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Chemistry 101 - Introduction to Chemistry I

Lanthanum 57 **La** 138.90547	Boron 5 **B** 10.811

Boron 5 **B** 10.811	Oxygen 8 **O** 15.9994	Oxygen 8 **O** 15.9994	Potassium 19 **K** 39.0983

Jay C. M^{c}Laughlin, Ph.D.

Colorado Northwestern Community College

Credit: `https://www.pxfuel.com/en/free-photo-qngaq\`

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credit: `https://commons.wikimedia.org/wiki/File:Marie-Curie.jpg`

> “Life is not easy for any of us, but what of that? We must have perseverance and above all confidence in ourselves. We must believe that we are gifted in something, and that this thing, at whatever cost, must be attained.”
>
> — Marie Curie

Disclaimer
This is the first draft of a laboratory manual designed to meet the needs of the CHE 101 - Introduction to Chemistry I in Colorado.

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List of Experiments

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Experiment 1
Lab Safety

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Understand basic rules of laboratory safety.
2. Learn the number 1 rule.
3. Pass safety quiz with 80% or higher

Discussion

Science is about discovery, and one way you will be making discoveries is through science investigations. Some of these investigations make use of equipment and chemicals that must be used safely. Accidents in a chemical laboratory usually result from improper judgment on the part of the victim or one of his or her neighbors. Learn and observe the safety and laboratory rules listed below. The science department continues to do its part in assuring that you will have a safe experience, but you must do your part, too.

Please read the following safety contract, than sign it signifying that you understand it, so that we can be sure that everyone is committed to safe laboratory practice. Every student must also take an exam over these rules and score 80% in order to be allowed in the laboratory.

General Rules

1. Follow all written and verbal instructions carefully. If you do not understand a direction or part of the procedure, **if you don't know ask the teacher** before proceeding.

2. Act appropriately at all times in the laboratory. The same high level of care and attentiveness is required as might be found in a wood working shop or any place where there are potential hazards. The laboratory is not the place for horseplay, jokes, or pranks. In other words, think before acting. Perform procedures with deliberation; do not rush.

> If you don't know, ask.
> you will be a fool for the
> moment, but a wise man
> for the rest of your life.
> -Senica the Younger

3. When first entering a science room, do not touch any equipment, chemicals, or other materials in the laboratory area until you are instructed to do so.

4. **Eating, drinking, and smoking**: Smoke outside. Eating and Drinking are only allowed in the "classroom" side of the room, never in the "lab" side.

5. Be prepared for your investigation
    (a) Come to class on time and prepared. For reasons of safety, you may not be allowed to attend lab if you are late. This will result in a zero for that lab.

    (b) Read all procedures carefully **before** coming to class. Make a note of any questions you have and ask them in class before beginning the experiment.

    (c) Complete any pre-lab assignment before coming to class

6. **Never work alone:** Never work in the laboratory without the instructor present including setting up equipment. Simply wait until instructed to begin.

Figure 1.1: Never work along in lab. Your instructor should **always** be present. credit: https://inside.ewu.edu/ehs/laboratories-and-shops/laboratory-safety/

7. Notify the instructor **immediately** of any unsafe conditions you observe.

8. Neighbors: Be aware of your lab neighbors' activities; you may be a victim of their mistakes. If you observe improper techniques or unsafe practices, advise your neighbor and your instructor if necessary.

9. **Perform only authorized experiments:** Unless authorized to do so by the instructor, a student will be subject to immediate and permanent expulsion from the lab if found attempting to conduct unauthorized experiments or attempting variations of the experiment in the lab manual. Set up and use the prescribed equipment as directed in the laboratory instructions or by your instructor.

10. Do not remove any chemicals from the lab: This will result in permanent expulsion from the lab.

11. **Clean and Orderly work-space:** Keep your work-space orderly.
    (a) Place tall items, such as graduated cylinders, toward the back of the workbench so they will not be overturned by reaching over them.

    (b) Clean up all chemical spills, scraps of paper, and broken glassware immediately.

    (c) Keep drawers closed while working and the aisles free of any obstructions, including chairs.

    (d) Keep the work area clear of everything except for materials, data sheets, and instruction sheets.

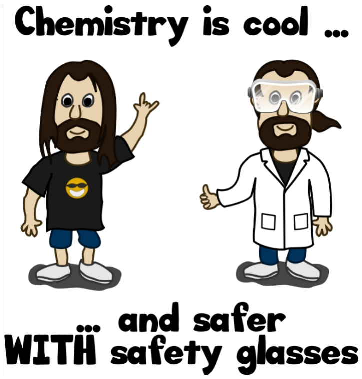
    (e) Never place coats, books, and other belongings on the laboratory bench where they will interfere with the experiment and are likely to be damaged.

    (f) Keep aisles clear. Do not crowd around hot plates, balances, or laboratory stations.

12. When using chemicals, keep hands away from face, eyes, mouth and body. Avoid rubbing your eyes unless you know your hands are clean. Wash your hands with soap and water after performing all investigations. Clean all work surfaces and equipment at the end of the experiment, and return all equipment to the proper storage area.
13. 1Do not leave an investigation unattended. For example, do not wander from your assigned area to talk with friends if you are in the middle of a procedure.
14. Students are not to enter science storage areas except under the direct instruction or supervision of the instructor.
15. If you have an allergy or sensitivity to chemicals, please inform your teacher.
16. **Before you leave:**
    (a) Clean your workspace – Return all glassware and equipment, turn off gas and water, wipe off your desktop.
    (b) Always wash your hands, before leaving the lab since toxic chemicals may be transferred to the mouth at a later time.
    (c) Notify your instructor.

Clothing

1. **Safety Goggles (REQUIRED):** Because the eyes may be permanently damaged by spilled chemicals and flying broken equipment, be sure to wear Z-87 safety goggles or safety glasses (State law) whenever anyone is working in the lab. Not wearing safety goggles will result in loss 10% on the lab grade for the day. Multiple infractions will result in a cumulative loss.
2. **Suitable clothing (RECOMMENDED):** Wear clothing that will protect you against spilled chemicals or flaming liquids. Hard-soled, close toed, covered footwear and pants/skirts that cover your legs.



credit: `https://openclipart.org/detail/251205/chemistry-is-cool-by-b-lachner`

Accidents and Injury

1. Notify your instructor at once in case of **ALL** (no matter how apparently minor) accidents or personal injuries to you or your neighbor. If your instructor is temporarily absent, notify the instructor in the adjoining lab. If you are asked to go to the hospital, you must go. Refusal to do so is

grounds for being dropped from the course.

2. **Chemicals:**
   (a) Chemical splashes in your eye(s): If you get anything in your eye, use the eye wash immediately, and then report it to your instructor. Use your hands to hold your eye open so that it can be rinsed thoroughly. Rinse for 15 minutes. Note: Eye washing with a contact lens in place will not clear a splashed chemical from the eye. The contact must · be removed for effective cleansing. It is advisable for those wearing contacts to switch to glasses for the lab period.

credit: `https://commons.wikimedia.org/wiki/File:Salle-blanche-rince-oeil-rot.jpg`

   (b) Chemical splashes on the skin, flush with water. Inform the instructor immediately.
   (c) Chemical spills on lab benches or floor. Inform the instructor immediately.
3. Report any broken glassware to the instructor.
4. **Fires:** If it is a small, contained fire, such as in a flask or beaker, cover the container with a piece of ceramic, cutting off the supply of oxygen to the fire and thus putting it out. If it is open fire, such as a large chemical spill on a lab bench, the correct extinguisher should be used. Inform your instructor immediately.
5. **Evacuation:** If we are forced to evacuate the building due to fire, chemical spills, or other reason, meet on the step in front of the McLaughlin Building. Do not leave campus, attendance will be taken to be sure that everyone is safe.
6. Be familiar with the location of exits and all safety equipment (Fire extinguisher, fire blanket, eyewash, medical kits, phones, fire alarms, safety shower).

Handling Chemicals

1. **All** chemicals in the laboratory are to be considered dangerous. Do not touch, taste, or smell any chemicals unless specifically instructed to do so.
2. **Smelling Chemicals:** When smelling chemical, do not inhale a large amount. Waft a small amount towards you as you move the container slowing towards you.
3. **Tasting Chemicals:** Don't ever taste a chemical or place it in your mouth.
4. **Double check** all labels on chemicals before using them. If you are uncertain a chemical is the one required for the experiment, **ask your instructor**.
5. Never return unused chemicals to their original containers.

6. When transferring chemicals from one container to another, hold the containers away from your body.

7. Acids require special care. You will be shown the proper way to dilute acids. Dilute concentrated acids and bases by pouring the reagent into water (room temperature or lower) while stirring constantly. Never pour water into concentrated acids; the heat of solution will cause the water to boil and the acid to splatter.

8. Never dispense flammable liquids anywhere near an open flame or source of heat.

9. Mixing chemicals: Always add a reagent slowly–never "dump" in. Some reactions give off a lot of heat, and unless adding slowly, can become too vigorous and out of control. If you make a mistake and choose the wrong chemical, adding slowly decreases the possibility of causing a serious accident.

10. Dispose of all chemical waste properly. Your instructor will tell what materials can be poured down the drain and what materials must be placed in a waste container. In general, solid chemicals, metals, matches, filter paper, and all other insoluble materials are to be disposed of in the proper waste containers, not in the sink.

Chemical Waste Disposal Guideline

Innocuous aqueous waste
- **Acid** (pH<4)
- **Alkali** (pH>10)
- Harmless soluble inorganic salt
- Alcohol containing salt
- Hypochlorite solution
- Fine (tlc grade) silica and alumina

These chemicals should be washed down with excess water.

Organic Solvent
- **Chlorinated**

Example: DCM, Chloroform, Chlorobenzene etc.

- **Non-Cholronated**

Example: THF, ethyl acetate, hexane, toluene, methanol, etc.

Red List
- Compounds with transitional metals
- Biocides
- Cyanides
- Mineral oils and hydrocarbons
- Poisonous organosilicon compounds
- Metal phosphides
- Phosphorus element
- Fluorides and nitrites.

Solid Waste
- Lightly contaminated

Example: Gloves, empty vials/centrifuge .

Broken Glassware

Broken glassware are usually collected in plastic-lined cardboard boxes for landfilling. Due to contamination, they are usually not suitable for recycling.

Figure 1.2: Dispose of waste in the properly marked containers. If you don't know ASK! credit: `https://en.wikipedia.org/wiki/Chemical_waste/media/File:Chemical_Waste_Disposal_Guideline.jpg`

Glassware

1. General: Always check glassware for cracks or chips before using. If it is cracked or chipped ask your instructor for a new piece.

2. 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.
3. When working with glass tubing, follow the specific instructions of your teacher. Puncture wounds by broken glass are some of the most common laboratory injuries.
4. Never use dirty glassware. Improperly cleaned glassware may give incorrect results for chemical tests. Always clean glassware before using it.
5. Do not immerse hot glassware in cold water; it may shatter.

Electrical Equipment

1. When removing an 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.
2. Report damaged electrical equipment immediately. Look for things such as frayed cords, exposed wires, and loose connections. Do not use damaged electrical equipment.
3. **Electronic Balances:** Do not spill chemicals (liquid or solids) on the scales.
    (a) Never pour liquids over the scale. Use a tared beaker and pour the liquid into it while the beaker is on the lab bench.
    (b) Measure out solids carefully.
    (c) Clean scales after each use.
4. If you do not understand how to use a piece of equipment, ask the instructor for help.

Heating Substances

1. Be careful around any heating device, but be especially cautious around open flames. Take care that hair, clothing and hands are a safe distance from the flame.
2. Never heat heavy glassware such as graduated cylinders, suction flasks, or reagent bottles since they might shatter. Only heat beakers, Erlenmeyer flasks, test tubes (without corks), and crucibles.

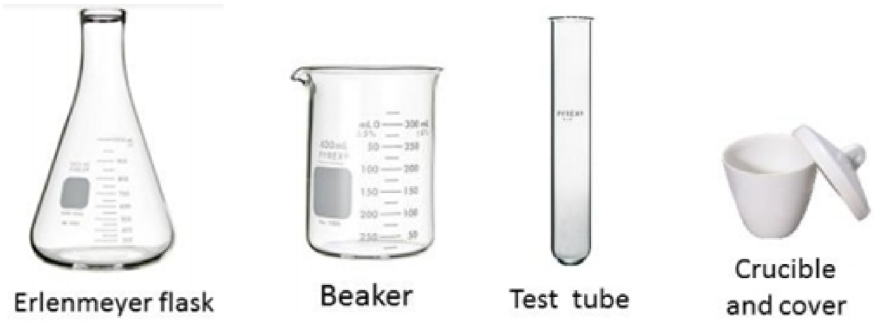
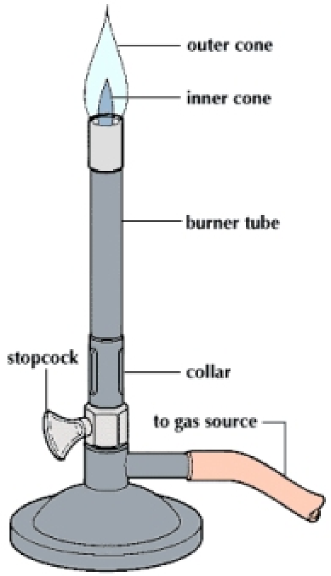


Figure 1.3: Only heat the following objects. credit: author

3. Do not put any substance into a flame unless specifically instructed to do so.
4. Do not reach over any heating device.
5. Do not leave a heating device unattended, and turn off a heating device when not in use. Never point a test tube towards a laboratory neighbor or yourself when heating or carrying out a chemical reaction. (It is like a loaded gun).
6. Heated metals and glass remain very hot for a long time. They should be identified (set on a ceramic pad) as being hot and set aside to cool.
7. Never look into a container that is being heated.
8. Be careful not to place hot equipment on flammable material.
9. Use tongs or heat-protective gloves if necessary.
10. Bunsen Burners:
    (a) The burner should be burning only for the period of time in which it is actually utilized.
    (b) Before lighting your burner carefully position it on the desk away from flammable materials, overhanging reagent shelves, flammable reagents such as acetone, toluene, and alcohol on neighboring desks.
    (c) Be careful not to extend your arm over a burner while reaching for something. Keep long hair tied back so that it cannot fall forward into a flame. Keep beards away from flames.
    (d) Always adjust your Bunsen burner to provide a small blue flame by adjusting the oxygen (by turning the barrel) and the gas (by turning the screw at the base). NEVER adjust the flame at the stop-cock. Be sure to turn off the Bunsen burner at the stop-cock (not at the base).



credit: unknown

Miscellaneous Procedures

1. Boiling Water: Add boiling chips to minimize bumping. Never fill the beaker over 75% full. Do not excessively heat the water. Make sure the container does not boil dry.
2. **Round objects roll.**
3. Detecting Temperature Changes: Carefully hold your hand near the object and feel for heat. If not heat is detected lightly touch the object. If no heat is detected you may hold the object normally. **Hot objects are hot.**
4. Fume Hoods. Any experiment involving the use of or production of poisonous or irritating gases

must be performed in a hood. Be sure the fume hood is turned on and that air is flowing from the room into the fume hood.

5. **Assemble safe apparatus:** Always assemble an apparatus as outlined in your instructions. Obtain instructor approval after apparatus is assembled before using it. Makeshift equipment and poor apparatus assemblies are the first steps to an accident.

Hazardous Materials (Chemical Safety and Waste Disposal)

1. General: Treat all chemicals as Hazardous unless you know for sure it is not.
2. MSDS: Material Safety Data Sheets are available for all chemicals used in lab. You may request to see them at any time (preferably prior to lab, to avoid unnecessary disruptions).
    (a) PEL/TLV's - Permissible Exposure Limit and Threshold Limit Value
    (b) Physical and Chemical Properties - BP, MP, VP, Solubility, Evaporation Rate
    (c) Fire and Explosion Hazards - Flash point, Auto ignition, Flammability limits, recommended extinguishing media
    (d) Health Hazard Data - LD50, LC50

Figure 1.4: Internationally accepted safey symbols. credit: Globally Harmonized System of Classification and Labeling of Chemicals, United Nations New York and Geneva, 2005

3. Classes of Hazardous Materials: Chemicals are often grouped into different classes of hazards materials. Learn each class.
    (a) Avoid Contact – general irritant
    (b) Carcinogen – cause cancer
    (c) Corrosive – destroy living tissue/equipment
    (d) Danger Unknown - Unknown
    (e) Explosive – may explode under some conditions
    (f) Flammable – vapors are flammable
    (g) Irritant – may irritate eyes, skin, respiratory tract
    (h) Lachrymator – as above but more severe

   (i) Mutagen – can cause genetic damage

   (j) Peroxide Former – very explosive

   (k) Poison – serious effects on body, often lethal

   (l) Stench – they smell, open only in hood

   (m) Teratogen – cause defects in fetus and embryo

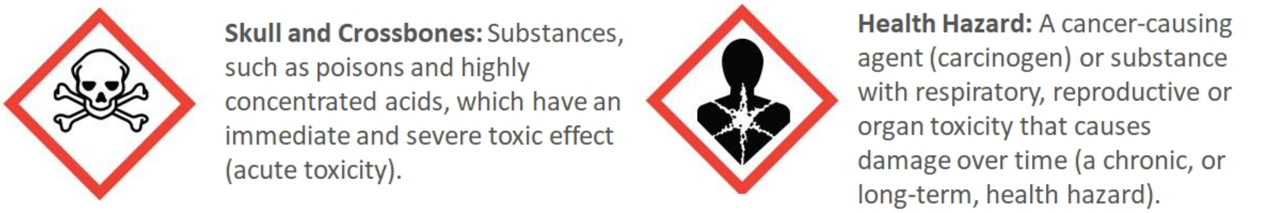
   (n) Toxic – hazardous to health if exposed

4. Types of Hazards:

   (a) **Acute Poisoning:** Rapid absorption leading to sudden and severe effects. Ex: CO, Cyanide poisoning

   (b) **Chronic Poisoning:** Prolonged or repeated exposure over long times, symptoms may not be immediate. Ex: Pb or Hg poisoning

   (c) **Substances in combination:** Two or more substances that interact leading to a larger effect. Ex: Alcohol + Chlorinated Solvents or Depressants



Skull and Crossbones: Substances, such as poisons and highly concentrated acids, which have an immediate and severe toxic effect (acute toxicity).

Health Hazard: A cancer-causing agent (carcinogen) or substance with respiratory, reproductive or organ toxicity that causes damage over time (a chronic, or long-term, health hazard).

Figure 1.5: Acute and chronic poisoning symbols. credit: author

5. Read the label: Read the label carefully, read it twice, before taking anything from a bottle. Many chemicals have similar names and chemical formula's, such as sodium sulfate and sodium sulfite. Using the wrong reagent can spoil an experiment or can cause a serious accident.

6. Excess Chemicals: Avoid using excessive amounts of reagent.

   (a) Never obtain more than called for in the experiment.

   (b) Do not return any excess chemical to the reagent bottle; share it with another student or dispose of it according the instructions given in the lab.

   (c) If you are uncertain how to dispose of an excess of a specific chemical, consult your instructor.

7. Matches/Burning Splits: Never throw lighted matches or wooden splits into a sink it may ignite a discarded flammable liquid. Wet the match or splint and dispose of in the properly labeled container.

8. Waste Chemical: Excess chemicals, or completed experiments should be discarded as instructed

in the lab manual. This will generally consist of placing the chemicals in the properly labeled container in the hood.

(a) Sinks and Drains: Never dispose of chemicals down the drain unless explicitly told to do so by the instructor or laboratory instructions.

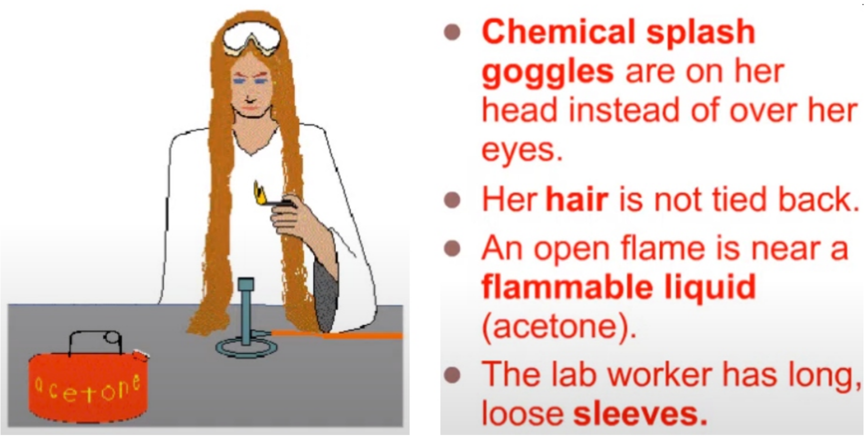
(b) Liquids: labeled containers in the hood.

(c) Solids: labeled container in the hood.

(d) Paper Towels: Trash can.

(e) Glass: Tubing waste or broken glass should be placed in the broken glass container.

The most important rule: "IF YOU DON'T KNOW ASK!"



- **Chemical splash goggles** are on her head instead of over her eyes.
- Her **hair** is not tied back.
- An open flame is near a **flammable liquid** (acetone).
- The lab worker has long, loose **sleeves.**

Figure 1.6: Quick safety quiz. For more examples visit the link. credit: `https://www.wisc-online.com/learn/abe-ell/science/sce0919/introduction-to-safety-in-the-chemistry-lab-v`

Your signature on this contract indicates that you have read this Safety Contract and that you subscribe to our efforts to insure the safety of students and staff during science laboratory activities. No student will be permitted to perform laboratory activities unless this contract is signed by both the student and instructor.

STUDENT'S PRINTED NAME

STUDENT'S SIGNATURE DATE

TEACHER'S SIGNATURE DATE

.

Experiment 2
Determination of an Unknown Anion

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Difference between qualitative and quantitative measurements.
2. Five signs that a chemical reaction has occurred.
3. Logical reasoning - inclusive, exclusive and indeterminate results.

Discussion

You accidentally forgot to label a bottle of clear liquid, how can you determine its contents? Or perhaps you find a mysterious bottle on the shelf, how can you identify it? We will explore using the chemical properties of compounds to identify an unknown liquid based on its characteristic chemical reactions. We will become familiar with the signs and observations that a chemical reaction has occurred. We will also explore several logic concepts including inclusive, exclusive, and indeterminate results.

Physical and Chemical Properties

Physical properties of substances is a characteristic of a substance that can be observed or measured without changing the identity of the substance, no chemical bonds are broken or formed. Physical properties include color, density, hardness, and melting and boiling points.

Chemical property describes the ability of a substance to undergo a specific chemical change. The substance is chemically changed, chemical bonds are broken and reformed resulting in a new substance with different physical and chemical properties.

In this experiment you will be given aqueous samples of sodium chloride ($NaCl$),sodium iodide (NaI), sodium hydrogen carbonate ($NaHCO_3$), sodium carbonate (Na_2CO_3), sodium hydrogen phosphate (Na_2HPO_4), and sodium sulfate (Na_2SO_4). You will test these compounds for characteristic reactions with nitric acid (HNO_3), barium nitrate ($Ba(NO_3)_2$), silver nitrate ($AgNO_3$), nitric acid (HNO_3), ammonium hydroxide (NH_4OH), and an acid-base indicator called thymol blue.

You will also be given an unknown sample that will be identical to one of the aqueous samples. After testing the known compounds, you will subject your unknown substance to those same tests. You will be able to determine the unknown sample's identity by matching its characteristic reactions with those of the known samples.

This experiment emphasizes the importance of accurate observations, properly recording them, and the use of deductive logic to draw conclusions from those observations. If you make accurate observations and record them faithfully, you will be able to identify the unknown compound without difficulty. Accurate observations are the only criterion for success. No prior knowledge of the chemistry that you will observe is necessary or assumed.

Signs of a Chemical Reaction

We will be observing the characteristic chemical properties of compounds through a series of tests with different reagents and observing if a chemical reaction occurs or not, and if so what specifically occurs. There are several signs that a chemical reaction has occurred. These signs are:

1. Color changes, that are not the result of simple dilution.
2. Evolution of a gas forming many bubbles rapidly.
3. Formation of a precipitate, normally a solid which will settle to the bottom of a test tube, but occasionally a finely divided precipitate that does not settle and often appears as only a milky color.
4. Disappearance of a precipitate.
5. Endothermic or Exothermic reactions (creation of heat or cold).

There are several cases where it is difficult to tell if a chemical reaction occurred. When dissolving a solid or precipitate, only a portion is dissolved, thus leaving some behind, and fooling a student into incorrectly reporting that no reaction occurred, or that a precipitate formed. This can be avoided by realizing that the original compounds are generally crystalline, and will sit at the bottom of a test tube (or settle out quickly), while a precipitate will have a milky appearance, and tend to spread evenly throughout the solution. A less common mistake is mistaking a lightening in color of a solution as a color change when it is simply due to the dilution of the original color. In general color changes will be fairly drastic, red to blue, green to yellow, not blue to light blue. Another common mistake is observing one or two bubbles due to the mixing or stirring of a liquid solution and reporting it as the formation of a gas. Generally the formation of a gas will be very rapid and vigorous, and should be very noticeable.

