Goodwin College Chemistry Lab

Student Manual

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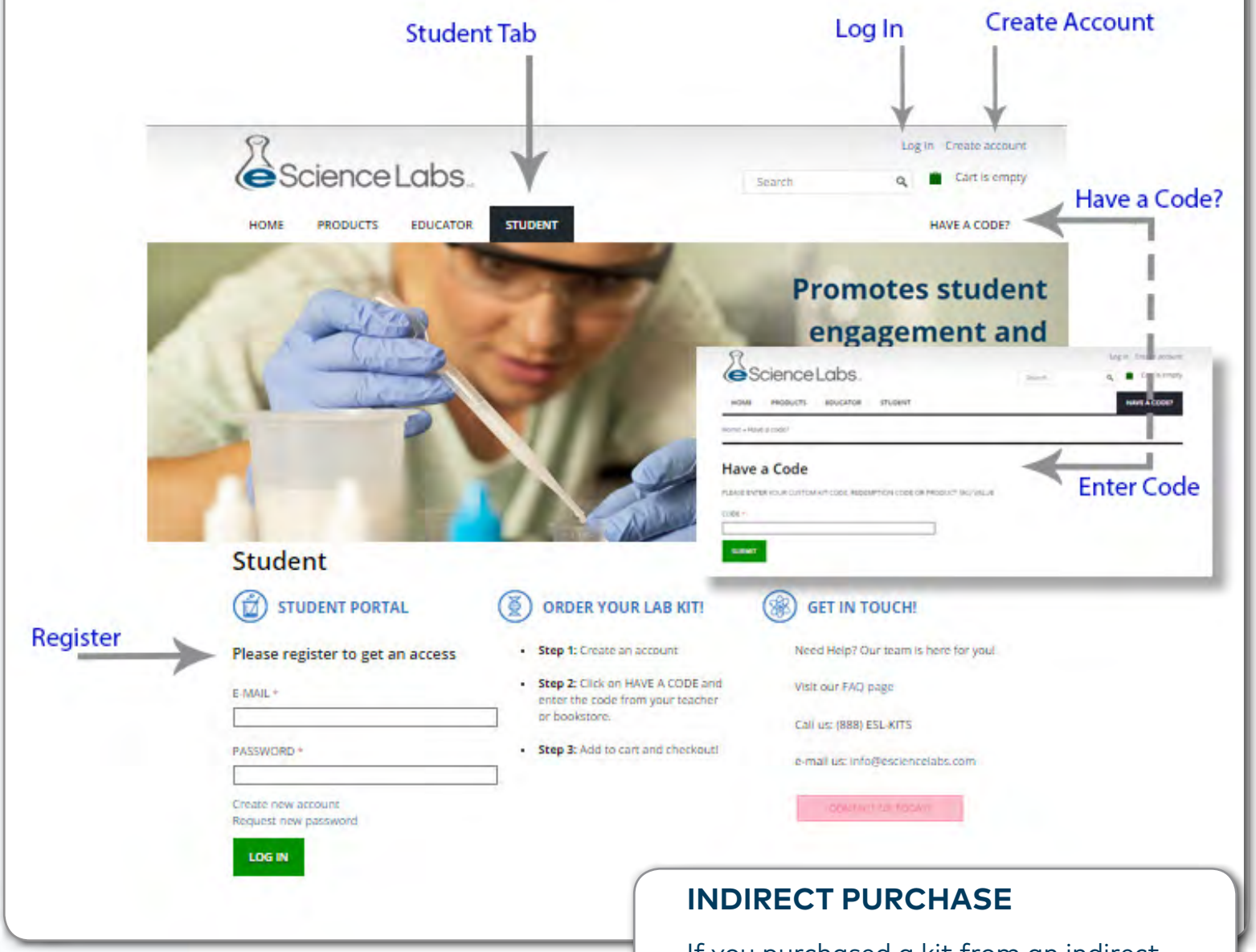

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The experiments included within this lab manual are suitable for supervised or unsupervised learning environments. eScience Labs assumes full liability for the safety and techniques employed within this manual provided that all users adhere to the safety guidelines outlined in the mandatory eScience Labs Safety Video, Preface, and Appendix. All users must understand and agree to the eScience Labs safety guidelines prior to beginning their lab experiments. eScience Labs does not condone use of the lab materials provided in its lab kits for any use outside of the curriculum expressly outlined within the lab manual.

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ACKNOWLEDGEMENTS

The Second Edition Introductory Chemistry Lab Manual is a collaborative development, which we are delighted to provide to the higher education curriculum. A key contributor of this manual is Dr. Jamille Jojo of Ivy Tech Community College. Dr. JoJo brought rigor, intelligence, and industry applicability to the authentic laboratory experiences contained with in this kit.

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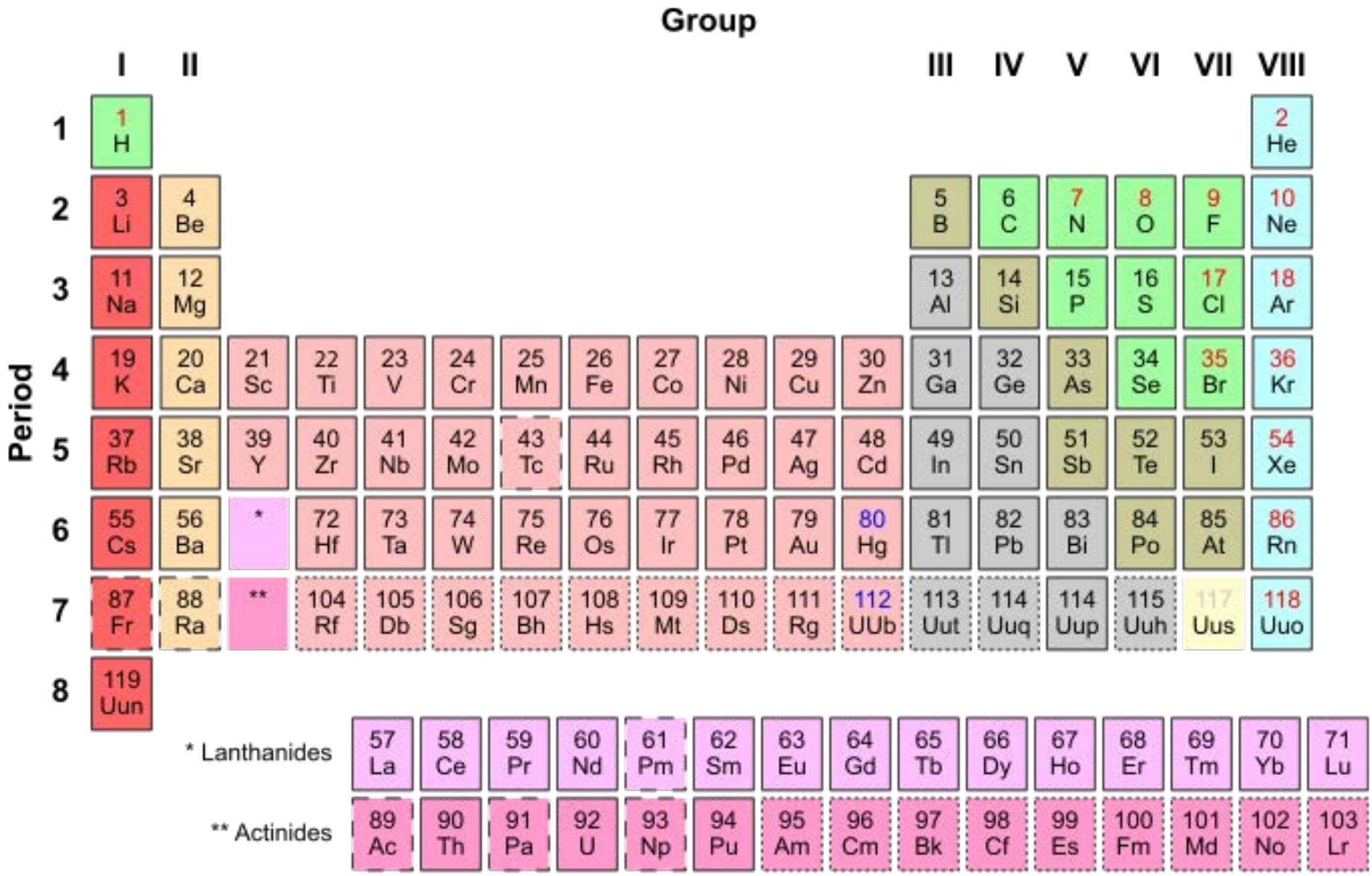

Alkali metals	Alkaline earth metals	Lanthanides	Actinides	Transition metals
Poor metals	Metalloids	Nonmetals	Halogens	Noble gases

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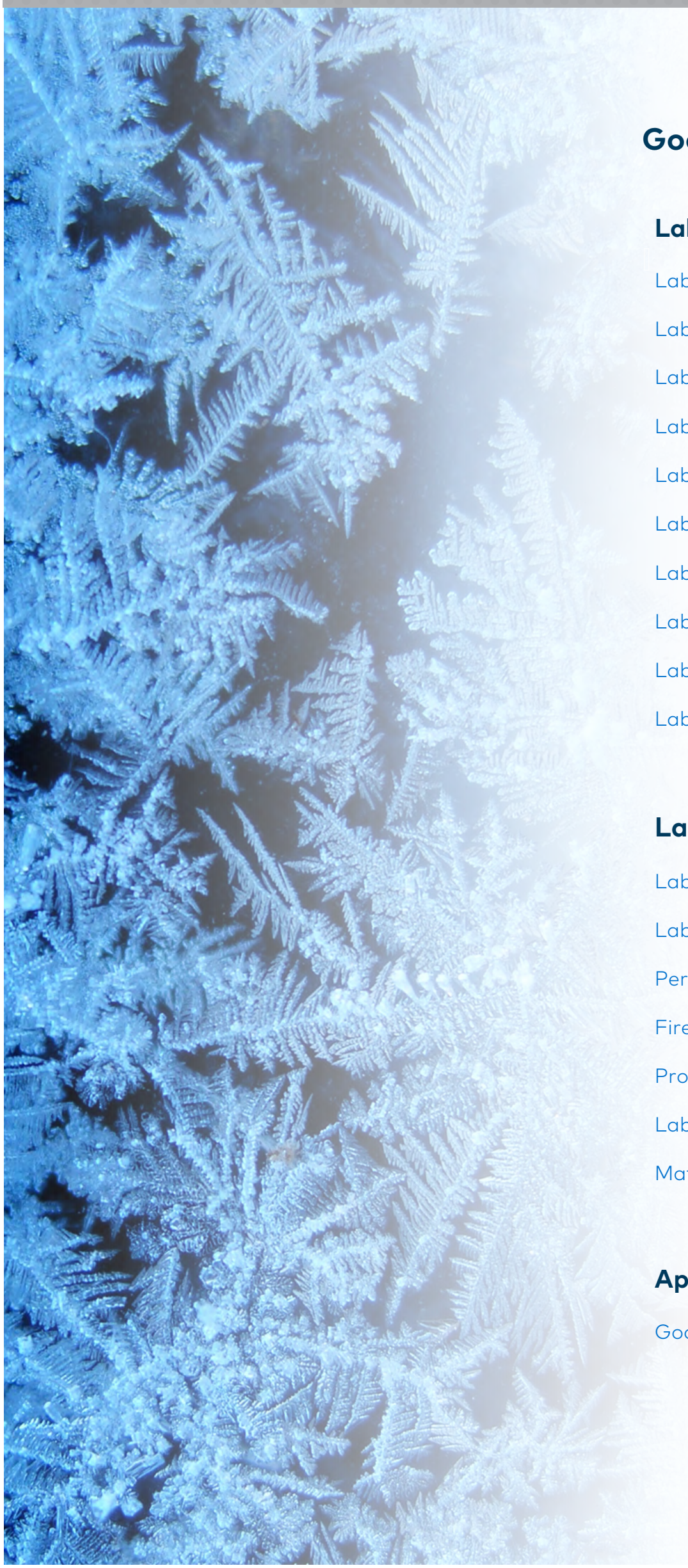
MAIN MENU



Goodwin College Chemistry Lab

Lab Menu

Lab Safety:

Appendix:

TIPS FOR USING YOUR CHEMISTRY LAB KIT

To successfully use and get the most learning and enjoyment out of your Chemistry lab kit here are some tips before you begin:

- Refer to your course's syllabus for the due date and sequence of your lab assignments. Your instructor may not assign all of the labs or may not assign them in the order they are presented in the lab manual, on the Student Portal, or your school's Learning Management System (LMS).

- As soon as you receive your kit, compare the contents of your kit to the inventory insert that comes in your kit. If you are missing any items or any item was broken during delivery call eScience Labs at 1-888-375-5487 or email Help@esciencelabs.com immediately for a replace-ment.

- There are a few materials that you will need to complete your experiments that are not provided in your kits such as distilled water. Review the Time and Materials section (after these tips) for a list of these items.

- Before performing an experiment, read through the procedure thoroughly so that you know what to expect when you begin to handle materials.

Enjoy your authentic chemistry lab experience!

eScience Labs®

LAB SAFETY

Always follow the procedure in your laboratory manual and these general rules:

LAB PREPARATION

- Please thoroughly read the experiment procedure before starting.
- If you have any doubt as to what you are supposed to be doing and how to do it safely, please STOP and then:
- Double-check the manual instructions.
- Check the eScience Labs website for updates and tips.
- Contact us for technical support by phone at 1-888-ESL-Kits (1-888-375-5487) or by email at Help@esciencelabs.com.
- Read and understand all labels on chemicals.
- If you have any questions or concerns, refer to the Material Safely Data Sheets (MSDS) available at the eScience Labs website. The MSDS (also called the SDS: safety data sheets) lists the dangers, storage requirements, exposure treatment and disposal instructions for each chemical.
- Consult your physician if you are pregnant, allergic to chemicals, or have other medical conditions that may require additional protective measures.

PROPER LAB ATTIRE

- Remove all loose clothing (jackets, sweatshirts, etc.) and always wear closed-toe shoes.
- Long hair should be pulled back and secured. All jewelry (rings, watches, necklaces, earrings, bracelets, etc.) should be removed.
- Safety glasses should be worn at all times. In addition, wearing soft contact lenses while conducting experiments is discouraged, as they can absorb potentially harmful chemicals.
- When handling chemicals, always wear the protective glasses, gloves, and apron provided in your safety kit.
- Do not eat, drink, chew gum, apply cosmetics, or smoke while conducting an experiment.

PERFORMING EXPERIMENTS

- Work in a well ventilated area and monitor experiments at all times, unless instructed otherwise.
- When working with chemicals:
- Never return unused chemicals to their original container to avoid contamination.
- Never place chemicals in an unmarked container to avoid identification or proper disposal problems.
- Always put lids back onto chemicals immediately after use to avoid contamination or potential hydration problems.
- Never ingest chemicals. If this occurs, seek immediate help. Call 911 or "Poison Control" 1-800-222-1222
- Never leave a heat source unattended.
- If there is a fire, evacuate the room immediately and dial 911.

LAB SAFETY

LAB CLEAN-UP AND DISPOSAL

- If a spill occurs, consult the safety data sheets (SDS) to determine how to clean it up.
- Never pick up broken glassware with your hands. Use a broom and a dustpan and discard in a safe area.
- Do not use any part of the lab kit as a container for food.
- Safely dispose of chemicals. If there are any special requirements for disposal, it will be noted in the lab manual.
- Each time you complete an experiment, wash hands and lab equipment thoroughly with soap and water.

Above all, use common sense. Read the manual carefully and pay close attention to the safety concerns prior to starting an experiment.

DISTILLED WATER

In addition to isopropyl alcohol, there is also a large amount of distilled water required for your experiments. The high volume required makes it necessary for you to purchase your own bottle of distilled water.

Please make this purchase prior to beginning your labs to ensure that you are prepared at all times. You will need approximately 1600 mL to complete the full set of lab experiments. A gallon of distilled water contains approximately 3785 mL, and will provide more than enough water to complete your labs. Distilled water can be purchased at most grocery stores for approximately $1.

MATERIALS SAFETY DATA SHEET

The SDS for every chemical in this kit is provided at eScience Labs MSDS. eScience Labs highly recommends that you download and print all of the SDS prior to starting your experiments. By doing so, you will ensure that you have all of the safety and cleanup information you need should you spill or encounter an accident during an experiment.

If you have questions about any of this information, please email eScience Labs at info@esciencelabs.com or call 888-ESL-KITS.

GREEN CHEMISTRY

- Green chemistry is division of chemistry which focuses on sustainable, high-quality, and environmentally-friendly experimental techniques. It proposes alternative procedures and materials which minimize the chemical impact on human health, reduce risk in the laboratory, and aims to eliminate the potential for environmental contamination. Green chemistry is becoming more and more common throughout chemistry. Familiarity of the tenets of this philosophy will help you perform experiments, as well as answer post-lab and real-world questions. The primary tenets are:

- Minimize or prevent waste whenever possible.
- Use catalysts rather than stoichiometric quantities to avoid using unnecessary chemical amounts.
- Avoid using chemical derivatives to decrease total waste created.
- Ensure efficient chemical processes.
- Use eco-friendly, biodegradable chemicals when possible.

LAB SAFETY

- Prepare for an experiment by collecting needed analytical tools prior to beginning the experiment. This reduces the possibility for hazardous waste development.
- Integrate recycled or upcycled materials whenever possible.
- Implement energy-efficient laboratory tools.

eScience Labs supports the green chemistry philosophy by integrating it into the core of all experimental procedures included in the manual. In doing so, students and teachers are offered a safe, reliable, and sustainable avenue for chemistry education. Additional green chemistry information is available online at http://www.epa.gov/greenchemistry/

FIRE HAZARDS

Several of the experiments require use of a Sterno® or flame to create the desired chemical reaction. Please be extra careful when working with flames to prevent burning yourself, lab-ware, or chemicals.

Thank you,

The eScience Team

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Lab 1

Introduction & Laboratory Safety Procedures

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LAB 1

Introduction to Laboratory Safety & Procedures

Learning Objectives

- **Understand the importance of safety in the chemistry laboratory.**
- **Learn the chemistry safety rules.**
- **Understand what to do in case of a chemistry laboratory accident.**
- **Demonstrate the safety rules by creating a safe chemistry laboratory environment.**

INTRODUCTION

The eScience Labs chemistry lab manual provides an authentic laboratory experience for distance and face-to face learners. Due to the unsupervised environment, safety is a top priority. Throughout the use of this manual, you will be asked to participate in some activities which may require the use of uncommon chemicals and equipment. These items can become potentially hazardous or dangerous if they are not handled appropriately and with proper caution. To ensure a safe chemistry laboratory experience, a list of safety rules have been developed and provided for you in this lab (as well as in subsequent lab procedures). For your personal safety, these rules should be followed at all times. In addition, you will be asked to sign a safety contract stating that you agree to abide by the eScience Labs chemistry lab safety rules. This activity will be completed in Exercise 1.

Figure 1: A good chemistry student like yourself will always do their research first, formulate a hypothesis, and never make a guesstimate.

Characters in Chemistry

Carl Sheele | 1742 - 1786

A brilliant pharmaceutical chemist who discovered many chemical elements, such as oxygen, molybdenum, tungsten, manganese, and chlorine. However, Scheele had a bad habit of sniffing and tasting any new substances he discovered. It is thought that his cumulative exposure to arsenic, mercury, lead, and hydrofluoric acid shortened his life.

Different guidelines exist for different actions within a laboratory. The following pages provide you with an overview of these guidelines. Keep in mind that if you are ever unsure, assume a chemical or tool may be potentially dangerous and handle it with care.

Introduction to Laboratory Safety and Procedures

LAB PREPARATION

- Please thoroughly read the lab exercise before starting.
- If you have any doubt as to what you are supposed to be doing and how to do it safely, please STOP and then:
- Double-check the manual instructions.
- Check the eScience Labs website for updates and tips.
- Contact us for technical support by phone at 1-888-ESL-Kits (1-888-375-5487) or by email at Help@esciencelabs.com.
- Read and understand all labels on chemicals.
- If you have any questions or concerns, refer to the Safety Data Sheets (SDS) available at the eScience Labs website. The SDS lists the dangers, storage requirements, exposure treatment, and disposal instructions for each chemical.
- Consult your physician if you are pregnant, allergic to chemicals, or have other medical conditions that may require additional protective measures.
- Never return unused chemicals to their original container to avoid contamination.
- Never place chemicals in an unmarked container to avoid identification or proper disposal issues.
- Always put lids back onto chemicals immediately after use to avoid contamination or potential hydration problems.
- Never ingest chemicals. If this occurs, seek immediate help.
- Call 911 or "Poison Control" 1-800-222-1222
- Never pipette anything into your mouth.
- Never leave a heat source unattended.
- If there is a fire, evacuate the room immediately and dial 911.

PROPER LAB ATTIRE

Introduction to Laboratory Safety and Procedures

Lab Clean-up and Disposal

- If a spill occurs, consult the MSDS to determine how to clean it up.
- Never pick up broken glassware with your hands. Use a broom and a dustpan and discard in a safe area.
- Do not use any part of the lab kit as a container for food.
- Safely dispose of chemicals. If there are any special requirements for disposal, it will be noted in the lab manual.
- When finished, wash hands and lab equipment thoroughly with soap and water.
- Once an experiment is concluded all of the remaining chemicals can be disposed of by pouring them down the drain with copious amounts of water. However, if the chemical is an acid or a base you will need to neutralize it before pouring it down the drain. A neutralization reaction is when an acid and a base react to form water and a salt and involves the combination of H^+ ions and OH^- ions to generate water. The neutralization of a strong acid and strong base has a pH equal to 7 (neutral and therefore safe).

Above all, use common sense. Read the manual carefully. Pay close attention to the safety concerns prior to starting an experiment.

Accidents and Injuries

- Cuts are to be washed with water and bandaged. Major cuts or other serious occurrences are to be seen by a medical professional.
- Burns are to be treated with ice water or cold tap water. Ointment is not to be applied. Serious burns are to be seen by a medical professional.
- If a chemical should splash in your eye(s) or on your skin, immediately flush with running water for at least 20 minutes.

Handling Chemicals

- All chemicals in the laboratory should be considered dangerous. Do not touch, taste, or smell any chemicals unless specifically instructed to do so.
- If you are instructed to smell a chemical you will use the technique called wafting. This technique consists of waving your hand over the container in such a way as to gently push the vapor towards you while keeping the chemical container away from your face (Figure 2).
- Always read the label twice before taking anything from a bottle to be sure you have the right substance. Serious accidents can occur if the wrong chemical is used. A good rule of thumb is, "READ THE LABEL TWICE; DO THE EXPERIMENT ONCE".
- When transferring reagents from one container to another, hold the containers away from your body.
- Handle flammable hazardous liquids over a pan to contain spills. Never dispense flammable liquids anywhere near an open flame or source of heat.
- Always pour acids into water. If you pour water into acid, the heat of reaction will cause the water to explode into steam, sometimes violently, and the acid will splatter.
- Never remove chemicals or other materials from the laboratory area.
- Take great care when transferring acids and other chemicals from one part of the laboratory to another. Hold them securely with both hands and walk carefully.

Introduction to Laboratory Safety and Procedures

Figure 2. Remember to always use the wafting technique to smell chemicals while conducting experiments.

Handling Glassware and Equipment

- Never handle broken glass with your bare hands. Use a brush and dustpan to clean up broken glass. Place broken or waste glassware in the designated glass disposal container.
- Fill wash bottles only with distilled water and use only as intended, e.g., rinsing glassware and equipment, or adding water to a container.
- When removing electrical plug from its socket, grasp the plug, not the electrical cord. Hands must be completely dry before touching an electrical switch, plug, or outlet.
- Examine glassware before each use. Never use chipped or cracked glassware. Never use dirty glassware.
- If you do not understand how to use a piece of equipment, ask for help. Do not immerse hot glassware in cold water; it may shatter.

Heating Substances

- Exercise extreme caution when using any type of heat source (e.g., gas burner, stove, or Sterno®). Take care that hair, clothing, and hands are a safe distance from the flame at all times. Do not put any substance into the flame unless specifically instructed to do so. Never reach over an exposed flame. Light burners only as instructed.
- Never leave a lit burner unattended. Never leave anything that is being heated or is visibly reacting unattended. Always turn the heat source off when not in use.
- Heated metals and glass remain very hot for a long time. They should be set aside to cool and picked up with caution. Use tongs or heat-protective gloves if necessary.
- Never look into a container that is being heated.
- Do not place a hot apparatus directly on the laboratory desk. Always use an insulating pad. Allow plenty of time for the hot apparatus to cool before touching it.
- In addition to these general guidelines, ALWAYS abide by any additional safety procedures provided by your instructor at the time of an activity.

Lab Safety Rules

- Conduct yourself in a responsible manner at all times in the area that you will be preforming the lab work.
- Follow all written and verbal instructions carefully. If you do not understand a direction or part of a procedure, ask before proceeding.
- Do not eat, drink, smoke, or chew gum in the area where you will be performing your lab work.
- Do not use any laboratory equipment as containers for food or beverages.
- Perform only those experiments issued in your eScience Labs kit. Unauthorized experiments are prohibited.
- Read all procedures thoroughly before starting the lab.
- Observe good housekeeping practices. Work areas should be kept clean and tidy at all times. Bring only your laboratory instructions, worksheets, and/or reports to the work area. Other materials (books, purses, backpacks etc.) should be stored in another area.
- Know the locations and operating procedures of all safety equipment including the first aid kit, eyewash station (sink), safety shower, fire extinguisher, and fire blanket. Know where the fire alarm and the exits are located.
- Always work in a well-ventilated area.
- Be alert and proceed with caution at all times in the area where you will be preforming the lab work.
- Never point a test tube or any vessel that you are heating at yourself or anyone else as it may explode.
- Do not "shake down" thermometers or use it as a stirring rod. Thermometers break easily.
- Dispose of all chemical waste properly. Solid chemicals, metals, matches, filter paper, and all other insoluble materials are to be disposed of in the proper waste containers.
- Labels and equipment instructions must be read carefully before use. Set up and use the pre-scribed apparatus as directed in the laboratory instructions.
- Keep hands away from face, eyes, mouth and body while using chemicals.
- Wash your hands with soap and water after performing all experiments. Clean (with detergent), rinse and wipe dry all work surfaces (including the sink) and apparatus at the end of the experiment. Return all equipment clean and in working order to the proper storage area.
- Experiments must be personally monitored at all times.

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Lab 1: Pre-Lab Questions

1. Mapping Laboratory Safety: Use the space provided to draw a map identifying your safety elements and where you will complete your lab experiments. If you do not possess a safety item indicate where the closest one that you can use is located. Use the list below to identify items on your map.

a. Clean, dry, uncluttered, laboratory work space

b. Window for ventilation

c. Trash can

d. Sink (eye wash station)

e. First Aid Kit

f. Fire Extinguisher

g. Phone

h. Storage location for your eScience Lab materials when not in use

i. Computer

j. Blanket that can be used to smother a fire (fire blanket)

k. Shower that can be used to wash off spilled chemicals (safety shower)

l. Exit door (Diagram on your map the route you will use to leave the building encase of an emergency).

m. Smoke detector

n. Dust pan and broom

Lab 1: Pre-Lab Questions

2. What would you do if you spilled a tiny amount of hydrochloric acid on your hand?

3. Describe a possible danger than can occur if you put food, soda, coffee, pencil erasers, etc. in your mouth in while conducting your chemistry experiments?

4. What information is given on a Material Safety Data Sheet?

5. List the contents from your first aid kit that you will use to treat a minor cut.

6. List the contents from your first aid kit that you will use to treat a minor burn.

7. List optimum clothing, shoes, jewelry and hair arrangement for your chemistry labs.

Lab 1: Pre-Lab Questions

8. What is the date that your fire extinguisher was last examined?

9. How old are the batteries in your smoke detector?

Exercise 1

Lab 1
Lab Safety Contract

eScience Labs feels confident that with the knowledge of the safety procedures you will be able to create and maintain a safe laboratory experience. With your cooperation you will be able to eliminate, prevent, and correct possible hazards.

Please complete the questions below and, if you comply with the safety rules, sign this agreement. If you do not comply, please note this within your agreement. Note that if you do not choose to comply, you will not be included in eScience Labs' safety and liability coverage.

Remember to return this safety contract to your Chemistry Instructor.

Questions

1. Do you wear contact lenses? ☐ YES ☐ NO

2. Are you color blind? ☐ YES ☐ NO

3. Do you have any allergies? ☐ YES ☐ NO

If YES, list specific allergies:_____

Student Agreement

I, _____ (your name) have read and choose to comply with all of the safety rules set forth in the Safety Procedures. By signing this contract I realize that I must obey these rules to insure my own safety. I will cooperate to the fullest extent with my instructor and the protocols in my lab kit to maintain a safe lab environment. I will also closely follow any oral and written instructions additionally provided by my instructor as part of a specific activity. I am aware that any violation of this safety contract that results in unsafe conduct in the laboratory or misbehavior on my part, may result in consequences as specified by my instructor.

_____
Student Signature

_____
Date

_____
Course Name

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Experiment 1

Lab 1
Neutralization of Acids and Bases

Experiment Inventory

Materials

(8) pH test Strips (with color reference chart)

5 mL 4.5% Acetic Acid (vinegar), $C_2H_4O_2$

0.5 g Sodium Bicarbonate (baking soda), $NaHCO_3$

Permanent Marker

Distilled Water

Labware

(2) Pipettes

(2) Weigh Boats

(1) 10 mL Graduated Cylinder

(1) 100 mL Graduated Cylinder

(1) 250 mL Beaker

Note: You must provide the materials listed in red.

EXPERIMENT 1: NEUTRALIZATION OF ACIDS AND BASES

In this experiment you will learn how to properly neutralize and dispose of acidic and basic solutions. The procedure will be part of lab clean up processes during your introductory chemistry experience

PROCEDURE

- [ ] 1. Use the permanent marker to label two of the weigh boats as A and C.
- [ ] 2. Use the permanent marker to label a 250 mL beaker B.
- [ ] 3. Set the containers in the order A, B, and C.
- [ ] 4. Use your 10 mL graduated cylinder to measure and pour 5 mL of water into weigh boat "A".
- [ ] 5. Add 0.5 g sodium bicarbonate to beaker "B".
- [ ] 6. Use the 100 mL graduated cylinder to measure and pour 100 mL of water into beaker "B". Gently pipette the solution up and down until the sodium bicarbonate is fully dissolved in the water. Leave this pipette in beaker B to mix the solution in future steps.
- [ ] 7. Use the 10 mL graduate cylinder to measure and pour 5 mL acetic acid solution to weigh boat "C".
- [ ] 8. Use the pH test strips to determine if the substances in containers A - C are acidic, basic, or neutral. Briefly dip an unused strip of the pH paper in each of the weigh boats. Look at the pH color reference chart to determine the acidic, basic, or neutral nature of the solutions tested. Record your color results immediately after dipping, for an accurate reading, in **Table 1** of your experiment data sheet.
- [ ] 9. Pipette 1 mL of acetic acid from weigh boat "C" into beaker "B". Ensure the transfer pipette does not touch the solution in beaker "B".
- [ ] 10. Gently pipette the solution in beaker B with its pipette to mix.
- [ ] 11. Test the pH of beaker "B" using new pH paper and immediately record your result in **Table 2** of your data sheet.
- [ ] 12. Repeat Steps 9-11 four more times until all the acetic acid has been added to beaker "B".

