1.
3. Use the short-stem pipet labeled "Well D1" to add 2 drops of 0.1 M HCl from well D1 into the distilled water in well C1.
4. Use a clean toothpick to stir the solution in the well.
5. Measure the pH of the solution in well C1 and record in Data Table 5.
6. Repeat steps 3–5 four more times, adding 0.1 M HCl to well C1. Complete Data Table 5.

   Note: Use a clean toothpick and a new strip of pH paper after each addition of 0.1 M HCl.
7. Measure the pH of the distilled water in well C6 and record in Data Table 6: Adding 0.1 M NaOH from D6 to C6.
8. Use the short-stem pipet labeled "Well D6" to add 2 drops of 0.1 M NaOH from well D6 into the distilled water in well C6.
9. Use a clean toothpick to stir the solution in the well.
10. Measure the pH of the solution in well C6 and record in Data Table 6.
11. Repeat steps 8–10 four more times, adding 0.1 M NaOH to well C6. Complete Data Table 6.

    Note: Use a clean toothpick and a new strip of pH paper after each addition of 0.1 M NaOH.

### Data Table 5: Adding 0.1 M HCl from D1 to C1

| Number of Drops | pH of Solution |
|---|---|
| 0 | |
| 2 | |
| 4 | |
| 6 | |
| 8 | |
| 10 | |

### Data Table 6: Adding 0.1 M NaOH from D6 to C6

| Number of Drops | pH of Solution |
|---|---|
| 0 | |
| 2 | |
| 4 | |
| 6 | |
| 8 | |
| 10 | |

## Cleanup

1. Properly dispose of the chemicals, pipets, and used pH paper. All lab liquids/waste should be soaked up with paper towels and put in the trash.
2. Using the test tube cleaning brush, thoroughly wash the well plate with soap and water and allow to dry.
3. Return cleaned equipment to the lab kit for future use.

## Exercise 1 Questions

1. Describe the purpose of a buffer.

2. Write the chemical equations for the neutralization reactions that occurred when HCl and NaOH were added to the buffer solution.

3. How do the results in Data Tables 1 and 2 support the role of a buffer?

4. Describe the buffer capacity of the acetic acid buffer solution in relation to the addition of both concentrated and dilute acids and bases. Reference the results in Data Tables 1–4 in your answer.

5. Did distilled water act as a buffer in this experiment? Use your data to support your answer.

6. An acetic acid buffer solution is required to have a pH of 5.27. You have a solution that contains 0.01 mol of acetic acid. What molarity of sodium acetate will you need to add to the solution? The $pK_a$ of acetic acid is 4.74. Show calculations in your answer.

# Appendix