Standard nomenclature for these observations are shown below in Table 2.1. The exact nomenclature will vary from book to book, and person to person, but these general rules work well.

Observation	Nomenclature	Example
Color change	color 1 $\longrightarrow$ color 2	red $\longrightarrow$ blue
Gas formation	(g) - descriptor	(g) - vigorous
Formation of a precipitate	ppt - descriptor	milky white ppt
Disappearance of a precipitate	ppt dissolves (descriptor)	ppt dissolves (2 drops)
Exothermic or Endothermic	hot/cold	hot
No reaction	NR or x	NR or x

Table 2.1: Typical nomenclature used to describe chemical properties.

Inclusive and Exclusive Reasoning

When making observations and applying them to determining the identity of an unknown that may be a single chemical or a mixture of two or more chemicals, it is often necessary to determine whether a test provides an inclusive, exclusive or indeterminate answer. A test would be considered inclusive if the result indicated that a specific substance must be in the unknown mixture, an exclusive test would indicate that a specific substance must be excluded from the mixture. A third situation does exist which is when a test is indeterminate about whether a substance is in the mixture, neither including it or excluding it. Consider the following set of results in Table 2.6.

Compound	**Reagent 1**	**Reagent 2**	**Reagent 3**
A	ppt occurs	no reaction	hot
B	ppt occurs	ppt occurs	hot
C	no reaction	ppt occurs	(g) - No temp change
D	(g)	no reaction	hot
Unknown	ppt occurs	ppt occurs	hot

Table 2.2: Reaction Data for Several Compounds

Assuming you have an unknown mixture of two of the compounds A, B, C, and D and mixed the unknown with each of the Reagents 1-3. Table 2.3 summarize the results. The results from reagent 1 are inclusive in that that you must have either compound A and/or B in your mixture, excludes compound D because no gas is observed and is indeterminate for compound C because it could be masked by A or B. The results from reagent 2 are inclusive in that you must have either B and/or C, but you can't exclude A or D being present because their presence could be masked by B or C. Reagent 3 is inclusive for A, B, and D but excludes C since no gas was formed. Combining our results we can state that our unknown can not contain C and D and therefore must be a mixture of compounds A and B.

Another way to view the results above would be to create a small table showing all possible combinations of the solutions and crossing off combinations that are excluded based on your observations. In this problem, there are 4 solutions, taken 2 at a time means that there are 6 possible combinations (as shown in the last column below). Combining the results from all 3 tests, the only solution not crossed off is the mixture of compounds A and B.

Reagent	**Inclusive**	**Exclusive**	**Indeterminate**	**Combinations**
Reagent 1	A and/or B	D	C	(a,b) (a,c) ~~(a,d)~~ (b,c) ~~(b,d)~~ ~~(c,d)~~
Reagent 2	B and/or C	none	A and D	(a,b) (a,c) (a,d) (b,c) (b,d) (c,d)
Reagent 3	A, B, D	C	none	(a,b) ~~(a,c)~~ (a,d) ~~(b,c)~~ (b,d) ~~(c,d)~~

Table 2.3: Inclusive, Exclusive, and Indeterminate results

Procedure

In this experiment you will be observing the reactions of 6 known compounds and one unknown compound with different reagents. Record all observations you make in the appropriate place. The tests are separated into 4 different tests (Parts 1-4), you may perform the tests in any order. You may work in pairs to perform Parts 1-4, but each individual should obtain a different unknown to identify.

Read the entire procedure before beginning the experiment.

Part 1 - Testing for Gas Evolution (Station 1)

1. Label 6 clean test tubes with the identity of each known compound and place 1 ml (about 20 drops) of each known liquid in each using a disposable pipet.
2. Add 5 drops of 3 M Nitric Acid to each test tube and record your observations. Be very careful with the Nitric Acid, it will eat holes in your clothes!
3. Discard the contents of the test tube in the sink and rinse with distilled water. You may use these same test tubes for the rest of the experiment.

Part 2 - Barium Nitrate and Nitric Acid Test (Station 2)

1. Label 6 clean test tubes with the identity of each known compound
2. Using the known solutions available in the laboratory place 1 mL (20 drops) of each known compound into your labelled test tubes.
3. Add 3 drops of ammonia to each test tube and stir with a clean stirring rod. Remove the stirring rod and touch a drop of each solution to a strip of red litmus paper. The paper should turn from red to blue, if it does not add drops of ammonia to the solution until it does. Do not record any results for this step, we are simply preparing the solution for the next step.
4. Add 5 drops of the Barium Nitrate solution to each test tube. Mix each tube gently (as demonstrated in class) to obtain a homogeneous solution. Record your observations in the column labelled Barium Nitrate.
5. The next step only applies to those test tubes that contained precipitates. Add 5 drops of 6 M Nitric Acid to each of these test tubes. Using a clean stirring rod, test each solution with a strip of blue litmus paper. If the paper does not turn from blue to pink, continue adding drops until it does. Examine each test tube carefully and note which of the precipitates has dissolved. Record your observations in the Barium Nitrate + Nitric Acid column.
6. Discard the solutions in the test tubes in the Waste Container labelled “Barium or Ba^{+2} Waste” which can be found in the hood.
7. Wash your test tubes and rinse them with distilled water.

Part 3 - Silver Nitrate and Nitric Acid Test (Station 3)

1. Label 6 clean test tubes with the identity of each known compound
2. Using the known solutions available in the laboratory place 1 mL (20 drops) of each known compound into your labelled test tubes.
3. Add 5 drops of 0.1 M Silver Nitrate solution to each test tube. Mix each tube gently (as demonstrated in class) to obtain a homogeneous solution. Record your observations in the column labelled Silver Nitrate.
4. The next step only apply to those test tubes that contained precipitates. Add 10 drops of 6 M Nitric Acid to each of these test tubes. Using a clean stirring rod, test each solution with a strip of blue litmus paper. If the paper does not turn from blue to pink, continue adding drops until it does. Examine each test tube carefully and note which of the precipitates has dissolved. Record your observations in the Silver Nitrate + Nitric Acid column.
5. Discard the solutions in the test tubes in the Waste Container labeled “Silver or Ag^+ Waste” which can be found in the hood.
6. Wash your test tubes and rinse them with distilled water.

Part 4 - Thymol Blue test (Station 4)

Thymol blue is a strong dying agent, it will stain clothing and fingers so be careful with it.

1. Label 6 clean test tubes with the identity of each known compound
2. Using the known solutions available in the laboratory place 1 mL (20 drops) of each known compound into your labeled test tubes.
3. Add 5 drops of thymol blue to each test tube.
4. Dilute each solution by adding 2.5 mL of water.
5. Record your observations.
6. Carefully add 1 drop of 6 M Nitric Acid to each test tube.
7. Stir each tube gently and record your observations in the column labeled Thymol Blue.
8. Discard the solutions in the test tubes in the sink.
9. Wash your test tubes and rinse them with distilled water.

Part 5 - Identifying an unknown solution

1. Perform each of the above tests (Parts 1-4) on your unknown solution and record your results.
2. You should be able to identify your unknown solution after comparing the results of your unknown with those of the known solutions. If any ambiguities occur or you are unable to identify your

unknown repeat as many of the tests as you find necessary. You may consult your instructor as to which tests are most likely to be in error.

3. The Thymol Blue test is very sensitive to the concentration of the solutions tested. If your results for your unknown do not match any of the known compounds, you may try adding a little more of your unknown to the solution to make it more concentrated, and/or add water to the solution to make it more dilute.
4. Identify your unknown, and provide an explanation as to why you believe it to be the compound you chose.

Results

	Station 1	**Station 2**	
Compound	**Gas Evolution**	**Barium Nitrate**	**Barium Nitrate + Nitric Acid**
NaCl			
NaI			
$NaHCO_3$			
Na_2CO_3			
Na_2HPO_4			
Na_2SO_4			
Unknown#			

Table 2.4: Results: Part 1-2

	Station 3		**Station 4**
Compound	**Silver Nitrate**	**Silver Nitrate + Nitric Acid**	**Thymol Blue**
NaCl			
NaI			
$NaHCO_3$			
Na_2CO_3			
Na_2HPO_4			
Na_2SO_4			
Unknown#			

Table 2.5: Results: Part 3-4

I was given unknown NUMBER __________I think it is ____________________because ...(Explain)

Post Lab Questions

1. An unknown sample contains two out of the following five compounds. Use information recorded below to determine the identity of the unknown. For each of the results below state which compounds are Inclusive (could be in your unknown), Exclusive (can't be in your unknown), or are Indeterminate (may or may not be in your unknown). Finally, cross out any combinations (solution) that are excluded by the test by crossing out the number/solution in the chart.

Compound	**Reagent 1**	**Reagent 2**	**Reagent 3**	**Reagent 4**	**Reagent 5**
1	ppt	NR	NR	(g)	NR
2	ppt	ppt	NR	NR	ppt
3	NR	ppt	Hot	(g)	NR
4	NR	NR	Hot	(g)	NR
5	ppt	NR	Hot	NR	ppt
Unknown	ppt	ppt	Hot	(g)	NR

Table 2.6: Reaction Data for Several Compounds

1(a) Reagent 1 Results: A white fluffy ppt was formed.

Inclusive:

Exclusive:

Indeterminate:

(1,2) (1,3) (1,4) (1,5)
(2,3) (2,4) (2,5)
(3,4) (3,5)
(4,5)

1(b) Reagent 2 Results: A chunky yellow ppt is formed.

Inclusive:

Exclusive:

Indeterminate:

(1,2) (1,3) (1,4) (1,5)
(2,3) (2,4) (2,5)
(3,4) (3,5)
(4,5)

1(c) Reagent 3 Results: The reaction mixture was warm to the touch.

Inclusive:

Exclusive:

Indeterminate:

(1,2) (1,3) (1,4) (1,5)
(2,3) (2,4) (2,5)
(3,4) (3,5)
(4,5)

1(d) Reagent 4 Results: Vigorous gas was evolved.

Inclusive:

Exclusive:

Indeterminate:

(1,2) (1,3) (1,4) (1,5)
(2,3) (2,4) (2,5)
(3,4) (3,5)
(4,5)

1(e) Reagent 5 Results: No reaction occured

Inclusive:

Exclusive:

Indeterminate:

(1,2) (1,3) (1,4) (1,5)
(2,3) (2,4) (2,5)
(3,4) (3,5)
(4,5)

2. Which 2 compounds are in the unknown? Explain.

3. Based on chemical properties, is it possible to be 100% sure that a chemical unknown you test which matches the properties of an existing compound is indeed that compound? Explain.

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Name: ________________________ Class: _________________ Date: _________________

Prelab Questions

1. What is the difference between an inclusive, exclusive and indeterminate test results?

2. List the five signs that a chemical reaction has occurred.

3. List two typical mistakes that a student might make when deciding whether a chemical reaction occurred or not.

4. When adding Thymol blue (a dark blue liquid) to a clear liquid, the resulting solution is light blue. Did a reaction occur? Explain.

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Experiment 3
Measurements - Length

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Understand Accuracy and Precision.
2. Identify what type of measurements need to be made to determine accuracy and precision.
3. Understanding how to take and report measurements with the proper number of significant figure.
4. Make length measurements with a variety of measuring devices.
5. Use of Dimensional Analysis in routine calculations.

Discussion

Accuracy and Precision

Through observation of the world around us both qualitatively and quantitatively we as scientists seek to formulate hypothesis, theories, rules and laws to describe how matter behaves. Our ability to acquire accurate and precise data is crucial to our success as scientists. In this (and the next few) experiments we will learn how to obtain correct measurements that are as accurate and precise as possible.

The **Accuracy** of a measurement or calculation is how close the result is to the correct value. This is most often determined by making measurements on known materials and comparing the measured value to the accepted (or literature) value. The correct values used to be tabulated in the *Handbook of Chemistry and Physics* but usually these values can now be found on the internet, though one should always used a reputable site like NIST - National Institute of Standards and Technology or check several site to see that the correct values agree with each other.

The **Precision** of a measurement or calculation is how reproducible the result is if the experiment is performed multiple times or how close repeated measurements are to each other. In practice this means that you must take many measurements of an observed quantity to determine how precise your measurement is.

An Analogy for Accuracy and Precision

The term precision and accuracy are often confused by students. A representation of the differences can be seen in Figure 3.1. Precision is how close each × is to each other, while accuracy is a measure of how close the × is to the center of the target. The worst measurements are represented by the first target which is neither precise nor accurate. The best measurements are represented by the fourth target which is both precise and accurate. Targets two and three represent precision or accuracy but not both at the same time and either target could be an acceptable measurement depending on the requirements of the measurement. We always strive for both accuracy and precision but quite often there is not time in lab to do both, in which case it is important to know whether you measured something accurately or precisely.

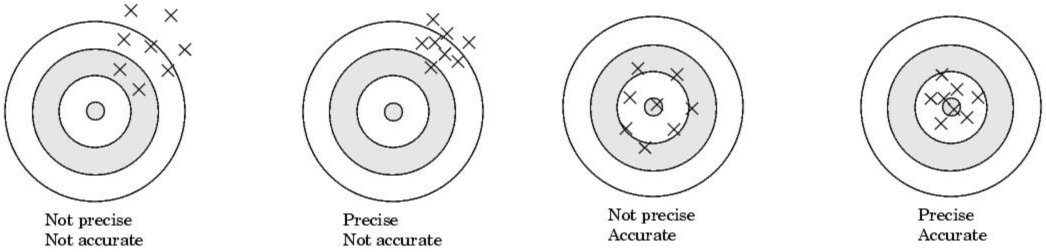


Figure 3.1: Targets illustrating the 4 possible combinations of accuracy and precision. Credit: unknown

Making Measurements in Lab

Scientific measurements always contain some amount of error in them. Even something as simple as counting the number of people the class room could result in errors if someone is counted twice or is out of the room using the bath room. Another example would be measuring the height of a person, did they have their shoes on or off, did they stand straight or slouch, are they bald or have lots of hair.

Even the devices used to make measurements are subject to error, often making it impossible to obtain the exact same measurement repeatedly. For example when you step on your bathroom scale it rarely reads the same weight. To indicate how well a measurement is made, record all the digits of the measurement using the markings that are known exactly and one further digit that is estimated or uncertain. The estimated or uncertain digit is obtained by noting the smallest unit of measurement given and then estimating between the values. These digits are collectively referred to as **significant figures**. Significant figures represent how well a measurement is made, the more significant figures in the measurement the better it is considered. Significant figures do not represent how accurate or precise a measurement is only how well you can make the measurement.

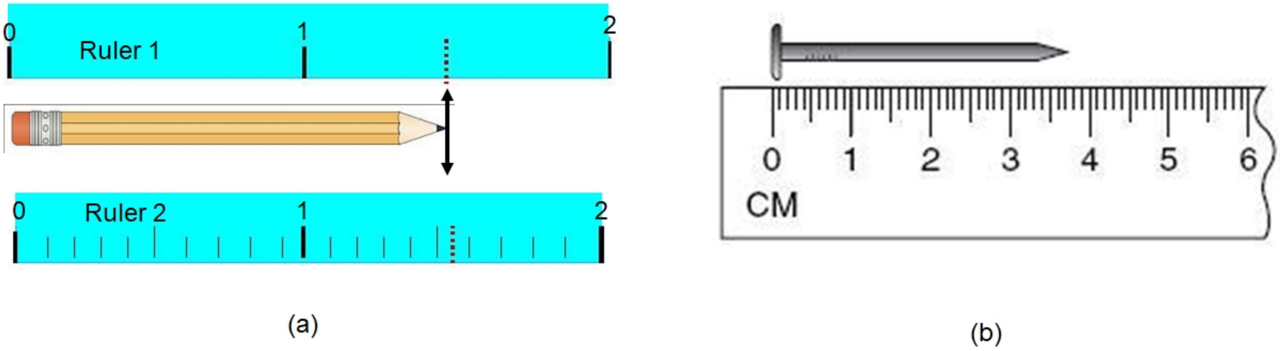


Figure 3.2: (a) Measuring a pencil with two different rulers will result in a different number of significant figures. (b) Anther example of how to read a ruler. credit: (a) author (b) unknown

For example in Figure 3.2 shows how one would report a measurement depends on what ruler is used to measure the pencil. Using the Ruler 1, the first digit is certainly 1, but we must estimate the tenths place. The length of the pencil, using Ruler 1, and written using the proper number of significant figures, would then be 1.5 inches, where the tenths place is an estimate. The second ruler has additional markings and allows a measurement with more significant figures to be made. The pencil is certainly between 1.5 and 1.6 inches, and we then must estimate the hundredths digit. The length of the pencil using Ruler 2 would be written as 1.55 inches.

In another example, the ruler shown in Figure 3.2 has large marks every 1 cm, and smaller marks every mm. The length of the nail is certainly between 3.7 and 3.8 cm, thus indicating at least 2 significant figures. However, you can add one more digit to your reported value by estimating the value between the lines. Thus, you could report the length of the nail as 3.75 cm, where the last digit represents your "estimated" value. The first rule is one should always try to "‘read between the lines"’. A second rule to remember when making measurements is that if the length of the nail was exactly on the 4 cm line, you should indicate the number of significant figures to which you can make this measurement by adding zero’s to your value and record the length as 4.00 cm.

The precision in a measurement is often quoted in terms of how inaccurate the estimated digit is. For example in Figure 3.2 the precision would be stated to be ± 0.05 cm. The ± indicates that the measurement could be shorter or longer then the stated value. The 0.05 cm indicates two things, first the number of places past the decimal point it is possible to read (in this case 2) and the precision to which the last digit can be estimated (in this case you can probably distinguish if a length was 1/2 way between the two millimeter marks (0.05 cm), but can not guess if it was 0.01 or 0.02 cm).

Key Idea: Making Measurements

Remember, **ALWAYS** read between the lines, and report measurements to as many decimal places as possible, using zeros to indicate additional precision if needed.

Procedure

To get you started performing measurements we will begin with a few simple exercises. Be sure to record data in the appropriate places in the results section as you complete each measurement. **All** results should be written with the proper number of significant figures and units. **ALL** calculations should be shown in the space provided.

A. Small Ruler

1. Measure the length of the following lines as accurately as possible using **centimeters**.
   Line 1:
   Line 2:
2. Make sure to record to the proper number of significant figures and include units in your answer.

B. Meter Stick

1. Measure the length and width of a piece of paper as accurately as possible using a meter stick in **centimeters**.
2. Measure the height of the lab bench.

C. Wood Block

1. Write down the Identity (ID) of the block.
2. Measure the Length, Width, Height of the block.
3. Calculate the volume of the block (Volume = L × W × H). Be sure to use the proper number of Significant Figures and units in your answer.
4. Check your answer with the instructor.

Results

All results should be written with the proper number of significant figures and include units. **ALL** calculations should be shown in the space provided.

A. Small Ruler

1. Length can be measured with the smaller ruler with what uncertainty? (±): ____________
2. Length of the first line ____________
3. Length of the second line ____________

B. Meter Stick

1. Length can be measured with the meter stick with what uncertainty? (±): ____________
2. Length of a piece of paper ____________
3. Width of a piece of paper ____________
4. Height of the Lab Bench. ____________

C. Wooden Block

1. Identity of the wooden block ____________
2. Length ____________
3. Width ____________
4. Height ____________
5. Volume ____________
6. Correct Answer (Check with instructor) ____________

Questions

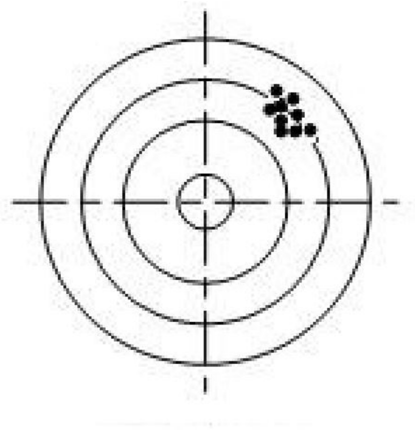
1. What **two** rules should one remember when measuring (in general, and specifically when using a ruler)?

2. The directions for an experiment state "cut a piece of chromatography paper 5.8 cm long".

   (a) Give the minimum length of paper that would satisfy these directions. 2(a) ____________

   (b) Give the maximum length of paper that would satisfy these directions. 2(b) ____________

3. Discuss the precision **and** accuracy of the shooter given the following target. Explain.



4. An alien from mars was measuring the length a football field using a meter stick. The alien said the room is 9100 cm long. What is wrong with the alien's answer? What should the alien's answer have been?

Credit: https://freesvg.org/angry-alien

Experiment 4
Measurements - Temperature

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Deepen our understanding of accuracy and precision.
2. Use of Percent Error.
3. Determine if a measurement is accurate or precise.
4. Understand the difference between Fahrenheit (°F), Celsius (℃), and Kelvin (K) scales
5. Convert between °F, ℃, and K scales.

Discussion

Temperature Measurements

Temperature is commonly measured in one of three units, degrees Fahrenheit (°F), degrees Celsius (℃) and Kelvin (K) as shown in Figure 4.1. In the United states the use of Fahrenheit is considered standard, while the rest of the world uses the metric Celsius scale. The Kelvin scale is an absolute scale which scientists frequently use and is used when discussing gas laws, thermodynamics and kinetics. For now the use of the Celsius scale will be considered our standard unit, with the introduction of the Kelvin scale discussed later.

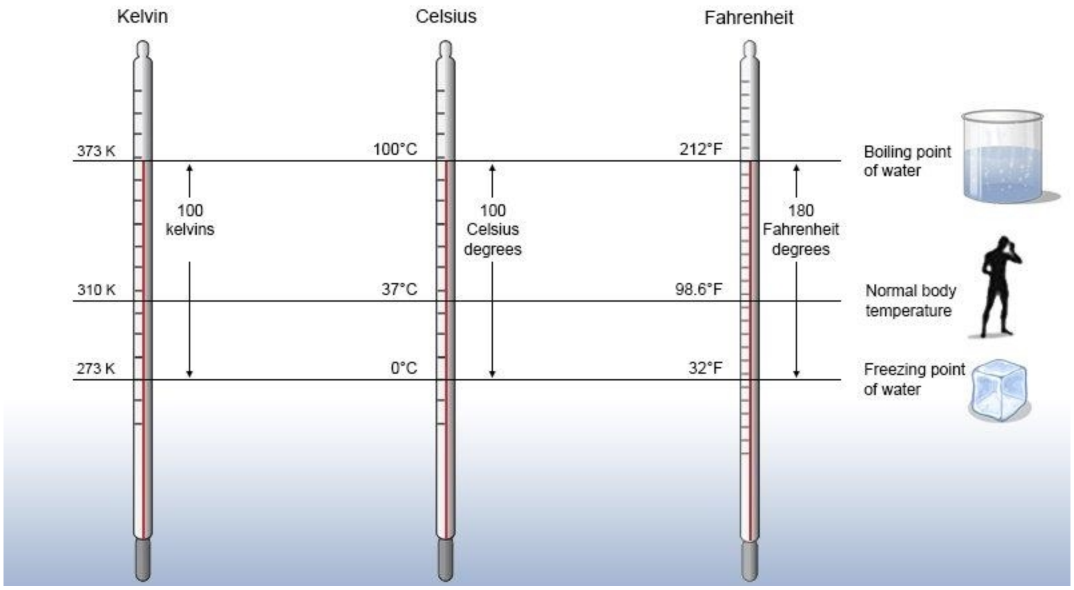


Figure 4.1: The relationship between the three most common temperature scales: degrees Fahrenheit (°F), degrees Celsius (℃) and Kelvin (K) credit: `https://commons.wikimedia.org/wiki/File:Temperature-Scale_diagram.jpg`

The main difference between the three temperature scales is the choice of standards for the freezing

$$
\begin{aligned}
{}^{\circ}\mathrm{F} &= 1.8 \times {}^{\circ}\mathrm{C} + 32 \\
{}^{\circ}\mathrm{C} &= \frac{({}^{\circ}\mathrm{F} - 32)}{1.8} \\
K &= {}^{\circ}\mathrm{C} + 273.15
\end{aligned}
$$

Figure 4.2: Equations for converting between the three most common temperature scales.

and boiling points and the size of the degree. The Fahrenheit scale was based on the freezing point of a salt water solution (0 °C) and the average human body temperature (originally set at 96 °F). The scale results in the freezing point of pure water as 32 °F and boiling water as 212 °F. There are 180 degrees separating the two values.

The Celsius scale sets the freezing point and boiling point of pure water at 0 °C and 100 °C and is based on the metric system resulting in the 100 units being placed between the two standards.