Data Sheet

Lab 1
Experiment Data Sheet

Table 1: Initial pH Test Results

Container	Chemical Contents	pH Results	Additional Observations
A			
B			
C			

Table 2: Neutralization

Amount of Acid (mL)	pH Results
1	
2	
3	
4	
5	

Lab 1: Post-Lab Questions

1. Most of the chemicals included in your Introductory Chemistry Lab kit can be discarded down a drain with copious amounts of water. Describe a situation in which you would need to neutralize (pH 7) a chemical before discarding down a drain.

2. Why should one add acid to water rather than add water to acid when preparing solutions?

Lab 1: Post-Lab Questions

3. At what point was the solution in beaker "B" neutralized?

4. Address the following scenarios: if a stronger solution of sodium bicarbonate was used in beaker B, would it require more or less acetic acid to neutralize and why? If a weaker solution of sodium bicarbonate was used in beaker B, would that solution require more or less acetic acid to neutralize it and why?

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Lab 2

Introduction to Science

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LAB 2:

Introduction to Science

Learning Objectives

- Understand the scientific method, including identifying variables and controls
- Practice making conversions including significant digits and scientific notation

INTRODUCTION

What is science? Science is a systematic approach to solving problems and explaining phenomena that occur in the universe. A scientific investigation (Figure 1) typically begins with observations. Once enough observations are collected, a testable statement called a **hypothesis** can be constructed to explain the observations. Experiments are then performed to gather data that will be interpreted to either support or reject the hypothesis. If supported, more tests are performed to further support the hypothesis and gather a large body of evidence. A large body of evidence gathered by scientific investigations from multiple, reliable sources that explain a phenomenon is called a **scientific theory**. A large body of evidence may also support a **scientific law**. A scientific law summarizes a phenomenon that always occurs under certain conditions. For example, Newton's Law of Gravity mathematically summarizes the natural attractive force between two objects in the universe.

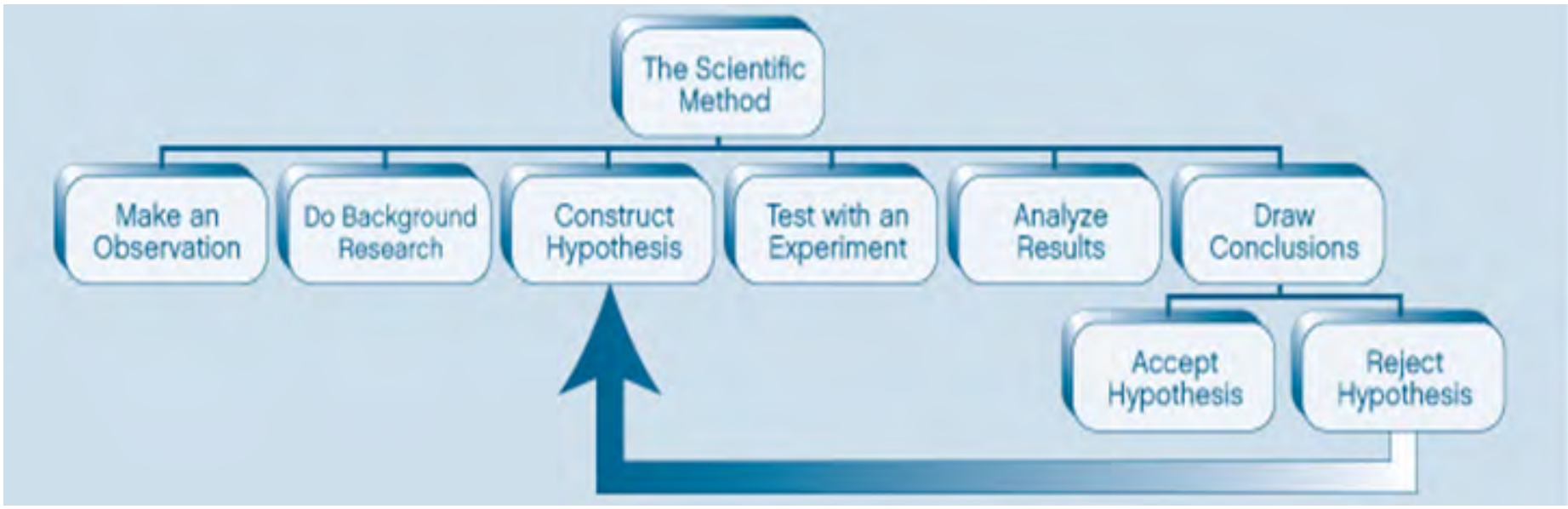

Figure 1: A scientific investigation can begin with an observation, which may lead to an experiment. If the hypothesis tested in an experiment is not supported by results, it can be revised and tested again.

OBSERVATIONS

Scientific investigations typically begin with observations. For example, suppose you observe a plant growing towards a window. This observation could be the first step in designing an experiment. Remember that observations may spark scientific investigations, but are used to analyze data. Observations can be quantitative (measurable) or qualitative (immeasurable, observational). **Quantitative** observations allow scientists to record findings as data and leave little room for subjective error.

Consider a plant growing towards the window. You can make quantitative observations that there are 15 leaves and each is 3.7 cm in diameter. Both the quantity and the diameter can be precisely measured. You can also make qualitative observations that they are green, circular, or smooth. The color and texture are not numerically measured, and may vary based on the individual's perception or background.

Quantitative observations are generally preferred in science because they involve "hard" data. Because of this, many scientific instruments, such as microscopes and scales, have been developed to alleviate the need for qualitative observations. Rather than observing that an object is large, we can now identify specific mass, shapes, structures, etc.

There are still many situations, as you will encounter throughout this lab manual, in which qualitative observations are useful. **Qualitative** observations cannot be measured. Instead, they rely on human sensory perceptions. The nature of these observations makes them more subjective and susceptible to human error. However, qualitative observations are still able to provide useful information. For example, noticing the color change of a leaf or the change in smell of a compound are important observations, and can provide a great deal of practical information.

Introduction to Science

DEVELOPING A HYPOTHESIS

Once observations are made, the next step is to develop a **hypothesis**. A hypothesis is a testable statement describing what the scientist predicts will happen in an experiment. In other words, it is a proposed explanation for an event based on observations. For every hypothesis, a scientist also develops a null hypothesis, but only one can be supported by results. A null hypothesis is also a testable statement that, if supported by results, indicates the hypothesis is not supported. For example, hypotheses can be created to explain factors that affect plant growth (Figure 2):

Figure 2: **Factors such as sunlight, soil composition, or humidity can affect plant growth**

Supported	Not Supported
Hypothesis: If plants are grown in soil with added nutrients, then they will grow at a faster rate than plants grown without added nutrients	**Null Hypothesis:** If plants are grown in soil with added nutrients, then they will grow at the same rate as plants grown without added nutrients

If plants grow at a faster rate when nutrients are added, then the hypothesis is accepted and the null hypothesis is rejected.

TESTING A HYPOTHESIS

There are often many ways to test a hypothesis. However, three guidelines should be followed during an experiment for results to be valid.

1. The experiment must be replicable.
2. Only one variable should be tested per experiment.
3. Controls must be included.

Experiments must be **replicable** to create valid theories. In other words, the procedure must always be diligently recorded, and an experiment must provide precise results over multiple trials (Figure 3). **Precise results** are those that have very similar values (e.g., 85, 86, and 86.5) over multiple trials (Figure 3). In contrast, **accurate results** are those that demonstrate what you expected to happen (e.g., you expect the test results of three students' tests to be 80%, 67%, and 100%).

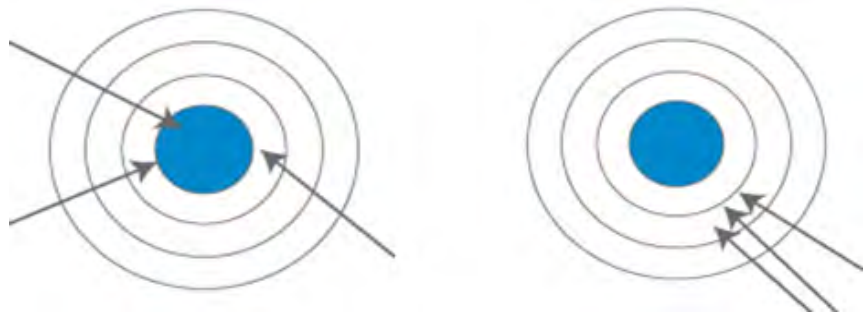

Figure 3. **Left: Accurate results all hit the bulls-eye on a target. Right: Pre-cise results may not hit the bulls-eye, but they all hit the same region.**

The following example demonstrates the significance of experimental replicability. Suppose you conduct an experiment and conclude that ice melts in 30 seconds when placed on a burner, but you do not record your procedure or define the exact variables included. The conclusion that you draw will not be recognized in the scientific community because other scientists cannot repeat your experiment and find the same results. What if another scientist tries to repeat your ice experiment, but does not turn on the burner or uses a larger ice chunk? The results will not be the same, because the experiment was not repeated using the same exact procedure. In order for results to be valid, repeated experiments must follow the original experiment exactly. Using this technique, multiple trials performed in this manner should yield comparable results.

VARIABLES

Variables are defined, measurable components of an experiment. Controlling variables in an experiment allows a scientist to quantify changes that occur. This allows for focused results to be measured and refined conclusions to be drawn. There are two types of variables: independent and dependent variables.

Independent variables are variables that scientists select to change within the experiment. For example, the time of day, amount of substrate, etc. can all be independent variables. Independent variables are also used by scientists to develop hypotheses.

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The “if” part of the hypothesis describes the independent variable and how the scientist will manipulate it. The dependent variable is reflected in the “then” part of the hypothesis. For example, if there is a change in the independent variable, then the dependent variable will also change. Independent variables are always placed on the x-axis of a chart or graph. For example, the independent variable in the hypothesis, “If plants are grown in soil with added nutrients, then they will grow faster than plants grown without added nutrients,” is soil with added nutrients. Experiments can only have one independent variable. In this way, scientists can determine if altering the independent variable is the reason for obtaining a different result. Scientists would not be able to conclusively determine which change affected the data if more than one independent variable is changed in an experiment.

Dependent variables are variables which are observed in relationship to the independent variable. Any changes observed in the dependent variable are caused by the changes in the independent variable. In other words, they depend on the independent variable. Common examples of this are: reaction rate, color change, etc. There can be more than one dependent variable in an experiment. Dependent variables are placed on the y-axis of a chart or graph.

CONTROLS

A **control** is a sample of data collected in an experiment that is not exposed to the independent variable. The control sample reflects the factors that could influence the results of the experiment, but do not reflect the planned changes that might result from manipulating the independent variable. Controls must be identified to eliminate compounding changes that could influence results. Often, the hardest part of designing an experiment is determining how to isolate the independent variable and control all other possible variables. Scientists must be careful not to eliminate or create a factor that could skew the results. For this reason, taking notes to account for unidentified variables is important. This might include factors such as temperature, humidity, time of day, or other environmental conditions that may impact results.

There are two types of controls: positive and negative. **Negative controls** are data samples in which you expect no change to occur. They help scientists determine that the experimental results are due to the independent variable, rather than an unidentified or unaccounted variable. For example, suppose you need to culture (grow) bacteria and want to include a negative control. You could create this by streaking a sterile loop across an agar plate. Sterile loops should not create any microbial growth; therefore, you expect no change to occur on the agar plate. If no growth occurs, you can assume the equipment used was sterile. However, if microbial growth does occur, you must assume that the equipment was contaminated prior to the experiment and must redo the experiment with new materials.

Alternatively, **positive controls** are data samples in which you do expect a change. Let’s return to the growth example and create a positive control. To do this, you now use a sterile loop to streak a plate with a bacterial sample that you know grows well on agar (such as E. coli). If bacteria grow, you can assume that the bacteria sample and agar are both suitable for the experiment. However, if bacteria do not grow, you must assume that the agar or bacteria have been compromised, the agar is inhibiting growth, or the bacteria in the sample are not viable.

COLLECTING AND PRESENTING DATA

A scientific investigation also requires data collection. This may reflect what occurred before, during, or after an experiment. Collected data help reveal experimental results. Data should include all relevant observations, both quantitative and qualitative.

After results are collected, they can be analyzed. Data analysis often involves a variety of calculations, conversions, graphs, tables, etc. A common task a scientist faces is unit conversion. Units are often displayed in an increment that must be converted. For example, suppose half of your data is measured in seconds, but the other half is measured in minutes. It will be difficult to understand the relationship between the data if the units are not equivalent. Figure 4 shows a sample calculation of converting seconds to minutes.

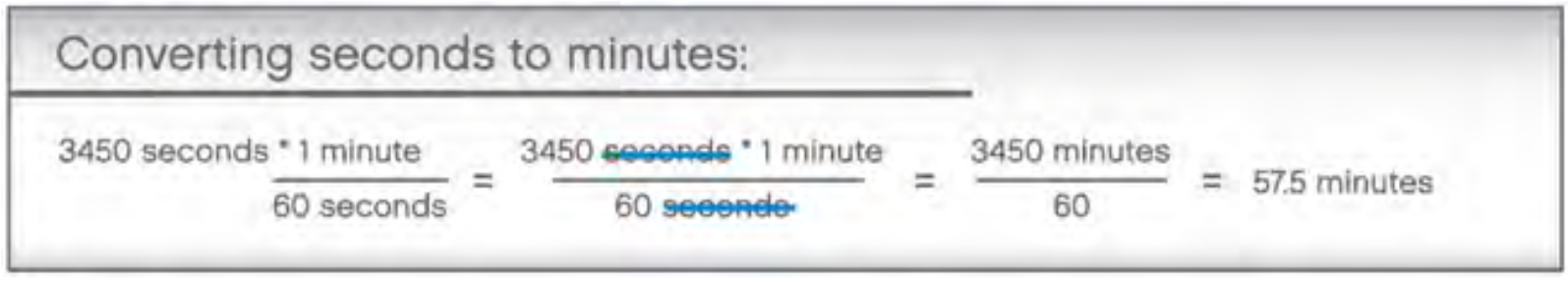

Figure 4: To convert the unit of time seconds into minutes, multiply 3450 seconds by 1/60 minutes, because there are 60 seconds in one minute.

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SIGNIFICANT DIGITS

When calculating a unit conversion, significant digits must be accounted for. Significant digits are the digits in a number or answer that describe how precise the value actually is. Consider the rules in Table 1.

Table 1. Significant Digits Rules

Rule	Examples
Any non-zero number (1 - 9) is always significant.	• 45 has two significant digits • 3.99 has three significant digits
Any time a zero appears between significant numbers, the zero is significant.	• 4005 has four significant digits • 0.3400000009 has ten significant digits
Zeros that are ending numbers after a decimal point or zeros that are after significant numbers before a decimal point are significant.	• 45.00 has four significant digits • 15000.00 has seven significant digits
Zeros that are used as placeholders before other digits are NOT significant digits.	• .0000000897 has three significant digits
A zero at the end of a number with no decimal can be a significant digit. *To avoid uncertainty, numbers can be written using scientific notation.	• 50 cm exactly has two significant digits (not rounded) • 6200 can have 2, 3, or 4 significant digits (e.g., 6.2×10^3 has 2, 6.20×10^3 has 3, and 6.200×10^3 has 4)

Addition and subtraction problems should result in an answer that has the same number of significant decimal places as the least precise number in the calculation. Multiplication and division problems should keep the same total number of significant digits as the least precise number in the calculation. For example:

- **Addition problem: 12.689 + 5.2 = 17.889 → round to 17.9 (the least precise number has one decimal place)**
- **Multiplication problem: 28.8 × 54.76 = 1577.088 → round to 1580 (three significant digits)**

Scientific notation is another common method used to report a number. Scientific data is often very large (e.g., the speed of light) or very small (e.g., the diameter of a cell). Scientific notation provides an abbreviated expression of a number so that scientists don't get caught up counting a long series of zeroes.

There are three parts to scientific notation: the base, the coefficient and the exponent (Figure 5). Base 10 is almost always used and makes the notation easy to translate. The coefficient is always a number between 1 and 10, and uses the significant digits of the original number. The exponent indicates whether the number is greater or less than 1 and can be used to "count" the number of digits the decimal must be moved to translate the number to regular notation. A negative exponent indicates movement of the decimal to the left, while a positive one indicates movement to the *right*.

For example, the number 5,600,000 can be written in scientific notation as 5.6×10^6. The coefficient is 5.6, the base is 10, and the exponent is 6. If you multiply 5.6 by 10 six times, you will arrive at 5,600,000. Note the exponent, 6, is positive because the number is larger than one. Alternatively, the number 0.00045 must be written using a negative exponent. To write this number in scientific notation, determine the coefficient. Remember that the coefficient must be between 1 and 10. The significant digits are 4 and 5. Therefore, 4.5 is the coefficient. To determine the exponent,

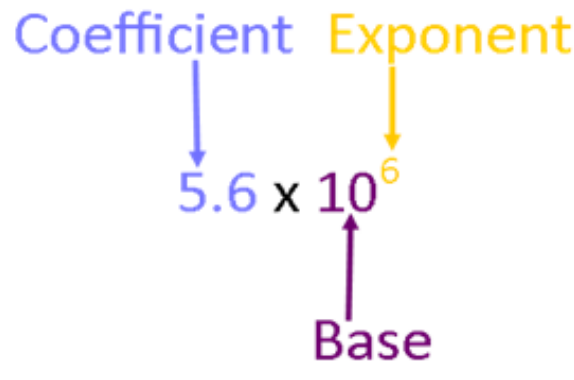

Figure 5: The exponent equals the number of decimal places moved until the coefficient is a number between 1 and 10.

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count how many places you must move the decimal over to create the original number. Moving to the left, you have 0.45, 0.045, 0.0045, and finally 0.00045. Since you moved the decimal 4 places to the left, the exponent is -4. Written in scientific notation, you have 4.5×10^{-4}.

Although these calculations may feel laborious, a well-calculated presentation can transform data into a format that scientists can more easily understand and evaluate. Some of the most common methods of data presentation are tables and graphs.

Table: An organized summary of the data collected. Only include information relevant to the hypothesis (e.g., don't include the color of the plant if it isn't relevant to what is being tested). Always include a clearly stated title, label your columns and rows, and include the units of measurement.

Graph: A visual representation of the relationship between the independent and dependent variable. Graphs are useful in identifying trends and illustrating findings. The independent variable is always graphed on the x-axis (horizontal), with the dependent variable on the y-axis (vertical).

Line Graph: Shows the relationship between variables using plotted points that are connected with a line. There must be a direct relationship and dependence between each point connected. More than one set of data can be presented on a line graph. Figure 6 uses the data from Table 2.

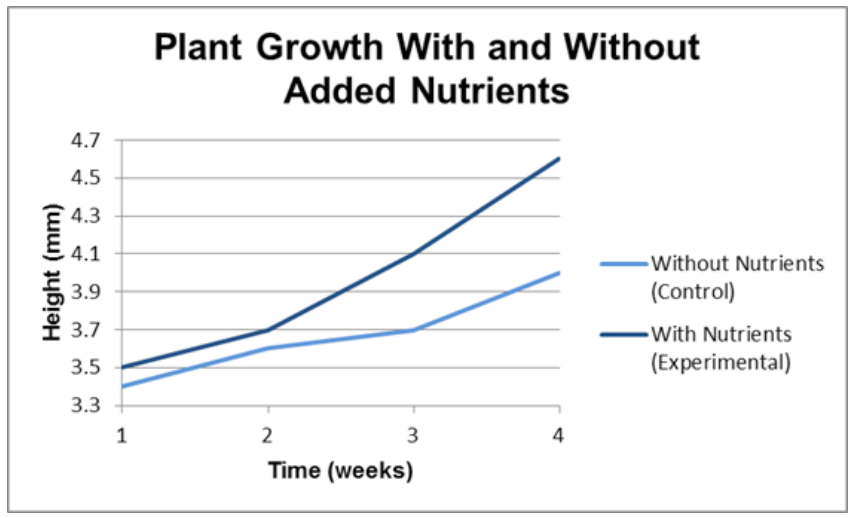

Figure 6. Sample **line graph. Line graphs are best used to show how changes occur for a variable over time.**

Bar Graph: Compares results that are independent from each other, as opposed to a continuous series. Since the results from our previous example are continuous, they are not appropriate for a bar graph. Figure 7 shows the top speeds of four cars. Since there is no relationship between each car, each result is independent and a bar graph is appropriate.

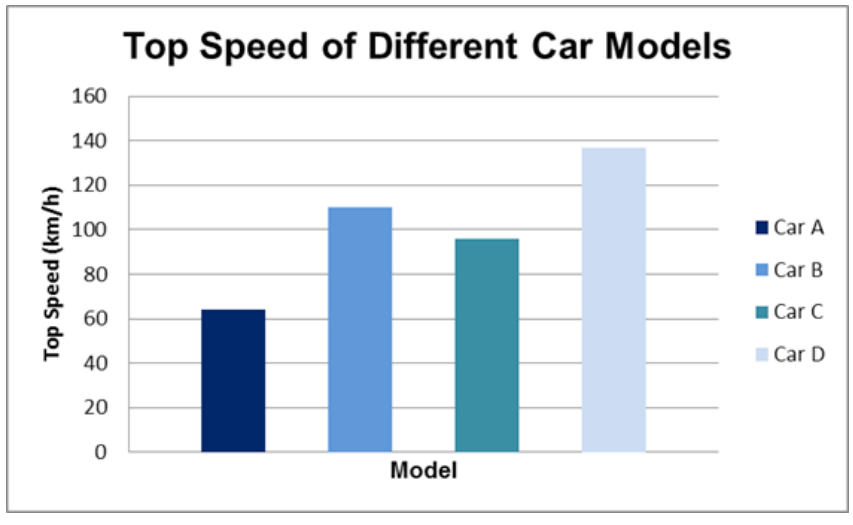

Figure 7. Sample bar graph. Bar graphs are best for demonstrating comparisons between categories or trial events.

Interpretation: After compiling the data, scientists **analyze** the data to determine if the experiment supports or refutes the hypothesis. If the hypothesis is supported, you may want to consider additional variables that should be examined. If your data does not provide clear results, you may want to consider running additional trials or revising the procedure to create a more precise outcome. For instance, was the amount of water and sunlight consistent between groups of plants - or, were all four cars driven on the same road?

Table 2. Plant Growth with and without Added Nutrients

Variable	Height Wk. 1 (mm)	Height Wk. 2 (mm)	Height Wk. 3 (mm)	Height Wk. 4 (mm)
Control (without nutrients)	3.4	3.6	3.7	4.0
Independent (with nutrients)	3.5	3.7	4.1	4.6

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PERCENT ERROR

One way to analyze data is to calculate **percent error**. Many experiments perform trials that calculate known values. When this happens, you can compare experimental results to known values and calculate percent error. Low percent error (< 5%) indicates that results are probably accurate, and high percent error (> 20%) indicates that results may be inaccurate. The formula for percent error is:

$$Percent\ Error = \frac{|Experimental\ Value - Actual\ Value|}{Actual\ Value} \times 100$$

Note: The brackets flanking the numerator indicate **absolute value**. This means that the number in the equation is always positive.

WRITING A LAB REPORT

Scientific investigations provide a great foundation to conduct scientific reasoning. The more data and observations a scientist makes, the more they will be able to accurately reason through natural phenomena. Scientific reasoning helps society think through difficult concepts and determine solutions. For example, scientific reasoning can be used to create a response to the changing global climate, develop medical solutions to health concerns, or even learn about subatomic particles and tendencies.

In general terms, a **lab report** is a scientific paper describing the premise of an experiment, the procedures taken, and the results of the study (Figure 8). It provides a written record of what took place to help others learn and expedite future experimental processes. Though most lab reports go unpublished, it is important to write a report that accurately characterizes the experiment performed. Table 3 summarizes the components of a typical lab report.

Figure 8. Lab reports are an important part of science, providing a way to report conclusions and ideas.

Table 3: Lab Report Components

Lab Report Section	Purpose
Title	A short statement summarizing the topic.
Abstract	A brief summary of the methods, results and conclusions. It should not exceed 200 words and should be the last part written.
Introduction	An overview of why the experiment was conducted. It should include: · Background - Provide an overview of what is already known and what questions remain unresolved. Be sure the reader is given enough information to know why and how the experiment was performed. · Objective - Explain the purpose of the experiment (i.e. "I want to determine if taking baby aspirin every day prevents second heart attacks.") · Hypothesis - This is your "prediction" as to what will happen when you do the experiment.
Materials and Methods	A detailed description of what was used to conduct the experiment, what was actually done (step by step), and how it was done. The description should be exact enough that someone reading the report can replicate the experiment.
Results	Data and observations obtained during the experiment. This section should be clear and concise. Tables and graphs are often appropriate in this section. Interpretations should not be included here.
Discussion	Data analysis interpretations and experimental conclusions. · Discuss the meaning of your findings. Look for common themes, relationships, and points that perhaps generate more questions. · When appropriate, discuss outside factors (e.g. temperature, time of day, etc.) that may have played a role in the experiment. · Identify what could be done to control for these factors in future experiments.
Conclusion	A short, concise summary that states what has been learned.
References	You should reference any articles, books, magazines, interviews, newspapers, etc. that were used to support your background, experimental protocols, discussions and conclusions. Your references should be written in the APA format. The following web site is a helpful citation maker http://www.citationmachine.net/apa/cite-a-book

Lab 2 Exercises

EXERCISE 1: UNIT CONVERSION

Use Figure 9 to convert each value into the designated units. Submit this exercise to your professor when completed.

1. 46.0 g = _____ kg

2. 57 seconds = _____ minutes

3. 13.5 cm = _____ inches

4. 47 °F = _____ °C

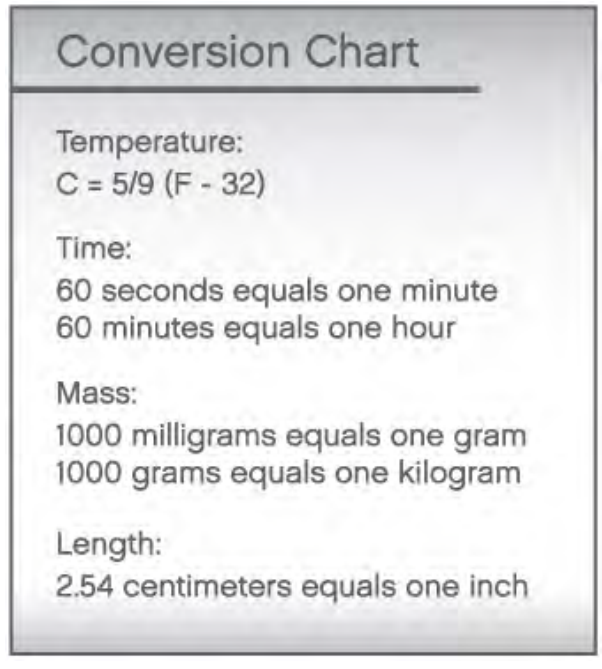

Figure 9. **Conversions for temperature, time, mass and length.**

EXERCISE 2: SIGNIFICANT DIGITS AND SCIENTIFIC NOTATION

Part 1: Determine the number of significant digits in each number and write out the specific significant digits. Submit Part 1 and Part 2 to your professor when completed.

1. 405000

2. 0.0098

3. 39.999999

4. 13.00

5. 80,000,089

6. 55,430.00

7. 0.000033

8. 620.03080

Part 2: Convert each regular number into scientific notation.

1. 70,000,000,000

2. 0.000000048

3. 67,890,000

4. 70,500

5. 450,900,800

6. 0.009045

7. 0.023

Lab 2 Exercises

EXERCISE 3: EXPERIMENTAL VARIABLES

Determine the variables tested in each of the following experiments. If applicable, determine and identify any positive or negative controls. Submit this exercise to your instructor upon completion.

1. A study is being done to test the effects of habitat space on the size of fish populations. Different sized aquariums are set up with six goldfish in each one. Over a period of six months, the fish are fed the same type and amount of food. The aquariums are equally maintained and cleaned throughout the experiment. The temperature of the water is kept constant. At the end of the experiment the number of surviving fish are surveyed.

   A. Independent Variable:

   B. Dependent Variable:

   C. Controls:

2. To determine if the type of agar affects bacterial growth, a scientist cultures on four different types of agar. Five petri dishes are set up to collect results:

   • One with nutrient agar and *E. coli*

   • One with mannitol-salt agar and *E. coli*

   • One with MacConkey agar and *E. coli*

   • One with LB agar and *E. coli*

   • One with nutrient agar but NO *E. coli*

   A. Independent Variable:

   B. Dependent Variable:

   C. Controls:

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Labs

Lab 3

Measuring Heats of Ractions

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LAB 3

Measuring Heats of Reactions

Learning Objectives

- **Calculate thermal energy based on the equation Q = C x m x ΔT**
- **Construct and use a calorimeter to experimentally determine the change in enthalpy**
- **Determine the identity of an unknown metal by investigating its specific heat**

INTRODUCTION

Heat is not the same thing as temperature, even though the two words are commonly used to talk about the same thing. Heat is the name for the energy that transfers from one object to another because of a temperature difference between the objects.

Temperature is the property of matter which reflects the quantity of kinetic energy of the particles. There are several standardized scales used to measure this value (e.g., Kelvin, Celsius, and Fahrenheit). Temperature can be used to calculate heat by looking at the change in temperature (Figure 1). A good way to define temperature is by the following definition:

Temperature is a property which reflects heat energy, and can change when two objects with different heat come into contact in an effort to create thermal equilibrium.

Figure 1. Simple thermometers utilize thermodynamic properties to measure the temperature of the environment.

THERMAL EQUILIBRIUM

When two bodies (e.g., a piece of metal and water) are at different temperatures, they exchange energy in the form of heat. Through this process their temperatures will gradually equalize. When a common temperature is reached and heat is no longer transferred, the bodies are said to be in thermal equilibrium (Figure 2).

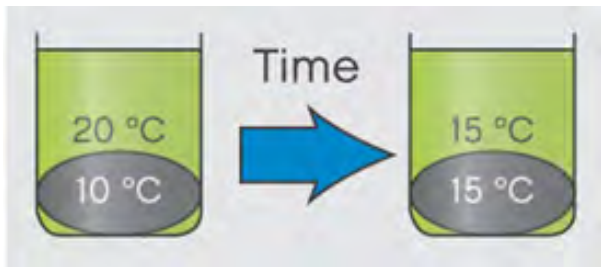

Figure 2. Given enough time the metal and the water will reach a state of thermal equilibrium and thus have the same temperature.

SPECIFIC HEAT CAPACITY

Specific heat is a physical property of matter. A physical property is a characteristic that can be observed or measured without changing the composition of the sample. All matter has a temperature associated with it. The temperature of matter is a direct measure of the motion of the molecules (kinetic energy of the particles which comprise it). The greater the motion, the higher the temperature. Every substance also has the ability to absorb heat. If enough heat is absorbed, the temperature of that substance will rise.