The Kelvin scale is based on the motion of atoms with 0 K (also known as absolute zero) being the temperature at which no molecular motion occurs. The size of the Kelvin unit is based on the metric system, thus the same size as the Celsius degree. The freezing point of water on the Kelvin scale is 273.15 K, while the boiling point is 373.15 K.

Conversions between the three temperature scales can be done using the equations in Figure 4.2. You do not have to memorize the formula, they are given on your cheat sheet. Other mathematical representations of the formula are possible with the factor of 1.8 often being replaced with the fraction 9/5, which is the ratio of the 180 Fahrenheit degree's to the 100 Celsius degree's. The addition or subtraction of 32 (and 273.15) is to correct for the different definitions of zero for each scale.

Error's in Measurements

In statistics **error** is the difference between the computed, estimated. exerimental or measured value and the true, specified or theoretically correct value. Errors are not mistakes. All measurements however are subject to a variety of errors which can result in incorrect values. Several examples are human error (simply being human is enough to get some things wrong), incorrect procedures, broken measuring instruments, and poorly calibrated instruments. The goal of a scientist is to minimize the possible sources of error in any measurement. The list of ways to do this is large, but the most common ones used in class will be by (1) following proper procedures, (2) calibrating instruments, (3) testing versus known quantities, and (4) making multiple measurements.

Percent Error

One measure of accuracy is to calculate the Percent Error in your measurement which represents the difference between the experimental or measured value and the theoretical or correct value scaled by the theoretical/correct value. It is given by the following formula:

$$\text{Percent Error} = \left| \frac{(\text{Experimental Value} - \text{Theoretical Value})}{\text{Theoretical Value}} \right| \times 100$$

For example know value for the boiling point of water is 100.0 ℃, but if you measure the temperature of boiling water in lab as 94.5 ℃ then the percent error in your measurement would be 5.5%. Make sure you try to calculation yourself and get the same result.

Calibration

Calibration is the validation of specific measurement techniques and equipment. At the simplest level, calibration is a comparison between measurements of known quantities and the value obtained using the measuring instrument. Many devices come pre-calibrated, but most electronic devices and thermometers do not come calibrated. The accuracy of the device is established by measuring known quantities and correlating this with the values measured on the laboratory equipment. This process is known as calibration. In general glassware for measuring volume is purchased already calibrated, thus we will only need to calibrate thermometers and balances (scales).

Your first task will be to check the accuracy of your thermometer by calibrating it to two known values. The two temperatures you will use are transition between ice and water (melting point) which occurs at 0 ℃ and the transition from liquid to gas (boiling point) which occurs at 100 ℃ at sea level. The melting point is not affected to a noticeable degree by elevation, in contrast, the boiling point is affected by elevation. Those curious may read about this in Hein Ch. 13.5 (p. 305), Chang Ch. 11.8 (p. 466) or in McMurry Ch. 10.5 (p. 396). The boiling point is the temperature at which the vapor pressure of a liquid is equal to the external pressure (generally atmospheric pressure) above the liquid. In Rangely CO, we are at an elevation of approximately 1 mile, this will result in the lowering of the boiling point of water to approximately 94.7 ℃ .

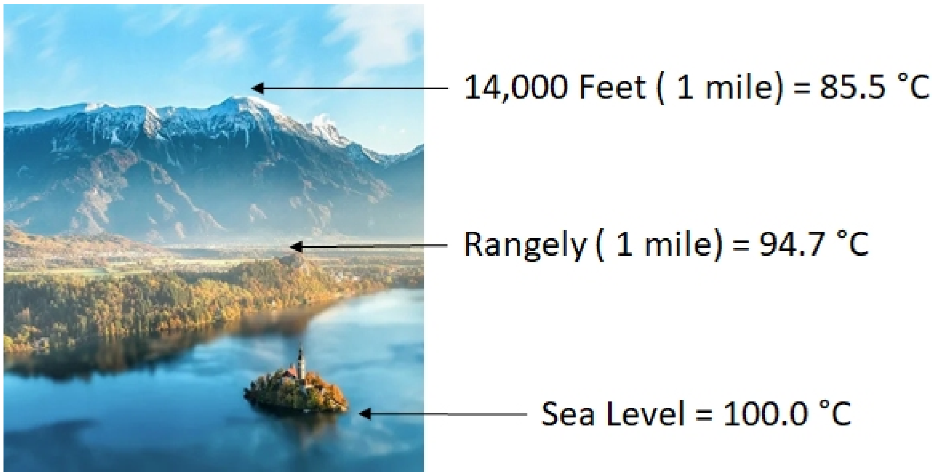


Figure 4.3: Boiling point of water at different elevations. credit: author

Thermometer Basics

The simple act of measuring temperature with a thermometer can easily involve many errors. When measuring the temperature of a liquid one can minimize error by observing the following procedures:

1. Hold the thermometer away from the walls of the container.
2. The bulb of the thermometer should be fully immersed in the water.
3. Allow sufficient time for the thermometer to reach equilibrium with the liquid.

4. Be sure the liquid is adequately mixed.
5. Always make readings at eye-level.
6. Sometimes a note card placed behind the thermometer makes the mercury easier to see.

Some additional pointers for using a thermometer:

1. Thermometers are round. Round objects on flat surfaces (lab bench tops) tend to roll. Always place your thermometer such that it will not roll off the lab bench top.
2. Thermometers are fragile, **do not** use them as stirring rods. Stirring rods should be used to stir liquids.
3. Always secure your thermometer with a clamp.
4. Many thermometers contain mercury (a toxic heavy metal), if you break a thermometer immediately inform your instructor so the proper hazardous waste procedures can be followed.

Proper use of Bunsen Burners

We will be using Bunsen burners for the first time in this experiment. Bunsen burners can be found in the tall drawer at the lab benches. Figure 4.4 below shows the proper set-up for a bunsen burner

1. Check the rubber hose attached to the Bunsen burner for cracks or splits. Report any bad hoses to your instructor so they may be replaced. Attach the hose to the stop-cock.
2. Check the turn screw at the bottom of the Bunsen burner, and leave it partially open (this controls the amount of gas burning).
3. Clear the area of flammable substances.
4. Turn the stop-cock perpendicular (90 degrees) to the pipe. Never adjust the amount of gas going to the Bunsen burner with the stop-cock, use the valve at the bottom of the burner. The stop-cock should always be either on or off, never between.
5. Adjust the flame using the screw at the bottom of the burner to adjust the gas flow, and the tube of the burner to adjust the oxygen flow. The goal is to form a small blue cone. The hottest part of the flame is the tip of the blue cone. It does not have to be large in order to provide the maximum amount of heat.
6. When done using a Bunsen burner turn it off at the stop-cock.
7. Allow the Bunsen burner to cool, then replace the Bunsen burner in the cabinet.

Procedure

A. Measure the melting point of ice

1. You will need a ring stand (in the cabinets under the lab benches), a ring clamp, wire pad, and a thermometer clamp.
2. Set up a ring stand as shown in Figure 4.5 and as can be seen at the instructor lab bench.
3. Have your instructor check your assembled apparatus before proceeding.
4. Fill a 500 mL beaker with approximately 75% ice and 25% distilled water.
5. Place the thermometer in the ice water solution and allow 3-5 minutes for the system to equilibrate.

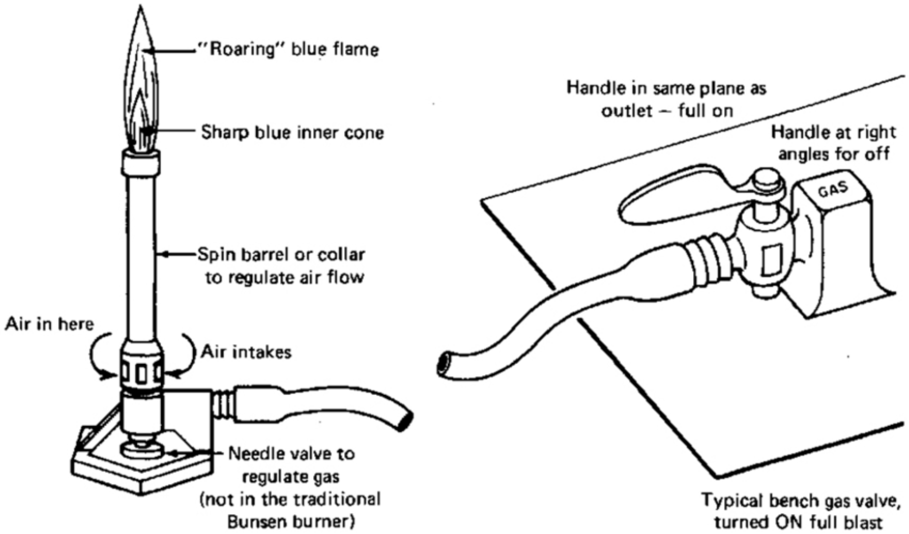


Figure 4.4: Typical bunsen burner. The flame on the bunsen burner is adjusted by changing the amount of air and gas. credit: unknown

6. Record the temperature on the results page.
7. If your results are not within 2 °C of the expected temperature, consult your laboratory instructor.

B. Measure the boiling point of water

1. Place approximately 350 mL of distilled water in a 500 mL beaker. (75% full)
2. Place 1-2 *boiling chips in the beaker.
3. Place the beaker on the ring stand as shown in Figure 4.5. Be sure to cover the ring stand with a piece of wire screening to set the beaker on.
4. Light the Bunsen burner and adjust the flame such that the blue cone is just touching the beaker.
5. Allow the water to boil for 3-5 minutes to establish a consistent temperature.
6. Obtain a small thermometer from your instructor.
7. Record the temperature of the water in the beaker.
8. If your results are not within 2 °C of the expected temperature of 96 °C consult your laboratory instructor.
9. Allow your water bath to cool before attempting to dispose of the water in the sink.

*Boiling results in the rapid formation of bubbles of air in a solution. In small containers (such as the beakers used in laboratory) or in pure liquids (such as the distilled water used in laboratory) the rapid formation of a large bubble can cause the beaker to "‘bump"’ or move abruptly causing it to be knocked off the ring stand. Boiling chips are used when boiling substances to prevent this from occurring by creating places for bubbles to form in a controlled manner.

C. Comparing Thermometers

1. Obtain a second (long) thermometer from your instructor.
2. The thermometer you received is **expensive**, take extremely good care of it to avoid the wrath of your instructor.
3. Record your observations of the thermometer in the results section.

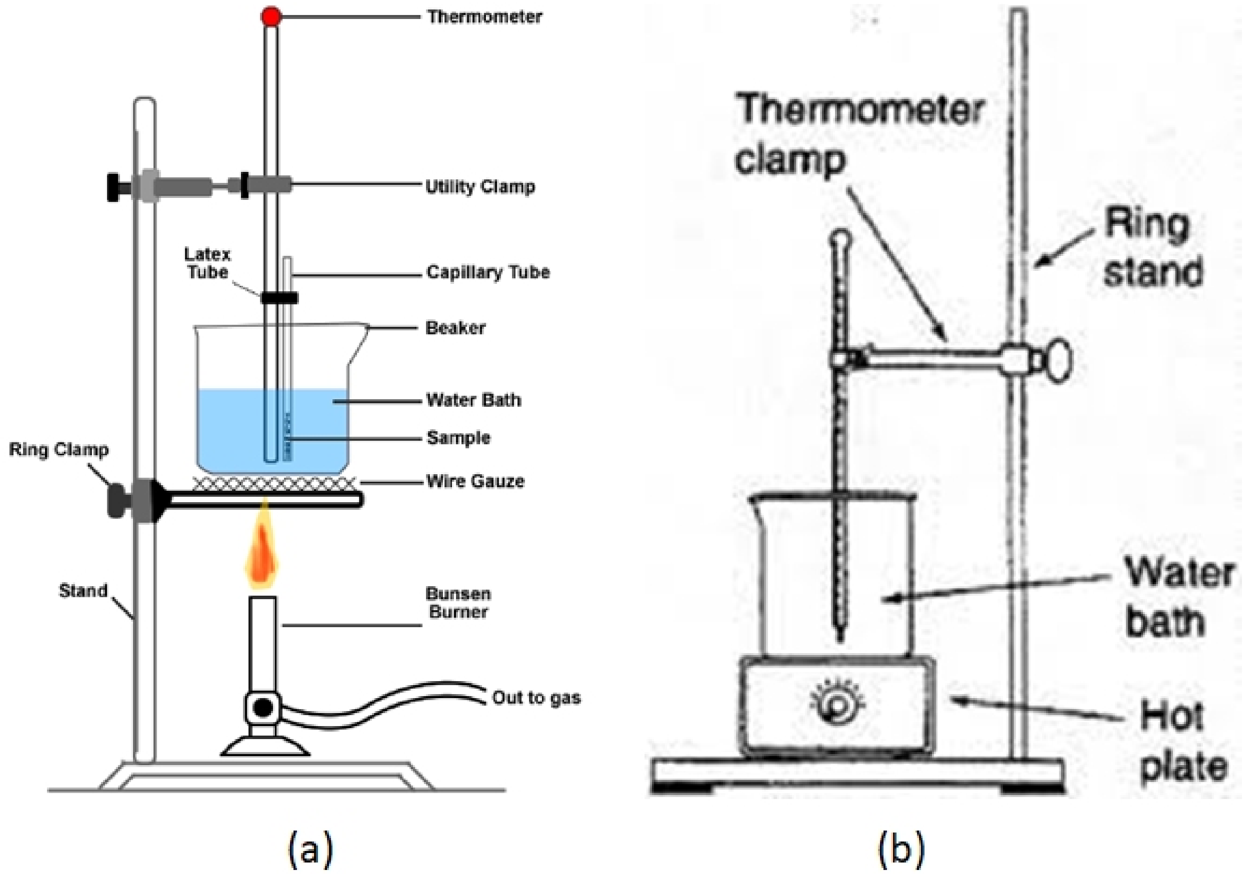


Figure 4.5: (a) Using a bunsen burner to boil water. (b) Using a hot-plate to boil water. credit: (a) `https://chem.libretexts.org/@go/page/95586` (accessed Jun 19, 2021) (b) unknown

Results

Cheap Thermometer

1. Temperature can be read with what uncertainty(±) (cheap thermometer): ______________
2. Temperature of ice-water bath (cheap thermometer): ______________
3. Percent Error in Freezing Point(show calculation below): ______________
4. Temperature of boiling water bath (cheap thermometer): ______________
5. Percent Error in Boiling Point (show calculation below): ______________

Expensive Thermometer

1. Temperature can be read with what uncertainty(±) (expensive thermometer) ______________
2. Temperature of ice-water bath (expensive thermometer): ______________
3. Percent Error in Freezing Point(show calculation below): ______________
4. Temperature of boiling water bath (expensive thermometer): ______________
5. Percent Error in Boiling Point (show calculation below): ______________

Questions

1. What are 4 most common ways to try to control or eliminate errors when performing experiments?

2. Are we testing the accuracy or the precision of the thermometer in this experiment? Explain.

3. Why is the second thermometer much more expensive than the first thermometer (used to measure the freezing point and boiling point of water)? Explain.

4. What should you do if you break a thermometer?

5. What basic rules should you follow when using bunsen burners?

6. Perform the following three conversions. Make sure to show your calculation, and include the correct number of Significant Figures and Units for each.

    (a) Convert 220 °C to °F. 6(a) ____________

    (b) Convert 15.4 °F to °C. 6(b) ____________

    (c) Convert 400. °F to Kelvin. 6(c) ____________

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Experiment 5
Measurements - Mass

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Understand the difference between Accuracy and Precision.
2. Understand how electronic instruments are calibrated.
3. Measure a requested quantity of a solid chemical.

Discussion

The *mass* of an object indicates the amount of matter present in the object. The *weight* of an object is a measure of the attraction that Earth has for the object. Because this attraction is proportional to the mass of the object, we will use the terms *mass* and *weight* interchangeably. The two most common devices used to measure mass are discussed below.

Triple beam balance

Before the advent of electronic balances, the triple beam balance was routinely used to measure masses. This is a **direct** measurement of mass by comparing the mass in the pan to the masses on the other side of the scale. While very accurate, triple beam balances are slow to use. We will not be using the triple beam balance in class.

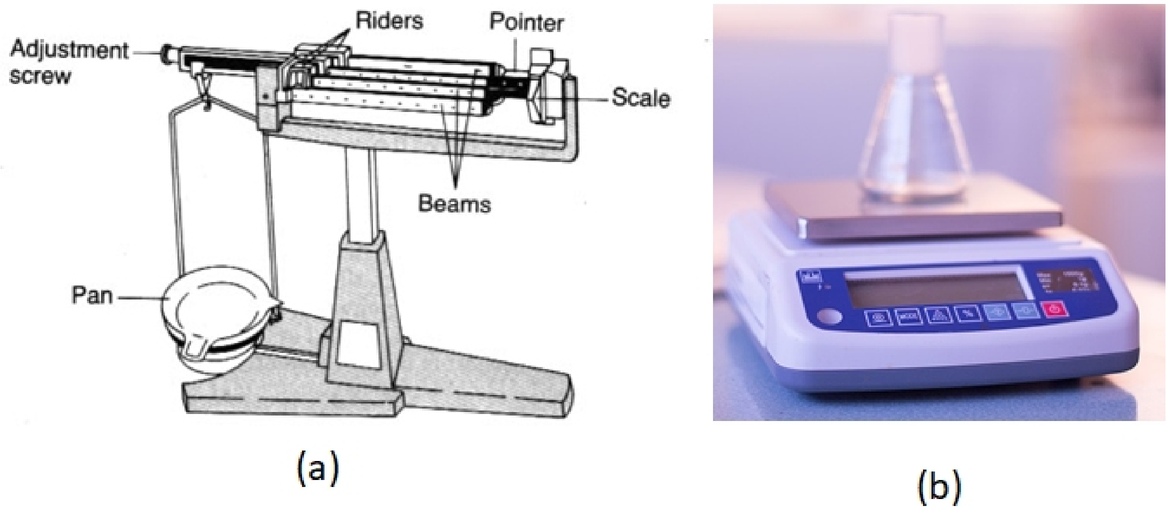


Figure 5.1: (a) Triple beam balance (b) Electronic balance. credit: (a) unknown (b) `https://commons.wikimedia.org/wiki/File:Electronic_scale.jpg`

Electronic balance The electronic balance provides an easy, precise and accurate way to measure mass. The precision and accuracy of electronic balances is generally a function of cost and can range from 0.1 gram (for 20 dollars) to 0.0001 gram (1000's of dollars). Electronic balances are an **indirect**

way to measure mass, the mass push’s down on the pan resulting in a change in magnetic field required to balance the weight. The balance must have calibration adjustments made to compensate for gravitational differences from changing locations and altitudes. This is achieved by **calibrating** the scale to a two known masses. The first mass used is generally zero (empty) and the second depends on the mass range of the scale (100.0 grams for out scales).

The electronic balances are an expensive piece of laboratory equipment and the utmost care should be taken when using them. Do not pour liquids over the scale, instead pour them into a pre-weighed or tared beaker. Abuse or misuse of the scales will result in removal and usage of the triple beam balances only.

Calibration

1. Calibration should be done the first time a measurement is made or anytime the scale is moved, bumped, left sitting for an extended period of time or if you looked at it cross-eyed. Subsequent measurements made should double check the calibration of your scale.
2. Turn the scale on by holding the ”ON” button for 5 seconds. The screen should read a model number and then read ”MENU”. Release the button.
3. The screen should then read ”CAL” which means the scale is in calibration mode.
4. With no mass on the scale push the ”ON/ZERO” button to calibrate the zero point on the scale.
5. After the scale is done blinking (1-2 seconds) the scale will prompt you to place the ”100 g” calibration mass on the balance. After placing the mass on the scale push the ”ON/ZERO” button.
6. The calibration is now complete.
7. The calibration mass should be within 0.005 grams of the correct value, if not inform your laboratory instructor.
8. Remove the mass and zero the scale.
9. The scale is now ready for making mass measurements.

Measurements

Mass measurements can be made one of two ways. The **method of differences** requires one to weigh the container used to mass a chemical, and then subtract that from the mass of the chemical plus container. This method is useful when sharing a scale, making measurements over a long period of time, or making measurements with liquids and will be used in some experiments.

For quick measurements or measurements that need to only be made once, the second method of **taring** is preferred. Using this method of weighing a chemical one places the empty weighing container on the scale and then **tares** or rezero the scale. This eliminates the mass of the weighing container and one is then able to measure only the mass of the chemical. The general procedure for making measurements in this manner is described below.

1. Tare the balance by placing the container you will measure your chemical in on the balance and pushing the "’Tare"’ or "‘Zero"’ button.
2. Remove the container from the scale and add the chemical to be massed.
3. Place the container + chemical to be measured in the balance pan. **NEVER** place chemicals directly on the balance pan, always use an appropriate container to hold the chemicals.

4. **NEVER** pour liquids directly over the scale, instead weight or tare the appropriate container, remove it from the scale, pour the liquid in, and place the container+liquid on the scale to be weighed.
5. **NEVER** overload the scale. The maximum mass placed on the scale should be 120 grams.
6. When measuring small masses (less then 1 gram) lower the plastic air shield to improve the accuracy of your measurements.

Measuring Requested Quantities

In most laboratory exercises you will be required to make many different measurements of volume, mass, and other quantities. Often, the precision required to produce a reliable results is quite different than the precision to which you can make the measurement. While you should always write down the measured value with as many significant figures as possible as possible, you do not always need exactly the specified value. For instance, if the instructions say to measure out 1.0 gram of a substance it is implied that the needed amount is 1.0 ± 0.1 grams, thus you should make sure your measurement is between 0.9 and 1.1 grams. It is not necessary (nor desirable) to try to measure exactly 1.000 grams of solid out, and often it is a waste of time to do so. When in doubt about how precise a measurement should be, ask your laboratory instructor.

Key Idea: Making Measurements

Always calibrate your scale before use. When measuring required quantities weigh out an approximate amount, but **always** record your result to as many significant figures as possible.

Procedure

1. Check the calibration of the balance using the instructions given in the discussion portion of the laboratory.
2. Record the mass written on the reference/known weight.
3. Record the measured mass of the reference/known weight. The reference/known mass should be within 0.005 g of the correct mass. If it is not, recalibrate the scale and inform the instructor.
4. Obtain an unknown mass from your instructor and record the identity of the unknown mass.
5. Record the mass of your unknown.
6. Verify you measured the correct mass by comparing your answer to the instructors answer. If your answers are more than 0.005 grams apart, you will need to redo the measurement.

Results

1. Mass can be read with what uncertainty (±): ____________
2. Weight of reference/known mass (value given on bottle): ____________
3. Measured weight of reference/known mass: ____________
4. Identity of unknown mass (A-Z or 1-10): ____________
5. Weight of unknown mass: ____________
6. Verify result with your Instructor: ____________
7. Mass of salt weighted out (student): ____________
8. Mass of salt weighted out (instructor): ____________
9. Percent error in measured mass of salt: ____________

Questions

1. In this experiment are we more concerned with Accuracy or Precision? Explain.

2. Why does an electronic balance need to be calibrated before each use? Explain.

3. Does the calibration mass have to weigh 100 grams? Explain.

4. The directions for an experiment state "weigh 15.50 grams of sodium chloride".

   (a) Minimum amount of sodium chloride that would satisfy the directions. 4(a) ____________

   (b) Maximum amount of sodium chloride that would satisfy the directions. 4(b) ____________

5. If asked to weigh out 6.70 oz of sodium chloride, how many grams should you weigh out? Explain. 5. ____________

.

Experiment 6
Measurements - Volume

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

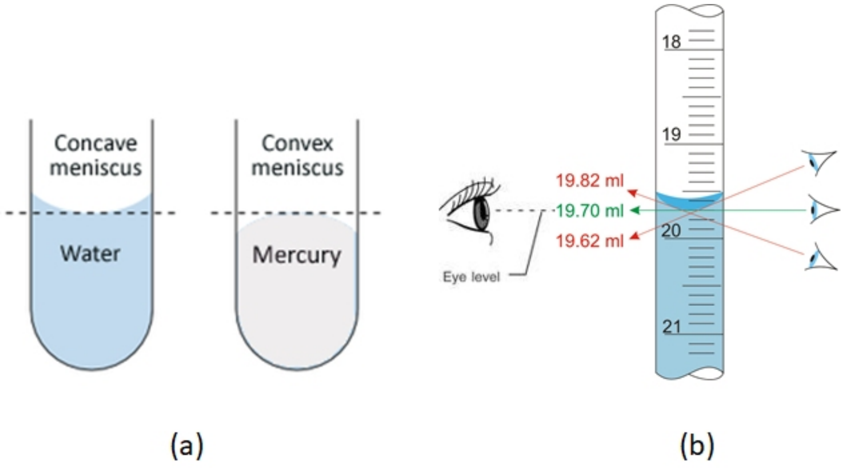
1. Understand the difference between Accuracy and Precision.
2. Calculate average and standard deviation
3. Measure volume using a variety of techniques

Discussion

The volume of a substance refers to the space it occupies. In chemistry labs liquids are most frequently measured by volume, though the may occasionally be measured by mass if the density is known.

Most liquids form a meniscus due to different attractive forces between the liquid and the liquid (cohesive) and the liquid and the container (adhesive). The formation of a meniscus is most apparent in narrow tubes. For water the menisus curves upwards (is concave) as the attractive forces between the water and the glass is larger than that between the water and other water molecules. For mercury it is the opposite, the meniscus curves downwards (concave) as mercury is more attracted to mercury than glass.

Volume measurements are commonly made by reading the point on the graduated scale that coincides with the bottom of the curved surface called the **meniscus** of the liquid as shown in Figure 13.2. A discussion of why a meniscus is formed can be found in Hein Ch. 13.5 (p. 305), Chang Ch. 11.3 (p. 443) or McMurry Ch. 10.4 (p. 391).



(a) (b)

Figure 6.1: (a) Two types of meniscus, (b) When measuring the volume of liquids always look at eye level and for water read the bottom of the meniscus. credit: (a) `http://water.usgs.gov/edu/meniscus.html` (b) E. Generalic, `https://glossary.periodni.com/glossary.php?en=meniscus`

Several methods for measuring liquid volumes are described below. Measurements in this laboratory will most often be done with graduated cylinders and pipets.

Beakers: Beakers are the most inaccurate way to measure volume, and are used only when the volume of a liquid is not important. They are inaccurate because of the large meniscus.

Graduated Cylinders: A graduated cylinder is a quick and easy (though less accurate) way to measure the volume of a liquid. They are useful for measuring any volume of liquid.

Plastic Pipets: Pipets come in two general types. Plastic pipets are used to quickly measure a volume of liquid when accuracy and precision do not matter. It is often assumed that 1 drop from a pipet is equal to 0.05 mL, thus there are 20 drops/mL.

Volumetric Flasks: These flasks are generally the most accurate measuring device, but are only designed to measure one specific volume. Common volumes are generally 100.0 mL, 200.0 mL, 250.0 mL, 500.0 mL, and 1000.0 mL.

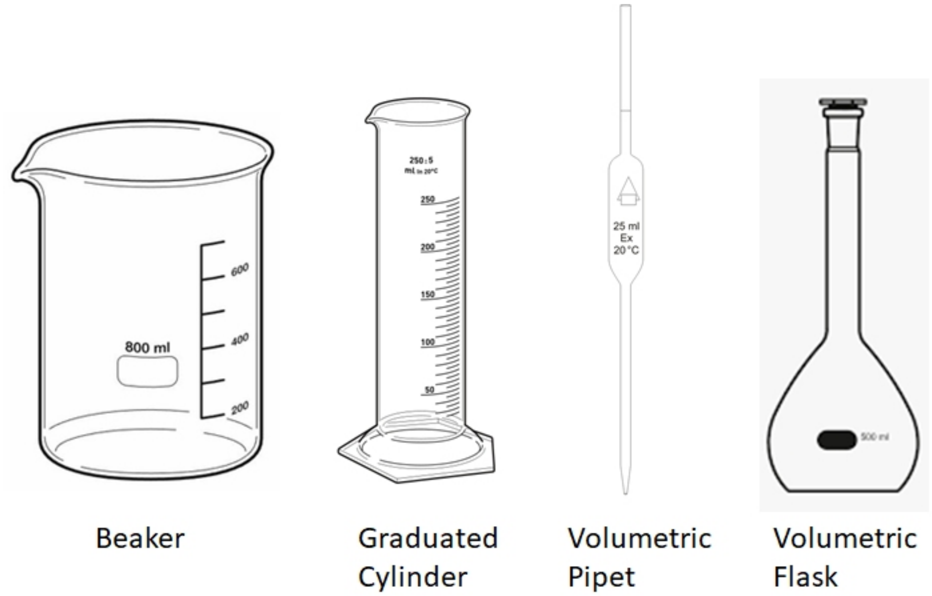
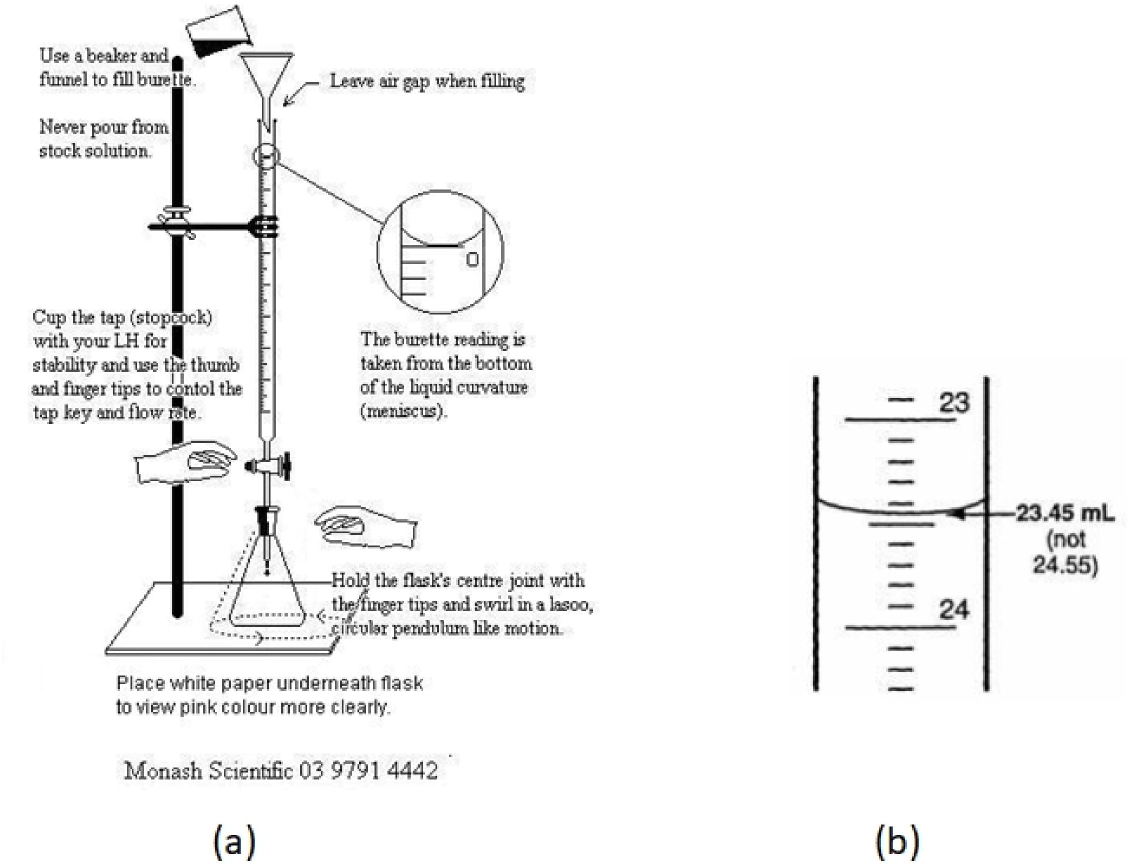


Figure 6.2: 4 different methods for measuring volume - beakers, graduated cylinders, volumetric pipets and volumetric flask. Each gives a different level of accuracy and precision. credit: author

Burets: A buret is used for measuring varying volumes of liquids and for delivering volumes of liquids accurately.

The proper use of burets is not obvious, the following steps should be followed:

1. Set up the buret as shown in Figure 13.1.
2. When reading the meniscus level your line of sight should be level with the meniscus to obtain the most accurate reading.
3. Note the numbering is reversed on a buret.
4. Add between 40 and 45 mL of distilled water to the buret using a funnel placed on the top of the buret. Be sure to check for air bubbles in the tip of the buret, if found you need to dispense your liquid until they are removed. Record your initial reading on your data sheet. Verify the reading.



(a) (b)

Figure 6.3: (a) How to setup and fill a buret. (b) Note the scale on a buret goes in the reverse direction. Be careful when making measurements. credit: (a) Monash Scientific 03 9797 4442 (b) `https://effectiveness.lahc.edu/`

5. Weigh an empty beaker.
6. Place an empty beaker below your buret and dispense 10 mL of distilled water into it. Record your final reading. Verify the reading.
7. Record the mass of water dispensed.
8. Calculate the volume of water dispensed into your beaker. (Show work)
9. If you have any troubles with this portion of the lab, or your mass and volume measurements disagree too much, consult your instructor, and do a second trial if needed.
10. Any water used in this part of the experiment may be disposed of down the sink.

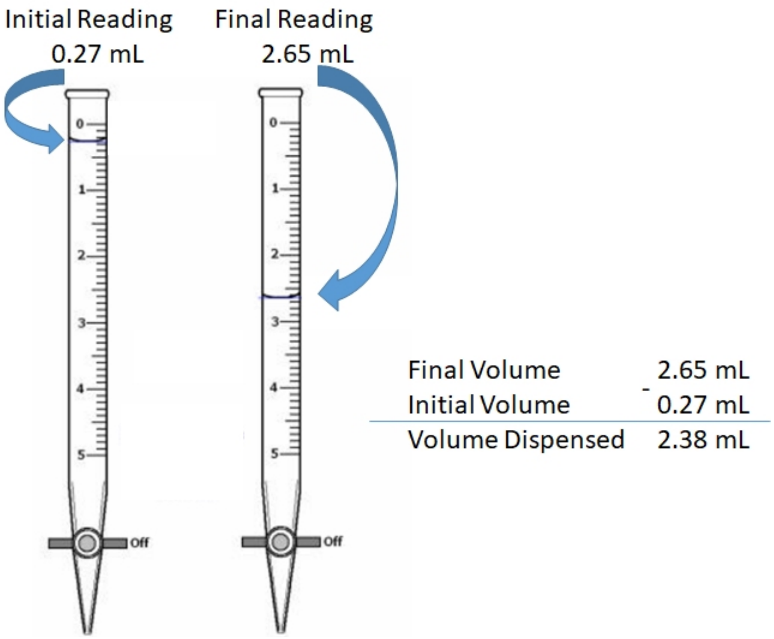
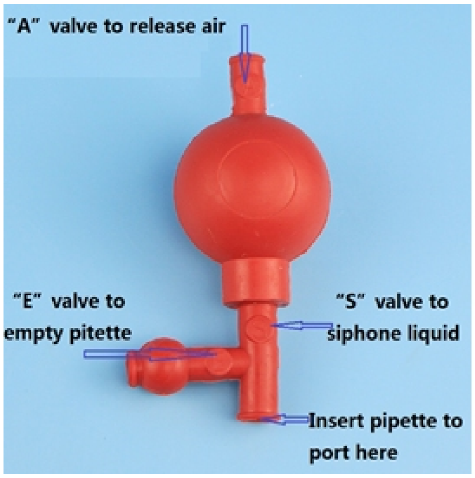


Figure 6.4: How to read a buret. Take the final volume - Initial volume = Volume dispensed. credit: author

Volumetric Pipets: Glass pipets are commonly called volumetric pipets because they are a simple, yet very accurate way to measure volume. Volumetric pipets are so accurate because the long neck decreases the error in measuring volume of the meniscus. They are only designed to measure one specific volume and commonly come in sizes 5 mL, 10 mL, 25 mL, and 50 mL. The accuracy of a volumetric pipets is commonly taken to be ± 0.01 mL. The pipet is read by placing the bottom of the meniscus on the measuring line. They are designed to leave one drop of liquid in the tip.

The proper use of pipets is not obvious, the following steps should be followed:

1. Place the pipet or suction bulb over the top of the pipet.
2. Evacuate air from the suction bulb by squeezing open the top valve (Labeled A) and gently squeezing air out of the bulb. When the bulb is empty release the top valve.
3. Place the pipet in the liquid being measured. Do not place the bottom of the pipet on the bottom of the container as this may damage the pipet.
4. Slowly open the valve between the suction bulb and the pipet (Labeled S) and allow liquid to fill the pipet to just past the reference line. Release the valve. Be careful not to overfill the pipet or get liquid into the bulb, as this may ruin the bulb.
5. Use the side valve (Labeled E) to lower the level of liquid in the pipet until the bottom of the meniscus is on the reference line. Close the valve.
6. Remove the pipet from the liquid being measured and place it over the container you are placing the liquid into. Use the side valve to release the liquid.
7. There will be a tiny amount of liquid left in the tip of the pipet, this is normal and the pipet is calibrated to leave this amount in the tip. If you blow out this liquid, your measurement will be inaccurate.



A – used to empty bulb of Air
S – used to draw liquid into pipet
E – used to release liquid from pipet
I – Insert pipet at bottom

Figure 6.5: How to use a pipet bulb. Insert pipet into bulb, remove (Air) from pipet, (S)iphon liquid into pipet, (E)mpty pipet. credit: author

Density

Mass measures the amount of matter in an object, while volume measures that space the object takes up. Both mass and volume are **extensive** properties, meaning that they depend on the size of an object. The larger the object the larger its mass and volume. Extensive properties are not very useful for determining the identity of the object since they are size dependent.

Density is the ratio of an objects mass to its volume. Density is an **intensive** property and does **not** depend on the size of the object. The density of any substance is the same no mater the size of the object. Density is also relatively unique for each substance (best seen by examining the density of different elements in a future experiment).

The density of water (by definition) is 1.000 grams/mL. We can use this fact to verify the volume of water measured by different laboratory equipment by comparing the volume measured to the mass measured which should be the same.

Accuracy and Precision

Recall in earlier labs we learned that the accuracy of a measurement is determined by how close the measurement is to the true value and we learned how to calculate the Percent Error in the measurement. To determine the precision of a measurement, multiple trials are required. The average value will be used when calculating the percent error. The **average** is given in Figure 6.6 in two forms, on the left is the formal mathematical expression, on the right is the more standard form.

$$\text{Average} = \frac{x_1 + x_2 + x_3 \ldots.}{n}$$

$$\bar{x} = \frac{1}{n} \times \sum_{i=1}^{n} x_i$$

$\bar{x} = \text{Average}$

$i = \text{individual measurement}$

$n = \text{number of measurements}$

Figure 6.6: (a) Standard Definition of Average. (b) Formal Mathematical Definition of an Average.

How close each measurement is to each other gives an estimate of the precision (reproducibility). Mathematically this is determined by calculating the standard deviation of the measurement. This is beyond the scope of this class, however you can still estimate which set of measurements is most reproducible and thus the most precise.

The following example shows the calculation of the average and standard deviation for an example set of data. A student measures the change in temperature of a sample 5 times with the following results:

Trial	**ΔT (℃)**
1	25.0
2	24.0
3	23.0
4	23.5
5	26.0

$$\bar{x} = \frac{25.0 + 24.0 + 23.0 + 23.5 + 26.0}{5} = \frac{121}{5} = 24.3$$

Figure 6.7: Example Calculation for average temperature change measured in an experiment.

Based on the example above, one would report the measured value as 24.3 ℃.

Procedure

1. Use your 100 mL graduated cylinder to measure out 10 mL of distilled water.
2. Weigh the mass of your water on the electronic balance. The graduated cylinder will probably overload your balance, so use a small beaker to measure the mass of the water. **DO NOT** pour water over the balances, take the beaker off the balance before pouring the water in. Remember to tare the beaker before making your measurement. Record your result.
3. Repeat the above measurement (steps 1 and 2) 2 more times. Record your results.
4. Calculate the average for your three measurements.
5. Repeat steps 1-4 using a 10 mL graduated cylinder.
6. Repeat steps 1-4 using a 10 mL volumetric pipet.
7. Repeat steps 1-4 using a buret.

Results

A. 100 mL Graduated Cylinder

1. Volume can be read to what uncertainty? (±): ____________
2. Weight of water in trial 1: ____________
3. Weight of water in trial 2: ____________
4. Weight of water in trial 3: ____________
5. Average weight of water in trials 1-3: $\left(\frac{\text{Trial 1}+\text{Trial 2}+\text{Trial 3}}{3}\right)$ ____________

B. 10 mL Graduated Cylinder

1. Volume can be read to what uncertainty (±): ____________
2. Weight of water in trial 1: ____________
3. Weight of water in trial 2: ____________
4. Weight of water in trial 3: ____________
5. Average weight of water in trials 1-3: $\left(\frac{\text{Trial 1}+\text{Trial 2}+\text{Trial 3}}{3}\right)$ ____________

C. 10 mL Volumetric Pipet

1. Volume can be read to what uncertainty (±): ____________
2. Weight of water in trial 1: ____________
3. Weight of water in trial 2: ____________
4. Weight of water in trial 3: ____________
5. Average weight of water in trials 1-3: $\left(\frac{\text{Trial 1}+\text{Trial 2}+\text{Trial 3}}{3}\right)$ ____________

D. Buret

1. Volume can be read to what uncertainty (±): ______________

2. Initial volume of water in the buret (trial 1): ______________
3. Final volume of water in the buret (trial 1): ______________
4. Difference in volume (trial 1): ______________
5. Weight of water (trial 1:) ______________

6. Initial volume of water in the buret (trial 2): ______________
7. Final volume of water in the buret (trial 2): ______________
8. Difference in volume (trial 2): ______________
9. Weight of water (trial 2:) ______________

10. Initial volume of water in the buret (trial 3): ______________
11. Final volume of water in the buret (trial 3): ______________
12. Difference in volume (trial 3): ______________
13. Weight of water (trial 3:) ______________

14. Average volume of water in trials 1-3: $\left(\frac{\text{Trial 1}+\text{Trial 2}+\text{Trial 3}}{3}\right)$ ______________

15. Average weight of water in trials 1-3: $\left(\frac{\text{Trial 1}+\text{Trial 2}+\text{Trial 3}}{3}\right)$ ______________

Questions

1. When making volume measurements one should read the level of water at the top, middle, or bottom of the meniscus?

2. What mass should your water have if you measured out 10.0 mL of water? Show your calculation or Explain.

3. In this lab are we measuring Accuracy or Precision? Explain.

4. Based on your measurements made in class, which measurement device was the most accurate? Explain.

5. Based on your measurements made in class, which measurement device was the most precise? Explain.

6. What sources of error could lead to imprecise and inaccurate measurements in measuring volume with the graduated cylinders? (Give 2 sources of error) Exlain.

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Experiment 7
Measurements - Density

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Understand the difference between Accuracy and Precision.
2. Make all measurements with the correct number of Significant Figures.
3. Calculate average, standard deviation, and percent error.
4. Measure density by calculating volume of a regular object.
5. Calculate density by volume displacement method.

Discussion

The density of an object is the ratio of the mass to volume, and is unique for any given compound or element. By measuring the density of an object, it is often possible to determine the nature of an unknown compound. A discussion of density can be found in Hein Ch. 2.9 (p. 37), Chang Ch. 1.7 (p. 18), or McMurry 1.10 (p. 16).

Density is defined as the ratio of an objects mass to its volume. Mathematically we can use the equation below where D = Density(in g/mL), M = mass (in grams) and V = volume (in mL).

$$D = \frac{M}{V} \tag{1}$$

While the equation form is useful and is taught in many schools, we should try to think of Density as simply another conversion factor that allows one to convert between mass and volume. Using Density as a conversion factor will also help us if a problem is given in non-standard units, for instance lb/ft^3.

For example we can ask what is the volume of 25.0 grams of gold. Using the equation method we would solve:

$$19.3\ \mathrm{g/mL} = \frac{25.0\ \mathrm{g}}{X\ \mathrm{mL}} \tag{2}$$

If instead we think of it as a conversion factor we would solve:

$$\frac{25.0\ \mathrm{g}}{1} \times \frac{1\ \mathrm{mL}}{19.3\ \mathrm{g}} = 1.30\ \mathrm{mL} \tag{3}$$

The mass of an object is relatively easy to measure using a balance. Volume is more difficult and is generally measured in two ways. Regular objects (such as cubes, cylinders, and spheres) can be measured with a ruler and the volume calculated mathematically as shown in Figure 7.1. The volume

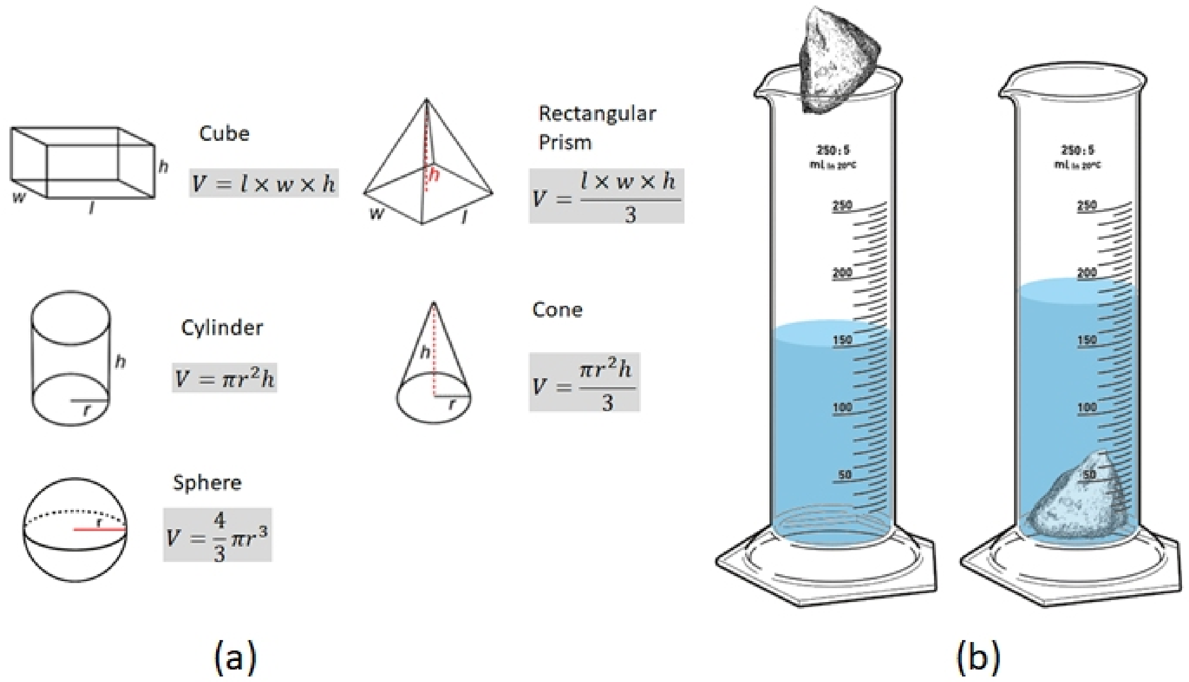

Figure 7.1: (a) Formula's for finding volume of regular objects. (b) Method of water displacement for finding volume of irregular objects. credit: (a) Author (b) `https://en.wikipedia.org/wiki/Displacement_(fluid)`

of irregular objects is generally measured by the water displacement method, the object is submerged in water and the change in volume of the water is the volume of the object. In this lab both methods will be used.

Because most substances change in volume when heated or cooled due to expansion or contraction, the density of an object is also temperature-dependent. Usually the variance is quite small for instance, the density of water at 3.98 ℃ is 1.0000 g/mL which decreases to 0.99987 g/mL at 0 ℃. We will generally assume the density of water is 1.00 g/mL at room temperatures.

The density of the samples used in class can be found in the table below. Additional values can be found on the back of your Cheat Sheet.

Object	**Density (g/ml)**	**Object**	**Density (g/mL)**
Water (4 ℃)	1.000	Aluminum	2.70
Copper	8.96	Lead	11.34
Mercury	13.55	Gold	19.3
Brass	8.5	Iron	7.86
Nickel	8.90	Steel	7.8
Tin	7.28	Manganese	7.20
Zinc	7.10	Woods Alloy	9.63

Table 7.1: Density of Common Metals

Procedure

A. Density of a regular object

1. Obtain a regular shaped object from your laboratory instructor. Record its identity.
2. Weigh the object, record your result.
3. Measure the objects length, width, and height.
4. Calculate the volume of the object. Be sure to show your calculation.
5. Calculate the density of the object. Be sure to show your calculation.

B. Density of an irregular object

1. Obtain an irregular shaped object from your laboratory instructor. Record its' identity.
2. Weigh the object and record your result.
3. Fill a burret half full of water. Record the volume.
4. Carefully (slide it down the side of the tube, do not drop it in) place your irregular object in the burret and record the volume.
5. Calculate the volume of the item.
6. Calculate the density of the item using the appropriate formula.
7. Compare the density of the item with the value given by your instructor, if your measurement is within 0.2 g/mL of the correct value move onto the next section, if not, redo this section.

C. Density of an unknown object

1. Obtain an unknown object from your instructor. Record its identity.
2. Determine the density of the object using the methods described above.
3. Do 3 trials and average your result.
4. Compare your experimental density with the density of the known elements provided in the discussion section.
5. Write down the identity of the known element which most closely matches your measured density.