Specific heat capacity is the amount of heat per unit mass required to raise the temperature of a substance by one degree Kelvin (Figure 3). The SI unit used to measure specific heat capacity is joules per kilogram Kelvin (J/kg K). However, you may also see specific heat capacity expressed as calories per gram degrees.

Measuring Heats of Ractions

Figure 3. Water and sand have different specific heat capacities.

Energy can be calculated via the relationship between heat and temperature change. This relationship can be described by the equation:

$$q = C_s \times \Delta T$$

or

$$q = m \times C \times \Delta T$$

In this equation:

- **q = the energy in joules or calories (amount of heat being transferred to or from a substance)**
- **C (or C_p) = specific heat (heat capacity)**
- **C_s = heat capacity**
- **m = mass in grams**
- **Δ = symbol that indicates change**
- **T = temperature [°C + 273 = Kelvin (K)]**
- **ΔT = change in temperature (Final temperature minus Initial temperature = ΔT)**
- **x = multiplication symbol**

Every substance has a different specific heat capacity (Table 1). For example, the specific heat of water is one calorie/gram °C, or, 4.186 joules/gram °C. This value is higher than most other common substances. As a result, water plays a very important role in temperature regulation and is the reason water is used as a medium in calorimetry (Figure 4).

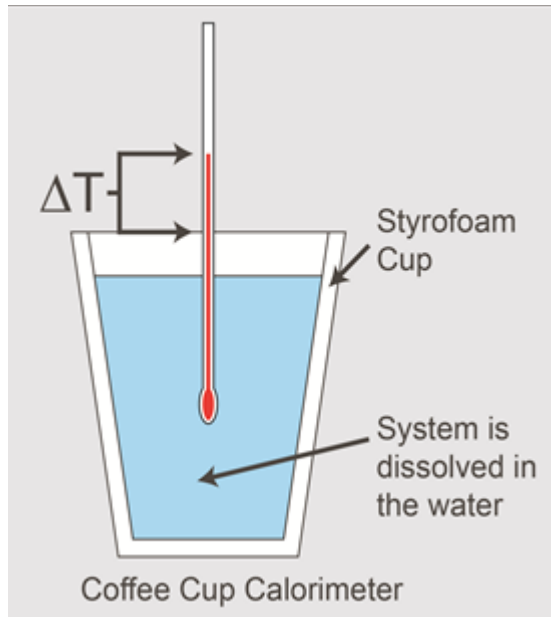

Figure 4. **A calorimeter can be as simple as a polystyrene (Styrofoam™) cup with a lid. The cup is filled with a known volume of water and a thermometer is inserted through the lid of the cup so that its bulb is below the water surface.**

Measuring Heats of Ractions

Table 1: Specific Heat Capacity of Common Substances

Substance	Specific Heat Capacity C_{sp} (J/g∘C)
$H_2O_{(l)}$	4.184
Ice at 0 °C	2.010
Steam at 100 °C	2.010
Aluminum	0.900
Chromium	0.448
Copper	0.385
Lead	0.160
Magnesium	1.017
Manganese	0.479
Tin	0.213
Zinc	0.388

CALORIMETRY

Scientists can measure the energy content of food by burning the food (combustion reaction). To a scientist, a calorie is the heat flow (amount of energy) needed to raise the temperature of 1 gram of water 1°C (1°C + 273 = 274 Kelvin (K)). A **calorimeter** is typically used to determine the amount of calories in a substance (Figure 4). The calorimeter uses the **Law of Conservation of Energy**, which states that energy is never created or destroyed but is transferred between objects. Calorimetry is a way to measure the heat that is generated or consumed by a substance during a chemical reaction or physical change. If heat is absorbed, it is an **endothermic process**. If heat is generated, it is an **exothermic process**. Most reactions involve some amount of heat transfer. Therefore, calorimetry has industrial applications in pharmaceutical, chemistry, and biological fields.

Student labs that explore calorimetry often use a Styrofoam™ calorimeter due to its ability to minimize heat exchange with the outside environment and reliable insulation. This device is also used to contain the reaction and provide an environment with either constant pressure or constant volume. The heat capacity (C) of a calorimeter (the amount of heat required to raise the temperature of a calorimeter by one Kelvin) should also be determined prior to the experiment. This can be done by transferring heat into a calorimeter and measuring how much the internal temperature increases. The formula to determine this value is:

$$C = \frac{Amount\ of\ Energy\ Inpuy\ (Joules)}{Resulting\ Temperature\ Increase\ (Kelvin)}$$

ENTHALPY

The term **enthalpy** (abbreviated as H) is composed of the prefix en, meaning "to put into," and the Greek word thalpein, meaning "to heat". Enthalpy is the heat content of a chemical system. It is difficult to measure the total enthalpy of a chemical system, therefore, the **enthalpy change** (ΔH) is more commonly calculated. Enthalpy change is calculated by subtracting the sum (∑) of the products' standard heat of formation from the sum (∑) of the reactants' standard heat of formation.

$$\Delta H = \Sigma H\ (products) - \Sigma H\ (reactants)$$

Measuring Heats of Ractions

The change in enthalpy, ΔH, is specified per mole of substance in the balanced chemical equation for the reaction. The units are typically given as kJ mol^{-1} (kJ/mol) or sometimes as kcal mol^{-1} (kcal/mol). Energy changes are measured under standard temperature and pressure under laboratory conditions, defined as 25 °C and 1 atm.

Enthalpy changes for a given reaction (ΔH_{rxn}) are positive (greater than zero) if a reaction is **endothermic** (energy consuming, Figure 5), and negative (less than zero) if a reaction is **exothermic** (energy generating, Figure 6).

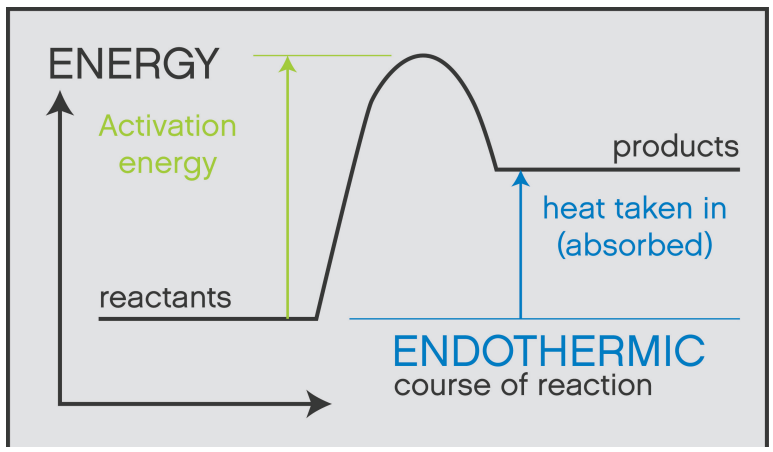

Figure 5. **If you were to graph an endothermic reaction, this is what it would look like when ΔH rxn > 0.**

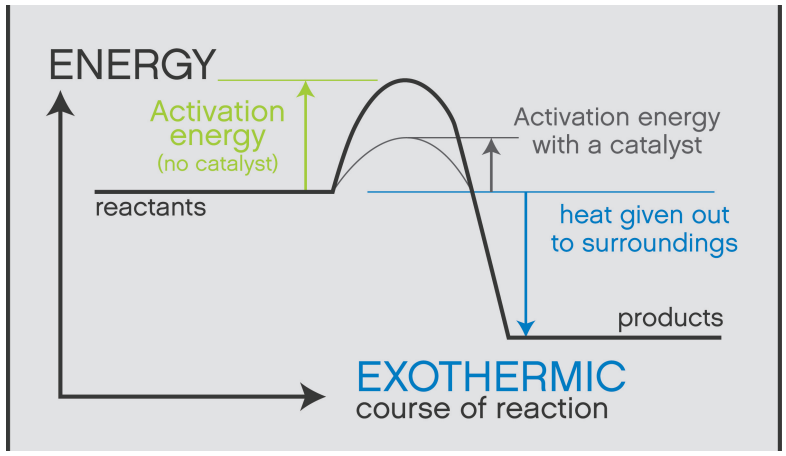

Figure 6. **If you were to graph an exothermic reaction, this is what it would look like when** ΔH rxn **< 0**

FIRST LAW OF THERMODYNAMICS

One of the basic assumptions of **thermodynamics** is the idea that we can arbitrarily divide the universe into a **system** and its s**urroundings**. The boundary between the system and its surroundings can be as real as the walls of a beaker that separates a solution from the rest of the universe (Figure 7).

To better understand how to measure the **heats of reaction**, imagine that a preweighed amount of metal is heated to a known temperature, and is then quickly transferred into a calorimeter that contains a measured amount of water at a known temperature (Figure 8).

Energy in the form of **heat** flows from the metal to the water, and the two eventually equilibrate at some temperature between the initial temperatures of the water and metal.

Assuming that no heat is lost from the calorimeter to the surroundings, and that a negligible amount of energy is absorbed by the calorimeter walls, the amount of energy that flows from the metal as it cools is equal to the amount of energy absorbed by the water. In other words, the energy that the metal loses is equal to the energy that the water gains (Figure 9). As discussed, when heat energy flows into a substance, the temperature of that substance will increase. The quantity of heat energy (q) required to cause a temperature change in any substance is equal to the specific heat capacity (C_{sp}) of that particular substance times the mass (m) of the substance times temperature change (ΔT), as given in this equation:

$$q = m \times C_{sp} \times \Delta T$$

Figure 8. **An unknown type of metal will be provided to perform your first calorimetry experiment.**

Since the metal loses energy (T_{final} is less than $T_{initial}$; therefore ΔT is negative because ΔT = $T_{final} - T_{initial}$) q_{metal} is negative. The water in the calorimeter gains energy (T_{final} is greater than $T_{initial}$; therefore ΔT is positive) and Q_{water} is positive. The following equation can be written since the total energy is always conserved:

$$q_{metal} + q_{water} = 0$$

Rearranging this equation gives (note the negative sign):

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

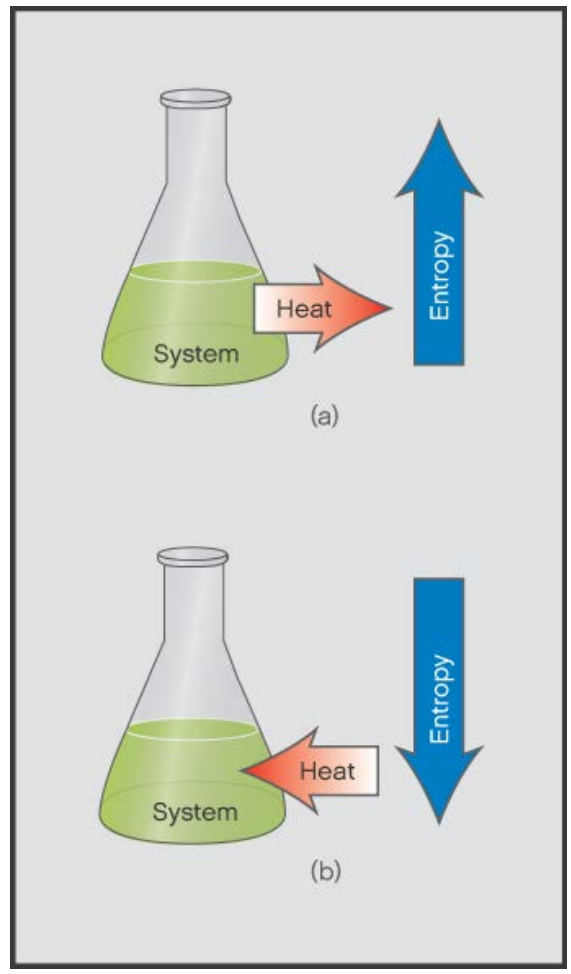

Figure 7. **The first law of thermodynamics states that energy of the universe is constant. Energy can be transferred from the system to its surroundings, or vice versa, but it can not be created or destroyed.**
(a) As heat leaves the system, entropy increases in the surroundings.
H final is less than H initial. Therefore, delta H is negative (less than zero), and the reaction is exothermic.
(b) As heat enters the system, entropy decreas-es in the surroundings.
H final is greater than H initial. Therefore, delta H is positive (greater than zero), and the reaction is endothermic.

Measuring Heats of Ractions

Figure 9. Heated metal will lose heat when put in a calorimeter. The lost heat is primarily absorbed (gained) by the water, while a small amount may also be absorbed by the calorimeter. Thus, Q_{copper} is negative, but Q_{cal} and Q_{water} are positive.

Did You Know?

The Eiffel Tower was built by the engineering firm of Gustave Eiffel, to be used as the entrance arch to the World's Fair held in Paris in 1889. The Eiffel Tower is composed of puddling iron (wrought iron), not steel, as many of today's buildings. Depending on the temperature, the top of the tower may shift up to 18cm (7.1 inches) away from the sun due to thermal expansion caused by the sun's rays!

In addition to the Eiffel Tower, Eiffel's company designed and built the interior frame of the colossal neoclassical sculpture on Liberty Island in New York Harbor, known as the Statue of Liberty

Lab 3: Pre-Lab Questions

1. Water and steam are both 100 °C when water is boiling, but a burn from steam is worse than a burn from the water. Hypothesize why this is true.

2. A 10 g ice cube, initially at 0 °C, is melted in 100 g of water that was initially 20 °C. After the ice has melted, the equilibrium temperature is 10.93 °C. Calculate:

    a. The total heat lost by the water (the specific heat for water is 4.186 J/g/K).

    b. The heat gained by the ice cube after it melts (the specific heat for ice is 2.093 J/g/K).

    c. The heat it took to melt the ice (Hint: it takes 334 J of heat energy to melt 1 g of ice).

3. Inside a calorimeter is 100 g of water at 39.8 °C. A 10 g object at 50 °C is placed inside of a calorimeter. When equilibrium has been reached the new temperature of the water and metal object is 40 °C. What type of metal is the object made from?

Experiment 1

Lab 3
Determination of Specific Heat of a Metal

Experiment Inventory

Materials

(2) Styrofoam™ 8 oz. Cups

Styrofoam™ 8 oz. Cup Lid

Unknown Metal

Access to Graphing Software

Camera/Smart Phone

Computer Access

Distilled Water

Wooden Toothpick (if using a microwave)

Paper Towels

Labware

(1) 100 mL Graduated Cylinder

(1) 500 mL Glass Beaker

(1) Test Tube Clamp

(1) Thermometer

(1) Scale

(1) Insulated Glove

(1) Stove-Top or Microwave

(1) Cooking Pot

(1) Stopwatch / Clock

Note: You must provide items listed in red.

EXPERIMENT 1: DETERMINATION OF SPECIFIC HEAT OF A METAL

In this experiment, you will determine the identity of an unknown metal by investigating its specific heat.

Procedure

☐ 1. In a cooking pot add 500 mL of tap water, place on the stovetop and bring the water to a boil. Set a stop-watch for 3 minutes. Once it has boiled for three minutes, reduce the heat from a boil to a simmer. (See Figure 10).

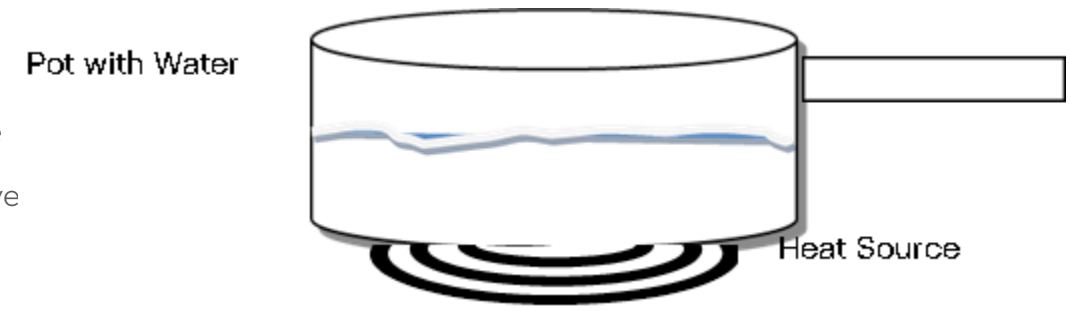

Figure 10: Sample set-up for Step 1. A pot with 500 mL of water rests on a stove-top burner.

Note: If you do not have a stove, heat the 500 mL of water to boiling in a microwave safe container. To avoid the risk of super-heating water (water that does not look boiling, but explodes when agitated), place a microwave safe object in the water and heat the water at 1 minute increments (such as a wooden toothpick). At the end of every minute, tap the microwave safe container before removing the container from the microwave to test if your water is super-heated. Superheated water can be extremely dangerous and will end in hospitalization. Use extreme caution if using the microwave method to boil water. Once boiling, carefully remove the container using the provided insulated glove.

- [ ] 2. Using your 100mL graduated cylinder measure 50 mL of room temperature tap water into one of the Styrofoam™ cups.
- [ ] 3. Record the mass of the water in **Table 2**.
- [ ] 4. Put the two Styrofoam™ cups together by placing the Styrofoam™ cup containing 50 mL of room temperature tap water inside the empty Styrofoam™ cup.
- [ ] 5. Place your cups upright inside of the 500 mL beaker. The beaker is used to provide vertical support for the Styrofoam™ cups. Place the beaker + cups on the counter.
- [ ] 6. Cover the cup containing the 50 mL of tap water with the Styrofoam™ cup lid.
- [ ] 7. Insert the thermometer into the hole in the lid. This apparatus is your calorimeter
- [ ] 8. Record the temperature of the water (this is the initial temperature) in **Table 3**.
- [ ] 9. On your scale, press the button on the right hand side ("0/T"). Your scale should now read 0.0 g. Make sure you are wearing your safety gloves. Pickup your unknown metal with the test tube clamp and place this on top of the scale. Measure and record the weight of the unknown metal strip in **Table 2**.
- [ ] 10. Put your insulated glove on over your safety glove, then pick up the unknown metal strip with the test tube clamp and hold the clamp and unknown metal in the simmering water for five minutes to ensure that the clamp and metal reach the same temperature as the water.

Hint: Do not drop the unknown metal into the pot of hot water and leave it there. It will be difficult to retrieve the heated metal from the hot water and any attempt to do so could cause severe burns.

- [ ] 11. After five minutes has past, quickly transfer the hot unknown metal into the calorimeter. Remember to quickly replace the lid with the thermometer in it. Be careful to ensure that the thermometer is not touching the metal, only the water.
- [ ] 12. In **Table 3**, record the temperature of the water every minute for five minutes. Gently swirl the contents of the cup right before recording the temperature. Gently swirl the contents of the cup and record the temperature every minute for five minutes.
- [ ] 13. Repeat Steps 1 - 10 two more times. Calculate the average heat capacity of the unknown metal. Begin by calculating q for water then use that information to calculate heat capacity for the metal using the equations:

$$q_{water} = C_{sp\ water} \times m_{water} \times \Delta T_{water}$$

$$q_{metal} = C_{sp\ metal} \times m_{metal} \times \Delta T_{metal}$$

- [ ] 14. Use **Table 1** as a reference for determining your unknown metal.

Lab 3
Experiment 1 Data Sheet

Table 2: Mass

	Mass (g)
Water	
Unknown Metal Strip	

Table 3: Specific Heat Data

Time (minutes)	Temperature (°C)		
	Trial 1	Trial 2	Trial 3
Initial			
5 minutes			
6 minutes			
7 minutes			
8 minutes			
9 minutes			
10 minutes			
Average Specific Heat Capacity of the Unknown Metal:			

Lab 3: Post-Lab Questions

1. Why is $\Delta T_{metal} < 0$?

2. Why is $\Delta T_{water} > 0$?

3. A metal sample weighing 43.5 g and at a temperature of 100.0 °C was placed in 39.9 g of water in a calorimeter at 25.1 °C. At equilibrium the temperature of the water and metal was 33.5 °C.

   a. What was ΔT for the water? ($\Delta T = T_{final} - T_{initial}$)

   b. What was ΔT for the metal?

   c. Using the specific heat of water (4.184 J/g °C), calculate how much heat flowed into the water?

   d. Calculate the specific heat of the metal.

4. What is the average specific heat capacity of the unknown metal in this experiment?

5. What is the unknown metal? Use Table 1 for reference.

Lab 3
Determining the Energy in Food

Experiment Inventory

Materials

10 x 10 cm Aluminum Foil

Marshmallow (Large)

Matches

Small, Empty Aluminum Can

(soda can, soup can, etc.)

Safety Glass of Water

Tap Water

Labware

(1) 5 cm Petri Dish

(1) Thermometer

(1) Scale

(1) Insulated Glove

(1) 100 mL Beaker

(1) Butter Knife or Scissors

Note: You must provide the materials listed in red.

EXPERIMENT 2: DETERMINING THE ENERGY IN FOOD

Use calorimetric calculations to experimentally determine the energy content of a marshmallow.

LAB SAFETY: ***The underpad located in your safety kit is not an ideal surface for this experiment as it could catch fire.***

PROCEDURE

- [ ] 1. Obtain the mass of the marshmallow using the scale. Record this value in **Table 3**.
- [ ] 2. Measure 100 mL of room temperature tap water using the 100 mL glass beaker. Pour the tap water into the aluminum can. Then, place the thermometer in the can to measure the temperature of the tap water. Record this value in **Table 3**.
- [ ] 3. Use a butter knife or scissors to carefully cut the marshmallow into four equal quarters (Figure 11).
- [ ] 4. Using either the top or bottom of your Petri dish, completely cover the Petri dish in aluminum foil.

  Note: The aluminum foil is used to create a nonflammable surface. Be sure to completely cover all edges and area of the Petri dish.
- [ ] 5. Place the cut marshmallows into the dish so that the sides of the dish can contain the marshmallows (do not let the marshmallows roll around as they will soon be on fire).
- [ ] 6. Place the Petri dish with cut marshmallows on a flat, nonflammable surface, ignite all four marshmallow pieces with the matches. Dispose of the matches in your safety water (Figure 12).

***Figure 11:** Make sure you have safety water ready and your area cleared of flammable materials before you light the marshmallow fire.*

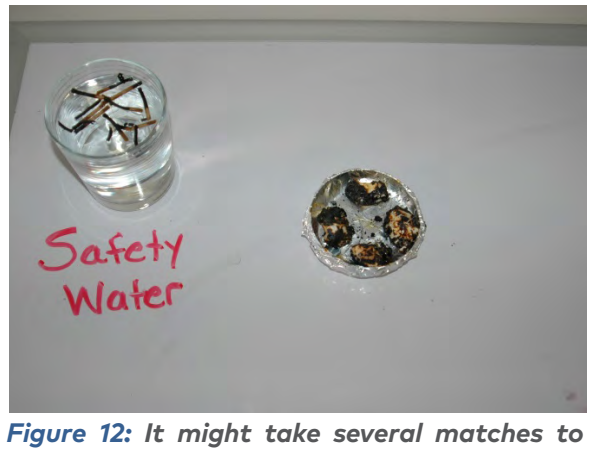

***Figure 12:** It might take several matches to light your marshmallow quarters.*

7. Wearing your insulated glove hold the aluminum can full of tap water over the center of the flame. Keep the bottom of the can close to the flame while ensuring that neither touch.

LAB SAFETY: ***Fire can be dangerous if not monitored properly. Do not stick any body part, loose clothing, or other flammable items in or near the flame. Do this exercise in a well ventilated area and have water nearby.***

8. The flame may burn out before burning through the entire marshmallow. If this occurs, relight the marshmallow. Be sure that you do not let the flame touch the aluminum can when lighting the marshmallow again. Continue this process until the marshmallow is burnt completely.
9. After all of the marshmallow is completely burnt quickly measure the temperature of the tap water inside of the can with the thermometer. Do not touch the thermometer to the sides or bottom of the can. Record the temperature in **Table 3**.

Lab 3
Experiment 2 Data Sheet

Table 4: Calorie Testing

	Initial	Final
Mass of Marshmallow		
Temperature		

Lab 3: Post-Lab Questions

1. Calculate the thermal energy in the burnt marshmallow based on the following information:

$$q = m \times C \times \Delta T$$

2. Is there a relationship between the final mass of the marshmallow and an increase in temperature of the water inside of the can? If so, explain your experimental observations.

3. What is the purpose of cutting the marshmallows into four quarters? How might the experimental results have changed if the marshmallow was burned in one piece?

4. Based on the design of the experiment and on your percent error, would you have designed this experiment differently? Explain why or why not, and if applicable, how you would design it differently.

5. Does toasting bread change the amount of calories in the bread (i.e., does a slice of toast have more/less calories that a slice of bread)? Explain your answer using the data in Table 4.

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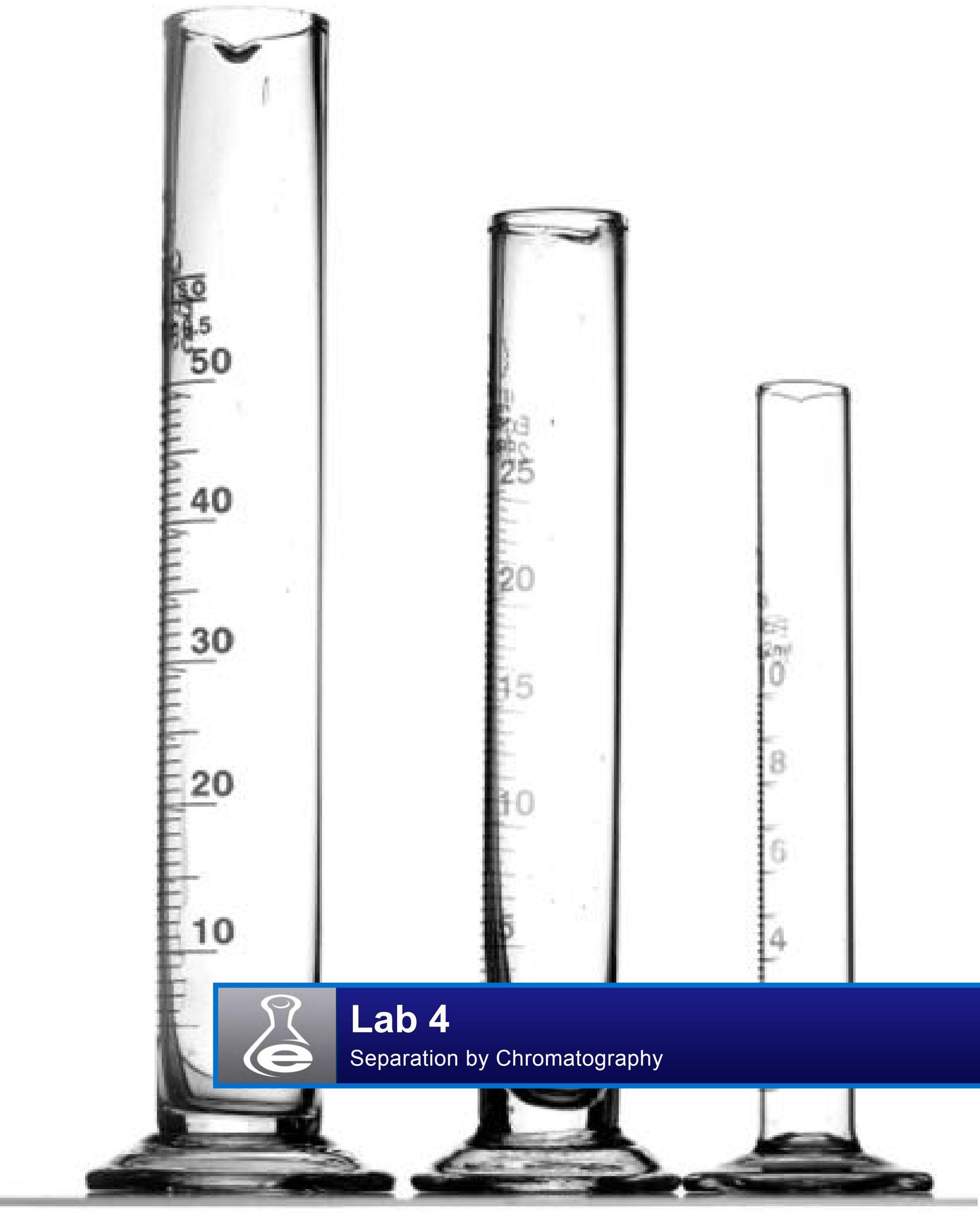

Lab 4

Separation by Chromatography

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LAB 4

Separation by Chromatography

Learning Objectives

- Interpret the results from a paper chromatography separation experiment
- Determine the solvent which will most effectively separate the analytes
- Calculate the R_f for colored dyes found in candy-coated chocolates

INTRODUCTION

How do green M&Ms get their color? Is the dye a mixture of blue and yellow dyes, or is it a pure green dye? What dye colors are used to make an orange M&M? These are questions that can be answered through the use of a chemical technique called chromatography.

Figure 1. M&Ms were developed in 1941. The red, green, and yellow colors were added in 1960, orange was added in 1976, and blue was added in 1995.

CHROMATOGRAPHY

Chromatography refers to a broad range of techniques used to separate or identify individual analytes (components) within a complex mixture. In each case, the mixture is added to an eluting solvent, and the analytes within the mixture migrate along an adsorptive material at different rates. The physical and chemical properties of each analyte affect the rate of migration. These different rates, consequently, separate the analytes.

Chromatography can be applied to any chemical or bioprocessing industry. In these industries, the need to separate or purify a substance is significant. One industry in which substance purification is essential is forensics, where chromatography is used to identify or isolate biological substances and develop evidence. It is also used to isolate proteins in many biotechnology industries, such as pharmaceuticals or drug synthesis (Figure 2). The petrochemical industry also uses chromatography to determine fuel or fuel additive purity. Chromatography is further used in environmental arenas to isolate trace pesticides or detrimental chemicals from ground water. Chemists often use paper chromatography to separate colored or dyed mixtures. The separated chemicals can then be used in more advanced chemical preparations or simply identified.

MOBILE PHASE AND STATIONARY PHASE

Chromatography is based on two phases: the mobile phase and the stationary phase. The **mobile phase** (eluting solvent) is the phase that moves up the chromatography paper. The mixture of analytes is placed in the mobile phase.

The **stationary phase** is the material held in place for the chromatography procedure. A good separation results when the components of a mixture have varying levels of affinity for the mobile and stationary phases. Think of the mobile phase as a moving stream and the stationary phase as the stream bed. If you were to toss in a leaf, a stick, and a large rock, what would happen? Each unique component would travel at different rates along the stationary phase, using the mobile phase as a vehicle. Many properties affect the affinity of a substance for the mobile or stationary phase, including polarity, solubility, particle size, and electrical charge. Chemists can use their knowledge of these properties to separate a mixture effectively.

Figure 2. One type of chromatography used in laboratories is gas chromatography (GC). This is a picture of an GC autosampler.

Separation by Chromatography

Terminology:

Analyte: An individual component within a mixture that is of interest in analytical chemistry. Analytes are separated out during the mobile phase. They are carried up the stationary phase in the eluting solvent (the carrier).