.

Results

A. Density of a regular object

1. Identity of regular object: ______________
2. Mass of regular object: ______________
3. Length of regular object: ______________
4. Height of regular object: ______________
5. Width of regular object: ______________
6. Calculate the volume of the regular object, show the formula used and the calculation: ______________
7. Calculate the density of the regular object: (Show formula and work) ______________
8. Correct value for density of regular object (ask instructor) ______________
9. Your answer should be within 0.05 g/mL of the correct answer. If your answer is not, redo steps 1-8.
10. Calculate the Percent Error in your measurement using the following formula: $\left|\frac{\text{Actual Density}-\text{Measured Density}}{\text{ActualDensity}}\right| * 100\ \%$ ______________

B. Density of an irregular object

1. Identity of irregular object: ______________
2. Mass of irregular object: ______________
3. Initial volume of water in the buret: ______________
4. Final volume of water in the buret after adding irregular object: ______________
5. Calculate the volume of the irregular object. (Show work): ______________
6. Calculate the density of the irregular object (Show work): ______________
7. Correct value for density (obtained from cheat sheet): ______________
8. Your answer should be within 0.2 g/mL of the correct answer. If your answer is not inform your instructor, redo steps 1-8.
9. Calculate the Percent Error in your measurement: ______________

C. Density of an unknown object (3 Trials)

		Trial 1	Trial 2	Trial 3
1.	Identity of unknown object:	________	________	________
2.	Mass of unknown object:	________	________	________
3.	Initial volume of water in the burret:	________	________	________
4.	Final volume of water in the burret after adding unknown object:	________	________	________
5.	Volume of unknown object (show work):	________	________	________
6.	Measured density of unknown object (show work):	________	________	________
7.	Average density of unknown object:		________	
8.	Using Table 1, what metal is the unknown object?		________	
9.	What is the percent error in your measurement?		________	

Questions

1. The density of a lead block is 11.34 g/mL. What is the volume of a lead block weighing 120.5 grams?

1. ______________

2. A lead fishing weight displaces 12.0 mL of water. What is the mass (in milligrams) of the lead weight?

2. ______________

3. A bar of gold weighs 250 lbs. What is its volume in ft^3?

3. ______________

4. An empty graduated cylinder weighs 82.450 g. When filled to 50.0 mL with an unknown liquid it weighs 110.810 g. What is the density of the unknown liquid?

4. ______________

5. Five liquids (Ethyl Alcohol, Glycerin, Sulfuric Acid, Vegetable Oil, and Water,) are poured into a large beaker and given time to separate. What is the order of the liquids in the beaker from bottom to top. Sketch and label a picture for your answer.

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Experiment 8
Graphing

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Graph experimental data.
2. Make a "good" looking graph.
3. Labeling and numbering of axis.
4. Adding a curve-fit to a graph.

Discussion

It is often said that a picture is worth 1,000 words, or for scientists we might rephrase it to say that a graph is worth 1,000 words. Graphs are most often used to express data in a clear, concise and meaningful way. Trends are readily apparent, and information is easy to extract. In today's laboratory we will learn how to properly present data in a graphical format. Graph terminology, basic graph features, basic graphing skills using excel, and general graph functions will be mastered.

> A picture is worth a thousand words.
> -unknown
>
> If a picture is worth a thousand words, please paint me the Gettysburg Address
> -Leo Rosten

Various types of graphs are possible, the most common type uses the Cartesian coordinate system to show the relationship between two variables, the independent (x) and dependent (y) variables. The independent (x) variable is generally controlled by the experimenter (or naturally increments) and is plotted on the x-axis (horizontal axis) while the dependent (y) variable changes as the independent variable changes and is generally plotted on the y-axis (vertical axis). The goal of any graph is to maximize the amount of data shown in the space provided.

A. Graph Terminology

The terminology used in making graphs may be new to many of you. Figure 8.1 shows features commonly found on graphs.

1. X-axis: Generally the independent variable is plotted on the horizontal axis.

2. Y-axis: Generally the dependent variable(s) are plotted on the vertical axis.

3. Title: A general description of what is being graphed.

4. Axis Title: Titles for X and Y axis. The title should include a Label for what you are plotting and the Units you are plotting with.

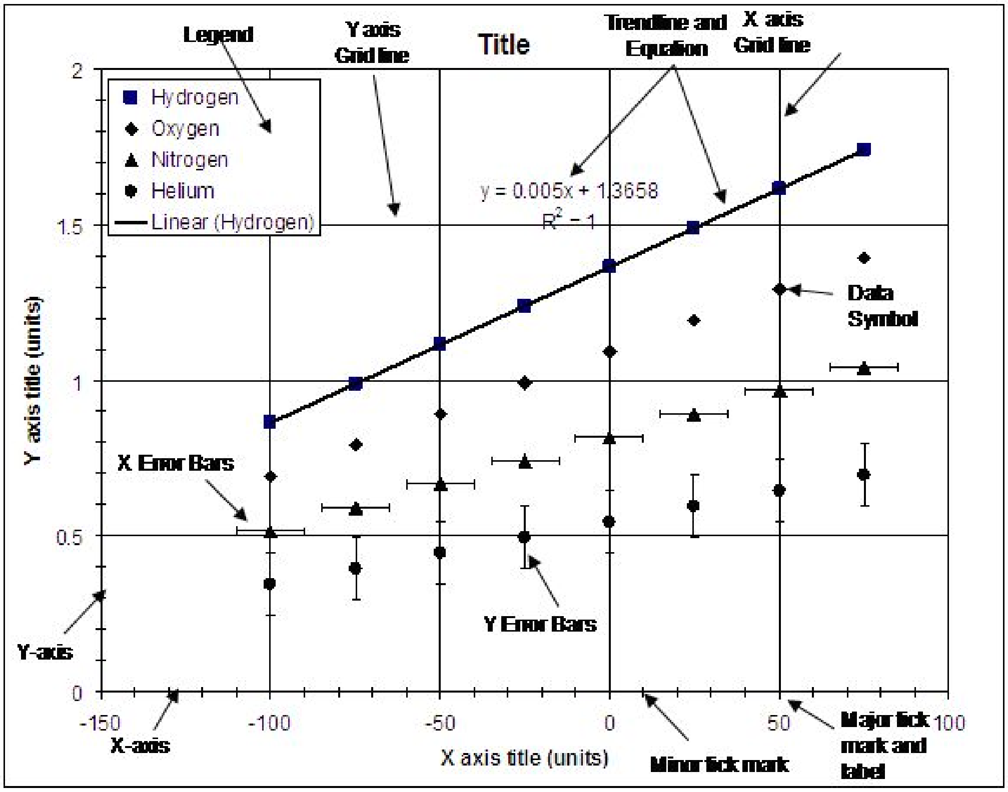


Figure 8.1: Anatomy of a graph. credit: author

5. Legend: Describes the independent variables being graphed. It includes the symbol representing each individual variable and a description or identifier.
6. Grid Lines: Lines on the graph drawn parallel to the X or Y axis. They generally are used as guides for the eye to make it easier to see the value of data points.
7. Tick Mark: Small hash mark on X or Y axis that denotes a specific value. Major: A label and grid line are normally drawn here. Minor: Normally just a tick mark on axis to indicate smaller units. Normally not labeled or given grid lines.
8. Tick Label: Label associated with tick marks on the X and Y axis.
9. Data Symbols: Symbols used to mark the data points on a graph.
10. Error bars: Horizontal bars (X error bars) or vertical (Y error bars) sometimes found attached to data points indicating the error involved in the measurement.

B. Good vs. Bad Graphs

A good graph is designed to convey a large amount of information in a small space. It is important to make full use of the space provided to maximize the information. An example of a good graph is shown in Figure 8.2. Several design goals are listed below.

1. Title - All graphs should have a title which conveys what is being graphed.

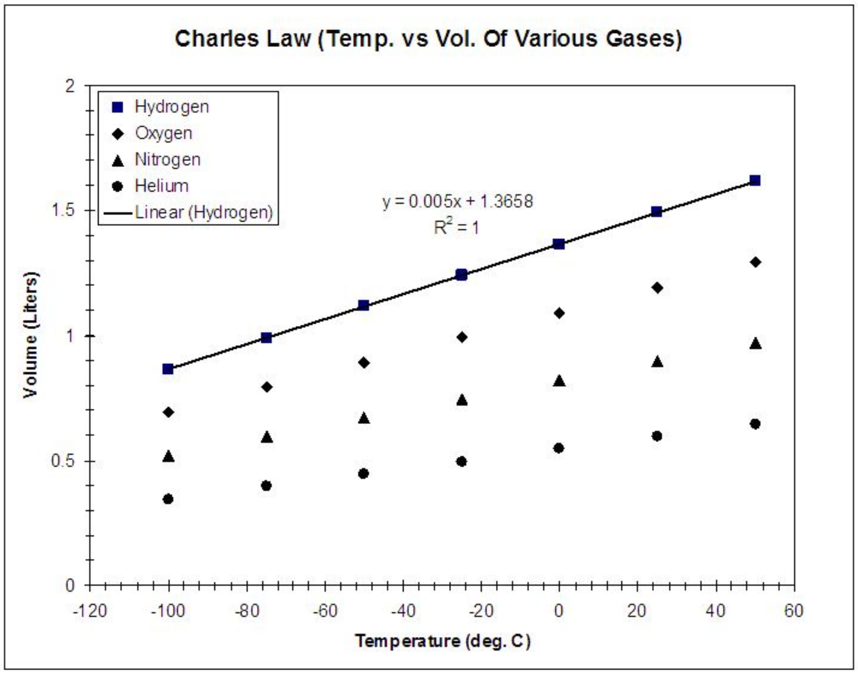


Figure 8.2: An example of a good graph. credit: author

2. Labels - All graphs should have both axis properly labeled with what is being plotted and the units.

3. The x and y axis should use convenient units (both major and minor). The X and Y axis labels are reasonably spaced (i.e. not too close together or too far apart).

4. Legend - All graphs should include a legend if more than one set of data is being plotted that provides information about the different data sets.

5. Text should be readable. Don't use fancy fonts or fonts too small to read.

6. Symbols used for data points should be:

    (a) Proper size (not too big or small)

    (b) Proper color (easy to distinguish). Generally use black and white only as people may not be able to print graphs in color.

    (c) Proper shape (don't use crosses or dashes and avoid open symbols when possible)

    (d) Avoid using duplicate shapes in different colors. It looks fine on a computer screen but if printed in black and white it will be impossible to distinguish between them.

7. Maximize the size of your graph. It is often convenient to drag the Legend inside the graph so that you can resize the graph and make it larger and easier to read.

We will look next at several bad graphs. Figure 8.3 has several things wrong with it.

1. No title
2. Missing label on Y axis
3. Missing units on X axis
4. No legend
5. X axis units are too close together, overlap, and are nearly impossible to read.
6. Y axis units are spaced so far apart they are useless

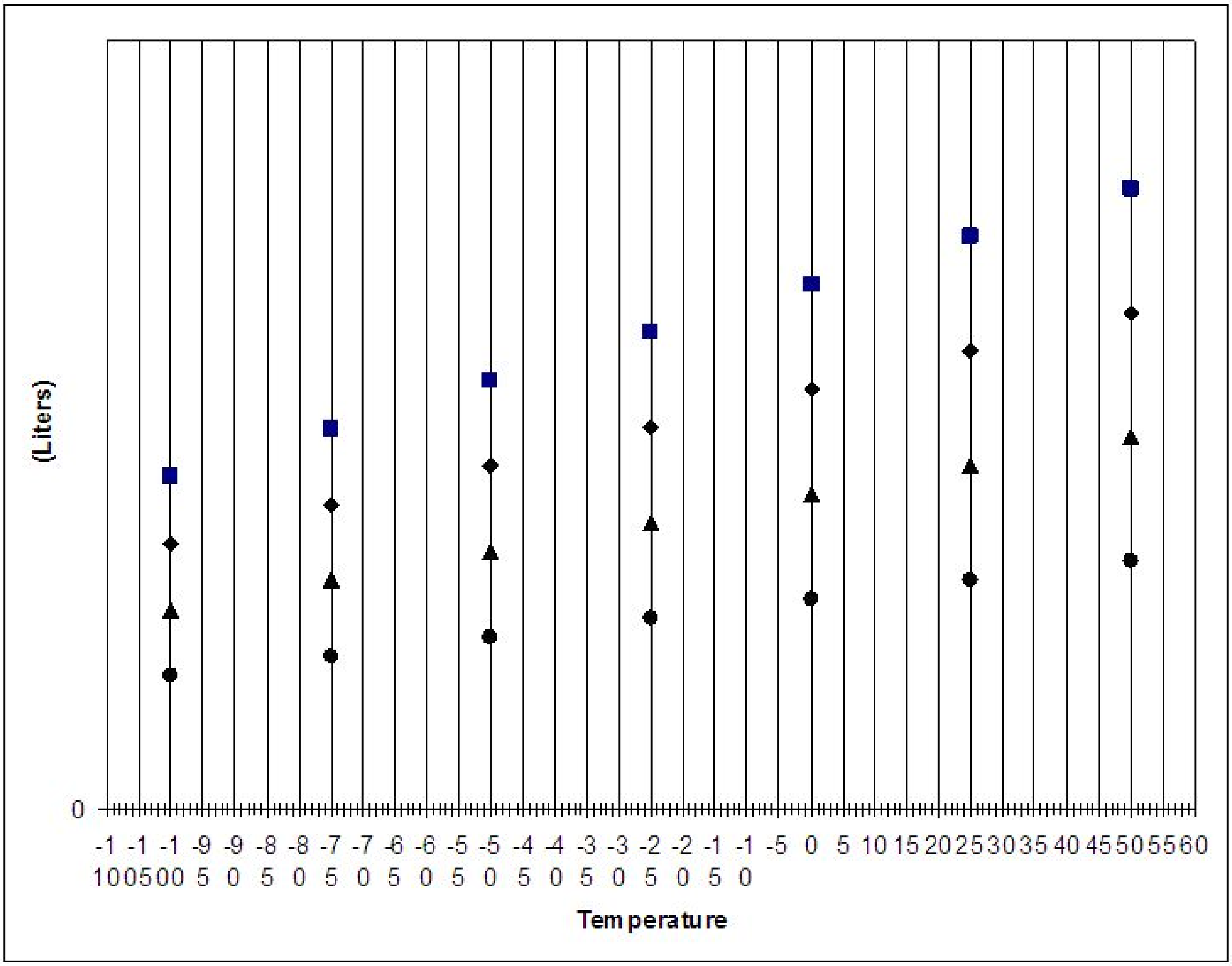


Figure 8.3: An example of a bad graph. credit: author

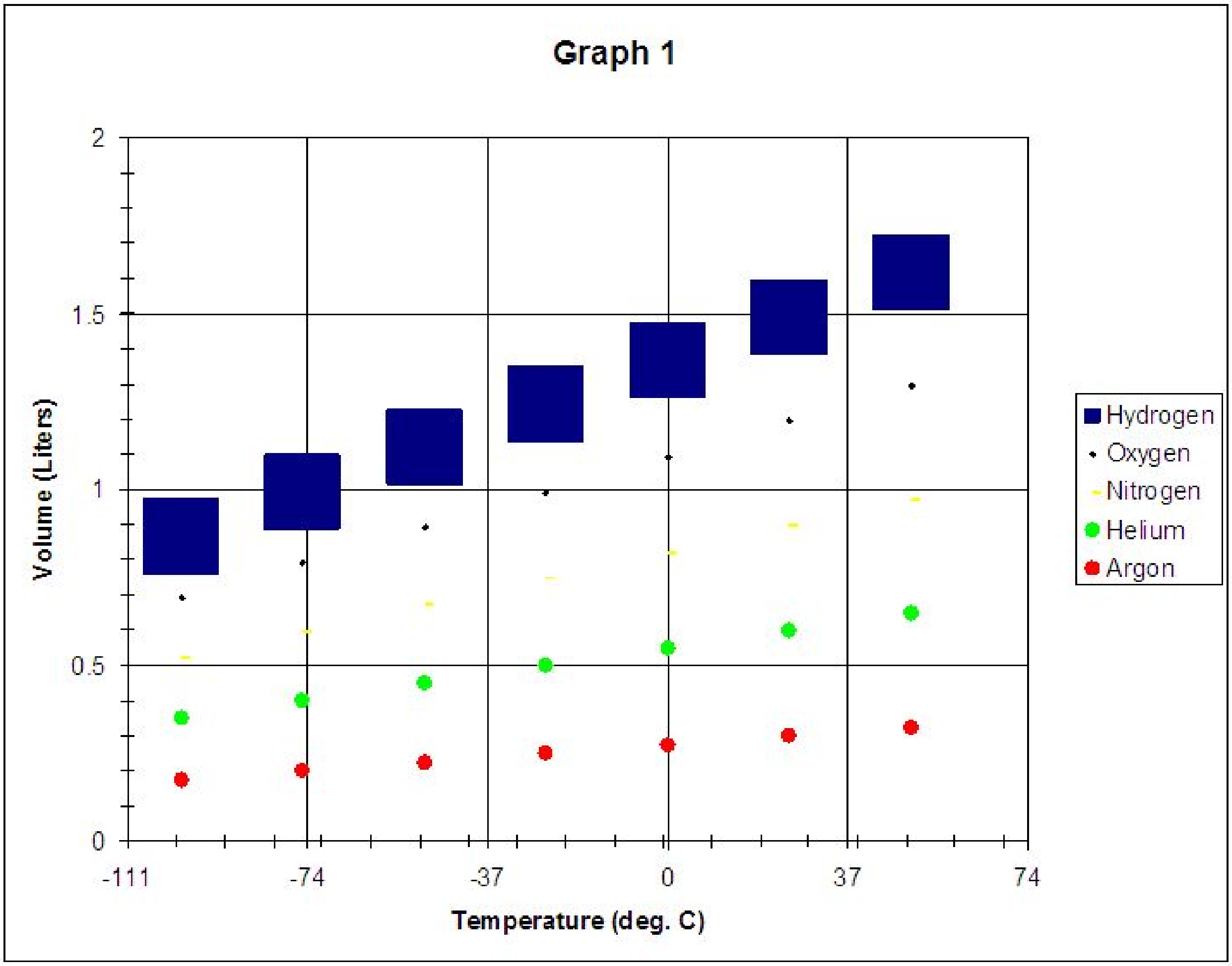


Figure 8.4: Another example of a bad graph. credit: author

Another example of a bad graph is shown in Figure 8.4.

1. Title is not descriptive
2. X axis uses odd spacing and makes no sense. The minor tick marks are not evenly divisible units of the major tick mark.
3. The symbols used for the data points are poor: Series 1 is too big. Series 2 is too small. Series 3 is a hard to see color, (yellow is most often a poor color choice) and the symbol "-" is almost impossible to see on the graph. Series 4 and 5 are the same symbol but different colors. It looks good on the screen but if you print it in black and white you will not be able to tell the difference between the symbols

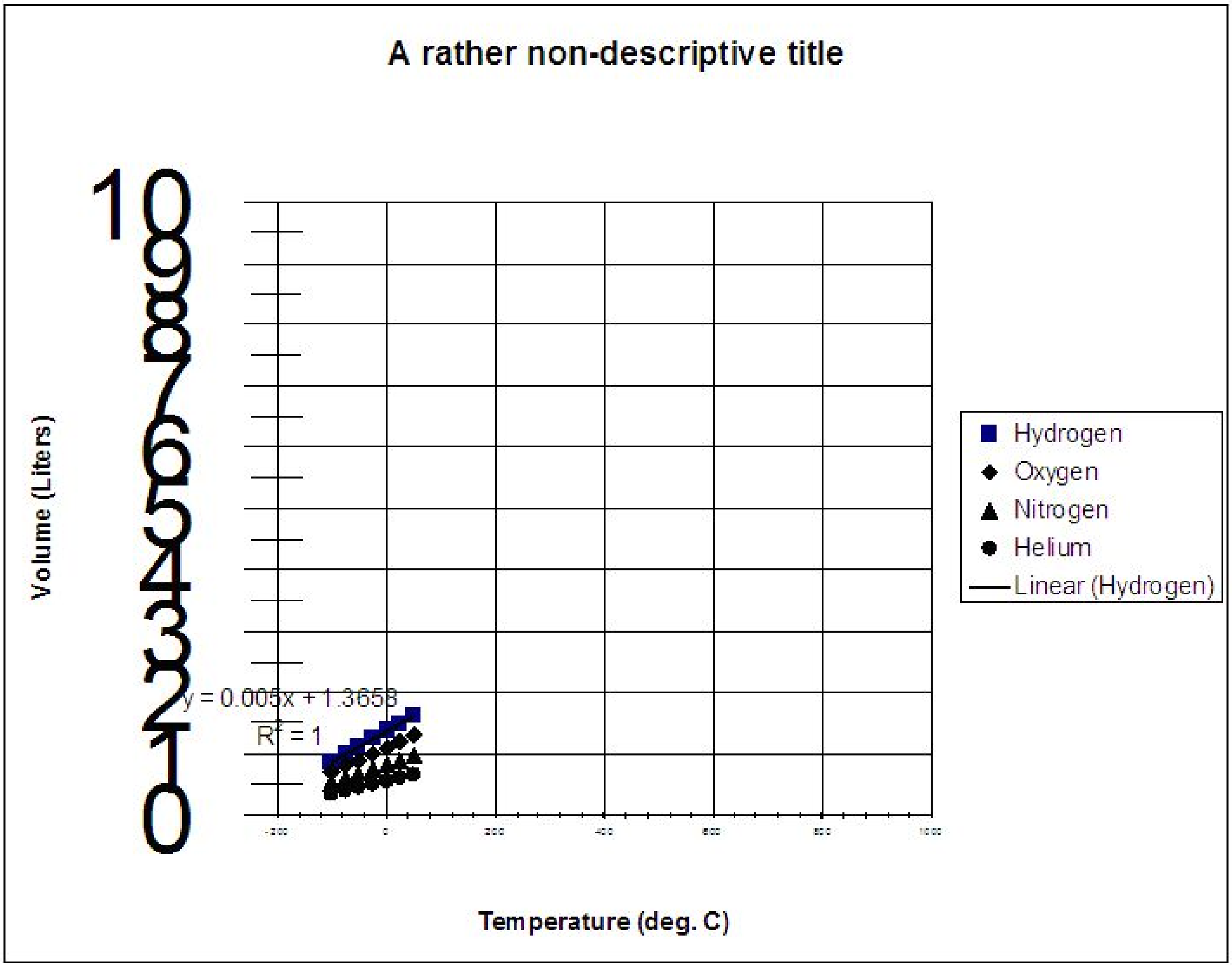


Figure 8.5: Another example of a bad graph. credit: author

One last example of a bad graph is shown in Figure 8.5.

1. The range of the X and Y axis is too large. The data does not span the range selected and ends up in only one small corner of your graph. This leaves a lot of empty unused space which is a waste.
2. The labels for the X axis are too small to read. The font size should be increased.
3. The labels for the Y axis are unnecessarily large and look out of place. The font size should be decreased.
4. The title is rather poor and non-descriptive.
5. The curve fit information overlaps the axis and is hard to read.

C. Graphing with a Computer

Increases in computing power in the last 20 years has lead to the practice of using computers to graph data. This generally speeds up the process. While much of the process is now automated, computers are not yet smarter than humans, and quite often make poor choices. The user needs to make corrections to graphs where needed. In this class we will be using Microsoft Excel to plot our data. It is not the best tool for the job (there are better professional software programs), however, it is sufficient for our purposes and is installed on all campus computers. In the next section we will complete a simple graph using Excel, followed by a general overview of the menu systems for modifying graphs, and some information on how to use the Curve Fitting features of Excel to draw lines through your data points.

D. Problem 1 - Computer Graph 1

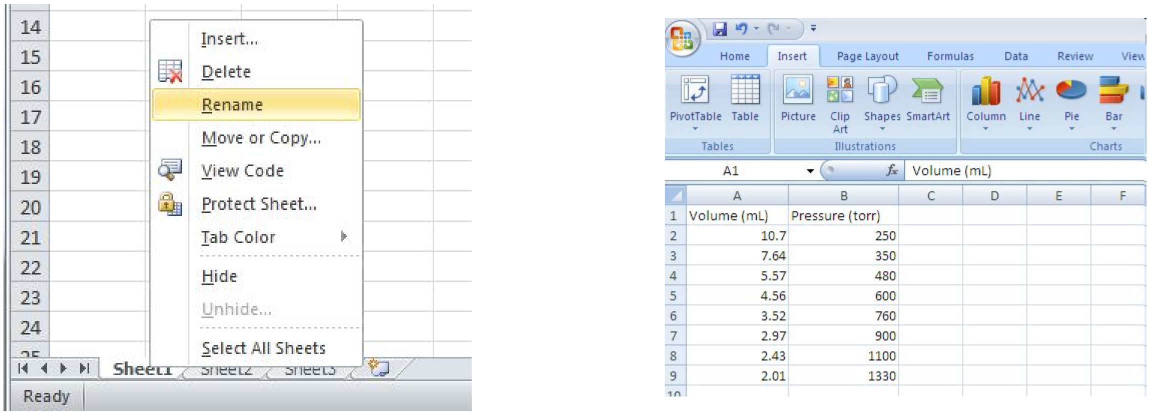
Using the following instructions we will now walk through how to make a graph using Excel using the data given below.

Volume (mL)	Pressure (torr)
10.70	250.
7.64	350.
5.57	480.
4.56	600.
3.52	760.
2.97	900.
2.43	1100
2.01	1330

Table 8.1: Data for Practice Computer Graph 1.

Procedure

1. Open the Excel program. In Excel rename the first tab (Sheet 1) to Problem 1. To do this right click on the Sheet 1 tab and choose the rename option. (Figure 8.6a).



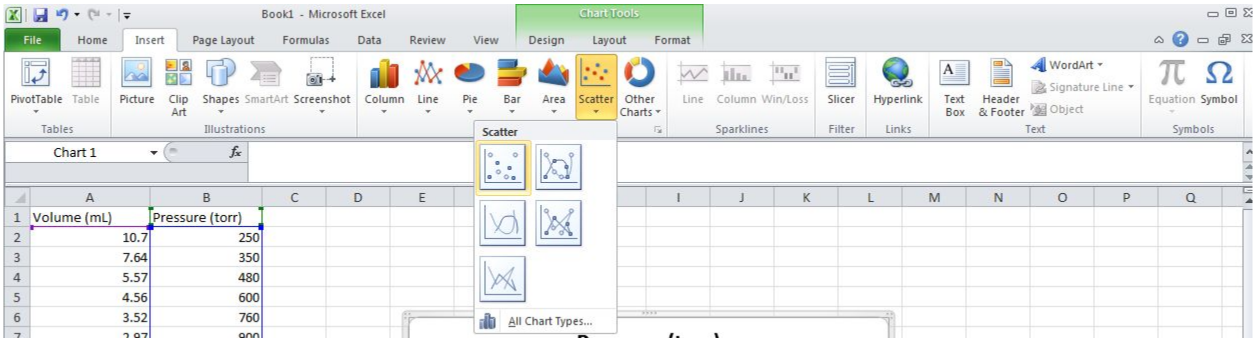
	A	B
1	Volume (mL)	Pressure (torr)
2	10.7	250
3	7.64	350
4	5.57	480
5	4.56	600
6	3.52	760
7	2.97	900
8	2.43	1100
9	2.01	1330

Figure 8.6: (a) Opening Excel (b) Entering data in Excel. credit: author

2. Add the data in Table 8.1 to the sheet. Go to Column A on the new Excel sheet that was opened and in cell A1 type "Volume (mL)" and in B1 type "Pressure (torr)". Adjust the width of the columns

if necessary. Now enter the data shown in the table. When finished entering the data highlight the cells that contain your data, the results should look similar to Figure 8.6.

3. Create a graph by choosing the **Insert** tab and then **Chart** from the ribbon menu. A variety of graph types are available to choose from, we will generally use the XY (Scatter) chart type with the first sub-type shown (default choice).



	A	B
1	Volume (mL)	Pressure (torr)
2	10.7	250
3	7.64	350
4	5.57	480
5	4.56	600
6	3.52	760

Figure 8.7: Chart Tool - Choose Scatter Plot. credit: author

4. After choosing the scatter plot option, you should be presented with a graph similar to that shown in Figure 8.8. To make the graph easier to work with, you may wish to enlarge the graph by clicking on the corners and dragging.

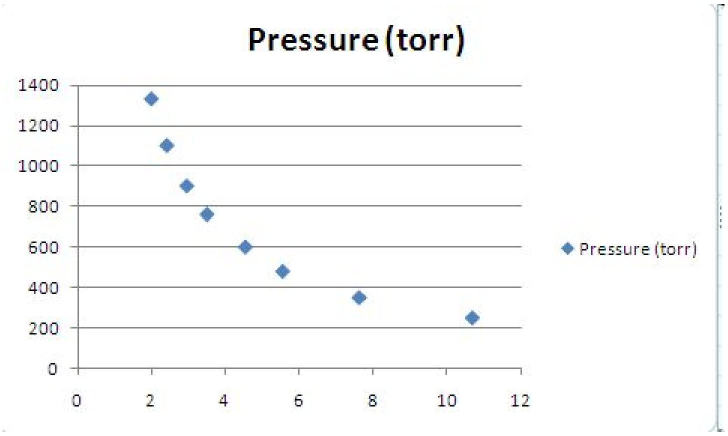
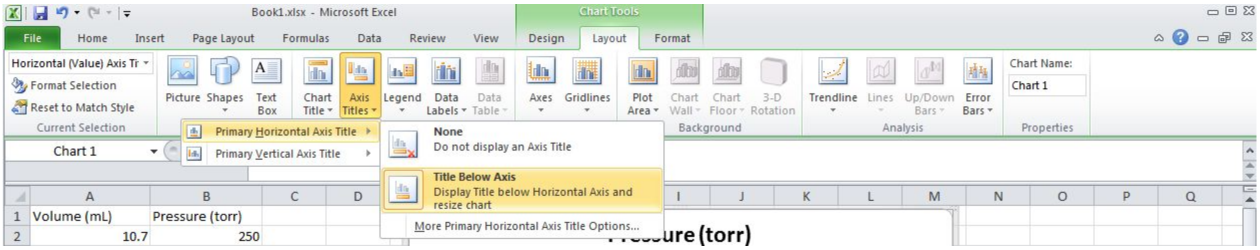


Figure 8.8: Typical Scatter plot generated. credit: author

5. Chart Tools: The current version of Excel uses a tabbed/ribbon menu system has placed many useful tools for formatting your graph in 3 locations as shown below in Figure 8.9 which shows the Layout tools and has the Axis Titles sub menu highlighted.



	A	B
1	Volume (mL)	Pressure (torr)
2	10.7	250

Figure 8.9: Scatter plot with new x and y axis marks. credit: author

   (a) Chart Design Tab: This tab contains many options for formatting the data used in the plot, and some layout features such as changing the shape and size of the data points on the graph.

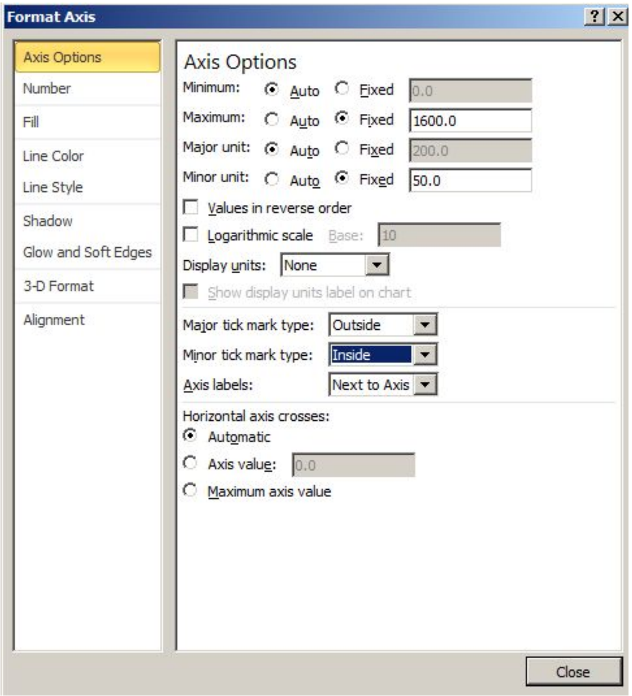
   (b) Chart Layout Tab: This tab contains many options for formatting specific features of your chart. You can add/change the Chart Title, Axis Titles, Legends, Data Labels, Grid Lines, Trendlines and many other features.

   (c) Chart Format Tab: This tab contains many options for formatting colors and fonts used in your chart. You can change the color, font, shape, alignment and add special effects to your graph.

6. There is no general rule for bringing up menu's in Excel (which is one reason it can be frustrating). In general double clicking on a feature will bring up the default menu, while right clicking (and holding) will bring up a list with many options. To master the use of Excel requires practice and patience.

7. There are several different menu's for changing the major chart options, formatting the axis, formatting data series and many other options. You can access these options by left-clicking on the desired feature and then right-clicking to bring up an options menu for that feature. For example, left-clicking on labels for the y-axis will bring up the format axis menu which allows you to set minimum and maximum values, spacings, tick marks and many other options. (See Figure 8.10.)

   While you are here you may as well change the maximum value on the y-axis to 1600, the minor unit to 50.0, and change the drop down menu for minor tick marks to show inside the graph.



Axis Options

Minimum: ● Auto ○ Fixed 0.0

Maximum: ○ Auto ● Fixed 1600.0

Major unit: ● Auto ○ Fixed 200.0

Minor unit: ○ Auto ● Fixed 50.0

☐ Values in reverse order

☐ Logarithmic scale Base: 10

Display units: None

☐ Show display units label on chart

Major tick mark type: Outside

Minor tick mark type: Inside

Axis labels: Next to Axis

Horizontal axis crosses:

● Automatic

○ Axis value: 0.0

○ Maximum axis value

Figure 8.10: The Format Axis Menu. credit: author

8. Now change the x-axis values (using what you have just learned) to label the axis every 2 units with a minor tick mark every 0.5 units. Change the labels on the axis to reflect the precision of your data (to the hundredths). This can be accomplished by changing the number format in your graph from "general" to "number" and choosing to display the numbers to 2 decimal places. The resulting graph should now look similar to what is shown below in Figure 8.11.

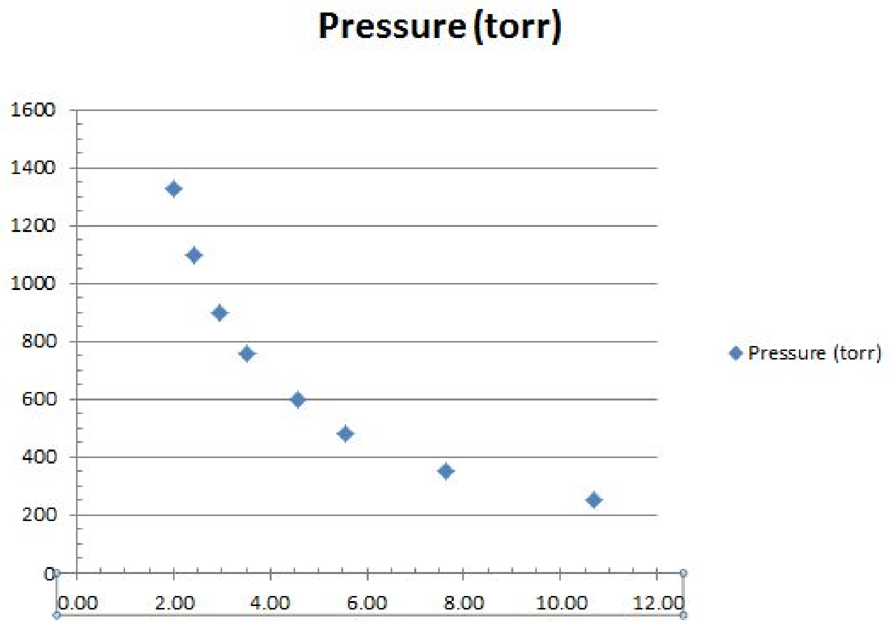


Figure 8.11: Scatter plot with new x and y axis marks. credit author

9. The next several items are some of the more important options you can change. Explore! It is the only way to really learn to properly use Excel to graph data.

10. Change the shape of your data points to be open circles with size = 8.

11. Add Axis titles to your graph. Label the x-axis Volume (mL) and the y-axis Pressure (torr).

12. Add a Chart title to your graph. Name the Graph: Computer Graph 1 - Volume vs. Pressure.

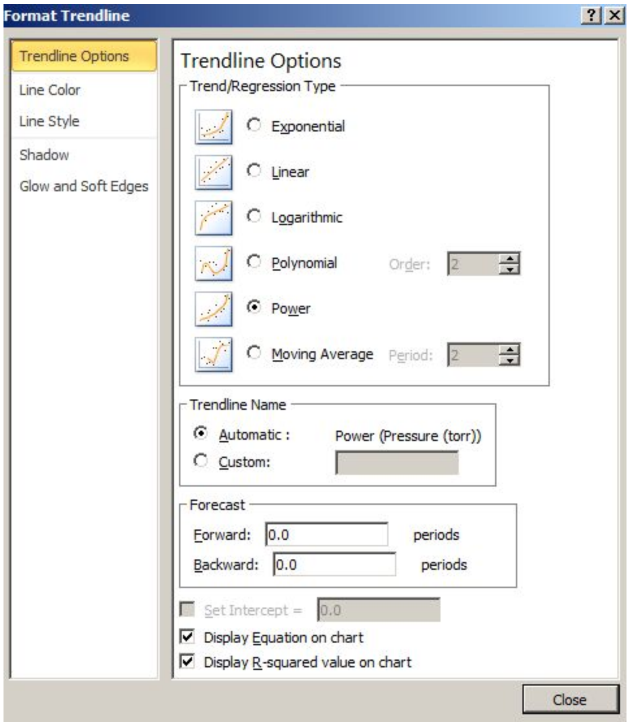
13. Remove the minor and major grid lines.

14. Add a border line around the entire graph.

15. Move the Legend inside the graph (it is a waste of space to leave it outside the graph).

16. Resize your graph to make maximum use of the space on the page.

17. Curve Fitting/Trend Lines: It is often desirable to either draw a line though your data to make it easier to follow separate data sets, or more importantly to fit a mathematical equation to your data in order to extract experimental information. The most common equation to fit to is that of a straight line with the formula $y = mx + b$ where m = slope and b = intercept. It is also useful to print the R^2 value of the equation, because it tells you how good of a curve fit it is, 1 = perfect correlation, 0 = no correlation. An R^2 value greater then 0.95 is generally considered good, anything lower indicates either poor data or the equation used to fit the line is not the correct one.
    (a) Open the **Add Trend line** menu (Figure 8.12) by selecting a data point on the graph and right clicking or chosing the Trendline option on the Chart Tools Layout menu.
    (b) Under the **Type** tab select **Power Series**.
    (c) Under the **Options** tab select the boxes to display the equation on the chart, and to display the R-squared value.
    (d) Move the equation and R-squared value to nice location that does not overlap anything.



○ Exponential
○ Linear
○ Logarithmic
○ Polynomial Order: 2
● Power
○ Moving Average Period: 2

● Automatic : Power (Pressure (torr))
○ Custom:

Forward: 0.0 periods
Backward: 0.0 periods

☐ Set Intercept = 0.0
☑ Display Equation on chart
☑ Display R-squared value on chart

Figure 8.12: Adding a curve fit (trend line). credit: author

18. Add your name and date to the upper right hand corner of your graph. You can use the **Insert tab or the Chart Tools Layout tab** which are selectable from the top menu. Select the **Add Text box** option.
19. Take one last look at your graph. It should look similar to the one shown in Figure 8.13. Are there any more changes you need to make? If so make them.
20. You should be done with your graph now. Show it to your instructor, and print a copy to be turned in with this lab.

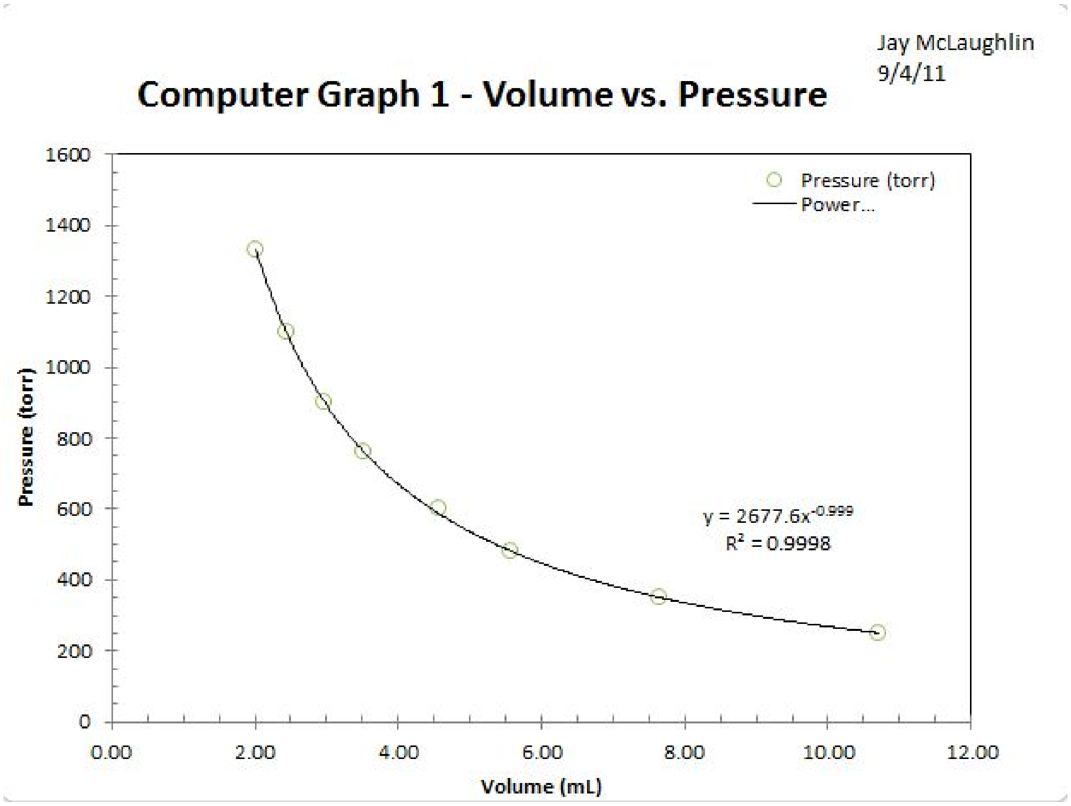


Figure 8.13: Completed computer graph. credit: author

E. Problem 2: Computer Graph 2

Using data provided in Table 8.2 you will make a graph following the instructions below. The data shows the temperature effect on the solubility (how many grams will dissolve in 100. g of water) of four salts. The temperature (x-axis) is measured in °C and the solubility (y-axis)in grams substance/100 g water.

1. Type the data set into Excel.
2. The degree symbol for °C is found on the top tab under **Insert** than the **Symbols** menu. Scroll to the top and click down three rows, it is located on the right side.
3. Using the mouse highlight the entire data set.
4. Choose the XY (scatter) plot and Click Next.
5. Click the Grid lines tab and add vertical grid lines. Click Next.

Now its time to make the graph look better. Make the following changes to your graph:

1. Change scales on axis. You may need to turn off the "auto" feature in Excel in order to set specific values.
    (a) Change the scale on the X axis to go from 0 to 100 with labels every 10 °C. Add minor tick marks every 5 °C.

Temperature (℃)	KCl (g)	NaCl (g)	KBr (g)	$BaCl_2$ (g)
0.0	27.6	35.7	53.5	31.6
10.	31.0	35.8	59.5	33.3
20.	34.0	36.0	65.2	35.7
30.	37.0	36.3	70.6	38.2
40.	40.0	36.6	75.5	40.7
50.	42.6	37.0	80.2	43.6
60.	45.5	37.3	85.5	46.6
70.	48.3	37.8	90.0	49.4
80.	51.1	38.4	95.0	52.6
90.	54.0	39.0	99.2	55.7
100.	55.6	39.8	104	58.8

Table 8.2: Data for Problem 4

   (b) Change the scale on the Y axis to go from 0 to 110 with labels every 10 grams and minor tick marks every 2 grams.

2. Remove the grid lines in the X and Y directions. (Hold Right Click choose Chart Options - Grid lines tab)

3. Add a border to your graph. (Hold Right Click choose Format Plot Area).

4. Change the data symbols to all be black and white and size 8. Change the shape so that □ = KCl, ◯ = NaCl , △ = KBr and •= $BaCl_2$.

5. Use the **Insert Tools** tab and then **shapes** draw an arrow pointing to the solubility of NaCl at 20 ℃ and use the Text Box tool to add a text at the end of the arrow saying "Solubility of NaCl at 20 ℃".

6. Curve Fitting: Right click on the KBr data and choose the "Add Trend line" option. Under the "Type" tab choose Linear. Under the "Options" tab click the check box's for Display Equation and Display R Value. Draw an arrow from the equation to the trend line.

7. Move the Legend to inside your graph in the lower right hand corner. Now drag your graph so it takes up most of the space in the graphing window. You always want to maximize the size of your graph so that it is easy to see and read your data.

8. Compare this graph to your first graph and make any changes required to make it a "good" graph.

F. Problem 3 Computer Graph 3

Using the data in the Table 8.3 make a presentable graph from the data. The data you have been given is experimental data for the Vapor Pressure of Water at Various Temperatures. The temperature is given in ℃ and the vapor pressure is measured in torr. Make sure you add a trend line to the graph using a 3rd order polynomial and make sure to include the equation and the R^2 value on the graph.

When you are done making your graph, email your instructor your Excel file with your name in it to `jay.mclaughlin@cncc.edu`. If you don't want to email a copy you may print a copy to hand in to your instructor.

Temperature (℃)	Vapor Pressure (torr)
0	4.6
5	6.5
10	9.2
15	12.8
16	13.6
17	14.5
18	15.5
19	16.5
20	17.5
21	18.6
22	19.8
23	21.2
24	22.4
25	23.8
26	25.2
27	26.7
28	28.3
29	30.0
30	31.8
40	55.3
50	92.5
60	149.4
70	233.7
80	355.7
90	525.8
100	760.0
110	1074.6

Table 8.3: Data for Problem 5

Experiment 9
Specific Heat

Jay C. McLaughlin
Colorado Northwestern Community College

- August 19, 2021

Name:

Date:

Key Objectives

1. Understand the equation $q = ms\Delta$T, what each variable represents and typical units.
2. Solve problems using $q = ms\Delta$T.
3. Difference between heat and temperature.
4. Use of Significant Figures in calculations.

Discussion

Calorimetry is the science of measuring heat. We will use a calorimeter to measure the amount of heat transferred from an unknown metal to water. By knowing the value of the specific heat of water we will then be able to calculate the specific heat of the unknown metal and identify it based on a comparison to known values. In addition we will be able to work on our graphing skills. A basic discussion of calorimetry can be found in Hein Ch. 4.5 (p. 71) and in McMurry Ch. 8.8 (p. 311).

The **Temperature** of an object is a measure of how rapidly the atoms composing an object are moving. The faster the atoms move, the higher the temperature. Temperature is an intensive property that does not depend on the amount of matter present, thus the temperature of a cup of coffee and a bathtub full of water can be the same. Temperature is generally measured in degrees Celsius (℃), Fahrenheit (℉), or Kelvin (K).

Heat is defined as the transfer of thermal energy, which is associated with the random motion of atoms and molecules, between two bodies that are at different temperatures. Heat is an extensive property that does depend on the amount of matter present, thus the heat content of cup of water is very different from the heat content of a bathtub full of water. Heat is generally measured in energy units as joules (J) or calories (cal) where 4.184 joules = 1 calorie. The calorie is defined as the amount of energy required to heat 1 gram of water 1 ℃.

The **Specific heat** of a substance is an intensive property that relates the temperature and heat of an object. It is defined as the quantity of heat required to raise 1 gram of a substance by 1 ℃ and has units of (J/g·℃, cal/g·℃, or J/mol·℃). The relationship between heat (q), specific heat (s), mass (m) and a temperature change (Δt or $t_f - t_i$) can be expressed mathematically as:

$$q = m \cdot s \cdot \Delta T \tag{4}$$

The change in temperature of a substance is often expressed using the greek letter delta (Δ) “which means the change in”. This can also be expressed as the final temperature (t_f) minus the initial (t_i) of an object. Knowing 3 of the variables in the equation above allows us to solve for the fourth variable. A fact we will make use of in this experiment.

Based on our definitions of specific heat and joules/calories, the specific heat of water is exactly 4.184 J/g·℃ or 1.00 cal/g·℃. Knowing this value we will be able to determine the specific heat of a known metal and an unknown metal. We will do so by heating the metal to a known temperature and then placing it in contact with water at a known temperature. The heat from the metal will be transferred to the water raising the temperature. Using the final measured temperature of the water we will be able to determine the amount of heat transferred from the metal to the water and thus calculate the specific heat of the metal. The following equations show the relationships that allow us to determine the specific heat of the metal. Variables with a subscript "w" refer to the water while a subscript "m" will refer to our metal.

$$q_w = m_w \cdot s_w \cdot (\Delta t_w) \tag{5}$$

$$q_m = m_m \cdot s_m \cdot (\Delta t_m) \tag{6}$$

We can combine the above equations into a single equation by realizing that all of the heat that the water absorbed had to come from the metal sample, thus:

$$q_w = q_m \tag{7}$$

$$m_w \cdot s_w \cdot (\Delta t_w) = m_m \cdot s_m \cdot (\Delta t_m) \tag{8}$$

Figure 9.1 shows the rough setup for the experiment and Table 9.1 shows the values measured in the experiment, because we know all the values for water, we can calculate q_w and using equation 7 we then know q_m and can calculate the specific heat of the metal.