Eluent or Eluting Solvent: The solvent which the mixture of analytes is placed in. This chemical acts as the carrier molecules within the mobile phase.

Immobilized Phase: A mobile phase that has become immobilized (movement is halted) on the stationary phase.

Mobile Phase: A mixture of the eluting solvent and the analyte components. This phase moves (usually up) the stationary phase. The physical and chemical properties of the analytes suspended in the mobile phase allow the analytes to separate during the mobile phase's movement.

Solvent Front: The edge of the mobile phase after it has completed its migration up the stationary phase (see Figure 3 for example)

Stationary Phase: A fixed substance or material. Used as a foundation for the mobile phase to travel upon. Chromatography paper is the stationary phase in paper chromatography.

R_F VALUE

One way to compare the movement of the analyte is to calculate the retention factor value, or **R_f** value. Assuming that the same mobile and stationary phase substances are used, the R_f value for a chemical will not change. Therefore, R_f values allow for individual analyte identification within a compound. The R_f value is determined by taking the distance travelled by the analyte and divide it by the distance travelled by the mobile phase. This is mathematically expressed as:

$$Rf = \frac{migration\ distance\ of\ substance}{migration\ of\ solvent\ front}$$

Where:

- **Migration Distance of Substance:** The distance the analyte travels up the stationary phase
- **Migration Distance of Solvent Front:** The total distance which the mobile phase travels up the stationary phase.

For example, suppose an analyte travels up the stationary phase 3.2 cm, but the mobile phase travels a total of 7.5 cm. Calculate the R_f value:

$$Rf = \frac{3.2\ cm}{7.5\ cm}$$

$$Rf = 0.43$$

Separation by Chromatography

CHROMATOGRAPHY FACTORS

Different types of materials used for mobile and stationary phases lead to different types of chromatography. Some of these include paper, ion-exchange, gas, high performance liquid, column, affinity, and thin layer chromatography. In paper chromatography, the type of paper used controls how fast the mobile phase moves.

There are also factors which influence how effective an eluting solvent is at separating the analyte components. For example, salt is an ionic eluting solvent, isopropyl alcohol is a non-polar eluting solvent, and water is a polar eluting solvent. By changing the concentration of salt, you change the ionic characteristics of the solvent. By altering the concentrations of water and alcohol in the solution, you change the polarity of the solution. All of these factors affect the overall affinity and specificity of the analyte separation.

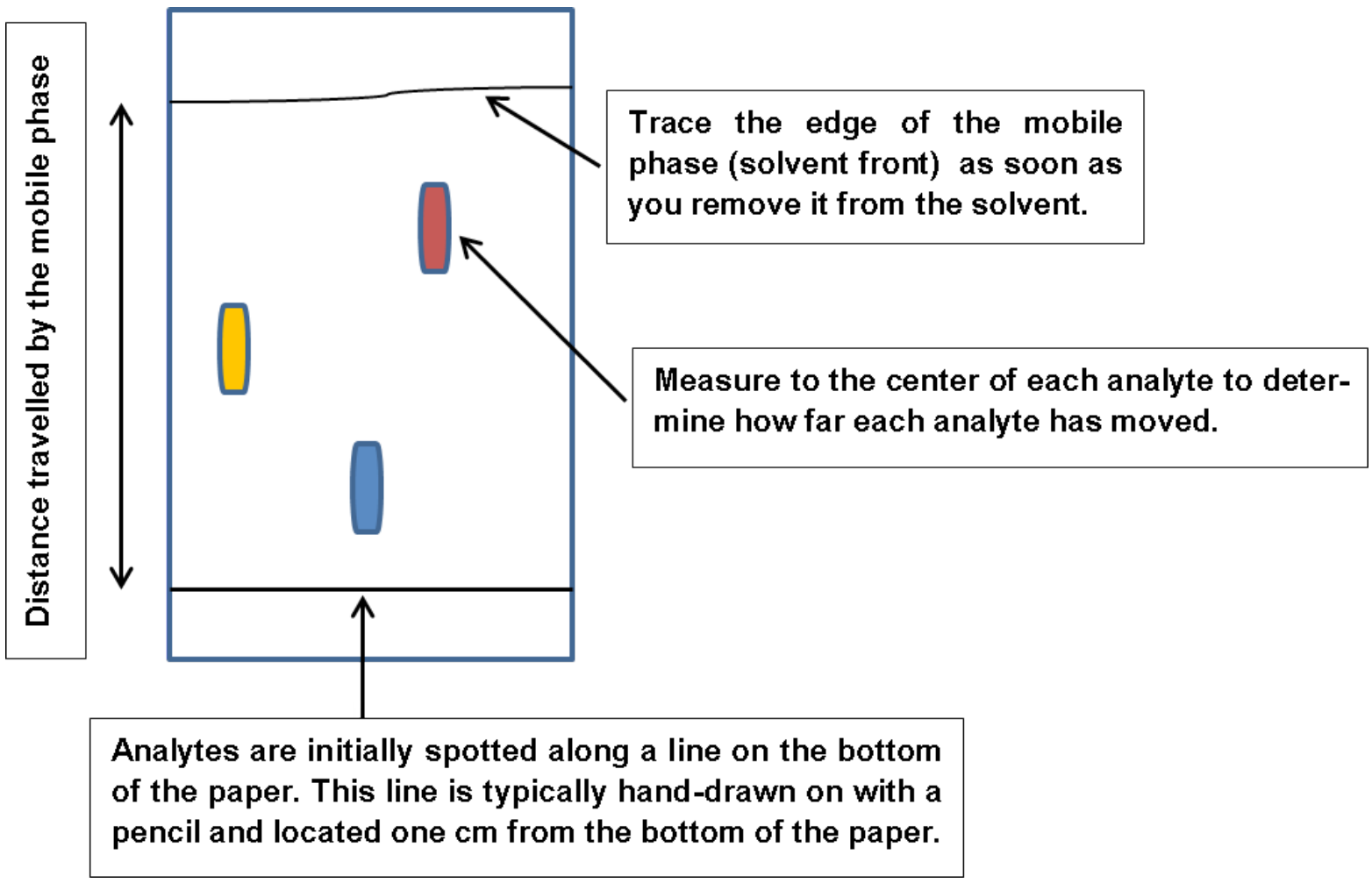

Figure 3: Sample paper chromatography results.

Lab 4: Pre-Lab Questions

1. Rank the following solutions from least polar to most polar.

   Rank on a scale of 1-4: 1 being the least polar and 4 being the most polar.

   _____ 50% ISOPROPANOL / H_2O

   _____ 25% ISOPROPANOL / H_2O

   _____ PURE WATER

   _____ ISOPROPANOL / H_2O

2. Identify the analyte, eluting solvents (mobile phases), and stationary phase used in Experiment 1.

3. Why would it be important to know which food dyes are in food?

Experiment 1

Lab 4
Paper Chromatography

Experiment Inventory

Materials

Eluting Solvents

- 30 mL 0.5% Sodium Chloride, NaCl solution
- 30 mL 0.2% Sodium Chloride, NaCl solution
- **30 mL Distilled Water**
- **30 mL Isopropyl Alcohol, C_3H_8O**

(6) M&Ms® (2 blue, 2 green, 2 red) - DO NOT EAT

(1) Pencil

Computer with Internet Access

Camera / Smart Phone

Tap Water

Labware

(3) 50 mL Beakers

(1) 100 mL Graduated Cylinder

(1) 500 mL Beaker

(3) Wooden Stir Sticks

(3) Capillary Tubes (packaged inside of a test tube)(1) Ruler

(2) 11 x 11 cm Pieces of Chromatography Paper

(1) Scissors

Note: You must provide the materials listed in red.

EXPERIMENT 1: PAPER CHROMATOGRAPHY

In this experiment, you will use paper chromatography to determine the Rf value for the dyes found in candy-coated chocolate pieces. The chromatography paper acts as the stationary phase for the procedure, and a variety of mobile phases will be tested. Multiple tests with different eluting solvents must be run to determine the best eluting solvent to separate the food dyes.

PROCEDURE

- [ ] 1. Put on your safety glasses and gloves (provided in your safety box).
- [ ] 2. Gather three 50 mL beakers, one for each color candy you will test.
- [ ] 3. Place 2 M&Ms® candies of one color into a 50 mL beaker.
- [ ] 4. Repeat for each color of the candy you will test. You should have two green candies in one beaker, two red candies in a second beaker, and two blue candies in the third beaker.

PART 1: PREPARATION OF THE ANALYTE

- [ ] 1. Use a pipette to add 1 mL of isopropyl alcohol to each 50 mL beaker.
- [ ] 2. Stir with a wooden stir stick for one minute (the colors may not appear to be very dark). Remove the candies. Be sure to use a clean stir stick each time you change beakers.
- [ ] 3. Allow the solutions to sit and concentrate while the stationary phase is prepared.

PART 2: PREPARATION OF STATIONARY PHASE

☐ 1. Cut each piece of chromatography paper in half. Direction doesn't matter, but keep it consistent for all papers.

☐ 2. Set up the chromatography paper according to Figure 4.

☐ 3. Using a pencil, mark the paper 1 cm from the bottom edge. Refer to Figure 4.

☐ 4. Using a capillary tube, place small spots of the analyte equal distance apart on the marked line. Since there are three colors to be tested, there will be three spots on the line (use one capillary tube per color; save the tubes for the additional trials).

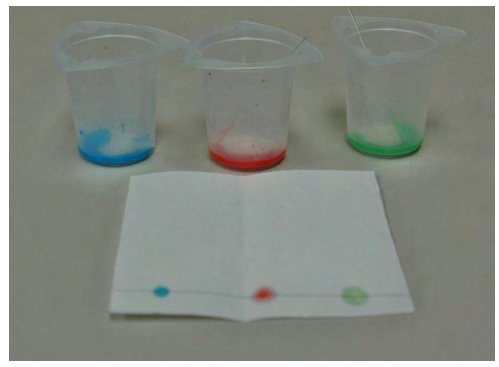

Figure 4. Chromatography set up.

Note: **Capillary tubes are extremely thin tubes. They are useful when working with very small amounts of a sample, and collect liquid samples through capillary action. To use the capillary tube, simply place the open end of the tube in the sample. The liquid molecules will be drawn into the tube and stick to the inner walls. Figure 5 provides a references for this process.**

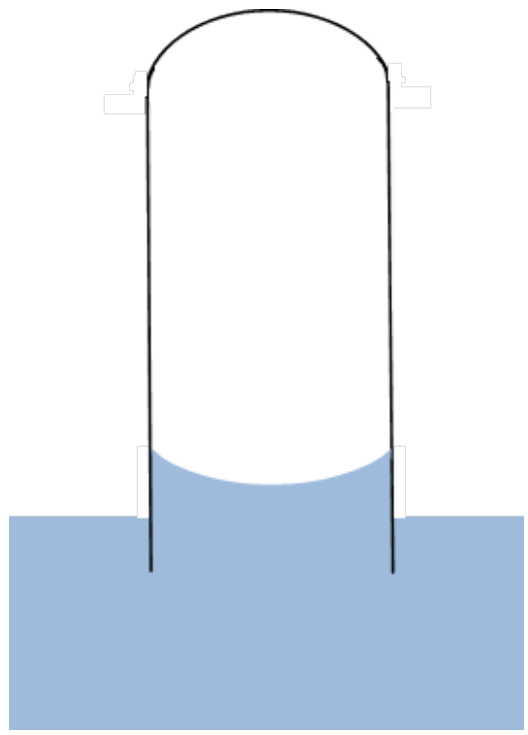

Figure 5. Capillary tubes use capillary action to pull up liquid.

☐ 5. Allow the spot to dry, and re-spot the analyte in the exact same area as done in Step 4. Repeat this process at least five times, or until the colored dots appear on the paper.

☐ 6. Use a 100 mL graduated cylinder to pour 20 mL distilled water (your eluting solvent) into the 500 mL beaker.

☐ 7. Place the paper vertically with the line-side down in the 500 mL beaker with the eluting solvent. Let it stand for 3 - 5 minutes (Figure 6).

☐ 8. Use a pencil to mark the edge of the solvent front (the edge of the mobile phase) and the location of the analytes with a pencil (see Figure 4 for reference). Measure the solvent and dye fronts based on farthest location from origin line. Record your data and any additional observations in **Table 1**.

Figure 6. Eluting solvent set up.

☐ 9. Use tap water to rinse the 500 mL beaker and then repeat steps 4-8 for the remaining eluting solvents (0.5% NaCl, 0.2% NaCl, and isopropyl alcohol, record your data in **Table 1**.

☐ 10. Line up the results of all your chromatography experiments and use your camera (or smart phone) to take a picture of the results. Be sure to correctly label the picture and send it to your instructor along with the answers to the lab questions.

Data Sheet

Lab 4
Experiment 1 Data Sheet

Table 1: Paper Chromatography Data and Observations

Solvent	Distance Traveled by Mobile Phase (mm)	Distance Traveled by Each Analyte (mm)	R_f Value	Additional Observations
1: Distilled Water		**Green:** **Blue:** **Red:**	**Green:** **Blue:** **Red:**	
2: 0.5% NaCl Solution		**Green:** **Blue:** **Red:**	**Green:** **Blue:** **Red:**	
3: 0.2% NaCl Solution		**Green:** **Blue:** **Red:**	**Green:** **Blue:** **Red:**	
4: 70% Isopropyl Alcohol		**Green:** **Blue:** **Red:**	**Green:** **Blue:** **Red:**	

Lab 4: Post-Lab Questions

1. Which solvent provided the best separation?

2. Explain which characteristics of the solvent were used to effectively separate the analytes.

3. Some children have reactions to Yellow 5 or Yellow 6 dye. Yellow 5 is a pale yellow color and Yellow 6 is more orange. Use the colors seen on the chromatograms to determine which M&Ms® candies you tested contain Yellow 5.

4. Chromatography has many applications. Research one application of chromatography and explain how it is used and what characteristic is utilized for the separation of the analyte(s).

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Labs

Lab 5

Electron Configuration

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LAB 5

Electron Configuration

Learning Objectives

- Identify elements using a flame test
- Determine the electron configuration of known elements
- Apply the concepts of quantized atomic energy
- Interpret the relationship between color and wavelength

INTRODUCTION

Chemistry is about matter, its properties, and how it is transformed during a chemical change, such as when iron rusts (Figure 1). When a chemical change occurs, the atoms of a molecule are rearranged to form new substances. The transformation occurs when the bonds that hold the atoms together in the substances break and new bonds are formed. When one bond is broken and a new bond is formed, the electrons which exist in a particular location in an atom are moved to a different location. These locations are called atomic orbitals. If we understand where electrons are in atoms, we can better understand what happens when a chemical transformation occurs.

Figure 1. Rust is formed by the chemical reaction of iron and oxygen in the presence of water or air moisture.

ATOMIC ENERGY

Atomic energy is the energy that an atom contains. It turns out that this energy is quantized, meaning there are specific energy levels. To understand what it means for something to be quantized, think of the word quantity. A quantity is a numerical description of something. The concept of energy being quantized is similar to the way we can use a ladder. We can only move up or down by stepping on the rungs, rather than the space in between. In a similar fashion, the energy of electrons can only move up or down in complete energy level changes. Energy is not available at values which fall in between two levels.

Figure 2. The energy levels in an atom are similar to the rungs of a ladder

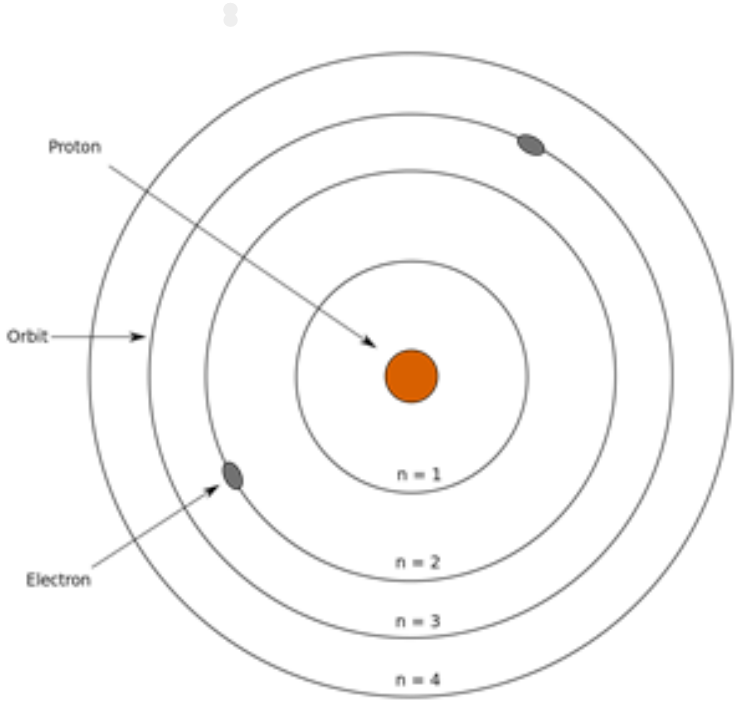

Figure 3. A simplified version of the different energy levels that are associated with different orbits.

ELECTRON CONFIGURATION

The word "configuration" means "the arrangement of the parts of something." **Electron configuration** is a shorthand notation system that is used for describing the locations of all of the electrons in an atom (Figure 3). Electrons generally do not orbit the nucleus, as scientists used to think. Rather, they hover, wiggle, jiggle and vibrate back and forth, and move in certain patterns based on three factors: (1) their energies, (2) their orientations (exactly where they fit in space relative to the other electrons in the atom) and (3) their angular momentum.

Electron Configuration

In order to visualize electron configuration, picture your foot on the rung of a ladder is like an electron closest to the ground (Figure 2). The electron configuration naming system describes the location of a specific electron in an atom. Each principal energy level has a sublevel or sublevels (s, p, d, or f). For example, the electron configuration for helium is $1s^2$ because it is in the first principal energy level (1), the first sublevel (s), and has two electrons (Figure 4).

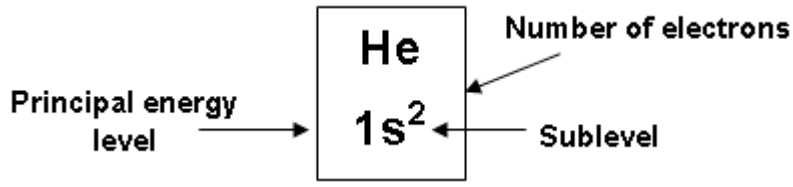

Figure 4. **The energy level, number of electrons, and sublevels for any given element can be determined by using the periodic table of elements.**

The lowest energy configuration for an atom is called its ground state (n=1). As energy is added from the ground state, electrons tend to fill energy levels and sublevels in a specific order (Figure 5). The first sublevel in any principle energy level is the "s" sublevel, which can hold up to two electrons. The next sublevel filled is the "p" sublevel, which can hold up to six electrons. Continuing, the "d" sublevel can hold up to 10 electrons, and "f" sublevel can hold up to 14 electrons. In the ground state, electrons will always occupy the lowest available energy level first, meaning that an electron will not exist in a higher energy level or sublevel when a lower one is vacant.

WRITING ELECTRON CONFIGURATIONS

Using the periodic table of elements, you can write the electron configuration for any atom. To do this, count the number of electrons, which is the same as the number of protons or the atomic number. Then use as many orbital sublevels as you need to hold all your electrons. Be sure to always fill the lowest energy levels first. For example, let's say you have an atom of lithium (Figure 6).

Figure 6. **From the periodic table of elements, lithium (Li) has atomic number of three, which means that it also has three electrons.**

Lithium's atomic number is three. So, an atom of lithium has three protons and three electrons. We would need the orbital sublevels to be able to hold three electrons. The electron configuration is written as: $1s^2 2s^1$. We can verify that this is correct because there are two electrons in the 1s sublevel and one electron in the 2s sublevel (2 + 1 = 3).

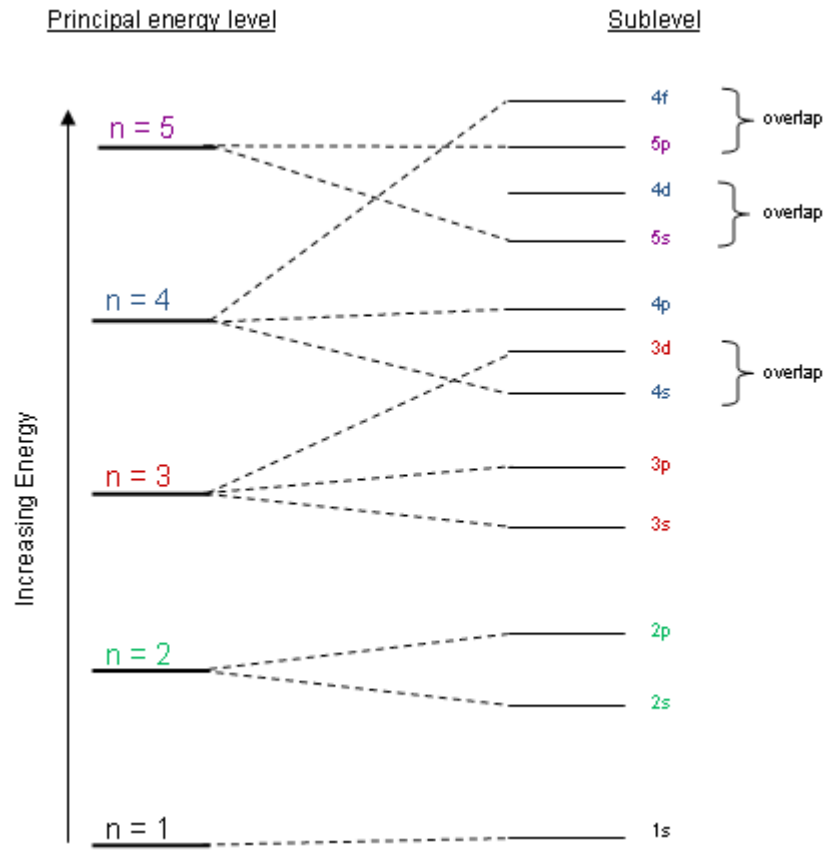

Figure 5. **Illustration for energy configuration in an atom.**

ENERGY OF ELECTRONS

The electrons in atoms have different amounts of energy proportional to the distance of their orbital from the nucleus. Electrons in low, close levels to the positive nucleus have lower potential energy, whereas those in higher energy levels that are farther away have more energy. In order for an electron to "jump" from a lower level to a higher one, it must **absorb** energy, often in the form of light. Conversely, when an electron "falls" from a higher level to a lower one, it gives off energy (**emits**), again in the form of a photon of light (Figure 7).

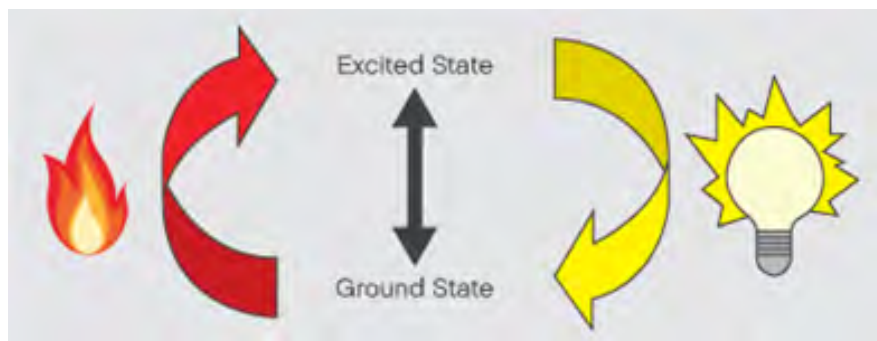

Figure 7. **Think of electrons as being lazy. They prefer to stay at the lowest possible energy level and will only go to a higher energy level (excited state) if they are given energy. However, as soon as possible, they will release their excited state energy and return to the lowest possible energy level.**

Electron Configuration

FIREWORKS

Electrons of different elements absorb and emit different amounts of energy. Recall that colors can be associated with discrete wavelengths. This means that by observing the color (wavelength) of light emitted by a substance when its atoms absorb and release energy, scientists can determine its chemical composition (Table 1). This phenomena helps explain how fireworks produce different colors of light. The main ingredients of fireworks are potassium chlorate or perchlorate, charcoal, and sulfur. When the atoms of different materials in a firework are excited by heat, they will absorb and release energy in particular wavelengths, producing a particular variety of colors (Figure 8).

Figure 8. The flame of burning lithium salt demonstrates the bright colors that can be achieved by burning different compounds.

Table 1: Wavelength Ranges for the Visible Spectrum

Color Shades	Wavelength Range (nm)
Red	650 - 750
Orange	595 - 650
Yellow	580 - 595
Yellow - Green	560 - 580
Green	500 - 560
Green - Blue	490 - 500
Blue - Green	480 - 490
Blue	435 - 480
Violet	400 - 435

Lab 5: Pre-Lab Questions

1. What is an electron configuration?

2. How is the light emitted by an atom related to its electron configuration?

Experiment 1

Lab 5
The Chemistry of Fireworks

Experiment Inventory

Materials

5 mL Calcium Chloride Solution, $CaCl_2$

5 mL Lithium Chloride Solution, LiCl

5 mL Potassium Chloride Solution, KCl

5 mL Sodium Chloride Solution, NaCl

Matches

Modeling Clay

3 Birthday Candles

Labware

(4) Inoculating Loops

Note: You must provide the materials listed in red.

EXPERIMENT 1: THE CHEMISTRY OF FIREWORKS

In this experiment, the flame from a birthday candle is the outside energy source. The flame emits a broad range of energy, but the electrons of the atom being heated will only absorb specific amounts of energy.

LAB SAFETY: *Wear your safety glasses and choose an area that is well-ventilated (yet not windy), and have a fire extinguisher nearby. Flames should be on a flat surface. Matches, chemicals and candle wax can cause fire or burns to skin, clothing, or lab materials. Do not place any body part, loose clothing, or other flammable items in or near the flame. Never leave any burning fuel unattended!*

PROCEDURE

Figure 9: Experiment set-up and flame technique.

- [ ] 1. Place a round piece of clay on the straight end of each inoculating loop. This will act as a holder.
- [ ] 2. Place in order the LiCl, NaCl, KCl, and $CaCl_2$ saturated solutions. Set one inoculating loop next to each sample.
- [ ] 3. Stabilize one birthday candle by placing the base of it in a piece of clay that is about 1 inch by 1 inch by 1 inch. The candle should stand freely and vertically (Figure 9). The other two candles are to be used if the first one burns down to the clay.
- [ ] 4. Use the matches to light the stabilized birthday candle. Hold the inoculating loop for the LiCl at the very end of the non-looped end in order to avoid burns. Heat the looped end in the candle flame until its loop is faintly orange and any coating is burned off (Figure 9).

CAUTION: ***Lit matches and candles can cause fire or burns to skin and/or clothing if the flame comes into contact with skin or clothes. Be sure you have your safety glasses on!***

- [ ] 5. Dip the loop into the LiCl solution.

CAUTION: ***The loop will remain extremely hot for several minutes following being in the flame. Do not touch the loop!***

- [ ] 6. Bring the looped end of the inoculating loop into the flame. Make observations about what is happening, especially any color changes.

Hint: The color change will be most apparent around the edges of the flame. You may have to try this a few times to see the color change.

- [ ] 7. Extinguish the candle and record your observations in **Table 2**.
- [ ] 8. Repeat the Steps 4 - 7 for each of the other three solutions. Use a different inoculating loop for each one.
- [ ] 9. To clean-up, you may throw away the inoculating loops after they have cooled to room temperature.

Lab 5
Experiment 1 Data Sheet

Table 2: Results of Firework Material Ignition

Substance	Observations
Lithium chloride	
Sodium chloride	
Potassium chloride	
Calcium chloride	

Table 3: Electron Configuration

Element	
K	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Li	
Na	
Ca	

Table 4: Color of Light Emitted by Salt Types

Salt	Color	Wavelength
LiCl		
NaCl		
KCl		
$CaCl_2$		

Lab 5: Post-Lab Questions

1. Write out the electron configurations of carbon and each of the metals of the salt compounds used. Record this in Table 3. Potassium is already done as an example for you. **Hint:** The periodic table is very helpful and can be used as guide.

2. Identify the color emitted by each of the salts along with the corresponding approximate wavelength using Table 1 in the Introduction. Record your results in Table 4.

3. Why does a salt compound give off light (or a colored flame) when burned?

4. Barium chloride emits a green color when flame tested. What can be said about the wavelength of light it emits?

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Labs

Lab 6

Discovering the 5 Types of Chemical Reactions

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LAB 6

Discovering the 5 Types of Chemical Reactions

Learning Objectives

- State observations that are evidence for a chemical reaction
- Classify a chemical reaction as one of the following types: combustion, decomposition, single replacement, double replacement, or synthesis
- Use saponification to produce a bar of soap

INTRODUCTION

Before modern manufacturing, if people needed a product, they had to make it themselves. For example, individuals routinely made their own soap. Whether people knew it or not, they were often using **chemical reactions** to obtain the products that they wanted. Chemical reactions are the main focus of chemistry, because it is what drives chemistry. Chemical reactions can be understood by examining what is being produced, what is changed, and what is broken down.

Figure 1. Chemical equations use a variety of reactants to produce different products. Hydrochloric acid (HCl) is a common reactant.

REACTION OVERVIEW

A chemical reaction is a process that is usually characterized by a chemical change, in which the starting materials (reactants) are different from the ending materials (products). Chemical reactions tend to involve the motion of electrons, leading to the formation and breaking of chemical bonds. Force and energy determine if the reaction occurs and how fast, while mathematics helps keep an inventory of the reactants and products.

Figure 2. Methane is a natural gas, and is used as a cooking fuel and for heating homes.

CHEMICAL EQUATIONS

Reactions are described with **chemical equations**. A chemical equation is the symbolic representation of the chemicals involved in a reaction. These chemical equations can be used to describe both physical processes and chemical reactions. All chemical equations consist of reactants, the starting materials, products, the ending materials, the state of matter that the materials are in, and a reaction arrow representing that a reaction has occurred (Figure 1).

Here is an example of a chemical equation:

$$CH_4\ (g) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g) + \text{Energy}$$

This equation represents the combustion of methane gas (CH_4 *(g)*). This equation can be interpreted as saying that when methane gas is combined with oxygen, it will undergo a chemical reaction to produce carbon dioxide (CO_2 *(g)*) and water (H_2O *(g)*). Energy in the form of heat and light are also produced (Figure 2).

Discovering the 5 Types of Chemical Reactions

LAW OF CONSERVATION OF MATTER

When writing a chemical equation, it is necessary to consider the amount of atoms on each side of the equation. This is necessary to be compliant with the **law of conservation of matter**. This law states that matter can not be created nor destroyed. Thus, in the chemical equation, you must have the same number of each type of atom on each side of the reaction. You may have noticed in the equation above that there are an unequal amount of atoms on each side of the reaction (e.g., four hydrogens on the reaction side and two hydrogens on the product side). Therefore, the chemical reaction must be balanced to adhere to the law of conservation of matter.