Variable	**Water**	**Metal**
m	measured	measured
s	measured	??
t_f	same value	
t_i	measured	measured
q	same value	

Table 9.1: Values measured in the experiment.

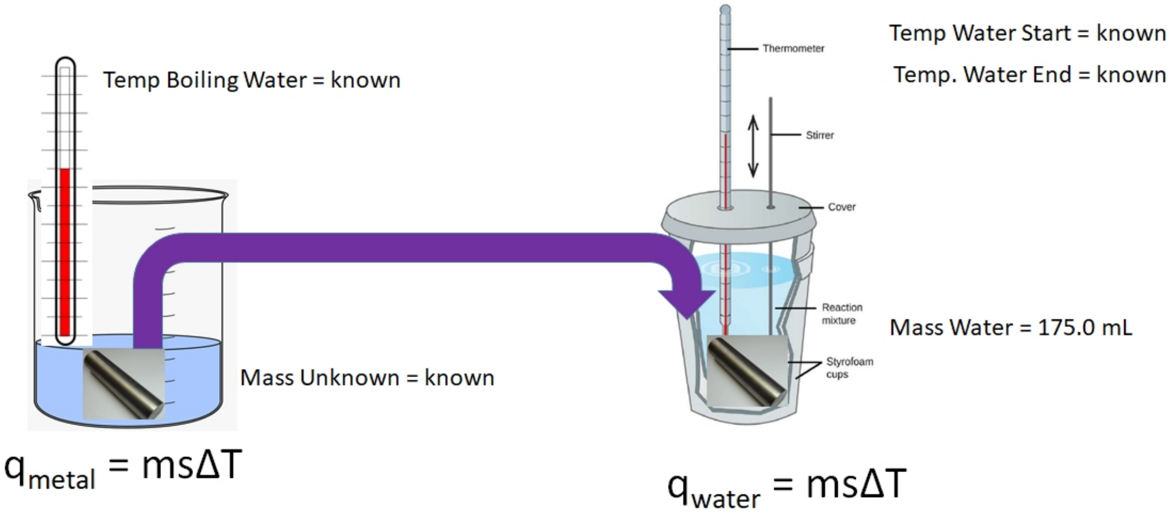


Figure 9.1: Experimental setup for finding the specific heat of an unknown metal. credit: author

Procedure

Read the entire procedure before beginning the experiment. Day one you will use a known sample (most likely Copper (Cu) to test the accuracy and precision of the measurement. Day two you will repeat the experiment on on unknown sample and calculate its specific heat and using Table 9.2 try to determine its identity.

1. Be sure to record all measured values with the appropriate number of significant figures and unit. Be sure to show all calculations required.

2. Obtain a known metal sample (most likely Copper (Cu)) from your instructor and record the name of the known in the results section.

3. Weigh the known metal sample. Record it in the results section.

4. Construct the apparatus shown in Figure 9.2a. Place the metal sample carefully in the beaker (do not drop it, you might break the beaker) before placing the thermometer in. Use the smaller thermometer and be sure to clamp the thermometer in place. Adjust the height above the flame to ensure rapid heating.

5. Fill the beaker with tap water so the height of the water is a least 2 inches higher than the top of the metal sample.

6. Begin heating the water in the beaker and continue with the next step(s). As you are working, check the water frequently and note when it starts to boil. Turn down the burner, but keep the water gently boiling. Do not do step 11 until the water has been boiling for about 10 minutes and the temperature of the metal object has stabilized.

7. Obtain a calorimeter (Figure 9.2b) and record the mass of the inner cylinder on your data sheet.

8. Add approximately 50 mL of water to the calorimeter (enough to cover the metal sample) by no

more than 1/2 inch. Use the minimum amount of water to cover the sample so the measured change in temperature will be as large as possible. Record the mass of the calorimeter + the water.

9. Use a cardboard cover to trap heat and slide the thermometer through it, clamp a large thermometer carefully and make sure the water will cover the mercury bulb. Adjust it so that it hangs suspended just above the bottom of the calorimeter, but will be covered by the water. Clamp it to a ring stand to prevent it from falling over.
10. Measure the initial temperature of the water in the calorimeter. Use a second ring stand to secure the thermometer in place.
11. After the water in the beaker has been boiling for 10 minutes and the temperature of the metal sample has been stable for 5 minutes record the temperature in the results section. Remove the thermometer from the test tube and set it aside.
12. Using tongs, transfer the metal sample from the boiling water to the calorimeter. It is important that the transfer take place quickly and that no water is transferred along with the metal. Be sure to remove the thermometer from the calorimeter and set it carefully aside so you don't break it when you place the metal sample in the calorimeter. Try to minimize splashing. Be sure not to confuse the two thermometers.
13. Once the metal sample is in the calorimeter, quickly place the cover on and the thermometer in the water. Stir gently by swirling the calorimeter. Watch the temperature of the water in the calorimeter, and note the peak temperature. If the temperature does not reach a maximum value and decrease then record the temperature when it has been stable for one minute.
14. Calculate the specific heat of your unknown sample. Compare the sample to the known value. If you are within 0.05 J/g·℃ of the correct value continue to the next step. If you were not accurate enough on the first trial of your known metal consult your instructor and perform a second trial.
15. Obtain an unknown sample and determine its specific heat. Do 2-3 trials (as time permits) on the unknown sample.
16. Ask your instructor for permission to start.

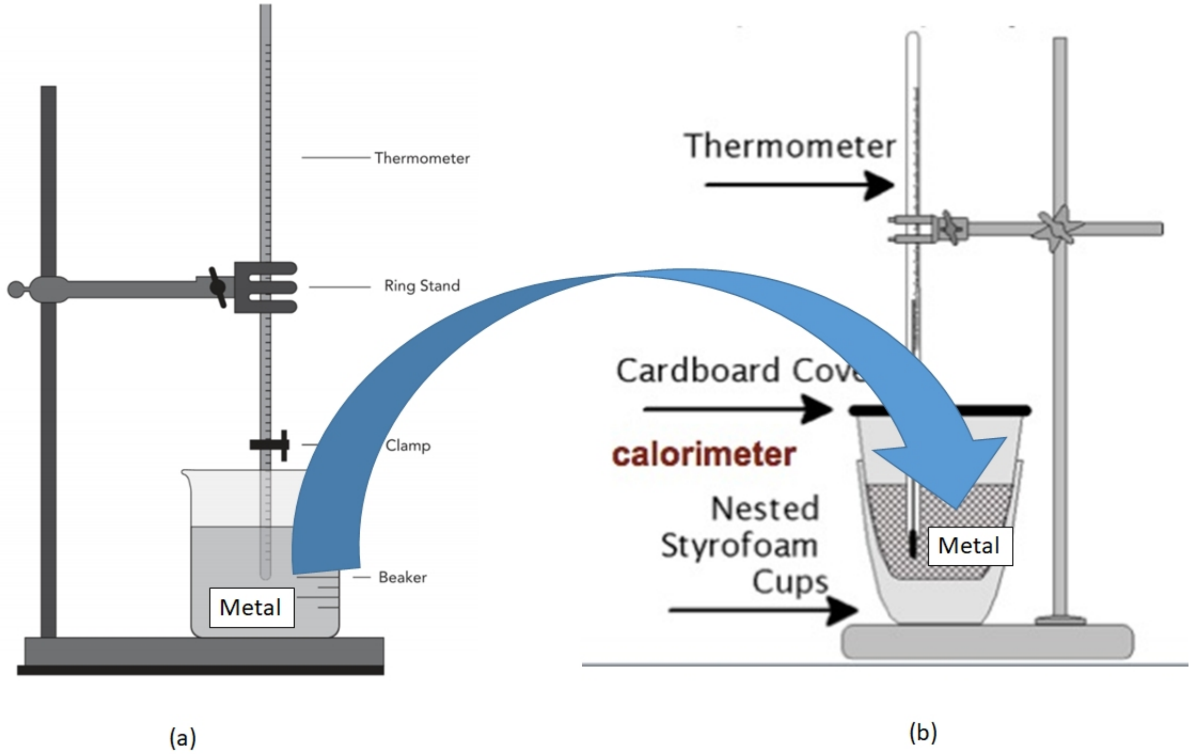


Figure 9.2: (a) Heat the metal in a beaker then transfer the metal to the (b) calorimeter and measure the temperature change. credit: (a) `https://seedsconnections.org` (b) `https://chemdemos.uoregon.edu/demos/Comparing-Specific-Heats-of-Metals`

Metal	Atomic Mass (amu)	Specific Heat (J/g·℃)
Aluminum	26.98	0.900
Brass	N/A	0.380
Copper	63.55	0.385
Gold	197.0	0.131
Iron	55.85	0.451
Lead	207.2	0.128
Magnesium	24.305	1.02
Silver	107.9	0.237
Tin	118.7	0.222
Zinc	65.39	0.390

Table 9.2: Specific Heat of Selected Metals

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Results

A. Known Metal Sample - Data

	Trial 1	**Trial 2**	**Trial 3**
1. Identity of known metal	___________	___________	___________
2. Uncertainty in mass measurements:	___________	___________	___________
3. Mass of metal sample:	___________	___________	___________
4. Mass of calorimeter:	___________	___________	___________
5. Mass of calorimeter + water:	___________	___________	___________
6. Uncertainty in temperature measurements (heating appratus):	___________	___________	___________
7. Uncertainty in temperature measurements (calorimeter):	___________	___________	___________
8. Initial temperature of water in calorimeter:	___________	___________	___________
9. Initial temperature of metal sample:	___________	___________	___________
10. Final temperature of water and metal:	___________	___________	___________

B. Known Metal Sample - Calculations

Show example calculations for Trial 1 in the space provided below.

	Trial 1	Trial 2	Trial 3
1. Mass of Water:	____	____	____
2. Change in temperature of water:	____	____	____
3. Change in temperature of metal:	____	____	____
4. Heat gained by water:	____	____	____
5. Heat lost by metal:	____	____	____
6. Specific Heat of Metal (Experimental):	____	____	____
7. Average Specific Heat:		____	
8. Theoretical specific heat of metal:		____	
9. Calculate the Percent Error using the following formula: $\left\lvert\frac{\text{Actual Value}-\text{Measured Value}}{\text{ActualValue}}\right\rvert * 100\ \%$	____	____	____

C. Unknown Metal Sample - Data and Results

	Trial 1	Trial 2	Trial 3
1. Identity of unknown metal	________	________	________
2. Mass of metal sample:	________	________	________
3. Mass of calorimeter:	________	________	________
4. Mass of calorimeter + water:	________	________	________
5. Initial temperature of water in calorimeter:	________	________	________
6. Initial temperature of metal sample:	________	________	________
7. Final temperature of water and metal:	________	________	________
8. Mass of Water:	________	________	________
9. Change in temperature of water:	________	________	________
10. Change in temperature of metal:	________	________	________
11. Heat gained by water:	________	________	________
12. Heat lost by metal:	________	________	________
13. Specific Heat of Metal (Experimental):	________	________	________
14. Average Specific Heat:		________	
15. Which element is your metal made from:		________	
16. Explain why you chose the above metal (as opposed to other possibilites):			
17. Theoretical specific heat of metal:	________	________	________
18. Calculate the Percent Error using the following formula: $\left\lvert\frac{\text{Actual Value}-\text{Measured Value}}{\text{ActualValue}}\right\rvert * 100\ \%$	________	________	________

Post Lab Questions

1. Why is it more accurate to weigh the mass of the calorimeter and the calorimeter + water instead of simply measuring the mass of water in a beaker and pouring it into the calorimeter?

2. Why do we want the water in the calorimeter to barely cover the metal sample (but not entirely fill up the calorimeter)?

3. Why do we want to be careful not to mix up which thermometer we use to measure the metal sample and which we use to measure the temperature in the calorimeter?

4. What effect does the initial temperature of the water in the calorimeter have on the change in temperature of the water after the hot metal sample is added, consider the following two scenario's.

   4(a) For instance what if you used water at 5 ℃ in your calorimeter instead of room temperature water? Explain.

   4(b) What if a mixture of ice and water was used? Explain.

5. Based on **YOUR** experimental data, which was better, the precision or the accuracy? Explain.

6. In this experiment you measured the specific heat of your known and unknown three times. Why did you do this?

7. Bob heats 25.00 grams of an unknown metal in a calorimeter from 18.0℃ to 23.0℃ using 500.0 J of heat. What is the specific heat of the metal?

7. ______________

8. Bob's brother heated up 2.5×10^3 L of Sulfuric Acid from 15.00℃ to 39.0℃. How many nJ of heat did he use? The specific heat of Sulfuric Acid is 2.59 $\mathrm{J\,/\,g}$ ℃.

8. ______________

9. Bob's sister uses 2.45×10^4 J of energy to heat a copper pan weighting 250 grams. If the final temperature of the pan was 600.℃ what was the initial temperature of the pan?

9. ______________

10. Bob's father heated a metal bar weighting 50.0 grams to 200.0℃ and placed it in a calorimeter containing 75.00 grams of water at an initial temperature of 30.0 ℃. If the final temperature of the metal object in the water was 45.0℃ what is the specific heat of the metal?

10. ______________

Name: ______________________ Class: __________ Date: __________

Prelab Questions

1. Why are we measuring the specific heat of a known compound (Cu - Copper) before we measure it for an unknown compound?

2. How much energy (in kJ) is required to heat a cup of coffee (mostly water) weighing 0.50 lbs from 25 ℃to 75 ℃?

3. What is the final temperature of a gold sample weighing 10.5 grams which absorbs 814 J of energy if it starts at 20.0 ℃? Explain.

4. A metal sample weighing 32.250 g was heated to 100.0 ℃ , then quickly transferred to a calorimeter containing 53.247 g of water at 18.0 ℃. The temperature of the water increased to 20.6 ℃. Calculate the specific heat of the metal in J/g·℃. (Hint: where did all the heat lost by the metal go?)

This is the back of the page, totally uninteresting.

Experiment 10
Double Displacement Reactions

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Five signs a reaction occurred.
2. Write complete Double Displacement chemical reactions
   (a) Balance charges.
   (b) Balance atoms.
   (c) Include states when known.
3. Recognize when no reaction (NR) occurs.
4. Use a solubility table to determine the state of compounds.

Discussion

Double displacement reactions are among the most common of the simple chemical reactions to study and understand. We will explore the driving forces behind the chemical reactions, and use observations made about individual reactions to write complete chemical equations.

In examining the characteristic chemical properties of several compounds we will recall that there are several readily apparent signs that a chemical reaction between two solutions has occurred. Those signs were:

1. Color changes, that are not the result of simple dilution.
2. Evolution of a gas forming many bubbles rapidly.
3. Formation of a precipitate, normally a solid which will settle to the bottom of a test tube, but occasionally a finely divided precipitate that does not settle and often appears as only a milky color.
4. Disappearance of a precipitate.
5. Endothermic or Exothermic reactions (creation of heat or cold).

Each of these signs will be used in this experiment to determine if a chemical reaction has occurred. For those experiments in which a reaction occurred we will write a balanced chemical reaction, including the states of products based on our observations.

Chemical reactions occur for a number of reasons. A detailed discussion is beyond the scope of this lab, however a brief discussion will provide insight into why each of the observable signs of a chemical reaction occur. **Enthalpy** is a measure of the energy flow in a reaction, while **Entropy** is a measure of the randomness or disorder in a system. Spontaneous chemical reactions seek to decrease the enthalpy of the products and increase the entropy of the products. One or both of these conditions must be met for the reaction to occur spontaneously. It is the interplay of these two mechanisms that will drive the chemical reactions we see.

When examining the physical states of the reactants and products entropy increase as we go from a solid to a liquid to a gas, because the motion (disorder) of the particles is increased. Enthalpy is more difficult to predict and will be covered in later lectures or experiments. As a general rule the formation small stable molecules (listed in Table 10.1 below or the formation of water (H_2O) is favorable and a reaction will occur. One can also use solubility tables to determine which solids are more stable and tend to form precipitates in solution.

The formation of precipitates can drive a reaction because the products are more stable and lower energy than the reactants. The production of gases is both an increase in the disorder in a system, and often results in the formation of small stable molecules. An exothermic reaction is one in which heat or energy is given off indicating that the products are lower in energy than the reactants. Color changes are a good sign that a chemical reaction has occurred, but are not directly tied to changes in enthalpy or entropy, and instead occur due to structural changes in how the atoms in the solution interact with light.

Identifying precipitates and gases

Recognizing precipitates and gases while performing a chemical reaction is very obvious, however when writing chemical reactions on paper, we need several aids to help us identify them.

There are several common gases it is worth memorizing the state of shown in Table 10.1

Common Gases
Elemental (11) - H_2, N_2, O_2, F_2 and Cl_2 + Noble Gases
Molecular (8) - CO, CO_2, NO, NO_2, N_2O_4, SO_2, SO_3, and NH_3

Table 10.1: Common gasses formed in chemical reactions.

Gases can also be formed by the decomposition of several common compounds formed in double displacement reactions. There are three common reactions that we will encounter and need to recognize given below. The formation of any of these compounds will result in the formation of a gas, indicating a chemical reaction has occurred.

$$\mathrm{H_2CO_3(d)} \longrightarrow \mathrm{H_2O(l) + CO_2(g)} \tag{9}$$

$$\mathrm{H_2SO_3(d)} \longrightarrow \mathrm{H_2O(l) + SO_2(g)} \tag{10}$$

$$\mathrm{NH_4OH(d)} \longrightarrow \mathrm{H_2O(l) + NH_3(g)} \tag{11}$$

Precipitates can be determined by consulting solubility tables. These are tables of common cations and anions arranged in columns and rows. The intersection of which gives the information about solubility. A soluble compound is one that dissolves in water, while insoluble compounds will not dissolve and thus form precipitates. A solubility table is included at the end of this experiment. Thus, if we examine the solubility of NaCl for example we note that it is labeled “(aq)” in the table indicating that no precipitate occurs, and thus no reaction is likely to have occurred. An example of when a precipitate occurs is the reaction of Ag + Cl to form AgCl which is labeled as “I” indicating that it would form a precipitate and a reaction would occur.

A. Exothermic and Endothermic Reactions

Recognizing temperature while performing a chemical reaction is very obvious, however when writing chemical reactions on paper, we need several aids to help us identify them. In this experiment we will only examine exothermic reactions, the classic example being the reaction of an acid and a base. The general reaction is shown below.

$$\text{Acid} + \text{Base} \longrightarrow \text{Ionic Salt} + \text{H}_2\text{O(l)} + \text{Heat} \qquad (12)$$

Table 10.2 lists the most common acids and bases. In order to write the chemical reactions it will be necessary to recognize them in chemical reactions.

The heat in an acid/base reaction occurs because of the formation of a slightly ionized compound (in this case water). Thus, we can add an additional sign that a reaction has occurred which is whenever water is formed in a chemical reaction. The heat is produced because water is a highly stable molecule, and when formed is lower in energy than the reactants, resulting in the release of heat. Other examples of slightly ionized compounds formed in double displacement reactions are acetic acid ($HC_2H_3O_2$), oxalic acid ($H_2C_2O_4$), and phosphoric acid (H_3PO_4).

Acids (start with H)	HF, HCl, HBr, HI, H_3PO_4, H_2SO_4, HNO_3, H_2CO_3, $HC_2H_3O_2$, $HClO_4$
Bases (contain OH)	NaOH, KOH, $Ba(OH)_2$, $Ca(OH)_2$, NH_4OH

Table 10.2: Common Acids and Bases.

Double Displacement Reactions

In the double displacement reactions in this experiment two aqueous solutions each containing an ionic compound will be mixed. The general reaction is given by:

$$\text{AB} + \text{CD} \longrightarrow \text{AD} + \text{CB} \qquad (13)$$

Each ionic compound (AB and CD) when dissolved in water will be present as a cation and anion pair. This implies that when the solutions are mixed we will have a mixture containing A^+, B^-, C^+, and D^- cations and anions. As the ions interact with each other there are six possible combinations of ions that could result in a reaction. The six outcomes are shown in Table 10.3.

Double displacement reactions are named because the reaction can be viewed as one cation displacing the other cation from the molecule and forming a new compound. These reactions are also often referred to as Double Replacement reactions since we can view one cation as replacing the other to form a new compound. A third way of viewing these reactions is illustrated below in Figure 10.1 where we can imagine that the cations simply swap places and each of the new combinations has the potential to cause a reaction to occur.

Combination	Result
A^+B^- C^+D^-	The ions recombine into the starting reactants. This can not result in a reaction because no change occurs.
A^+C^+ B^-D^-	This can not result in a reaction because like charges repel each other. No exchange of electrons can occur to form a compound.
A^+D^- C^+B^-	This may result in a reaction if a precipitate or gas is formed.

Table 10.3: Possible outcomes of a double displacement reaction.

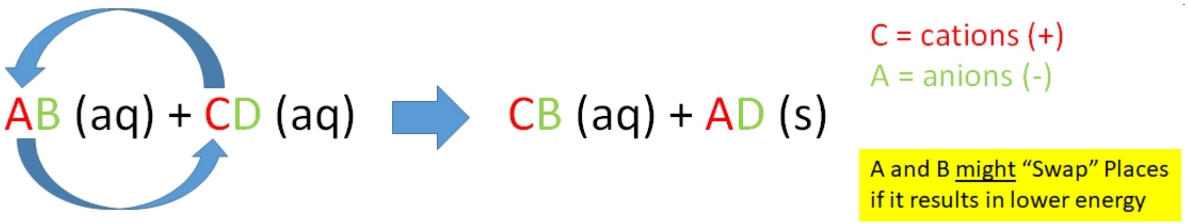

Figure 10.1: Double displacement reactions occur if the products formed by swapping cations are lower in energy than the reactants. credit: author

Example 1:

A solution of sodium chloride and lead (II) nitrate are mixed. A white precipitate is formed. Write the complete chemical reaction.

The precipitate is definite evidence that a chemical reaction occurred. We swap the sodium and lead cations to form new compounds, and checking the solubility tables we note that $PbCl_2$ is listed as (I), indicating it forms a solid precipitate, while $NaNO_3$ is listed as aqueous. We can then write:

$$\underline{\ 2\ }\mathrm{NaCl(aq)} + \underline{\ 1\ }\mathrm{Pb(NO_3)_2(aq)} \longrightarrow \underline{\ 1\ }\mathrm{PbCl_2(s)} + \underline{\ 2\ }\mathrm{NaNO_3(aq)} \tag{14}$$

Example 2:

A solution of sodium carbonate and sulfuric acid are mixed, resulting in the formation of a gas. Write the complete chemical reaction.

$$\underline{\ 1\ }\mathrm{Na_2CO_3(aq)} + \underline{\ 1\ }\mathrm{H_2SO_4(aq)} \longrightarrow \underline{\ 1\ }\mathrm{Na_2SO_4(aq)} + \underline{\ 1\ }\mathrm{H_2CO_3(aq)} \tag{15}$$

Examining the solubility tables indicates that neither product forms a precipitate. There is an acid in the reaction, however there is no base for it to react with, the experimenter did not note any temperature change. Gas was observed indicating that a chemical reaction must have occurred, and we must recall that H_2CO_3 readily decomposes into water and carbon dioxide. The final chemical reaction would then be:

$$\underline{\ 1\ }\mathrm{Na_2CO_3(aq)} + \underline{\ 1\ }\mathrm{H_2SO_4(aq)} \longrightarrow \underline{\ 1\ }\mathrm{Na_2SO_4(aq)} + \underline{\ 1\ }\mathrm{H_2O(l)} + \underline{\ 1\ }\mathrm{CO_2(g)} \tag{16}$$

Example 3:

A solution of magnesium hydroxide and hydrochloric acid are mixed, no visible sign of a reaction occurred but the test tube feels warm to the touch. The double displacement reaction is written as:

$$\underline{\ 1\ }\mathrm{Mg(OH)_2(aq)} + \underline{\ 2\ }\mathrm{HCl(aq)} \longrightarrow \underline{\ 1\ }\mathrm{MgCl_2(aq)} + \underline{\ 2\ }\mathrm{H_2O(l)} \tag{17}$$

Neither product is a precipitate or gas, however, we should recognize this as an acid/base reaction. Additionally we could note that water was formed in the reaction, indicating a reaction occurred. Both acid/base reactions and the formation of water result in the release of heat. We can write the final reaction then as:

$$\underline{\ 1\ }\mathrm{Mg(OH)_2(aq)} + \underline{\ 2\ }\mathrm{HCl(aq)} \longrightarrow \underline{\ 1\ }\mathrm{MgCl_2(a_1)} + \underline{\ 2\ }\mathrm{H_2O(l)} + \mathrm{Heat} \tag{18}$$

Example 4:

A solution of sodium chloride and potassium nitrate are mixed. Write the complete chemical reaction. No change is observed in the test tube.

$$\underline{\quad}\mathrm{NaCl(aq)} + \underline{\quad}\mathrm{KNO_3(aq)} \longrightarrow \underline{\quad}\mathrm{KCl(aq)} + \underline{\quad}\mathrm{NaNO_3(aq)} \tag{19}$$

We write the initial reaction by simply swapping the sodium and potassium cations. We then need to check the solubility tables for KCl and $NaNO_3$, from which we note both compounds form aqueous compounds. Neither compound is a gas or decomposes and the reactants are not acids and bases, thus we conclude that no reaction (NR) occurred and would write:

$$\underline{\quad}\mathrm{NaCl(aq)} + \underline{\quad}\mathrm{KCl(aq)} \longrightarrow \mathrm{NR} \tag{20}$$

Procedure

Obtain 12 medium sized test tubes and a test tube rack. Perform the following 12 reactions as described below. Each part of the experiment involves mixing equal volumes of solution (about 1 mL or 20 drop) in a standard sized test tube. If you are unsure of your observation you may add another milliliter of each solution. For each reaction record your observations at the time of mixing. Feel each tube to determine if heat was evolved. The following terminology should be used:

Observation	**Notation**
Precipitate	ppt - color
Gas	(g)
Heat	Heat
No reaction	NR

For each reaction write the complete chemical equation. Be sure to balance each product formed, and the overall chemical equation. Include the states of the reactants and products when known, if the state is not known leave empty bracket ( ). If the reaction is exothermic add “heat” as a product. If there is no evidence of a reaction write cross out the products and write “No Reaction” or “NR” as the products.