BALANCING CHEMICAL EQUATIONS

$$CH_4\ (g) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g) + \text{Energy}$$

Balancing a **chemical equation** refers to establishing the mathematical relationship between the amount of reactants and products. It takes practice in the form of trial and error to balance chemical equations. However, there are six basic steps:

1. Write the unbalanced equation.
2. Apply the law of conservation of matter by counting the number of atoms of every element on each side of the equation.
3. Save oxygen and hydrogen for last and balance the chemical formulas by placing coefficients in front of the formulas. You are allowed to add as many coefficients as it takes to balance the equation. However, you are never ever allowed to add subscripts! Adding subscripts will change the chemical formulas, and this will violate the law of conservation of matter.
4. Once you have one chemical balanced, move on to the next one.
5. Once you think you have the entire equation balanced count up all of the atoms again. Remember to use the distributive property to multiply your coefficients into all atoms in your formula.
6. Write your balanced equation

BALANCING CHEMICAL EQUATIONS EXAMPLE:

STEP 1:

Write the unbalanced equation.

$$CH_4\ (g) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g) + \text{Energy}$$

STEP 2:

Count the number of atoms of every element on each side of the equation (Table 1). Energy is not an atom, so although it is a product, it is not counted when balancing the equation.

Table 1: Unbalanced Atomic Count

Reactants	Products
1 C (carbon)	1 C (carbon)
4 H (hydrogen)	2 H (hydrogen)
2 O (oxygen)	3 O (oxygen)

STEP 3 AND STEP 4:

Notice that the carbon is balanced, so there is no need to add any coefficients. However, hydrogen and oxygen are not balanced. Use the least common multiple theory in order to determine which coefficient to place in front of the chemical. In the case of hydrogen, two is the least common multiple, and in the case of oxygen, two is also the least common multiple.

$$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g) + \text{Energy}$$

Discovering the 5 Types of Chemical Reactions

STEP 5:

Use the distributive property to multiply your coefficients into all atoms in your formula. For example, the coefficient of two in front of the hydrogen (product side of the equation) is multiplied with the subscript of two giving a count of four hydrogens, and this coefficient of two is also multiplied to the oxygen giving a count of two oxygens (Table 2).

Once you think you have the equation balanced count up all of the atoms again. Count the total number of each type of atom from both sides of the equation. For example, on the product side, you have two oxygens from the carbon dioxide and two oxygens from the water giving you the grand total of four oxygens (Table 2).

Table 2: Balanced Atomic Count

Reactants	Products
1 C (carbon)	1 C (carbon)
4 H (hydrogen)	4 H (hydrogen)
4 O (oxygen)	4 O (oxygen)

STEP 6:

Write your balanced equation.

$$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g) + \text{Energy}$$

CHEMICAL REACTIONS

There are an infinite number of chemical reactions. Chemists have divided these into broad classifications. The most important classifications are: **combustion**, **decomposition**, **single replacement**, **double replacement**, **synthesis**, **acid-base**, **and oxidation-reduction (also known as redox).** Keep in mind that some reactions will fall into more than one classification. For example, all single replacement reactions are also redox reactions.

Characters in Chemistry

Antoine Lavoisier | 1743 - 1794

One of the first breakthroughs in the study of chemical reactions resulted from the work of the French chemist Antoine Lavoisier between 1772 and 1794. Lavoisier found that mass is conserved in a chemical reaction. The total mass of the products of a chemical reaction is always the same as the total mass of the starting materials consumed in the reaction. His results led to one of the fundamental laws of chemical behavior: the law of conservation of matter.

He was certainly not the first to accept this law as true or to teach it, never the less he is credited as its discoverer (PD-US)

COMBUSTION REACTION

Combustion reactions are **exothermic** reactions that give off energy in the form of heat and light. These reactions occur when a fuel source (compound "A") combines with oxygen (O_2) to form carbon dioxide (CO_2) and water (H_2O). These reactions follow the general form:

$$A + O_2 \rightarrow CO_2 + H_2O + \text{Energy}$$

An example of this kind of reaction is the burning of naphthalene ($C_{10}H_8$, the active ingredient in mothballs):

$$C_{10}H_8\ (s) + 12\ O_2\ (g) \rightarrow 10\ CO_2\ (g) + 4\ H_2O\ (g) + \text{Energy}$$

To visualize a combustion reaction look at the cartoon (Figure 3):

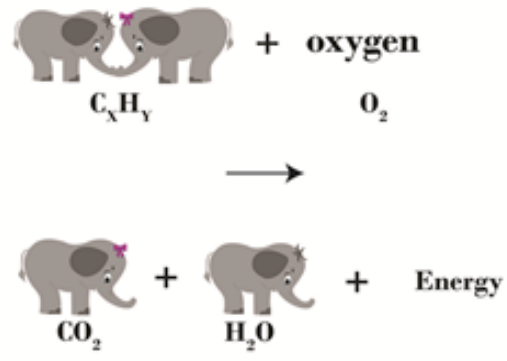

Figure 3. Visualizing a combustion reaction. During combustion, a compound (the pair of elephants) combines with oxygen which forms CO2 (pink elephant) and H2O (blue elephant) and also releases energy.

Figure 4. Wood and other fuels that undergo combustion reactions don't spontaneously catch on fire just because they are surrounded by oxygen. For the combustion reaction to happen, you have to ignite the fuel to a very high temperature. For example, wood will catch on fire whin it reaches a temperature of 300°F (150°C).

Discovering the 5 Types of Chemical Reactions

DECOMPOSITION REACTION

In a decomposition reaction, a complex molecule breaks down to make simpler ones. These reactions come in the general form:

$$AB \rightarrow A + B$$

One example of a decomposition reaction is when water is broken down into hydrogen gas and oxygen gas. The chemical equation for this decomposition reaction looks like:

$$2\,H_2O\,(l) \rightarrow 2\,H_2\,(g) + O_2\,(g)$$

To visualize a decomposition reaction look at the cartoon (Figure 5):

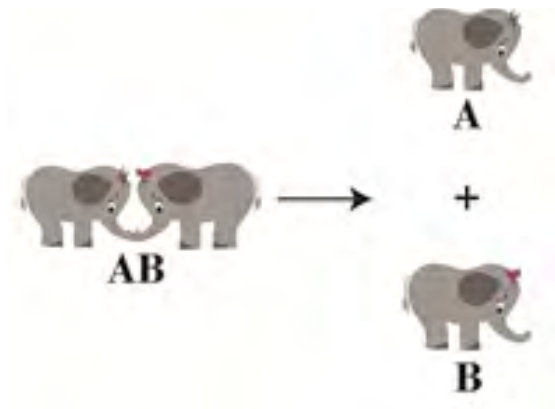

Figure 5. Visualizing a decomposition reaction. During a decomposition reaction the bonds of a compound (elephants AB) will be broken so each component exists individual constituents (elephant A and elephant B). This complex molecule breaks down into simpler molecules.

SINGLE REPLACEMENT REACTION

A single replacement reaction occurs when a single uncombined element switches places with another in a compound. The word "single" is used because there is only a single compound that is affected by the replacement. Two reactants yield two products. These reactions come in the general form of:

$$A+BC \rightarrow AC + B$$

A pure element is represented by the letter "A." The element represented by "A" breaks the "BC" bond, displacing "B" and creating "AC." One example of a single replacement reaction is when magnesium displaces hydrogen in water to make magnesium hydroxide and hydrogen gas:

$$Mg\,(s)+2\,H_2O\,(l) \rightarrow Mg(OH)_2\,(s) + H_2\,(g)$$

To visualize a single replacement reaction look at the following cartoon (Figure 6):

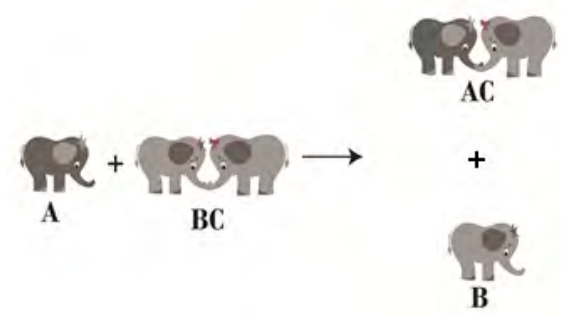

Figure 6. Visualizing a simple replacement reaction. In this reaction one molecule switches another out to form a new compound. Initially elephants B and C form a compound, but during a replacement reaction, elephant A is replaced for B. Now elephants A and C form a compound while elephant B stands alone.

DOUBLE DISPLACEMENT

A **double displacement** reaction occurs when two compounds switch places to form two new compounds. In the generic compound "AB," the "A" element, which is usually a metal, exchanges places with another metal in compound "CD." To the right side of the arrow shows their final arrangement. Keep in mind that the letters are just letters that represent any ion. The "B" is not boron and the "C" is not carbon.

$$AB+CD \rightarrow AD + CB$$

One example of a double displacement reaction is the reaction of lead (II) nitrate with potassium iodide to form lead (II) iodide and potassium nitrate:

$$Pb(NO_3)_2\,(aq) + 2\,KI\,(aq) \rightarrow PbI_2\,(s) + 2\,KNO_3\,(aq)$$

To visualize a double replacement reaction look at the following cartoon (Figure 7):

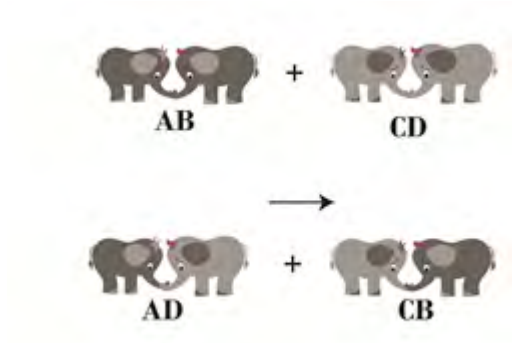

Figure 7. Double displacement reaction. Similar to a single replacement, compounds switch places. Initially elephants A and B and elephants C and D form two distinct compounds. In a double displacement reaction, two molecules switch to create two new compounds. Now, elephants A and D and elephants C and B form new compounds with each other.

SYNTHESIS

A synthesis reaction is when two or more simple compounds combine to form a more complicated one. These reactions come in the general form of:

$$A + B \rightarrow AB$$

A pure element is represented by the letters "A" and "B." The two elements join together to create an entirely new compound, represented by "AB." One example of a synthesis reaction is the tarnishing of a silver tea set:

$$Ag\,(s)+O_2\,(g) \rightarrow Ag_2O_2\,(s)$$

To visualize a synthesis reaction look at the following cartoon (Figure 8):

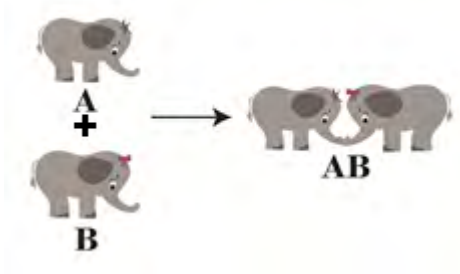

Figure 8. A synthesis reaction creates a complex compound from simple components. Elephants A and B, which can be considered simple components undergo a synthesis reaction to create AB, a more complex compound.

Discovering the 5 Types of Chemical Reactions

SYNTHESIS OF SOAP (SAPONIFICATION)

The process of making soap is called **saponification**. The word "soap" comes from either the Gallic (Gaulish) word "sapo" or a Germanic word, "saipa." Both sapo and saipa have their origins in the Latin word "sebum," meaning fat.

Saponification involves a synthesis reaction between a metallic alkali, such as lye (otherwise known as sodium hydroxide, NaOH), and an animal or vegetable fat (such as an oil) to produce soap. In saponification, the metallic alkali, in this case NaOH, breaks down the oil with which it is mixed (Figure 9). This particular synthesis reaction is **endothermic**, because it absorbs the surrounding heat.

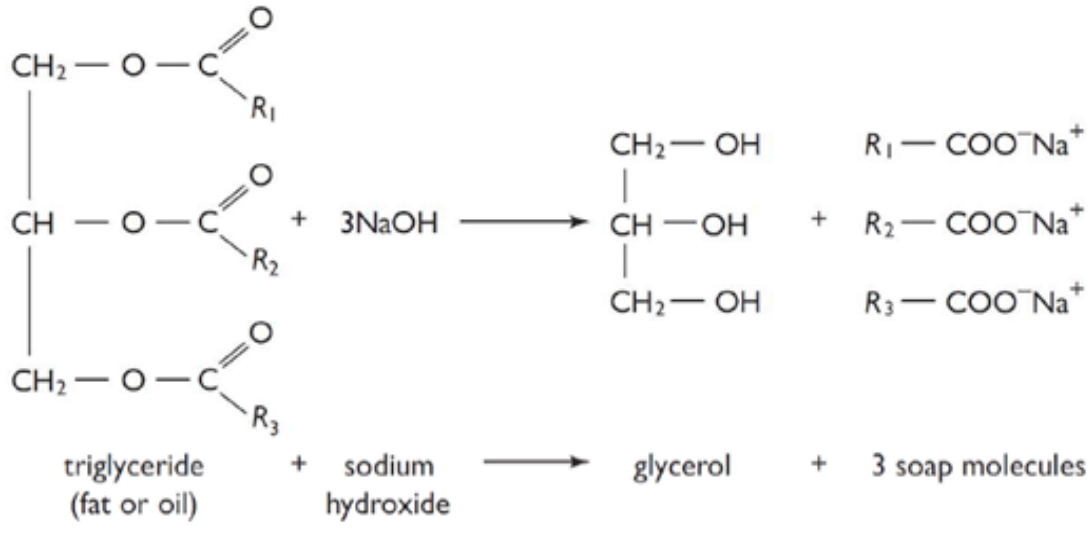

Figure 9. **Saponification is a synthesis reaction.**

REACTION CHECKLIST:

When determining what type of chemical reaction is taking place, follow the series of questions below. When you can answer "yes" to a question, then stop!

1. Does your reaction have oxygen as one of its reactants and carbon dioxide and water as products? If yes, then it's a combustion reaction.
2. Does your reaction have two (or more) chemicals combining to form one chemical? If yes, then it's a synthesis reaction.
3. Does your reaction have one large molecule falling apart to make several small ones? If yes, then it's a decomposition reaction.
4. Does your reaction have any molecules that contain only one element? If yes, then it's a single displacement reaction.
5. If you haven't answered "yes" to any of the questions above, then you've got a double displacement reaction.

Did You Know?

Amos Lattier used a single replacement reaction to build a battery out of 500 pounds of potatoes. Nails with zinc coating (galvanized nails) and the phosphoric acid in potatoes were the source of the electical power. The copper pieces in the potatoes provided places where the electrons from the zinc were passed on to the acid. His battery powered a sound system, and he drove around his neighborhood to show it off.

Lab 6: Pre-Lab Questions

Balance the following equations and determine which type of reaction is occurring.

1. $(NH_4)_2CO_{3\ (s)} \rightarrow NH_3 + H_2O + CO_{2\ (g)}$

2. $Zn_{\ (s)} + H_3C_6H_5O_{7\ (aq)} \rightarrow Zn_3(C_6H_5O_7)_{2\ (aq)} + H_{2\ (g)}$

3. $Zn(C_2H_3O_2)_{2\ (aq)} + Na_3PO_{4\ (aq)} \rightarrow NaC_2H_3O_{2\ (aq)} + Zn_3(PO_4)_{2\ (s)}$

Experiment 1

Lab 6
Types of Chemical Reactions

Experiment Inventory

Materials

0.5 g Ammonium Carbonate $(NH_4)_2CO_3$

4 g Sodium Bicarbonate, $NaHCO_3$, Baking Soda

2 mL 0.1M Sodium Phosphate Tribasic (Na_3PO_4)

2 mL Saturated Citric Acid Solution

2 mL 0.1M Zinc Acetate $(Zn(C_2H_3O_2)_2)$

Tea Light Candle

Permanent Marker

Zinc-coated (galvanized) Washer

Tap Water

Labware

(1) 10 mL Graduated Cylinder

(1) 50 mL Beaker

(1) 2 Glass Test Tubes

(1) Glass Stir Rod

(1) Metal Spatula

(1) Scale

(1) Test Tube Clamp

(1) Test Tube Rack

(1) Weigh Boat

(1) Butane Lighter

Note: You must provide the materials listed in red.

EXPERIMENT 1: TYPES OF CHEMICAL REACTIONS

In this experiment you will investigate different types of chemical reactions.

CAUTION: ***Long hair should be tied up and loose clothing restrained when around an open flame to prevent fire and burns. Be sure you are working in a well ventilated area and are wearing your safety goggles.***

PROCEDURE

PART 1: COMBUSTION

$$C_4H_{10}\,(g) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(g)$$

- [ ] 1. Light a butane lighter and observe the flame. The ignition of the flame is a reaction between butane and the oxygen in the air you breathe.
- [ ] 2. Record your observations in **Table 3**.

PART 2: DECOMPOSITION

$$(NH_4)_2CO_3\,(s) \rightarrow NH_3\,(g) + H_2O\,(g) + CO_2\,(g)$$

- [ ] 1. Using your permanent marker, label a clean, dry glass tube with 0.5 g ammonium carbonate [$(NH_4)_2CO_3$].
- [ ] 2. Make sure the scale is reading in the grams (g) mode. If you don't see a small g on the top right of your screen, push the button on the left hand side labeled M (for mode) until you see a g.

- [ ] 3. Place your weigh boat on the scale. Once you have the mass of your weigh boat, press the button on the right hand side (0/T). Your scale should now read 0.0 g.
- [ ] 4. Using your metal spatula, measure 0.5 g ammonium carbonate [$(NH_4)_2CO_3$] into your weigh boat.
- [ ] 5. Transfer the 0.5 g ammonium carbonate [$(NH_4)_2CO_3$] from your weigh boat into a clean dry glass test tube.
- [ ] 6. Place your test tube into the test tube rack.
- [ ] 7. Light the tea light candle using the butane lighter.
- [ ] 8. Take the test tube out of the rack and use a test tube clamp to hold the test tube at a 45 degree angle in the candle flame. Keep the open end of the tube pointed away from you.
- [ ] 9. Heat the test tube for a maximum of one minute. Record your observations in **Table 3**.
- [ ] 10. Place the test tube back into the test tube rack and allow the test tube to cool to room temperature before touching it.

PART 3: SINGLE REPLACEMENT

$$Zn\ (s) + C_6H_8O_7\ (aq) \rightarrow Zn_3(C_6H_5O_7)_2\ (aq) + H_2\ (g)$$

- [ ] 1. Place a test tube in the test tube rack.
- [ ] 2. Slightly tilt a clean test tube and slide a small zinc-coated (galvanized) washer down the side.
- [ ] 3. Use a 10 mL graduated cylinder to measure out 2 mL of saturated citric acid and carefully pour it into the test tube containing the zinc washer.
- [ ] 4. Observe the reaction for three minutes and record your observations in **Table 3**.
- [ ] 5. To clean up, separate the acid solution from the washer by pouring it into a 50 mL beaker while leaving the washer in the test tube. This is called decanting. Rinse the test tube containing the washer several times with water and add each rinse to the beaker.

  CAUTION: ***Do not pour the saturated citric acid directly down the drain. To neutralize the acid, add small amounts of baking soda directly from the packet to the solution in the beaker and stir with a stirring rod.***

- [ ] 6. Continue stirring and adding small amounts of baking soda until gas no longer forms. Pour the liquid down the drain and throw the washer in the trash.

PART 4: DOUBLE REPLACEMENT

$$3\ Zn(C_2H_3O_2)_2\ (aq) + 2\ Na_3PO_4\ (aq) \rightarrow 6\ Na(C_2H_3O_2)_2\ (aq) + Zn_3(PO_4)_2\ (s)$$

- [ ] 1. Thoroughly rinse your 10 mL graduated cylinder and allow it to air dry for one minute.
- [ ] 2. Thoroughly rinse your stirring rod and allow it to air dry for one minute.
- [ ] 3. Using your permanent marker, label a clean, dry glass test tube with the reaction that you are about to perform, i.e., 2 mL of 0.1 M zinc acetate ($Zn(C_2H_3O_2)_2$) + 2 mL of 0.1 M sodium phosphate tribasic (Na_3PO_4).
- [ ] 4. Using your clean, dry 10 mL graduated cylinder measure two mL of 0.1 M zinc acetate ($Zn(C_2H_3O_2)_2$) and add it to the test tube that you labeled.
- [ ] 5. Thoroughly rinse your 10 mL graduated cylinder and allow it to air dry for one minute.

- [ ] 6. Using your clean dry 10 mL graduated cylinder, measure 2 mL of 0.1 M sodium phosphate tribasic (Na3PO4) and add it to the test tube that you labeled.
- [ ] 7. Use your clean, dry glass stir rod to gently stir the test tube for one minute.
- [ ] 8. Record your observations in **Table 3**.
- [ ] 9. To clean up, pour the contents of the test tube down the drain with copious amounts of tap water.

Data Sheet

Lab 6
Experiment 1 Data Sheet

Table 3 Chemical Reaction Data

Reaction	Observation
Combustion	
Decomposition	
Single Replacement	
Double Replacement	

Lab 15: Post-Lab Questions

1. How could you verify that you produced carbon dioxide in your combustion reaction?

2. What indication did you have that NH_3 was produced in your decomposition reaction?

3. If you did not neutralize the saturated citric acid before you poured it down the drain describe the potential consequences to the plumbing and the environment.

Lab 6: Post-Lab Questions

4. Why is it important to balance chemical equations before performing the chemical reaction?

5. In your double replacement reaction, the more reactive metal "pushed" the other one out of its place. Without using the internet where could you look to determine which metal was more reactive?

Experiment 2

Lab 6
Synthesis of Soap

Experiment Inventory

Materials

20 g NaOH

10 mL of Vegetable Oil

Aluminum foil 10 cm x 10 cm square

Sterno® Cooking Fuel

Matches

Filter Paper (round)

Distilled Water

Safety Cup of Water

Labware

10 mL Graduated Cylinder

100 mL Graduated Cylinder

100 mL Glass Beaker

Funnel

Glass Stir Rod

Wire Mesh Stand

Hot Pad

Insulated Glove

Metal Spatula

Note: You must provide the materials listed in red.

EXPERIMENT 1: SYNTHESIS OF SOAP

In this experiment you will use the process of saponification to make your own bar of soap.

LAB SAFETY: ***Both the matches and the Sterno® can cause fire or burns to skin, clothing, or lab materials if the flame comes into contact with them. Do not stick any body part, loose clothing, or other flammable items in or near the flame. Be sure you have your safety glasses on and never leave any burning fuel unattended!! Do this exercise in a well-ventilated area and have a fire extinguisher nearby.***

PROCEDURE

- [ ] 1. Place a 100 mL glass beaker on the scale. Once you have the mass of your beaker, press the button on the right hand side (0/T). Your scale should now read 0.0 g.
- [ ] 2. Use your metal spatula to measure 20 g of NaOH in the 100 mL beaker. Your scale should now read 20 g.

  Warning: ***In the next few steps, you will want to avoid splashes, as NaOH will cause severe burns.***
- [ ] 3. Using your 100 mL graduated cylinder, measure 80 mL of distilled water.

  Warning: ***In the next step, your graduated cylinder will become warm to the touch.***
- [ ] 4. Very slowly pour the 20 g of NaOH into the 80 mL of distilled water and stir the solution with a stir rod until the NaOH is fully dissolved. You have just made a 20% NaOH solution.
- [ ] 5. Using your 10 mL graduated cylinder, measure 10 mL of vegetable oil.
- [ ] 6. Pour the vegetable oil into a clean dry 100 mL glass beaker.
- [ ] 7. Using your 100 mL graduated cylinder, measure 15 mL of the 20% NaOH solution.

- [ ] 8. Add the 15 mL of the 20% NaOH solution to the glass beaker containing your vegetable oil.
- [ ] 9. Using your glass stir rod, mix the solutions together (stir for at least one minute) then cover with the aluminum foil. The purpose of the aluminum foil is to protect you from the splattering that will occur in the preceding steps.
- [ ] 10. Choose an area that is well-ventilated (yet not windy) and place your Sterno® cooking fuel can on a flat, heat resistant surface.
- [ ] 11. Thoroughly read all of the directions on your Sterno® cooking fuel can, and then use your spatula to pry open the lid. Put the lid aside, being careful not to touch any of the contents. If you accidently touch the contents, wash your hands before you proceed with the experiment.
- [ ] 12. Touch a lit match to the Sterno® contents and drop the match into the contents. The contents will flame up immediately. Under bright lights you may not be able to see the flame, but know that it is there.
- [ ] 13. Use your metal spatula to slide the Sterno® cooking fuel can under the wire mesh stand.
- [ ] 14. Carefully place your beaker on top of the wire mesh stand.
- [ ] 15. Bring the mixture in the beaker to a boil.
- [ ] 16. Partially remove the aluminum foil covering so that you can stir the mixture. To avoid splattering, you must continuously stir vigorously.
- [ ] 17. Boil the mixture until all of the water has boiled off. When the reaction is almost complete, the mixture will get very thick, and you may need to remove it from the heat intermittently to keep it from burning to the bottom of the beaker.
- [ ] 18. Once the water has been removed, let the mixture cool slightly. Check the mixture. If there is a waxy solid forming, the reaction is complete. If the mixture is still the consistency of a thick syrup, you will need to boil it longer.
- [ ] 19. When the reaction is complete, use the insulated glove to place the hot beaker on a hot pad to cool. **Do not place the hot beaker directly on the counter because it can crack the glass!**
- [ ] 20. Extinguish your Sterno® cooking fuel can by placing the cover on top of the can. **Do not blow out the flame**. Once the can is cooled (look at the temperature indicator on the label) snuggly fit the cover onto the can and store it in a safe place away from heat or flames.

Cleaning the NaOH Out of Your Soap

- [ ] 1. Place a clean, dry 100 mL glass beaker on the scale. Once you have the mass of your beaker, press the button on the right hand side (0/T). Your scale should now read 0.0 g.
- [ ] 2. Measure 18 g of NaCl in the 100 mL beaker. Your scale should now read 18 g.
- [ ] 3. Using your 100 mL graduated cylinder, add 60 mL of distilled water to your NaCl and stir with a glass stir rod until completely mixed.
- [ ] 4. Once the soap mixture cools to room temperature, add 20 mL of the NaCl solution that you just made to the beaker with the soap in it.
- [ ] 5. Stir the mixture with your stirring rod and break up the larger chucks of soap.
- [ ] 6. Prepare a filtering funnel as shown in Figure 10. Fold a piece of filter paper in half twice to make 4 quarters, and place the paper in the funnel so that three quarters are open on one side and one quarter is on the opposite side. Seat the filter paper into the funnel by moistening the paper with a small amount of water.
- [ ] 7. Place the funnel inside of a 250 mL Erlenmeyer flask.

- [ ] 8. Place this mixture into the funnel to filter out the liquid.
- [ ] 9. Carefully scrape the soap off of the filter paper, put the soap back into the beaker, add 20 mL of NaCl solution, stir up the mixture, and break up the large chunks of soap. Then, filter this mixture into the funnel (remember to reform and put your filter paper back into the funnel). Repeat these steps with the remaining 20 mL of NaCl solution.

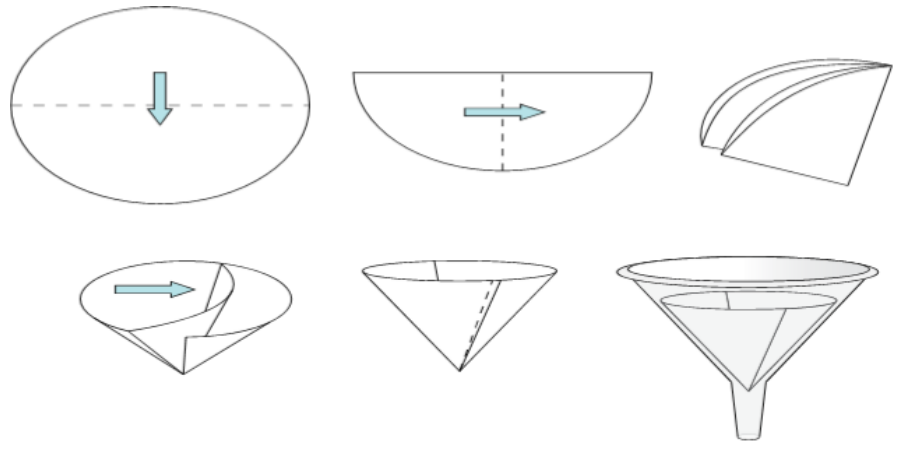

Figure 10. Step-by-step process of folding the filter paper so that it will fit into the funnel.

- [ ] 10. After your last washing and filtering, scrape the soap off of the filter paper onto a paper towel and pat the soap to remove the moister.
- [ ] 11. Press the soap into a shape such as a circle or a rectangle.
- [ ] 12. Evaporate any last remaining particles of water out of your soap by allowing it to naturally air dry. This will complete the saponification process and usually takes about two weeks, depending on your climate.
- [ ] 13. If desirous, you can allow your soap to air dry for a longer period of time (six to eight weeks). This longer evapo-ration time will insure a milder, longer lasting soap with a rich lather.

Lab 6: Post-Lab Questions

1. What is the purpose of continuously stirring the vegetable oil and the NaOH?

2. What would happen if you did not rinse out the excess NaOH from your soap?

3. What is the purpose of letting the soap cure before using it?

4. Could you use any kind of fat/oil in the saponification reaction? Justify your answer.

5. If you wanted to add a dye, or a perfume to your soap, at what point in the saponification would you add these extra ingredients?

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Lab 7

Using the Ideal Gas Law

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LAB 7

Using the Ideal Gas Law

Learning Objectives

- **Determine the relationship between pressure and temperature**
- **Understand how to use Charles's Law**
- **Understand how to use the Ideal Gas Law**

INTRODUCTION

Earth, fire, water, and air (Figure 1) made up the known elements in the ancient world that philosophers such as Aristotle and Plato observed. Aristotle believed that different combinations of these elements filled all of space, and that air, being a light substance, moved outwards away from the universe's center. Today, scientists know that Earth's gravity attracts particles and gases in the air towards the planet, making the air denser as it approaches the surface.

Figure 1. Air consists primarily of nitrogen (78%) and oxygen (21%).

PROPERTIES OF GASES

Isaac Beeckman (17th century) was one of the first people to realize that air had real substance. He speculated that air had weight, pressed down on objects on Earth, and was expandable. In this way, gases have several unique properties. For example, a gas expands spontaneously to fill its container, and the volume of a gas equals the volume of the container in which it is held. Gas is also highly compressible. When pressure is applied to a gas, its volume readily decreases. To further understand the properties of gases, it is important to appreciate temperature, pressure, and volume.

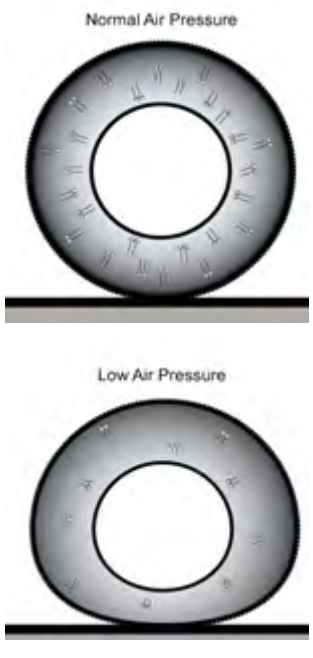

Figure 2. When temperature decreases, molecules move slower. This decrease in velocity reduces the force applied to the tire causing a less inflated tire.

TEMPERATURE

Temperature (T) is the property of matter which reflects the kinetic energy of the particles. There are several standardized scales used to measure this value (e.g., Kelvin, Celsius, and Fahrenheit). An increase in temperature means an increase in the motion of atoms in a material (Figure 2). Gas particles move past one another with little restriction, this is why gases occupy an entire container or space, while liquids and solids have a confined volume and surface area.

PRESSURE

Pressure (P) is defined as the amount of force gas molecules exert on the area of space they fill. If you have ever had low tire pressure due to cold weather, it was because the gas molecules that fill, were moving slower. Therefore, they did not exert as much force on the inside of the tire as they normally would (Figure 2). Changing a container's volume can also change pressure.

VOLUME

Volume (V) is simply the amount of space a gas occupies. Pressure decreases when volume increases because the molecules have more space in which they can travel. Pressure increases when volume decreases, because the molecules have less space in which they can travel (Figure 3).

Using the Ideal Gas Law

ATMOSPHERE

Atmosphere refers to the gases surrounding a star or planet held in place by gravity as seen in Figure 1. Atmosphere is also a **unit of pressure**. One atmosphere is defined as the pressure caused by the weight of air above a given point. Normal atmospheric pressure at sea level is 14.7 pounds per square inch (lb/in^2). In chemistry and in various industries, the standard pressure that is used is one atmosphere (1 atm).

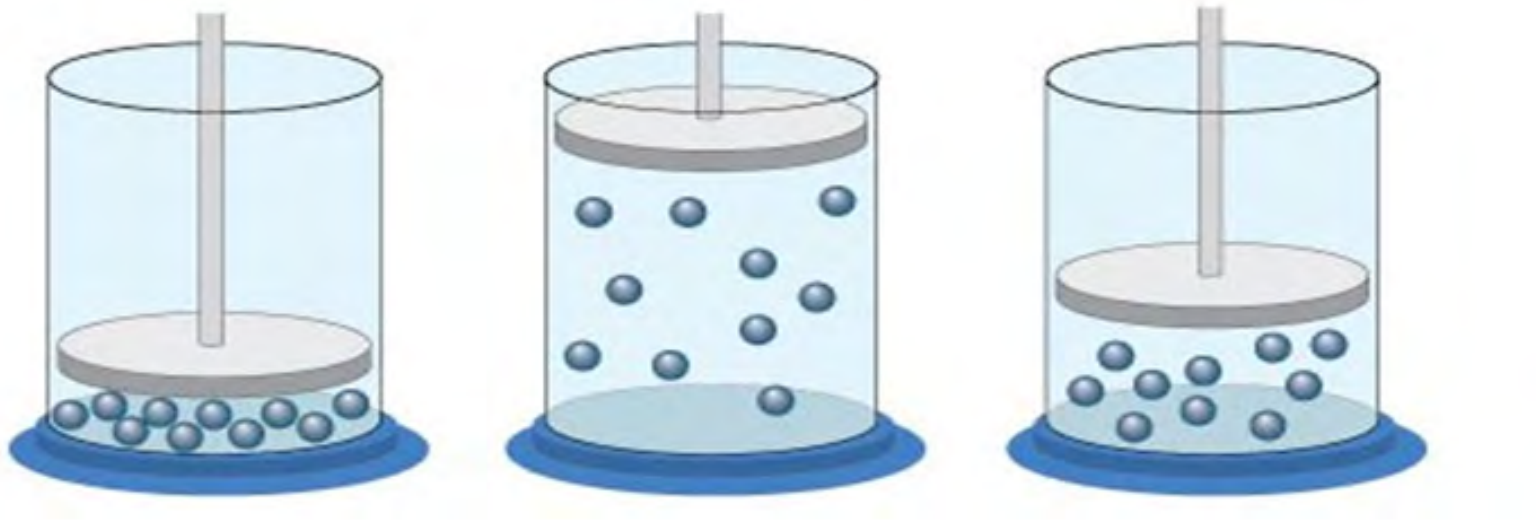

Figure 3. When the volume decrease, the pressure increases. Conversely, when the volume increases, the pressure decreases. This is what is known as an inverse relationship and is demonstrated in Boyle's Law (Figure 5).

GAS LAWS

There are three fundamental gas laws that describe the relationship between pressure, temperature, volume, and amount of gas. **Boyle's Law** tells us that the volume of gas increases as the pressure decreases. **Charles's Law** tells us that the volume of gas increases as the temperature increases. Finally, **Avogadro's Law** tell us that the volume of gas increases as the amount of gas increases. The **Ideal Gas Law** is the combination of the three simple gas laws.

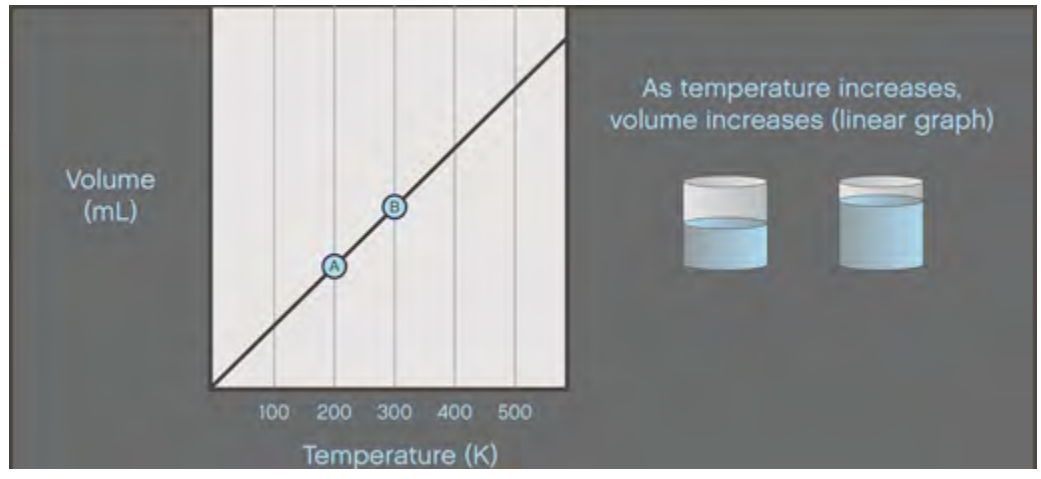

Figure 4. Point A on the graph represents the initial volume (V_1) while Point B represents the final volume (V_2). Notice that as the temperature increases the gas expands, thus increasing its volume.

CHARLES'S LAW

Charles's Law was developed in the 19th century by Jacques Charles, a French chemist. Charles investigated the variation of volume with temperature for a fixed mass of gas at constant pressure. He found that as temperature decreases, the volume of a gas also decreases (Figure 4). Further, as temperature increases, the volume of a gas also increases. The two variables are directly proportional. The mathematical expression of this is:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad or \quad V_1 T_2 = V_2 T_1$$

KELVIN

Charles is also credited with discovering that all gas law equations must be calculated in **Kelvin.** Kelvin is a temperature scale designed so that zero degrees K (written without the degree symbol) is defined as **absolute zero**. Absolute zero is a temperature in which all molecular movement stops. To convert from Celsius to Kelvin simply add 273. Note that you do not use the degree symbol when temperature is expressed in Kelvin.

$$K = ℃ + 273$$

Using the Ideal Gas Law

EXAMPLE USING CHARLES'S LAW:

The volume of a gas sample is 746 mL at 20 °C. What is its volume at body temperature (37 °C)? Assume the pressure remains constant.

STEP 1:

Use Charles's Law formula and write down all of the known values of the variables.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad or \quad V_1 T_2 = V_2 T_1$$

Initial Volume = V_1 = 746 mL
Initial Temperature = T_1 = 20 °C
Final Temperature = T_2 = 37 °C
Final Volume = V_2 = unknown variable to solve for

STEP 2:

Unit conversion.

$$T_1 = 20°C + 273 = 293\ K$$

$$T_2 = 37°C + 273 = 310\ K$$

STEP 3:

Convert the formula to solve for the desired variable, plug in the known quantities, and calculate the unknown variable.

$$V_2 = V_1 \frac{T_2}{T_1}$$

$$V_2 = 746\ mL \times \frac{310\ K}{293\ K}$$

$$V_2 = 786\ mL$$

STEP 4:

Check the logic of the answer. Is the answer reasonable? The final volume is larger than the initial volume. Since the temperature increased from 293 K to 310 K, it is reasonable that the volume of the gas increased.

Using the Ideal Gas Law

BOYLE'S LAW

Boyle's Law is the gas law that relates pressure and volume. Boyle's Law states that when temperature and amount of chemical are held constant, pressure and volume are inversely proportional. In other words, as the pressure increases, the volume decreases (Figure 5). This is mathematically expressed as:

$$P_1V_1 = P_2V_2$$

Boyle's Law further states that the product of pressure and volume is a constant. This constant is referred to as "k".

$$PV = k$$

For example, suppose an experiment begins with a pressure of 1 atmosphere (atm) and a volume of 10 L. In this case, k is equal to 1 atm x 10 L = 10 atm x 1 L. As the experiment proceeds, the pressure increases to 2 atm and the volume decreases to 5 L. As this happens, k will remain the same because 2 atm x 5 L is still equal to 10 atm x 1 L. Therefore, $P_1V_1 = P_2V_2$ is just another way of expressing k = k.

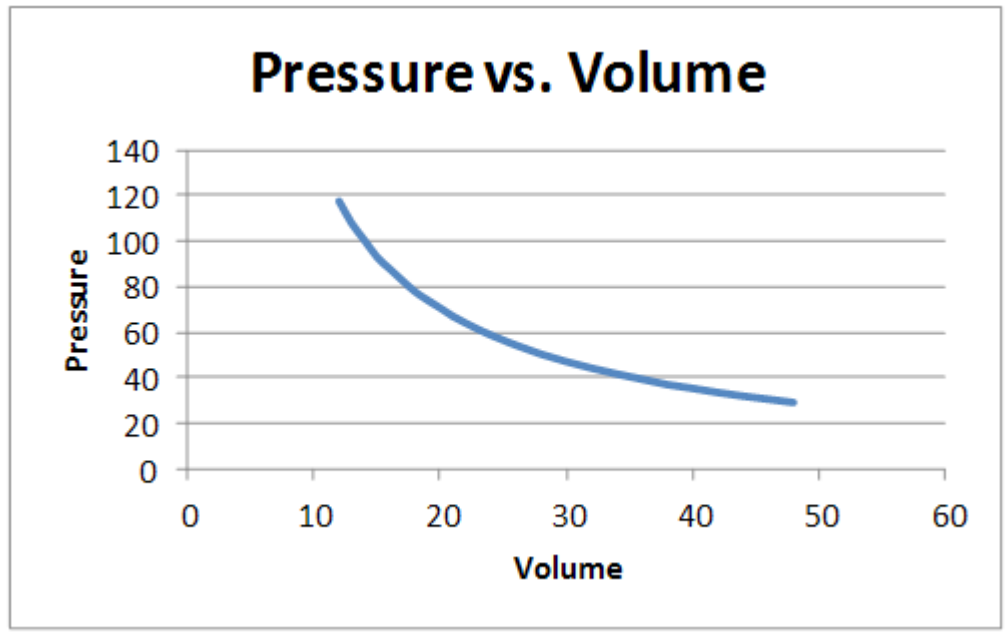

Figure 5. The relationship between the pressure and volume of an ideal gas. With a constant temperature and number of moles, pressure will decrease as volume increases.

AVOGADRO'S LAW

Avogadro's Law describes the relationship between volume (V) and the amount of gas (n) (measured in moles) when pressure and temperature are held constant (Figure 6). For example, 1L of carbon dioxide will have just as many molecules in it as a liter of nitrogen. The mathematical expression of this is:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

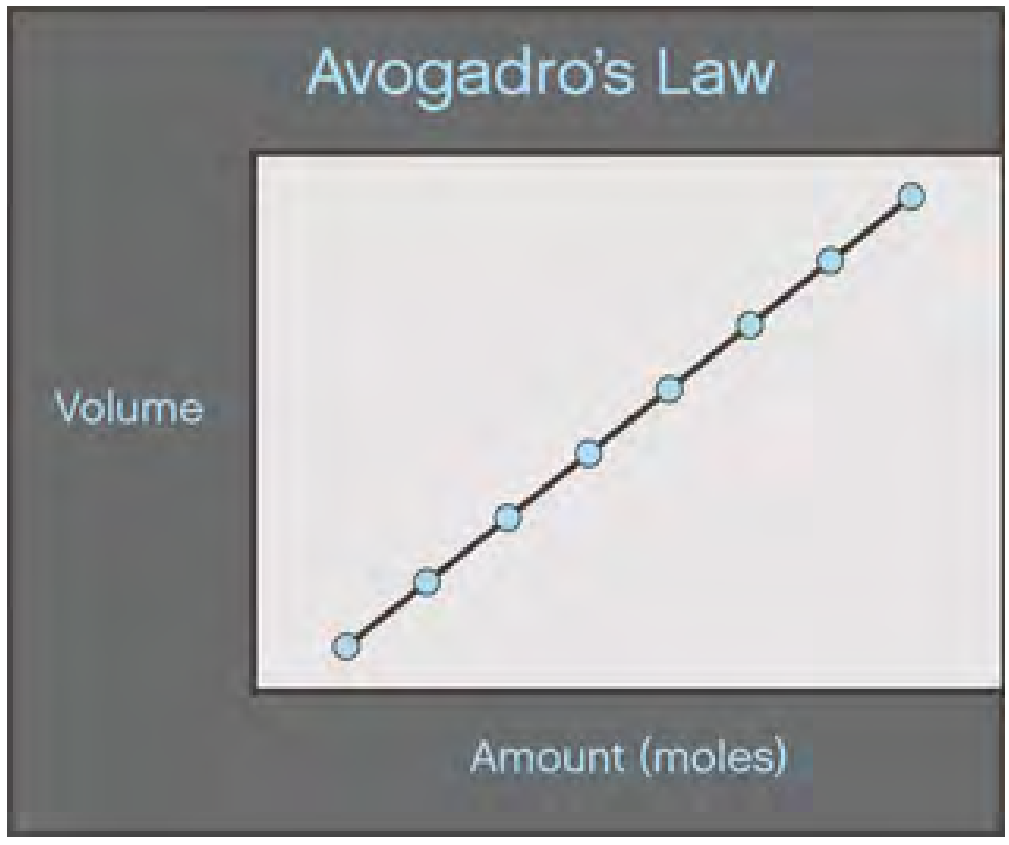

Figure 6. There is a direct relationship between volume and amount of a gas. As the volume increases, so will the number of moles of the gas.

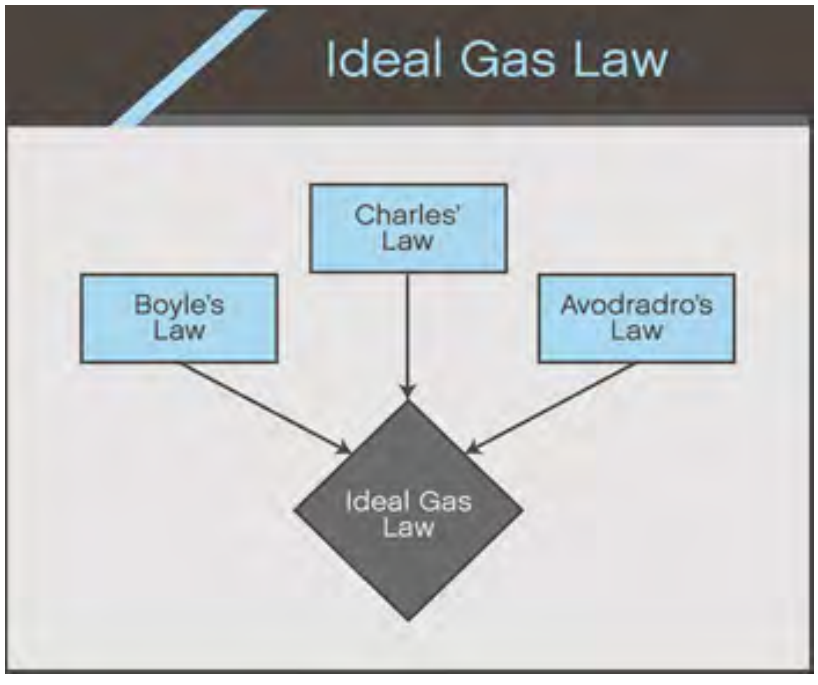

Figure 7. An ideal gas is an idealized model of a real gas.

IDEAL GAS LAW

The **Ideal Gas Law** is a combination of all the simple gas laws (Figure 7). The Ideal Gas Law describes the behavior of real gases as long as the pressure and temperature are not too extreme. The Ideal Gas Law is very valuable when dealing with gases since it establishes a relationship between temperature, pressure, volume, and amount of a gas. The behavior of an ideal gas, that is, the relationship of pressure (P), volume (V), and temperature (T) can be mathematically expressed as:

$$PV = nRT$$

Where:

- **P** = the gas pressure in atmospheres (atm)
- **V** = the volume of the gas in liters (L)
- **n** = the number of moles of the gas
- **R** = the constant value of 0.0821 L x (atm/mol x K)
- **T** = for the temperature of the gas in Kelvin

Using the Ideal Gas Law

EXAMPLE USING THE IDEAL GAS LAW:

5.0 g of neon is at 256 mm Hg and at a temperature of 35 °C. What is the volume?

STEP 1:

Use the Ideal Gas Law formula and write down all of the known values of the variables.

$$P = 256\ mm\ Hg$$
$$V = unknown$$
$$n = 5.0g$$
$$R = 0.0820574\ L \cdot atm \cdot mol^{-1} \cdot K^{-1}$$
$$T = 35°C$$

STEP 2:

Unit conversion (round up).

$$P = 256\ mm\ Hg \times \frac{1\ atm}{760\ mm\ Hg} = 0.3368\ atm$$

$$n = 5.0\ g\ Ne \times \frac{1\ mol}{20.1797\ g} = 0.25\ mol$$

$$T = 35°C + 273 = 308\ K$$

STEP 3:

Convert the formula to solve for the desired variable, plug in the known quantities, and calculate the unknown variable (round up).

$$V = \frac{nRT}{P}$$

$$V = \frac{0.25\ mol \times 0.08206 \frac{L\ atm}{K\ mol} \times 308\ K}{0.3368\ atm} = 19\ L$$

STEP 4:

Check the logic of the answer. Imagine 9.5 two liter bottles of a soft drink all stacked on top of each other. That is the volume (19 L) that was calculated.

Using the Ideal Gas Law

HYDROGEN PEROXIDE

Now that we know how to use the Ideal Gas Law, let us take a look at a couple of real world reagents that can be used in the application of the Ideal Gas Law. **Hydrogen peroxide (H_2O_2)** is a chemical used in industry and research for its oxidative properties and in medicine as an antiseptic. You may have a bottle of 3% hydrogen peroxide in your medicine cabinet. At room temperature, a water solution of hydrogen peroxide (H_2O_2) will undergo an extremely slow decomposition reaction to form molecular oxygen gas (O_2) and liquid water (H_2O). Since hydrogen peroxide forms oxygen gas when it decomposes, we can use the Ideal Gas Law to observe this decomposition reaction. However, since the reaction will decompose very slowly (weeks) we will use a catalyst.

Figure 8. Catalase catalyzes the decomposition of hydrogen peroxide to water and oxygen.

CATALASE

A **catalyst** is a substance that speeds up a reaction, without being part of the reaction. Yeast cells contain a catalyst called **catalase**. Catalase speeds up the decomposition of hydrogen peroxide to water and oxygen (Figure 8). Catalase is very effective at decomposing hydrogen peroxide. In fact, one molecule of the enzyme can catalyze the conversion of over 6,000,000 hydrogen peroxide molecules into water and oxygen every second.

Lab 7: Pre-Lab Questions

1. Write the balanced equation for the decomposition of hydrogen peroxide.

2. According to Charles's law, what is the relationship between temperature and pressure?

3. Atmospheric pressure depends on the altitude (or height) of your location. How should the air pressure change if you were in Denver, Colorado, which is 1.5 kilometers (1 mile) above sea level?

4. Considering that catalysts are not consumed in a reaction, how do you think increasing the amount of catalyst would affect the reaction rate for the decomposition of hydrogen peroxide?

Experiment 1

Lab 7

Charles Law

Experiment Inventory

Materials

(1) Styrofoam Cup

Graph Paper

Ice

Pencil

Tap Water

Labware

(1) 10 mL Sealable Syringe with Cap

(1) Thermometer

(1) Stopwatch or Clock

Note: You must provide the materials listed in red.

EXPERIMENT 1: CHARLES'S LAW

In this experiment, you will explore the relationship between temperature and volume, and connect this to Charles's Law.

PROCEDURE

- [ ] 1. Fill the syringe halfway with air, and seal the syringe by screwing the cap on. Record the volume (mL) of gas in the cylinder in **Table 1.**
- [ ] 2. Use the stopwatch and hold the top of the thermometer up in the air for 30 seconds. Record the temperature of the room in **Table 1.**
- [x] 3. Fill one Styrofoam® cup with hot water from the sink. Place the thermometer in the cup.

   LAB SAFETY: ***Avoid using extremely hot water. Handle hot water carefully.***
- [ ] 4. Fully submerge and hold the air-filled syringe in the hot water for thirty seconds. Use the stopwatch to monitor time.
- [ ] 5. Gently push the piston down until the plunger cannot move any farther. Observe the new volume inside of the syringe. Record this volume and the water temperature in **Table 1**.
- [ ] 6. Prepare a cup full of water mixed with ice. Repeat steps 4 - 5 for the cold water bath. Record the volume and temperature in **Table 1**.

Data Sheet

Lab 7
Experiment 1 Data Sheet

Table 1 Temperature vs. Volume of Gas Data

Temperature Conditions	Temperature (°C)	Volume (mL)
Room Temperature		
Hot Water		
Ice Water		

Lab 7: Post-Lab Questions

1. Use a pencil and graph paper to create a graph of temperature and volume data. Place temperature on the x-axis (in Kelvin) and volume (mL) on the y-axis. Leave room on the left side of your chart for temperature values below zero. You can also use a graphing program to create your graph.

2. What happened to the volume of gas when the syringe was exposed to various temperature conditions? Using the concepts explored in the Introduction, describe why this occurred, keeping in mind the definition of temperature.

3. Using a ruler, draw a straight line of best fit through your data points, extrapolating the line until it intersects the (negative) x-axis. Why can you assume a linear relationship (a straight-lined slope)?

4. At what temperature does your line intersect the x-axis? What volume corresponds to this temperature?

Experiment 2

Lab 7
Using the Ideal Gas Law

Experiment Inventory

Materials

10 mL Hydrogen Peroxide, H_2O_2

Yeast Packet

Paper Clip

Rubber Band

Clear Tape

Internet

100 mL Warm Tap Water

500 mL Distilled Water

Labware

(1) 10 mL Graduated Cylinder

(1) 100 mL Graduated Cylinder

(1) 250 mL Beaker

(1) 250 mL Erlenmeyer Flask

(1) 500 mL Beaker

(1) 24 in. Flexible Tubing

(1) 3 in. Rigid Tubing

(2) Pipettes

(1) Glass Stir Rod

(1) Thermometer

(1) Stopper with 1-hole

(1) Stopwatch / Clock

Note: You must provide the materials listed in red.

EXPERIMENT 1: USING THE IDEAL GAS LAW

In this experiment you will use a catalyst to observe the Ideal Gas Law.

PROCEDURE

☐ 1. Prepare the materials for the apparatus as shown in Figure 9. Insert the smaller rigid tubing into one end of the larger, flexible tubing. Insert the free end of the rigid tubing securely into the rubber stopper hole.

Caution: ***Be careful when working with rigid tubing. This tubing is made of glass and can break if excess force is applied.***

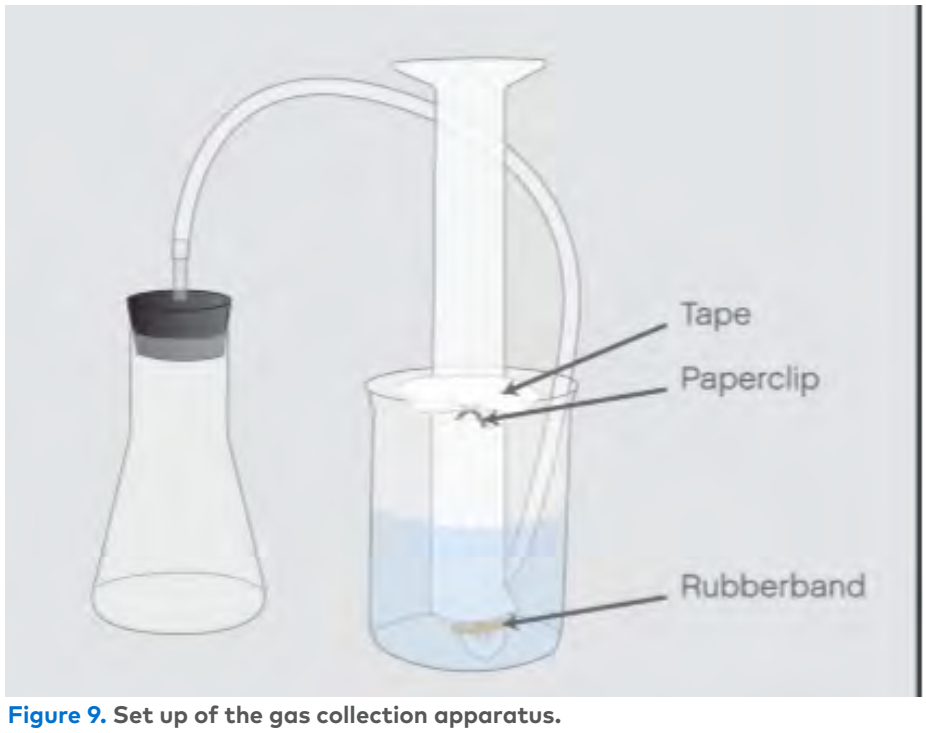

Figure 9. Set up of the gas collection apparatus.

- [ ] 2. Bend the free end of the flexible tubing into a U-shape, and use a rubber band to hold this shape in place. This will allow you to more easily insert the end of the flexible tubing into the inverted graduated cylinder. Make sure the tubing is not pinched and that gas can flow freely through it.
- [ ] 3. Fill the 100 mL graduated cylinder with distilled water to the 100 mL mark.
- [ ] 4. Fill the 500 mL beaker with 400 mL of distilled water.
- [ ] 5. Take the temperature of the water in the 500 mL beaker, and record it in Table 2. Also, use the internet to determine the barometric pressure in the room, and record it in **Table 2**.

> **Hint:** If necessary, use the regional pressure as a close substitute to the room pressure. This can easily be found online - if necessary, convert this value to atm.

- [ ] 6. Mix 100 mL of warm water (45 °C) and one packet of baker's yeast in a 250 mL beaker. This will activate the yeast from the dormant (dry) state. Be sure to mix well with a glass stir rod until the yeast is completely dissolved.

> **Hint:** Water must be between 42 - 47°C otherwise you will kill the yeast (no activation).

- [ ] 7. Use a 10 mL graduated cylinder and pipette to measure out 5 mL of hydrogen peroxide. Pour the hydrogen peroxide into the Erlenmeyer flask, and securely place the stopper with stopper tube into the top of the Erlenmeyer flask.
- [ ] 8. Clean the 10 mL graduated cylinder by rinsing it at least three times with tap water. Dispose of the rinse down the drain.
- [ ] 9. Cover the opening of the graduated cylinder with two or three fingers and quickly turn it upside down into the 500 mL beaker already containing 400 mL of water. DO NOT remove your fingers from the opening until the graduated cylinder is fully submerged under the water. If the amount of trapped air exceeds 10 mL, refill the cylinder and try again.
- [ ] 10. Insert the U-shaped flexible tubing into the beaker, and carefully snake it into the submerged opening of the graduated cylinder. You want as little air as possible to be in the graduated cylinder.
- [ ] 11. Secure the graduated cylinder to the beaker by bending a paper clip around the graduated cylinder and using a piece of tape to hold it in place.
- [ ] 12. With the cylinder vertical, record the volume of air inside (the line at which the water reaches in the cylinder) in **Table 2**.
- [ ] 13. Using the pipette, measure out 5 mL of yeast solution into the rinsed 10 mL graduated cylinder.

> **Note:** Do not immediately pour the yeast solution into the Erlenmeyer flask.

- [ ] 14. Remove the stopper (still connected to the hose) from the Erlenmeyer flask. Get the stopwatch or clock ready.
- [ ] 15. Quickly pour the 5 mL of yeast solution into the Erlenmeyer flask. Immediately place the stopper securely in the opening of the Erlenmeyer flask by twisting it down into the flask gently.
- [ ] 16. Start timing the reaction with the stopwatch or clock. Record the time in Table 3.

- [ ] 17. Swirl the Erlenmeyer flask to mix the two solutions together.
- [ ] 18. Bubbles will begin to form in the 100 mL graduated cylinder.

> **Hint**: If gas bubbles are not immediately visible, make sure the stopper is on tight enough and the tubing is not leaking.

- [ ] 19. Continue to swirl the Erlenmeyer flask and let the reaction run until no more bubbles form (to assure the reaction has gone to completion).

> **Hint:** Catalase in the yeast works best around the temperature of the human body. You can speed the reaction up by warming the Erlenmeyer flask with your hands.

- [ ] 20. Record the time when the reaction is finished in **Table 3** of the data section, along with the final volume of air in **Table 2**. Remember to read it at eye-level and measure from the bottom of the meniscus.
- [ ] 21. Pour all other liquids down the drain and clean the labware.

Lab 7
Experiment 2 Data Sheet

Table 2 Temperature, Pressure and Volume Data

Temperature of Distilled H_2O	Room (or regional) Pressure (atm)	Initial Volume of Air (mL)	Final Volume of Air (after reaction) (mL)	Volume of O_2 Collected (Final Volume - Initial Volume)

Table 3 Reaction Time Data

Time Reaction Started	Time Reaction Ended	Total Reaction Time

Lab 7: Post-Lab Questions

1. What would happen if you added more than 5 mL of yeast to the H_2O_2?

2. What would happen if you added more than 5 mL H_2O_2 to the 5 mL of yeast?

3. What was going on in the graduated cylinder as the H_2O was pushed out?

4. How would the number of moles (n) of O_2 change if your atmosphere was doubled and all other variables

Lab 7: Post-Lab Questions

stayed the same?

5. How would the number of moles (n) of O_2 change if your temperature was doubled and all other variables stayed the same?

6. In this experiment, the temperature of the gas evolved is equal to the temperature of the water in the beaker, which ideally should be the same as the air temperature. Explain how the volume of oxygen evolved would change if you used ice water instead of room temperature water. How would it change if you used boiling water?

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Labs

Lab 8

Titrations and Equivalence Points

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LAB 8

Titrations and Equivalence Points

Learning Objectives

- **Utilize the laboratory method of titration**
- **Quantitatively choose an indicator for a titration**
- **Determine the end point and the equivalence point of a titration**

INTRODUCTION

Titration is a laboratory method used to determine the unknown concentration of a reactant. Titration is used in laboratory medicine to determine unknown concentrations of chemicals in blood and urine. Titration may also be used to determine the amount of certain chemicals in food. Often, titration is used to determine fat content, water content, and concentrations of vitamins. This is a delicate procedure that requires patience and good observation skills. Usually, an indicator is used to determine the endpoint of the reaction.

CHEMICAL INDICATORS

Much like the way the human body indicates change, there are many **chemical indicators** used in chemistry. The characteristics of an indicator vary; however, the most common indicator is color. Color indicators are halochromic chemicals that reveal the pH of a chemical reaction through color changes (Figure 1). In other words, they provide a visual representation of the amount of hydrogen or hydronium ions which exist in a solution. They are commonly used because they are so easy to view and regulate in an experiment. For example, the chemical litmus (made available as litmus paper) is often used to provide an indication of a solution's pH. Imprecise readings may occur due to the fact that color observations may be subject to individual discrimination and ability. However, color scales are often provided for circumstances in which color differentiation is difficult. Additionally, pH meters can be used to determine the numeric pH value of a solution.

ACID-BASE TITRATION INDICATORS

Color indicators are used in acid-base titrations to reveal the exact amount of acidic and basic solution that is needed to create a neutral solution. This is useful because this allows chemists to determine molar concentration of an unknown solution (provided that the molar concentration of the other solution used in the titration is known). There is a large variety of indicators used in acid-base titrations, all of which change color at a different pH. It is therefore important to select an appropriate indicator for the chemical pair being titrated. Some complex titrations even require multiple indicators for easy visualization.

Figure 1. Swimming pool test kits are used as small titration experiments.

ENDPOINT AND EQUIVALENCE POINT

The point at which the indicator changes color in titration is known as the end point of the titration. This is the moment when color of the indicator changes. However, the **equivalence point** (or, neutralization point) is often what a chemist is trying to find (in order to determine molarity). The equivalence point is the point in which the combined acid-base solution reaches a neutral pH. Fortunately, if the correct indicator is chosen, the end point will occur at approximately the same moment as the equivalence point. Skillful technique (slow titration) can provide an accurate depiction of where this point occurs on a titration curve.

Titrations and Equivalence Points

INDICATOR SELECTION

To better grasp the concept of **indicator selection**, let's take a look at the titration curve in Figure 2. As you can see, the equivalence point (or, midpoint) occurs at a pH of approximately 7. Suppose Indicator A changes color (i.e., hits the end point) around a pH of 12 - 13. This would not be a good indicator for this titration because the end point is so far above from the equivalence point on the pH scale. Indicator C changes color around a pH of 3 - 4. This is also not a good indicator for this titration because the end point is so far below the equivalence point. Indicator B changes color around a pH of 7. This is a good indicator for this titration because the color change (end point) takes place at nearly the same pH as the equivalence point.

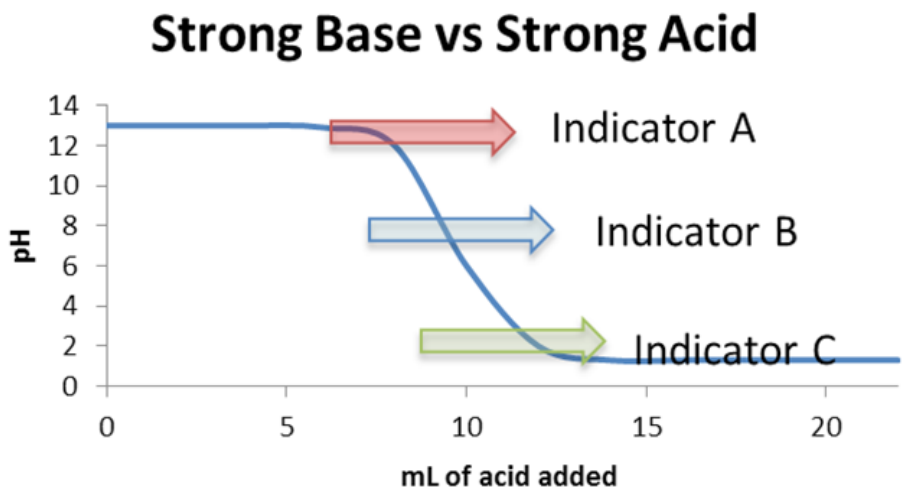

Figure 2. Indicators in acid-base titration.

TYPES OF INDICATORS

Reliable information exists to inform scientists of the different indicator qualities. This information, along with the types of chemical solutions being titrated and the known molar concentration of one of the solutions, helps scientists select the correct indicator for an experiment. Table 1 lists some of the most common indicators.

Table 1: Common Titration Indicators

	pH Range	Base Color (High on the pH Scale)	Acid Color (Low on the pH Scale)
Methyl Yellow	2.9 - 4.0	Yellow	Red
Methyl Orange	2.5 - 4.4	Yellow	Red
Congo Red	3.0 - 5.0	Red	Blue-Violet
Bromothymol Blue	6.0 - 7.6	Blue	Yellow
Phenol Red	6.4 - 8.2	Red	Yellow
Phenolphthalein	8.3 - 10.0	Fuschia	Colorless

INDICATOR EXAMPLE

Suppose your teacher gives you an assignment in which you need to determine the molar concentration of a hydrochloride solution (acidic). You choose to do this through an acid-base titration using a 0.1 M sodium hydroxide primary standard solution. To begin, you place the solution with the unknown molarity in a flask or beaker below the buret (Figure 3) or syringe. Then, you must add the indicator. Based on your knowledge of hydrochloride and sodium hy-droxide, you select phenolphthalein as your indicator and add a few drops of the indicator to the hydrochloride solution. Because hydrochloride is acidic, the phenolphthalein is in a low pH environment and it will remain in its colorless state.

Make sure the stopcock is closed (Figure 3), pour the sodium hydroxide into a buret, and begin to slowly titrate it into the hydrochloride by opening the stopcock. You continue in this fashion, looking for a pink-fuschia color change. As soon as you see a faint trace of pink appear in the hydrochloride solution, your reaction has reached the endpoint and the acid has been neutralized. You can then check on the amount of sodium hydroxide required to neutralize the solution, and use stoichiometry to determine the molarity using the equation:

$$N_a \times M_a \times V_a = N_b \times M_b \times V_b$$

Where:

- $\mathbf{N}_a$ = Moles of acid
- $\mathbf{N}_b$ = Moles of base
- $\mathbf{M}_a$ = Molarity of the acid
- $\mathbf{M}_b$ = Molarity of the base
- $\mathbf{V}_a$ = Volume of the acid
- $\mathbf{V}_b$ = Volume of the base

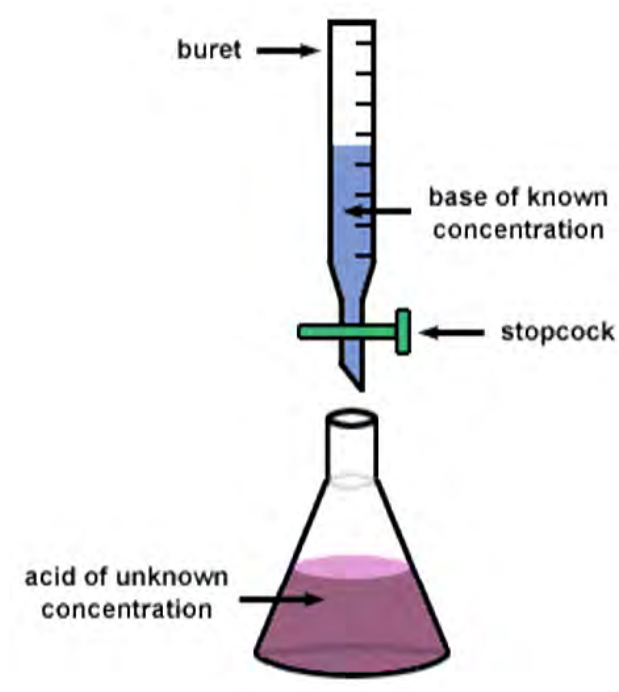

Figure 3. Sample acid-base titration set-up.

Lab 8: Pre-Lab Questions

1. What is the difference between an end point and an equivalence point?

2. Table 1 lists several indicators that are commonly used along with their equivalence points and the color change that accompanies the reaction. Review the information, and answer the questions below.

   a. Suppose a titration was performed in which a base of pH 6 was being titrated. The equivalence point of the titration was at pH near 8. What indicators should be added to the base solution before the titration is carried out?

   b. What is the color of the basic solution before any acid is added?

   c. What is the color of the solution after the equivalence point is reached at a pH of 8?

Experiment 1

Lab 8
Getting Aquainted with Indicators

Experiment Inventory

Materials

20 Drops of Bromothymol Blue

20 Drops of Methyl Orange

100 mL Sudsy Ammonia

100 mL 0.5 M Citric Acid, $C_6H_8O_7$

3 mL Turmeric Indicator

Graph Paper

Permanent Marker

30 mL Distilled water

Clear Tape

Labware

(1) 30 mL Syringe

(4) 100 mL Beakers (Plastic)

(1) 100 mL Graduated Cylinder

(1) Syringe Stopcock

(3) Pipettes

(1) Funnel

(1) Glass Stir Rod

(1) pH Meter

Note: You must provide the materials listed in red.

EXPERIMENT 1: GETTING AQUAINTED WITH INDICATORS

In this experiment, you will determine the pH of a solution at the equivalence point by graphing a titration curve.

PROCEDURE

PART 1: INDICATOR SET-UP

- [x] 1. Set up three 100 mL beakers. Use the permanent marker to label the beakers 1 - 3.
- [ ] 2. Use the graduated cylinder to measure and pour 20 mL of sudsy ammonia into each of the beakers. When measuring it is best to slightly tilt the graduated cylinder for the majority of the time while pouring the sudsy ammonia to avoid bubble formation. Be sure to return the graduated cylinder to a vertical position when you near the end of the pour to obtain an accurate measurement.
- [ ] 3. Use a pipette to transfer 5 drops of bromothymol blue into Beaker 1. Use a glass stir rod to mix the indicator into the solution. This should create a blue solution.
- [ ] 4. Use a new pipette to transfer 0.5 mL of the turmeric indicator into Beaker 2. Rinse the glass stir rod with tap water and use it to mix the indicator into solution. This should create a red or amber colored solution.
- [ ] 5. Use a new pipette to transfer 5 drops of methyl orange into Beaker 3. Rinse the glass stir rod with tap water and use it to mix the indicator into solution. This should create an orange solution.
- [ ] 6. Record the color of each beaker in **Table 2.**
- [ ] 7. Fasten the stopcock onto the threaded tip of the syringe. Make sure the stopcock is closed by rotating the handle on the stopcock to a position that is perpendicular to the syringe. Remove the plunger from the syringe.

- [ ] 8. Use a piece of tape and secure the syringe to the wall (Figure 4).
- [ ] 9. Place the funnel on top of the syringe (Figure 4).
- [ ] 10. Fill the syringe with 20 mL of 0.5 M citric acid.
- [ ] 11. Place Beaker 1 under the syringe.
- [ ] 12. Calibrate your pH meter. Refer to the Good Lab Techniques Appendix for detailed instructions.
- [ ] 13. Remove the cap from the pH meter and turn the pH meter on by moving the switch at the top to the right. Rinse the pH meter probe with 10 mL distilled water.
- [ ] 14. Place the pH meter into Beaker 1 and record the initial pH of the solution in **Table 2**. The pH meter reading may take up to 30 seconds to stabilize.

Figure 4. General set-up for titration and taking the pH reading.

- [ ] 15. Slowly open the stopcock by rotating the handle to a position that is parallel to the syringe. Slowly add the citric acid from the syringe, while continually swirling the solution in Beaker 1. Pay special attention to the color and pH of the solution.
- [ ] 16. As you notice the color of the solution begin to change, close the stopcock and take another pH measurement. When you re-open the stopcock, make sure the citric acid flow is slow; drop-wise flow is ideal.
- [ ] 17. Continue this process until the color has completely changed and the pH has leveled out. Record the final pH in **Table 2**.
- [ ] 18. Record the final color of the solution in **Table 2**. Also record the pH range at which the color change of the solution occurred.
- [ ] 19. Repeat Steps 10 - 17 for Beaker 2 and Beaker 3.
- [ ] 20. Rinse out Beakers 1, 2, and 3 and the graduated cylinder with tap water. Keep the syringe with the citric acid in the titration apparatus.

PART 2: TITRATION OF SUDSY AMMONIA WITH 0.5 M CITRIC ACID

- [ ] 1. Use a clean graduated cylinder to measure and pour 20 mL of sudsy ammonia into a clean 100 mL beaker. When measuring the sudsy ammonia, it is best to slightly tilt the graduated cylinder for the majority of the time while pouring the sudsy ammonia to avoid bubble formation. Be sure to return the graduated cylinder to a vertical position when you near the end of the pour to obtain an accurate measurement.
- [ ] 2. Add enough 0.5 M citric acid to the syringe so that the meniscus meets the 30.0 mL line of the syringe. Remove the cap from the pH meter, and turn the pH meter on by moving the switch at the top to the right. Rinse the pH meter probe with 10 mL distilled water.
- [ ] 3. Place your pH meter into the beaker with sudsy ammonia. Swirl it around and record the initial pH in Table 3. This establishes the initial reading of the solution.
- [ ] 4. Begin to add 0.5 mL (approximately 8 drops) aliquots of the 0.5 M citric acid from the syringe. Swirl the beaker to mix the solution and record the pH of the solution for each aliquot added in Table 3.
- [ ] 5. Slow down the flow of citric acid to a dropwise technique when you notice the pH begin to drop significantly. This should take approximately 2 - 4 mL.
- [ ] 6. Record the final pH of the solution when the reaction is complete (i.e., when the pH has dropped significantly).

- [ ] 7. Using the recorded data, make a plot of the titration curve. Include your graph in Post-Lab Question 1.
- [ ] 8. Answer Post-Lab Questions 2 and 3. You will need this information to proceed with Part 3 of the procedure.
- [ ] 9. Clean the beaker and the graduated cylinder.

PART 3: TITRATION OF SUDSY AMMONIA WITH 0.5 M CITRIC ACID USING THE APPROPRIATE INDICATOR

- [ ] 1. Place 20 mL of sudsy ammonia into a clean 100 mL beaker. When measuring the sudsy ammonia, it is best to slightly tilt the graduated cylinder for the majority of the time while pouring the sudsy ammonia to avoid bubble formation. Be sure to return the graduated cylinder to a vertical position when you near the end of the pour to obtain an accurate measurement.
- [ ] 2. Top off the 30 mL syringe with citric acid so that the meniscus meets the 30.0 mL mark.
- [ ] 3. Place the correct amount of the appropriate indicator you chose (the indicator you chose in Post–Lab Question 3) into the sudsy ammonia. Swirl the solution and record the initial pH and color of the solution in Table 3.
- [ ] 4. Slowly add the citric acid from the syringe in 0.5 mL (approximately 8 drops) aliquots, continually swirling the solution in the beaker. Pay special attention to the **color** and **pH** of the solution. Record the pH and color of the solution in Table 3.
- [ ] 5. As you notice the color of the solution begin to change, slightly close the stopcock to slow the citric acid flow. A drop-wise flow is ideal.
- [ ] 6. Record the final pH and color of the solution when the reaction is complete (i.e., when the color change is complete).

Lab 8
Experiment 1 Data Sheet

Table 2: Part 1 - Indicator, pH Range, and Color Change

Beaker	Indicator	Amount Needed for Reaction to Complete	pH Range	pH at Color Change	Color (Initial : Final)
1					
2					
3					

Table 3: Titration Data

Part 2: mL Citric Acid Added	Part 2: pH	Part 3: mL Citric Acid Added	Part 3: pH	Part 3: Color
0.0		0.0		
Total mL Added:		Total mL Added:		

Lab 8: Post-Lab Questions

1. Use the graph paper supplied to you to create a graph of the titration curve of 20 mL sudsy ammonia with 0.5 M citric acid. Be sure to title your graph and label your axes.

2. What is the pH of the solution at the equivalence point shown by the titration curve?

3. Which indicator shows a color change at about the same pH as the equivalence point?

4. Fill in the following data:

   a. Initial color of the solution:

   b. Amount of citric acid needed to reach the end point:

   c. Color of solution after endpoint was reached:

5. Is the indicator you chose a good indicator for this titration? Explain your answer.

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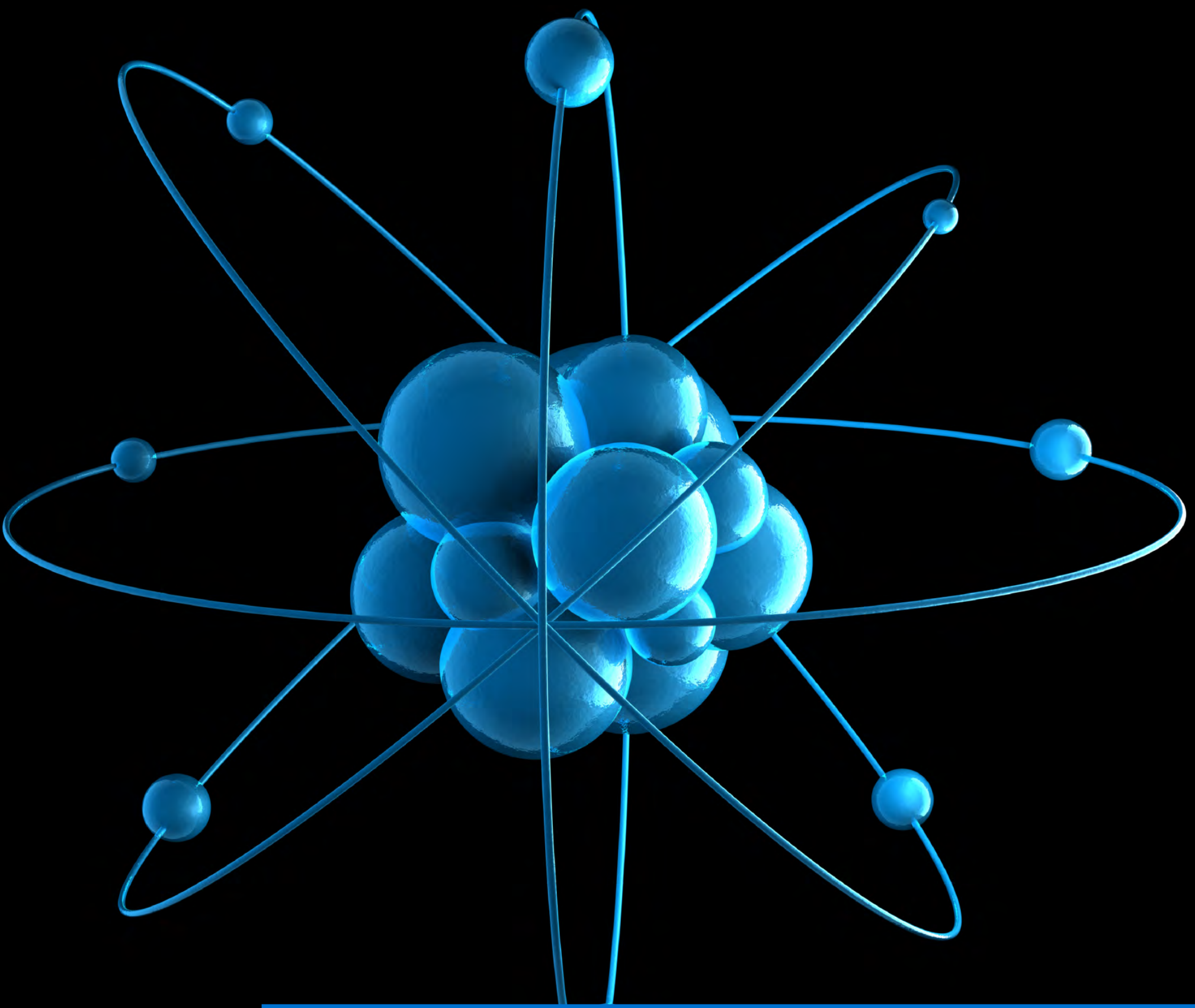

Lab 9

Nuclear Chemistry

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LAB 9

Nuclear Chemistry

Learning Objectives

- Apply the principles of radioactivity and nuclear chemistry through experimentation
- Create a graphical representation of the experimentally-derived radioactive decay
- Calculate the half-life for a given isotope

INTRODUCTION

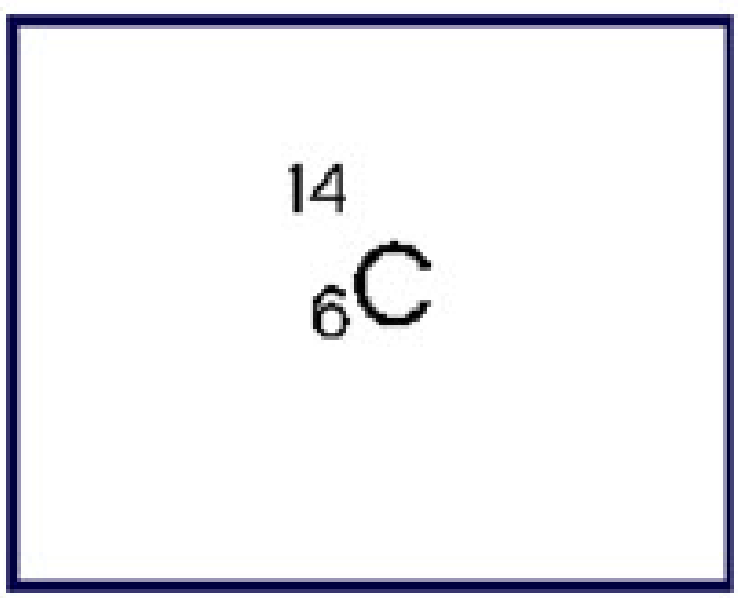
All matter consists of atoms. Most of an atom is actually empty space defined by an electron cloud and a much smaller nucleus of protons and neutrons. Most of the mass of an atom comes from the nucleus, while electrons are less than a thousandth of the mass of a proton or neutron.

Figure 1. If a nucleus was the size of a grain of sugar, the electron cloud would span 10 m from the grain in all directions!

STRONG FORCE

Protons and neutrons are attracted to each other by the strong force. The **strong force** is one of the four basic forces in nature, and is more than 100 times stronger than the electric force. However, it is only active in very short ranges, such as in the nucleus of an atom.

There are two main forces that contribute to the stability of the nucleus: the attractive strong force, which keeps the nucleus together, and the electrostatic repulsion between positively charged protons. If a nucleus has too much electrostatic repulsion and not enough attraction, it is unstable. Another term for an unstable nuclei is radioactive. An unstable nucleus will change to a more stable state by releasing radiation. Both protons and neutrons exhibit the attractive strong force for themselves and each other. Neutrons are neutral in charge so they do not contribute electrostatic repulsion. Neutrons, in a way, act like glue to keep the nucleus together. As nuclei get larger or have more protons, more and more neutrons are needed to keep the nucleus stable. For instance, carbon-12 has 6 protons and 6 neutrons, and lead-206 has 82 protons and 124 neutrons. Nuclear decay occurs in all nuclei with more than 83 protons. These atoms are both unstable and radioactive.



ISOTOPES

The number of protons in an element is constant and represented by the atomic number (See Figure 2). In contrast, the number of neutrons present can vary. Atoms with the same number of protons and electrons, but different numbers of neutrons, are called **isotopes**. Isotopes have nearly the same chemical properties, but the stability of the nuclei may differ. Nuclei that have too many or too few neutrons relative to the number of protons are considered unstable. The mass of an electron can be considered negligible compared to the mass of protons and neutrons. Therefore, the mass of an atom can be considered equivalent to the combined mass of protons and neutrons in the atom. The combined mass gives rise to the mass number.

Figure 2: The nucleus symbol includes the mass number (above the C) as well as the atomic number (below the C). How many neutrons does carbon-14 have?

Nuclear Chemistry

RADIATION

Unstable nuclei are constantly changing as a result of the energy imbalance within the nucleus. As unstable nuclei decay, they emit particles and electromagnetic energy, called radiation. **Radiation** is energy transmitted through space in the form of electromagnetic waves or energetic particles. As radioactive isotopes decay, they emit radiation only once. However, it may take several steps for an unstable atom to become stable, and radiation will be given off at each step. For this reason, radioactive sources become weaker with time. As more and more unstable atoms of a material become stable through successive radioactive decay, less radiation is produced by the material, and eventually the material will cease being radioactive and unstable.

ALPHA, BETA, AND GAMMA DECAY

Radiation is natural and categorized into three types, based on the decay product released: alpha, beta, and gamma. When alpha decay occurs, an alpha particle made of two protons and two neutrons is emitted from the decaying nucleus. The alpha particle has a charge of $^{+}2$ and atomic mass of 4. Therefore, when an atom loses an alpha particle, it undergoes a transmutation, and becomes another element. Alpha particles are the largest radia-tion particle and have the greatest electric charge, which makes them lose energy quickly when they collide with other matter. As a result, the alpha particles are the lowest penetrating form of radiation, stoppable by a single sheet of paper.

A second type of radiation occurs when a neutron in an unsta-ble nucleus transforms through a series of steps into an electron. This is called beta radiation. The electron that is lost is referred to as a beta particle. This particle is faster and more penetrating than an alpha particle, but can be stopped by a piece of aluminum foil. As with alpha radiation, the atom undergoes a transmutation when beta decay occurs, becoming an element with one more proton and an atomic number one greater than before.

The most penetrating form of radiation is gamma radiation. Gamma rays have no mass or charge, travel at the speed of light, and require thick, dense materials (such as lead or concrete) to stop their penetration. Gamma rays are emitted from the nucleus when alpha or beta decay occurs.

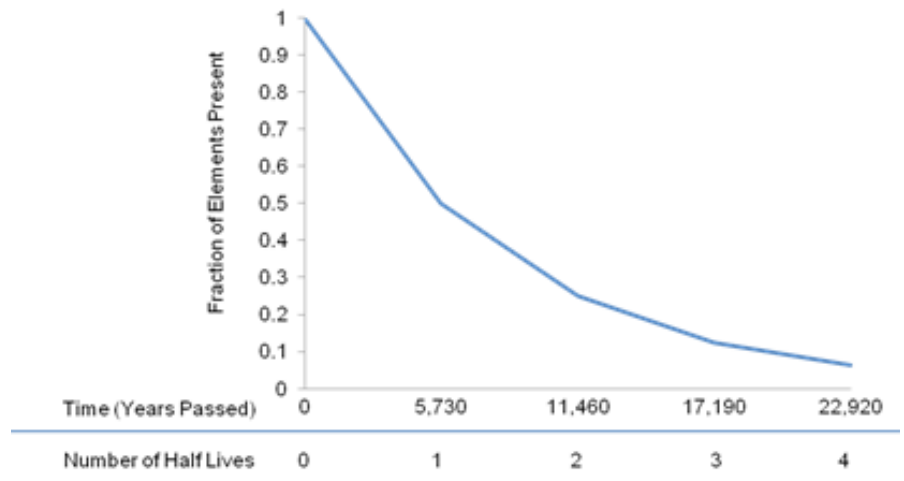


Figure 3: Carbon-14 has a half-life of 5,730 years. After 11,460 years (5,730 x 2) pass by, you might think that there are zero elements remaining. However, there are half as many as were present after 5,730 years passed.

Table 1. Half-Life of Radioisotopes

Radioisotope	Half-life
Polonium-215	0.0018 seconds
Bismuth-212	60.5 seconds
Sodium-24	15 hours
Iodine-131	8.07 days
Cobalt-60	5.26 years
Radium-226	1,600 years
Carbon-14	5,730 years
Uranium-238	4.5 billion years

HALF-LIFE

The behavior and effects of radioactive isotopes (radioisotopes) are influenced by the half-life of the isotope. The **half-life** of a radioactive isotope is the amount of time required for half the nuclei in the sample to decay (Figure 3). It also provides information about the frequency of radioactive emissions. Note that it does not represent a fixed number of atoms that disintegrate, but a fraction. A radioisotope with a long half-life will only infrequently emit radiation, while a short-lived radioactive isotope will emit radiation repeatedly over a short period of time. Half-life varies widely among the radioisotopes, from a fraction of a second to billions of years, as shown in Table 1.

Since the number of atoms present decreases by one half with the passing of each half-life, the fraction of atoms remaining can be calculated as:

$$undecayed\ atoms = \left(\frac{1}{2}\right)^{n}$$

where n is the number of half-lives that have passed. After one half-life, half of the atoms remain unstable (and undecayed), and the other half achieve stability. After two half-lives, 1/4 ($(½)^2$) of the atoms in the sample are undecayed. After three half-lives, 1/8 ($(½)^3$) of the atoms remain undecayed, and so on. This expression demonstrates how sequential decay events result in a reduction in the amount of unstable radioisotopes present. The decay pattern follows the characteristic curve demonstrated in Figure 3, which shows the decay rate of carbon-14.

Lab 9: Pre-Lab Questions

1. Define radioactive decay.

2. What is predictable about radioactive decay? What is unpredictable?

3. Describe how half-life is used to determine the geologic age of a rock.

Experiment 1

Lab 9
Estimating Half-Life

Experiment Inventory

Materials

(1) Bag of Skittles (approx. 60 candies)

Note: Do not eat candies.

(1) Styrofoam Cup

Camera / Smart Phone

Computer Access

Counter Top / Flat Surface

Note: You must provide the materials listed in red.

Labware

None

EXPERIMENT 1: ESTIMATING HALF-LIFE

In this experiment, you will be estimating the half-life of a radioactive isotope. While it would be nice to do an actual decay experiment, the time, money, and equipment required is unrealistic. Instead, Skittles® candies will be used to demonstrate the concept of half-life.

PROCEDURE

1. On a flat surface, count the number of Skittles® that were in your bag. Record this number in Trial 0 of **Table 2** as the number of Skittles® "S" Up (Parent Atoms) and record 0 as all other values in Trial 0
2. Place all the Skittles® in your cup. Place your hand over the opening of the cup and shake it several times.
3. Pour the Skittles® out on a flat surface, such as a table or countertop.
4. Count the number of Skittles® with the "S" facing up. Record this number in **Table 2**. These Skittles® represent the parent atoms. Place the parent atoms back into the cup.

   Note: The "S" on some of your Skittles® may be faded. Take care in counting which have an "S" and which don't.

5. Count the number of Skittles® with the "S" facing down. Record this number in **Table 2**. The Skittles® represent the daughter atoms. Set them to the side.
6. You should now have "parent atoms" in your cup, and a pile of "daughter atoms" on the side. Shake the cup several times.
7. Repeat Steps 2 - 5 until all your "parent atoms" have decayed. Record your data in **Table 2**.
8. Use your camera (or smart phone) to take a picture of the results. Be sure to correctly label the picture and send it to your instructor along with the answers to the lab questions.

Lab 9
Experiment 1 Data Sheet

Table 2: Radioactive Decay Data

Trial	Skittles® "S" Up (Parent Atoms)	Skittles® "S" Down (Daughter Atoms) for each Trial	Skittles® "S" Down (Daughter Atoms) Cumulative Total
0			
1			
2			
3			
4			
5			
6			
7			
8			
9			
10			

Lab 9: Post-Lab Questions

1. Create a graph using your data from Table 2 and a computer program such as Microsoft Excel®. If you do not have a graphing program installed on your computer, you can access one on the internet via the following links: http://nces.ed.gov/nceskids/createagraph/ or http://www.onlinecharttool.com. On the x-axis plot “Trial Number.” On the y-axis plot “Parent Atoms” and “Total Daughter Atoms.”

2. Suppose the isotope your Skittles® represented was uranium-238 and the trials represent the number of half-lives. How old was the sample at the end of your tests? Use Table 1 in the Introduction to help you answer this question. Include your calculations.

3. If 1/8 of a radioactive element remains after 600 years, what is that element’s half-life?

4. Identify and describe similarities and differences between this experiment and radioactive decay in nature.

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Lab 10

Organic Compounds

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LAB 10

Organic Compounds

Learning Objectives

- Understand the structure of organic compounds
- Visualize the structures of hydrocarbons by constructing molecular models
- Test for the presence of lipids in four unknown solutions using newsprint

INTRODUCTION

The term “organic” is frequently used to refer to foods or clothing. This application often connotes that the origin of those materials is of better quality and is less modified from its original, or natural, form. While this is a subjective interpretation of the word organic, it has been extrapolated from the original definition of organic, which refers to materials that are derived from living organisms.” In chemistry, we use the word organic to refer to carbon-based compounds. Life is based on organic compounds and, therefore, knowledge of these types of chemicals is beneficial.

There are many classes of organic compounds, based on the functional groups they contain. To give an introduction to organic chemistry, this exercise will focus on hydrocarbons, nucleic acids, carbohydrates, and lipids.

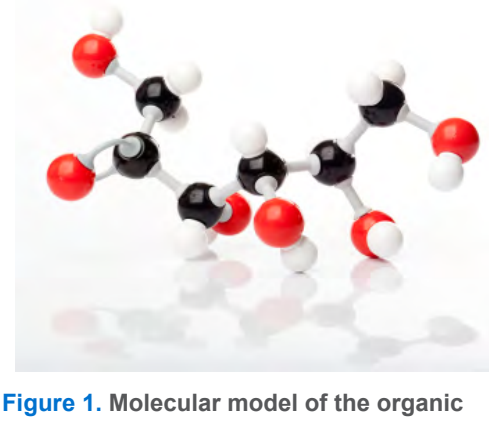
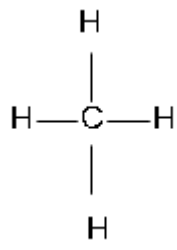


Figure 1. Molecular model of the organic compound, fructose.

HYDROCARBONS

Carbon is one of the most common elements found in a chemical compound. This is due to two factors. First, carbon contains four valence electrons in its outer shell. This allows it to form four strong covalent bonds with other elements. Second, it can form single, double, and triple covalent bonds with itself (forming long carbon chains) or other elements. This ability is a very rare amongst elements, thus making carbon a more commonly bonded element.

Hydrocarbons are compounds made up of hydrogen and carbon atoms. We know that carbon can form four covalent bonds and hydrogen can form one covalent bond. When a carbon element is bonded to four other elements, it is considered “saturated”. The simplest saturated hydrocarbon is methane:



Hydrocarbons contain single, double, or triple bonds, or a mixture of two or more of these. They have significant industrial value, as they provide much of our fuel resources used in everything, from a gas grills to airplanes.

The naming convention for hydrocarbons uses the root word to denote the number of carbons in the compound and attaches a suffix to indicate the type of bond used. For example, the suffix “–ane” indicates single bonds, “-ene” indicates one or more double bonds, and “–yne” indicates one or more triple bonds. In a compound containing more than two carbons, the double or triple bonds may be located on any of the carbons, but to keep this introduction to nomenclature simple, we will assume the double or triple bond is located between the first and second carbons of the compound. The prefixes and corresponding number of carbons are given in Table 1 on the following page.

Organic Compounds

Table 1. Common Carbon Molecules and Nomenclature

Number of Carbon Atoms	Root Word	Containing No Double or Triple Bonds (Saturated Hydrocarbons)	Containing 1 Double Bond, Located at the First Carbon	Containing 1 Triple Bond, Located at the First Carbon
1	Meth-	Methane, CH_4	n/a	n/a
2	Eth-	Ethane, C_2H_6	Ethene, C_2H_4	Ethyne, C_2H_2
3	Prop-	Propane, C_3H_8	Propene, C_3H_6	Propyne, C_3H_4
4	But-	Butane, C_4H_{10}	Butene, C_4H_8	Butyne, C_4H_6
5	Pent-	Pentane, C_5H_{12}	Pentene, C_5H_{10}	Pentyne, C_5H_8
6	Hex-	Hexane, C_6H_{14}	Hexene, C_6H_{12}	Hexyne, C_6H_{10}
7	Hept-	Heptane, C_7H_{16}	Heptene, C_7H_{14}	Heptyne, C_7H_{12}
8	Oct-	Octane, C_8H_{18}	Octene, C_8H_{16}	Octyne, C_8H_{14}
9	Non-	Nonane, C_9H_{20}	Nonene, C_9H_{18}	Nonyne, C_9H_{16}
10	Dec-	Decane, $C_{10}H_{22}$	Decene, $C_{10}H_{20}$	Decyne, $C_{10}H_{18}$

Table 1 also helps us understand the mathematical equation that can be used to calculate a hydrocarbon chemical formula. For example, $C_nH_{(2n+2)}$ describes the chemical formula for alkanes (structures containing only single bonds). C_nH_{2n} can be used for alkenes (structures containing a double bond) and $C_nH_{(2n-2)}$ can be used for alkynes (structures containing a triple bond). This holds true for all hydrocarbons. Note that the chemical formula (e.g., C_3H_8 for propane) lists the atoms as the ratio of C:H, while the structural formula (e.g., $CH_3CH_2CH_3$ for propane) denotes the relationship of each atom as it is arranged in the compound. The first experiment will explore these compounds and help you become familiar with the arrangement of the atoms.

Four types of molecules represent the majority of organic compounds: nucleic acids, proteins, carbohydrates, and lipids. In addition to the carbon framework, these biomolecules also often include hydrogen, oxygen, nitrogen, sulfur, and phosphorus. These molecules attach to the carbon-hydrogen backbone of the molecule as functional groups. Functional groups account for the differences in molecular properties between molecules and dictate their utility in cells.

NUCLEIC ACIDS

Nucleic acids are found within every living organism, and together with proteins, are the most important macromolecules in the cell. They carry the information necessary to make proteins, the building blocks of cells. Deoxyribonucleic acid (DNA) and ribonucleic acid (RNA) are the two types of nucleic acids. They consist of a long chain of five-carbon sugars with an organic, nitrogen-rich base attached to it. These are assembled in a specific sequence in order to store the hereditary information.

PROTIENS

Proteins are the most abundant macromolecule in a living system. The many proteins found within the body are all made from different combinations of amino acids (the monomers of proteins, Figure 2). From just 21 amino acids, thousands of proteins are made. Because each one has a unique chemistry and structure, there is ~~a~~ great diversity in the roles proteins play within cells. Muscles, bones, hormones, enzymes, cell-cell connections, and many more vital structures are all dependent on proteins. The shape of a protein determines its function. The primary structure of a protein is the amino acid sequence. Secondary structure results from hydrogen bonds forming with other amino acids within the same protein, resulting in two possible patterns: beta-pleated sheets and alpha-helices. The tertiary structure is the 3D shape of a single protein subunit. This determines how the protein will fold or coil with itself. When proteins are denatured, they can lose their 3D shape or even unfold completely. This can be achieved by changes in pH, temperature, and salinity, which renders the protein inactive. Finally, quaternary structure is the three dimensional shape of a protein with multiple subunits.

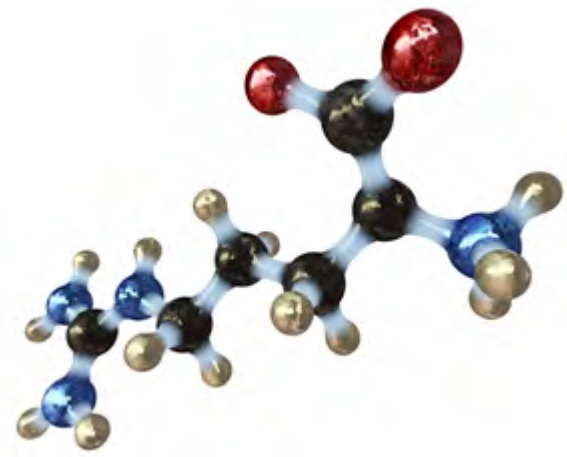


Figure 2: Amino acid arginine.

Organic Compounds

Table 2. The Building Blocks of an Organism

Monomer	Simple Polymer	Macromolecule
Amino acid	Peptide	Polypeptide/protein
Nucleotide	Oligonucleotide	Nucleic acid
Monosaccharide	Oligosaccharide	Polysaccharide/carbohydrate
Varies	Varies	Lipids (including fats and phospholipids)

CARBOHYDRATES

Commonly referred to as sugar, the name "**carbohydrate**" comes from its chemical formula, $C_n(H_2O)_n$, which gave early scientists the idea that these molecules were carbon hydrates, hence the name. This class of organic compounds serves as one of the body's primary sources of metabolic energy. This is partially because numerous C-H bonds form within carbohydrates and release energy when oxidation occurs, therefore making this molecule perfect for energy storage. Carbohydrates are also used as a structural material (for example, cellulose, Figure 3), as one of the three essential constituents of DNA and RNA, and as a vital component of ATP, the energy molecule that fuels cells.

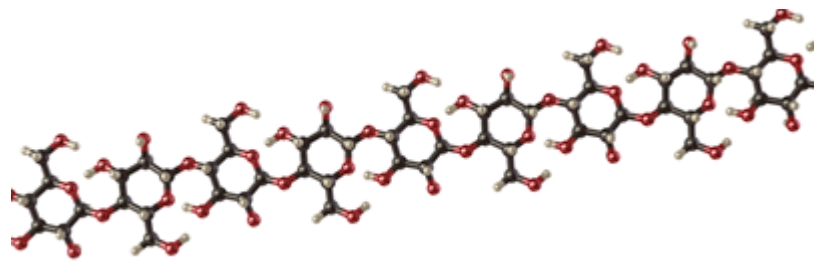


Figure 3: Cellulose model.

Simple sugars are called monosaccharides, and are often used as the building blocks for larger molecules. Disaccharides form when two monosaccharides link together, and can serve as transport molecules or provide nutrition. Polysaccharides are long chains of monosaccharides linked together with glycosidic bonds. Starch is an example of an energy storage polysaccharide, while cellulose is one that provides structure to cells. Starch can be separated into two fractions - amylose and amylopectin. Natural starches are mixtures of amylose (10-20%) and amylopectin (80-90%).

Three of the most common carbohydrates are actually a trio of related monosaccharides with the same chemical formula ($C_6H_{12}O_6$), but with a different atomic arrangement. These are called structural isomers, and are known as glucose (blood sugar), fructose (fruit sugar), and galactose. Galactose is part of a disaccharide called lactose. The inability to break down lactose (milk sugar) into glucose and galactose is the hallmark of lactose intolerance (Figure 4).

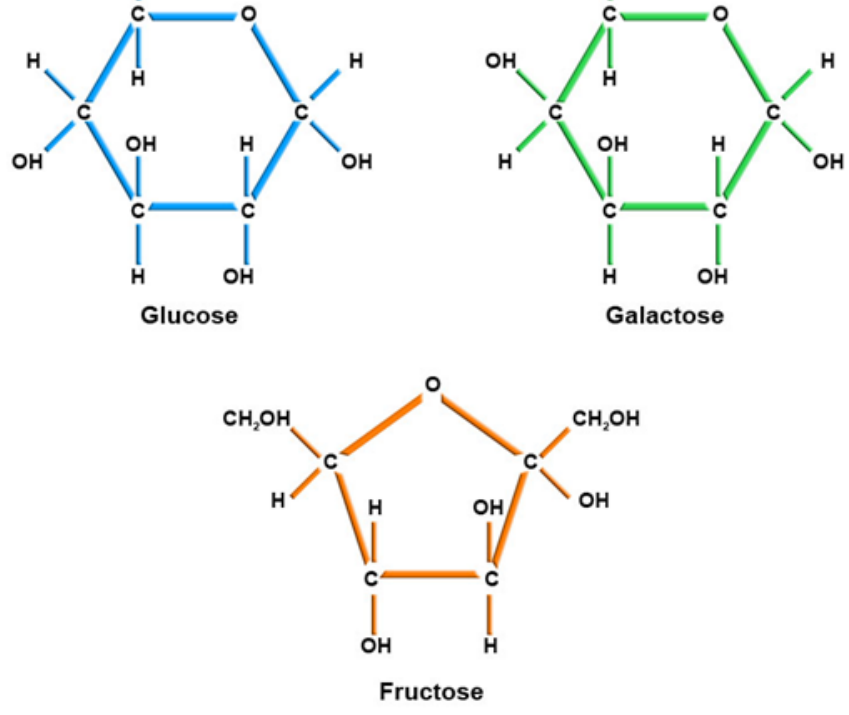


Figure 4: Glucose, galactose, and fructose are all examples of monosaccharides.

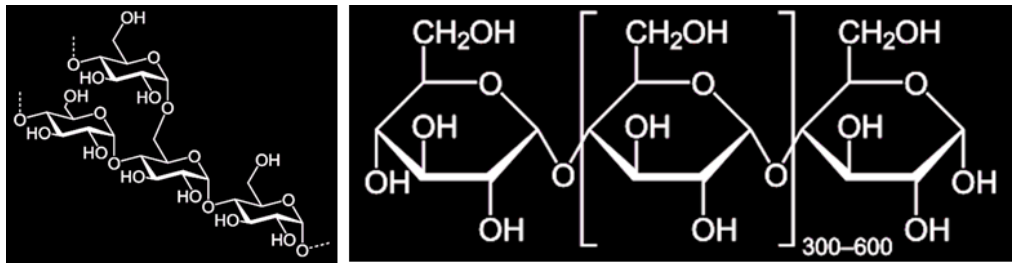


Figure 5: Amylopectin (left) is a polysaccharide found in plants. Amylose (right) is a large polysaccharade (as indicated by the brackets labeled "300-600" which indicates that the central unit can be repeated many times.

LIPIDS

Lipids are organic molecules made of long chains of hydrocarbons. There is no general chemical formula for all lipids, but the majority of lipids are triglycerides. Triglycerides are fats and oils which have the general structure shown in Figure 7, in which each R group is a fatty acid attached to a glycerol.

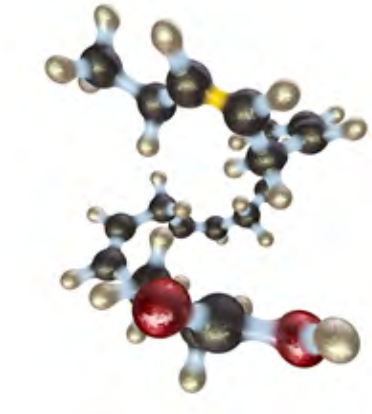


Figure 6: Omega-3 Fatty Acid

Organic Compounds

Many lipids are constructed with two main molecules – fatty acids and glycerol. Fatty acids are long chain hydrocarbons with a COOH (carboxylic acid) at one end. Glycerol is a simple alcohol containing three hydroxyl groups (-OH). Complex lipid molecules incorporate unique functional groups, in addition to fatty acids and glycerol. For example, the lipids that form biological membranes contain fatty acids, glycerol, and a phosphate group, giving them another name – phospholipids.

Numerous C-H bonds serve as a form of energy storage within the molecule. A saturated fatty acid is formed when single bonds comprise all bonds between carbon atoms in a fatty acid. Single bonds allow for the greatest number of hydrogen bonds to be formed. As a result, the fatty acid is considered to be saturated with hydrogen atoms. Unsaturated fats contain one or more double bonds between carbon atoms. A fatty acid is called a monounsaturated fatty acid if one double bond is present and polyunsaturated if there is more than one double bond.

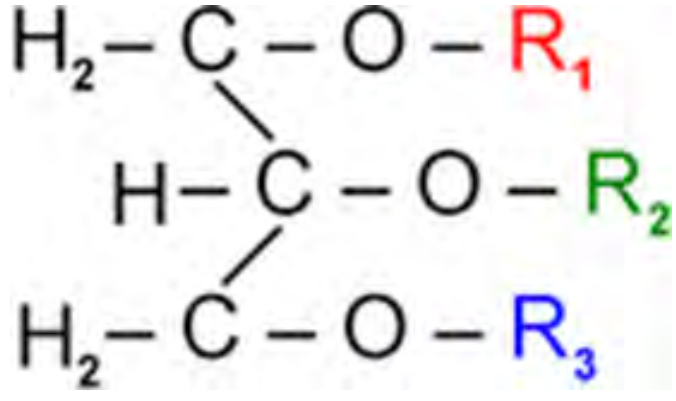


Figure 7: General chemical structure of tri-glycerides, a major class of lipids.

Lab 10: Pre-Lab Questions

1. Compare and contrast monosaccharides and a polysaccharides.

2. What is the difference between a carbohydrate and a lipid?

3. How are unsaturated fats different from saturated fats?

Experiment 1

Lab 10
Modeling Basic Organic Compounds

Experiment Inventory

Materials

60 Rasins - Do Not Eat

30 Marshmallows- Do Not Eat

80 Toothpicks

Camera / Smart Phone

Labware

None

Note: You must provide the materials listed in red.

EXPERIMENT 1: MODELING BASIC ORGANIC COMPOUNDS

In this experiment, molecular models of various hydrocarbons will be made to visualize the structures of the organic compounds listed in Table 1.

PROCEDURE

1. For each of the structures listed in **Table 1**, draw the structure named, count the number of carbon atoms and hydrogen atoms, count the number of C-H bonds, count the number of C-C bonds, and count the number of C=C bonds. List these numbers in **Table 3**.
2. Use your camera (or smart phone) to take a picture of the results. Be sure to correctly label the picture and send it to your instructor along with the answers to the lab questions.
3. Use the raisins (hydrogen atoms), marshmallows (carbon atoms), and toothpicks (each toothpick represents a single bond) to construct a model of each of the structures listed in **Table 1**.
4. Use your camera (or smart phone) to take a picture of the results. Be sure to correctly label the picture and send it to your instructor along with the answers to the lab questions.

Data Sheet

Lab 10
Experiment 1 Data Sheet

Table 3: Carbon Molecule Data

Compound	# H Atoms	# C Atoms	# C-H Single Bonds	# C-C Single Bonds	# C=C Double Bonds	# C≡C Triple Bonds	Chemical Formula	Structural Formula
Propane	**8**	**3**	**8**	**2**	**0**	**0**	C_3H_8	$CH_3CH_2CH_3$
Butane								
Heptane								
Ethene								
Butyne								
Hexene								
Methane								

Lab 10: Post-Lab Questions

1. Which organic compounds contained double bonds? Triple bonds?

2. Draw an organic molecule that was not included in Table 3 that contains a double bond. Name the compound.

Experiment 2

Lab 10
Testing for Lipids and Carbohydrates

Experiment Inventory

Materials

(1) Scoop of Unknown A

(10) Drops of Unknown B

(10) Drops of Unknown C

(1) Scoop of Unknown D

(3-4) Drops Iodine-Potassium-Iodide (IKI) Solution

Newspaper Print

Distilled Water

Permanent Marker

Camera / Smart phone

Labware

(1) 10 mL Graduated cylinder

(5) Glass Test Tubes

(10) Pipettes

(1) Test Tube Rack

(1) Spatula

Note: You must provide the materials listed in red.

EXPERIMENT 2: TESTING FOR LIPIDS AND CARBOHYDRATES

In the following experiment, you will test for the presence of different organic compounds. In part one of this experiment, you will test for the presence of lipids in four unknown solutions using newsprint.

PROCEDURE

PART 1: TESTING FOR LIPIDS

- [ ] 1. Put on your safety glasses and gloves (provided in your safety box).
- [ ] 2. Place five test tubes in the test tube rack.
- [ ] 3. Use a permanent marker to label each test tube as Unknown A, B, C, and D. Label the fifth test tube as “water.”
- [ ] 4. Use the spatula to place one scoop of the powdered Unknowns (A and D) into their respectively labeled test tube. Using a 10 mL graduated cylinder, measure and pour 5 mL of distilled water into test tubes A and D.
- [ ] 5. Place 10 drops of Unknowns B and C into their respectively labeled tubes.
- [ ] 6. Swirl tubes A, B, C and D to ensure that all diluted solutions are mixed thoroughly.
- [ ] 7. Use the 10 mL graduated cylinder to measure and pour 5 mL of water into the “water” test tube.
- [ ] 8. With a permanent marker label five small sections on your newspaper print with the letters A - D or the word “water” with a permanent marker.
- [ ] 9. Use separate clean pipettes to pipette one drop of Unknown A, Unknown B, Unknown C, and Unknown D onto each of the corresponding newsprint sections.

- [ ] 10. Pipette one drop of water (from the 10 mL graduated cylinder) onto the final section labeled “water.”
- [ ] 11. Observe and compare the translucence of each Unknown. Compounds containing lipids will have a higher translucence than water (i.e., the words under the drop on the newsprint will be harder to read with a lipid-containing sample).
- [ ] 12. Use your camera (or smart phone) to take a picture of the results. Be sure to correctly label the picture and send it to your instructor along with the answers to the lab questions.
- [ ] 13. Record the information, in order of translucence (0, +, ++, +++, or ++++) in Table 4. The Unknown which contains the most lipids should be recorded as ++++.

PART 2: TESTING FOR CARBOHYDRATES

- [ ] 1. Use a pipette to add three drops of IKI solution into each test tube, swirl the test tube to mix and observe any color change. If a carbohydrate (starch) is present in the unknown, the solution will turn a purple-black color upon addition of the IKI solution (see IKI Color Key).
- [ ] 2. Record your observations in Table 5. A positive result (visible color change) indicates that starch was present. A negative result (no color change) indicates that there was no starch.
- [ ] 3. Use your camera (or smart phone) to take a picture of the results. Be sure to correctly label the picture and send it to your instructor along with the answers to lab questions.

IKI Color Key

Yellow/No Color Change = No starch

Purple/Black = Starch

Lab 10
Experiment 2 Data Sheet

Table 4: Lipid Test Results

Unknown A	Unknown B	Unknown C	Unknown D	Water

Table 5: Starch Presence Results

Unknown A	Unknown B	Unknown C	Unknown D	Water

Lab 10: Post-Lab Questions

1. Which unknown(s) contained a lipid?

2. Which unknown(s) contained a starch?

3. Draw the chemical structure of a carbohydrate and a lipid of your choice. Name the structures you draw.

4. Why was water used as a negative control in the experiments in this lab?

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Appendix

Good Lab Techniques

Good Lab Techniques

GOOD LABORATORY TECHNIQUES

Science labs, whether at universities or in your home, are places of adventure and discovery. One of the first things scientists learn is how exciting experiments can be. However, they must also realize science can be dangerous without some instruction on good laboratory practices.

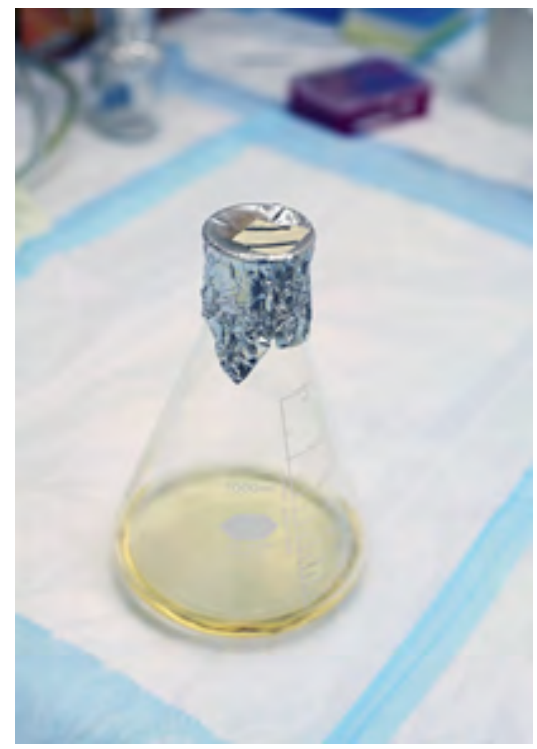

Figure 1. A bench coat or underpad helps prevent any spilled liquids from contaminating your work surface.

- Read the protocol thoroughly before starting any new experiment. You should be familiar with the action required every step of the way.
- Keep all work spaces free from clutter and dirty dishes.
- Read the labels on all chemicals, and note the chemical safety rating on each container. Read all Material Safety Data Sheets (provided on the eScience Labs website).
- Thoroughly rinse labware (test tubes, beakers, etc.) between experiments. To do so, wash with a soap and hot water solution using a bottle brush to scrub. Rinse completely at least four times. Let air dry
- Use a new pipette for each chemical dispensed.
- Wipe up any chemical spills immediately. Check MSDSs for special handling instructions (provided on the eScience Labs website).
- Use test tube caps or stoppers to cover test tubes when shaking or mixing – not your finger!
- When preparing a solution, refer to a protocol for any specific instructions on preparation. Weigh out the desired amount of chemicals, and transfer to a beaker or graduated cylinder. Add LESS than the required amount of water. Swirl or stir to dissolve the chemical (you can also pour the solution back • and forth between two test tubes), and once dissolved, transfer to a graduated cylinder and add the required amount of liquid to achieve the final volume.

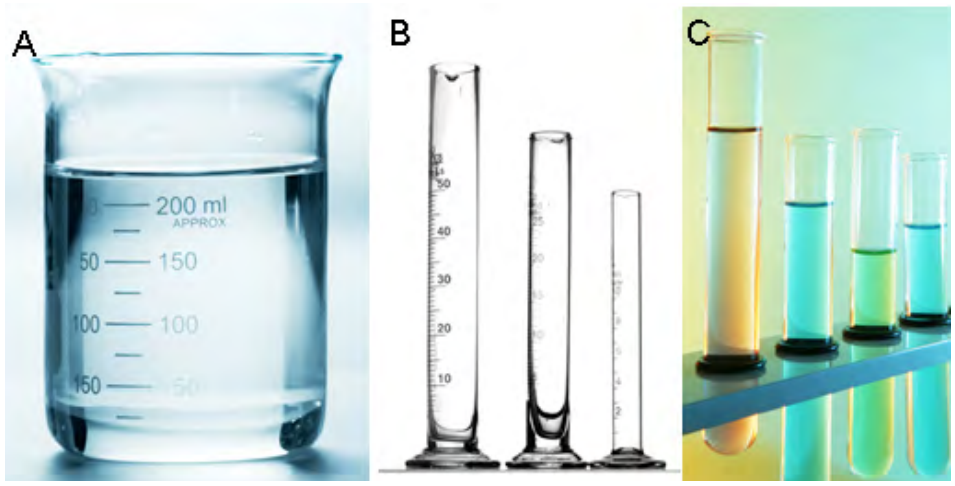

Figure 2. Special measuring tools can make experimentation easier and more accurate in the lab. A shows a beaker, B graduated cylinders, and C test tubes in a test tube rack.

Figure 3. Disposable pipettes aid in accurate measuring of small volumes of liquids. It is important to use a new pipette for each chemical to avoid contamination.

- A molar solution is one in which one liter (1L) of solution contains the number of grams equal to its molecular weight.

  For example:

$$1M = 110g\ CaCl \times \frac{110g\ CaCl}{1\ mol\ CaCl}$$

  (The formula weight of CaCl is 110 g/mol)

- A percent solution can be prepared by percentage of weight of chemical to 100ml of solvent (w/v) , or volume of chemical in 100ml of solvent (v/v).

  For example:

$$20\ g\ NaCl + 80\ mL\ H_2O = 20\%\ {}^{w}/_{v}\ NaCl\ solution$$

- Concentrated solutions, such as 10X, or ten times the normal strength, are diluted such that the final concentration of the solution is 1X.

  For example:

  To make a 100 mL solution of 1X TBE from a 10X solution:

$$10\ mL\ 10x\ TBE + 90\ mL\ H_2O = 100\ mL\ 1x\ TBE$$

Good Lab Techniques

Figure 4. When instructed to smell a chemical always use the wafting technique.

- Always read the MSDS before disposing of a chemical to insure it does not require extra measures. (provided on the eScience Labs website)
- Avoid prolonged exposure of chemicals to direct sunlight and extreme temperatures. Immediately secure the lid of a chemical after use.
- Prepare a dilution using the following equation:

$$C1V1 = C2V2$$

- Where C1 is the concentration of the original solution, V1 is the volume of the original solution, and C2 and V2 are the corresponding concentration and volume of the final solution. Since you know C1, C2, and V2, you solve for V1 to figure out how much of the original solution is needed to make a certain volume of a diluted concentration.
- Use only the chemicals needed for the activity.
- Keep lids closed when a chemical is not being used.
- When diluting an acid, always slowly pour the acid into the water. Never pour water into an acid, as this could cause both splashing and/or an explosion.
- Never return excess chemical back to the original bottle. This can contaminate the chemical supply.
- Be careful not to interchange lids between different chemical bottles.
- When pouring a chemical, always hold the lid of the chemical bottle between your fingers. Never lay the lid down on a surface. This can contaminate the chemical supply.
- When using knives or blades, always cut away from yourself.
- Wash your hands after each experiment

PH METER CALIBRATION PROCEDURES

1. Turn on the pH meter.
2. Immerse the portion of the pH meter previously covered by the black cap in room temperature distilled water. Not using room temperature distilled water will result in a failed calibration.
3. The pH meter should read 7.0 when in the distilled water. If it does, stop here and resume your lab. If it does not, see Step 4.
4. Locate the small screw on the back of the pH meter next to the clip.
5. Keep the pH meter partially submerged in the distilled water and use the supplied screwdriver in the pH meter kit to adjust the screw from Step 4 until the display shows a pH of 7.0.
6. Your pH meter is now calibrated.

POTENTIAL PROBLEMS

The pH meter display is faded:

1. Pull out the battery case. The battery case is the black strip at the top of the meter with the on switch. Gently pull this black strip away from the meter. If the terminals that hold the batteries are covered in rust, remove the batteries, and clean the terminals with sandpaper, a dry brillo pad, or a dry toothbrush that you are willing to discard.
2. Once you have cleaned the terminals, put the batteries back in place and try your pH meter again.
3. If this does not fix the problem please contact eScience Labs customer support at 888-ESL-KITS

The pH meter has measurements over 14:

1. Your pH meter is improperly calibrated. Please see the pH meter calibration procedures above.

www.eScienceLabs.com | 888.ESL.KITS