All reactions written should reflect observations (and vice versa), consult with your instructor about any discrepancies.

Perform the following twelve chemical reactions in any order.

1. Mix 0.1 M sodium chloride and 0.1 M potassium nitrate solutions.
2. Mix 0.1 M sodium chloride and 0.1 M silver nitrate solutions.
3. Mix 0.5 M sodium carbonate and concentrated (6 M) hydrochloric acid solutions.
4. Mix 20% sodium hydroxide and 6 M hydrochloric acid solutions.
5. Mix 0.1 M barium chloride and 9 M sulfuric acid solutions.
6. Mix 6 M ammonium hydroxide and 3 M sulfuric acid solutions.
7. Mix 0.1 M copper (II) sulfate and 0.1 M zinc nitrate solutions.
8. Mix 0.1 M sodium carbonate and 0.1 M calcium chloride solutions.
9. Mix 0.1 M copper (II) sulfate and 0.1 M ammonium chloride solutions.
10. Mix 20% sodium hydroxide and 6 M nitric acid solutions.
11. Mix 0.1 M iron (III) chloride and 6 M ammonium hydroxide solutions.
12. In the hood add 1 g of sodium sulfite to a test tube. Then add 2 mL of water and swirl vigorously to dissolve the solid (some small amount of solid may remain, this is acceptable). This produces an aqueous solution of sodium sulfite which we will consider as the reactant. Slowly (drop by drop) add 1 mL of 18 M sulfuric acid to the sodium sulfite solution.

☠ Dispose of mixtures 2, 5, 7, 9 in the “Experiment 6 Heavy Metal Waste” container. Dispose of all other mixtures in the sink and flush with water.

Name: ____________________ Class: ________________ Date: ________________

Results

For each reaction write the observed results from lab. Then write the complete chemical reaction. Be sure to balance each reactant and product compounds charges, and the overall chemical equation. Include the states of the reactants and products when known, if the state is not known leave empty bracket ( ). If the reaction is exothermic add "+ heat" as a product. If there is no evidence of a reaction cross out the products and write "No Reaction" or "NR" as the products. All reactions written should reflect observations (and vice versa), consult with your instructor about any discrepancies.

#	Observations	Complete Chemical Reaction
1.		
2.		
3.		
4.		
5.		
6.		
7.		
8.		
9.		
10.		
11.		
12.		

Table 10.4: Results

Post Lab Questions

1. Write the complete chemical reaction for the decomposition of sulfurous acid (H_2SO_3).

2. Using the criteria for double displacement reactions occurring, solubility tables, and the knowledge gained in this experiment predict whether a double displacement reaction will occur in each example below. If the reaction will occur complete and balance each compound, and equation, properly indicate the state of gases and precipitates, and indicate if the reaction is exothermic by including "heat" as a product. If no reaction will occur write "No reaction" or "NR" as the products. All reactants are assumed to be aqueous.

   (a) ___ $K_2S(aq) +$ ___ $CuSO_4(aq) \longrightarrow$

   (b) ___ $NH_4OH(aq) +$ ___ $H_2C_2O_4(aq) \longrightarrow$

   (c) ___ $KOH(aq) +$ ___ $NH_4Cl(aq) \longrightarrow$

   (d) ___ $Na_2CrO_4(aq) +$ ___ $Pb(C_2H_3O_2)_2(aq) \longrightarrow$

   (e) ___$(NH_4)_2SO_4(aq) +$ ___$NaI(aq) \longrightarrow$

   (f) ___$FeCl_3(aq) +$ ___ $NaOH(aq) \longrightarrow$

   (g) ___$KC_2H_3O_2(aq) +$ ___$CoSO_4(aq) \longrightarrow$

   (h) ___ $Na_2CO_3(aq) +$ ___ $HNO_3(aq) \longrightarrow$

   (i) ___ $ZnBr_2(aq) +$ ___ $K_3PO_4(aq) \longrightarrow$

   (j) ___ $KOH(aq) +$ ___ $H_3PO_4(aq) \longrightarrow$

Name: ______________________________ Class: ______________ Date: ______________

Prelab Questions

1. What are the 4 signs that a double displacement reaction has occurred in **LAB**.

2. Complete the table below listing all of the chemical compounds used in lab.

Name	**Formula**
Sodium Chloride	
Potassium Nitrate	
Silver Nitrate	
Sodium Carbonate	
Hydrochloric Acid	
Nitric Acid	
Sodium Sulfite	
	$BaCl_2$
	H_2SO_4
	NH_4OH
	$CuSO_4$
	$Zn(NO_3)_2$
	$CaCl_2$
	$FeCl_3$

Experiment 11
Single Displacement Reactions

Jay C. McLaughlin
Colorado Northwestern Community College

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Name:

Date:

Key Objectives

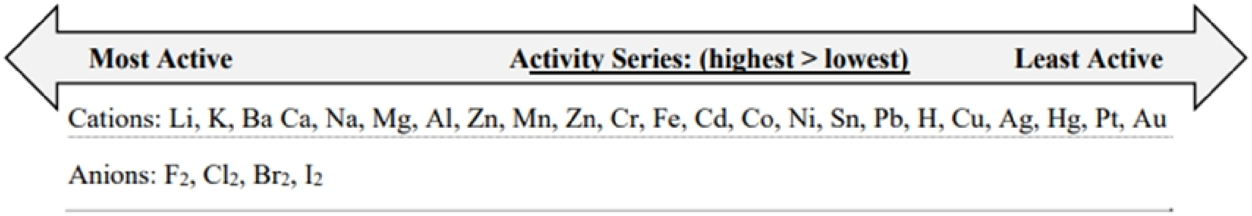
1. Understand Activity Series and use it to predict if a reaction occurs.
2. Write complete Single Displacement chemical reactions
   (a) Balance charges.
   (b) Balance atoms.
   (c) Include states when known.
3. Recognize when no reaction (NR) occurs.
4. Use a solubility table to determine the state of compounds.

Discussion

The chemical reactivity of elements varies over an immense range. Some, like sodium and chlorine are so reactive that they are never found in a pure state, and instead always form compounds. Others like xenon and platinum are nearly inert and can be made to form compounds with other elements only under special conditions. In this experiment we will explore the relative reactivity of the elements, and establish a reactivity order.

The reactivity of an element is related to its tendency to gain and lose electrons. An element which loses electrons in a reaction is said to be **oxidized**, and an element that gains electrons in a reaction is said to be **reduced**. In a chemical reaction, oxidation and reduction of elements is always paired, resulting in an exchange of electrons. It is this exchange of electrons which is the primary driving force for the reaction to occur. Oxidation and reductions reactions are very important industrially for the electroplating of metals, production of batteries and cathodic protection.

It is possible to arrange nearly all of the elements into a single series in order of their reactivities (ability to be oxidized or reduced). A series of this kind is commonly called an **activity series**.



Most Active **Activity Series: (highest > lowest)** **Least Active**

Cations: Li, K, Ba Ca, Na, Mg, Al, Zn, Mn, Zn, Cr, Fe, Cd, Co, Ni, Sn, Pb, H, Cu, Ag, Hg, Pt, Au

Anions: F_2, Cl_2, Br_2, I_2

Figure 11.1: The reactivity of elements decreases going from left (most active) to right (least active). credit: author

One key to successfully completing single displacement reactions is to know which elements are metals and form cations (+) and which elements are nonmetals and form anions (-). Metals are to the left of the **metalloids** (B, Si, Ge, As, Sb, Te and Po/At) except for Hydrogen which is considered a nonmetal. The nonmetals are to the right of the metalloids as shown in Figure 11.2.

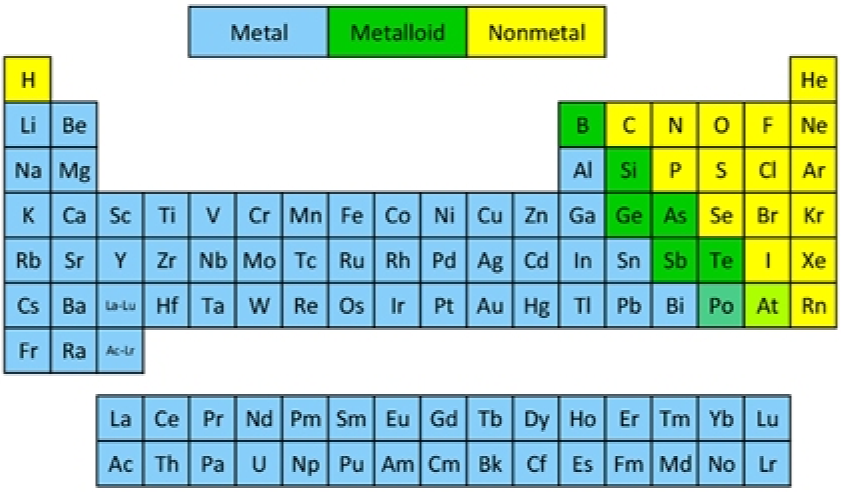


Figure 11.2: Periodic table illustrating the metalloids, metals and nonmetals. credit: unknown

A generalized single displacement reaction is given below. If the element A is more reactive then it will displace the less reactive element and form a new compound. If element A is less active then no reaction will occur. Which reaction occurs depends on if element A is a metal or nonmetal and thus more likely to form a cation or anion. If element A is a metal it will displace the metal/cation (B) in the compound or if element A is a nonmetal it will displace the nonmetal/anion (C) in the compound,

$$\mathrm{A + BC(aq) \longrightarrow B + AC(aq)} \qquad \text{or} \qquad \mathrm{A + BC(aq) \longrightarrow C + BA(aq)} \qquad (21)$$

A more visual representation of what occurs is shown in Figure 11.3, the Element (E) swaps with either the cation (if E is a metal) or the anion (if E is a nonmetal) as long as the Element (E) is more reactive than the cation or anion in the compound. If it is not no reaction will occur.

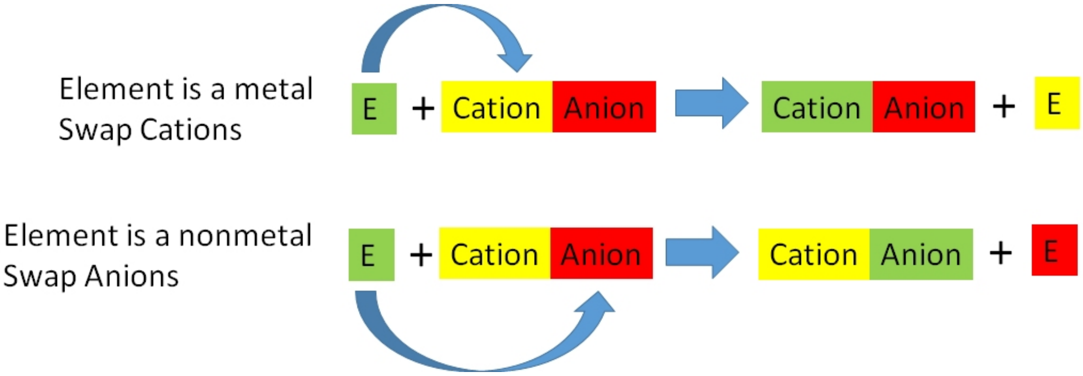


Figure 11.3: Another way of visualizing single displacement reactions. The nature of the element (metal or nonmetal) determine what elements swap (cations or anions). credit: author

One last way of thinking about single displacement reactions is illustrated in the cartoon in Figure 11.4.

A couple of examples will help to clarify what will occur in this experiment.

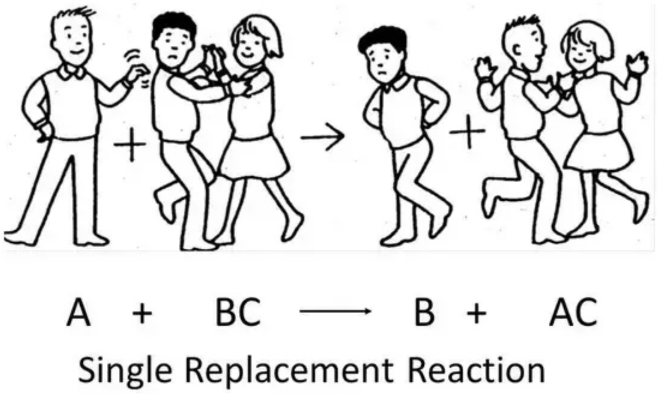


Figure 11.4: Cartoon illustrating the concept of single displacement reactions, the better dancer (A) displaces the bad dancer (B) and gets to dance with the girl (C). credit: unknown

Example 1

A few drops of mercury metal is reacted with a solution of copper (II) chloride. No reaction is observed. We can write the complete chemical equation as:

$$\mathrm{Hg(l) + CuCl_2(aq) \longrightarrow NR} \tag{22}$$

We conclude then that mercury is less reactive than copper (Hg $<$ Cu) because it will not displace copper from the compound. Similarly we could say that Cu is more reactive than Hg and thus desires to be in the compound more than Hg. (Cu $>$ Hg)

Example 2

A strip of metallic copper is immersed in a solution of mercury (II) chloride. After several minutes, the solution becomes pale green and the copper strip is coated with a metallic color. We can write the complete chemical reaction as:

$$\mathrm{Cu(s) + HgCl_2(aq) \longrightarrow Hg(s) + CuCl_2(aq)} \tag{23}$$

We conclude then that copper is more reactive than mercury and will displace mercury from compounds. (Cu $>$ Hg)

From the second example we can see that the oxidation number (relative charge on the atoms) that the copper lost two electrons and went from an oxidation number of 0 $\longrightarrow$ $+2$, thus it was oxidized. Mercury gained two electrons and went from an oxidation number of $+2$ $\longrightarrow$ 0, thus it was reduced. The chlorine remained unchanged in the reaction with an oxidation state of -1. We can summarize this by stating:

$$\mathrm{Cu^0(s) + Hg^{+2}(aq) \longrightarrow Hg^0(s) + Cu^{+2}(aq)} \qquad (24)$$

Procedure

An exciting new discovery has been made! Six new elements (5 metals and 1 gas) have been discovered. It is your job to create an activity series for these new elements. In this experiment you will be observing the reactions of various solid metals and aqueous solutions of a second metal. From this data we will be able to generate an activity series for the new elements.

The new metals are being named after their discoverer and have the temporary symbols of J1, J2, J3, J4, J5, J6. The charges of each metal have been determined to be: J1^{+2}, J2^{+2}, J3^{+1}, J4^{+3}, J5^{+2}, J6^{+1}.

Read the entire procedure before beginning the experiment.

1. The reactions may be done in any order desired.
2. Label 6 clean test tubes 1-6. To each test tube add about 3 mL of the aqueous solution.
3. Add the appropriate thin strip of the solid metal to the test tube indicated.
    (a) Tube 1: J3 (I) Nitrate + J4 strip.
    (b) Tube 2: J4 (III) Nitrate + J1 strip.
    (c) Tube 3: J1 (II) Nitrate + J2 strip.
    (d) Tube 4: J5 (II) Sulfate + J2 strip.
    (e) Tube 5: J6 (I) Sulfate + J2 strip.
    (f) Tube 6: J6 (I) Sulfate + J4 strip.
4. Observe the contents of each tube carefully and record any evidence of a chemical reaction. Some reactions may be slow to occur or difficult to detect. If no immediate evidence of a reaction occurs, set the tube aside for 10 minutes and then reexamine it.
5. After the reactions are complete:
    (a) Remove the metal strips from each test tube and place them in the labeled beaker.
    (b) Pour the solutions in each test tube into the “Experiment SD - Heavy Metals” waste container.
6. Write the complete chemical reaction (balanced, and include the states where known) for each test tube in the data table. If NO REACTION occurs write the reactants $\longrightarrow$ NR. Answer all of the questions before leaving. Ex: $\mathrm{J3_2SO_4(aq) + J_1(s) \longrightarrow}$

Results

TT #	**Observations**	**Complete Chemical Equation**
1.		
2.		
3.		
4.		
5.		
6.		

Table 11.1: Results

Post Lab Questions

1. Complete the following table by writing the symbols of the two elements whose reactivities are being compared in each test. Explain how you determined which element was the most active in the space below the table.

TT #	**1**	**2**	**3**	**4**	**5**	**6**
More Active Metal:						
Less Active Metal:						

Table 11.2: Relative Activity of Tested Metals

Explanation:

2. Arrange J1, J2, and J4 in order from highest activity to lowest activity (Ex: A $>$ B $>$ C).

3. Arrange J5, J6, and J3 in order from highest activity to lowest activity (Ex: A $>$ B $>$ C).

4. Arrange J3, J4, and J5 in order from highest activity to lowest activity (Ex: A $>$ B $>$ C).

5. Arrange all 5 metals (excluding J6) in order from highest activity to lowest activity (Ex: A $>$ B $>$ C). Explain your ordering based on the observations made in test tubes 1-6 or on the previous 3 questions.

6. On the basis of the reactions observed in test tubes 1-6 explain why it is not possible to place J6 in the activity series in the previous question?

7. Write a chemical reaction(s) (using chemicals available in the experiment) that would allow you to place J6 in the activity series. Only write the Reactants side of the reaction.

8. If the reaction written in the previous question results in NO REACTION, write the complete activity series for all 6 elements in order from highest activity to lowest activity.

9. Using the activity series in your book/cheat sheet complete the following reactions. If the reaction will occur complete and balance each compound, and equation, properly indicate the state of all products where known. If no reaction will occur write “No reaction” or “NR” as the products.

   (a) ___ $H_2(g)$ + ___ $AgNO_3(aq) \longrightarrow$

   (b) ___ $Pb(C_2H_3O_2)_2(aq)$ + ___ $Na(s) \longrightarrow$

   (c) ___ $H_2SO_4(aq)$ + ___ $Mg(s) \longrightarrow$

   (d) ___ $Pb(s)$ + ___ $Cu(NO_3)_2(aq) \longrightarrow$

   (e) ___ $AlBr_3(aq)$ + ___ $Cl_2(g) \longrightarrow$

   (f) ___ $ZnSO_4(aq)$ + ___ $Mg(s) \longrightarrow$

   (g) ___ $Cl_2(g)$ + ___ $NaI(aq) \longrightarrow$

   (h) ___ $Al_2(SO_4)_3(aq)$ + ___ $Na(s) \longrightarrow$

   (i) ___$MgF_2(aq)$ + ___$I_2(s) \longrightarrow$

   (j) ___ $Zn(s)$ + ___ $H_3PO_4(aq) \longrightarrow$

I thought of adding a bunch of extra questions to fill up the blank space, but was in a good/lazy mood when I wrote this so didn't add any!

Experiment 12
Stoichiometry

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Calculate Molecular Weight (MW).
2. Calculate Limiting Reactant (LR).
3. Calculate Excess Reactant left over (ER).
4. Calculate Theoretical Yield.
5. Calculate Percent Yield.

Discussion

A common chemical question is how much of a product can I make, or how much of a reactant do I need in a chemical reaction. In this lab you will learn how to answer both of these questions and a few more!

Stoichiometry calculations are about calculating the amounts of substances that react and form in a chemical reaction. The word “stoichiometry” comes from the Greek stoikheion "element" and metriā "measure."

Based on the balanced chemical equation, we can calculate the amount of a product substance that will form if we begin with a specific amount of one or more reactants. Or, you may have a target amount of product to prepare. How much starting compounds are needed to prepare this amount? These are practical calculations that are done frequently by chemists.

Kitchen Chemistry

Chemistry and eating have a lot in common (ok its a stretch). In Figure **??** you have 10 wieners + 8 buns and can make 8 hot dogs but you will have 2 left over weiners.

Figure 12.1: Why do wieners come in packages of 10 but the buns in a package of 8. That leaves 2 left over wieners. credit: author

In the previous example it was easy to do the math, lets take a look at a slightly harder example and include some math. Imagine you are making a sandwich using the following recipe:

$$\underline{\;2\;}\,\text{slice bread} + \underline{\;3\;}\,\text{slices ham} + \underline{\;3\;}\,\text{slices turkey} \longrightarrow \underline{\;1\;}\,\text{great sandwich} \tag{25}$$

If you have 12 slices of bread, 24 slices of ham and 36 slices of turkey how many sandwich's can you make? Since this is real life it seems easier than chemistry (hint: it might be, but you can take your experience in real life and apply it to chemistry!). What you (probably without thinking of it) would determine is how many sandwich's you can make with each ingredient: 12 slices of bread you can make 6 sandwich's, with 24 slices of ham you can make 8 sandwich's and with 36 slices of turkey you could make 12 sandwich's. Thus you can make 6 sandwich's before you run out of bread. That makes bread your **Limiting Reactant (LR)**! In other words, the reactant you run out of in a reaction is the Limiting Reactant.

If we imagined this as a chemist would you might do something like the following and deduce that the smallest number of sandwich's made tells us the Limiting Reactant.

$$\frac{\text{12 slices bread}}{} \times \frac{\text{1 sandwich}}{\text{2 slices bread}} = 6\ \text{sandwich}'\text{s} \tag{26}$$

$$\frac{\text{24 slices ham}}{} \times \frac{\text{1 sandwich}}{\text{3 slices ham}} = 8\ \text{sandwich}'\text{s} \tag{27}$$

$$\frac{\text{36 slices turkey}}{} \times \frac{\text{1 sandwich}}{\text{3 slices turkey}} = 12\ \text{sandwich}'\text{s} \tag{28}$$

We can now define a few more terms:

Excess Reactant (ER): The reactant that you have excess (or left over of).

Theoretical Yield: The maximum amount of a product you can make based on calculations.

Actual Yield: The actual yield of a product in the laboratory (or kitchen), What you actually measure/make.

Percent Yield: Is the percentage of the theoretical value you make in the lab or kitchen. It is essentially your "grade".

Mathematically we can write:

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100 \tag{29}$$

If we take our sandwich making example then we could say that the we have **Excess** ham (6 slices) and turkey (18 slices) and we can **Theoretically** make 6 sandwich's. If we dropped a slice of bread (yuck, no 5 second rule in chemistry lab) then our **Yield** of sandwich's would go down because we can only make 5 sandwich's.

Again switching to thinking like a chemist we might formulate the problem as follows:

To find the Excess Reactant left over we start with Limiting Reactant and calculate how much of the Excess Reactant we use and subtract it from what we started with to get the amount left over.

$$\frac{12\ \text{slices bread}}{} \times \frac{3\ \text{slices ham}}{2\ \text{slices bread}} = 18\ \text{slice}'\text{s ham} \therefore 24 - 18 = 6\ \text{slice}'\text{s ham left over} \tag{30}$$

$$\frac{12\ \text{slices bread}}{} \times \frac{3\ \text{slices turkey}}{2\ \text{slices bread}} = 18\ \text{slice}'\text{s turkey} \therefore 36 - 18 = 18\ \text{slice}'\text{s turkey left over} \tag{31}$$

We have already calculated our Theoretical Yield of sandwich's when we calculated the LR in equation 2-4 and found we can only make 6 sandwich's

Our Percent Yield (assuming we dropped a slice of bread would be:

$$\text{Percent Yield} = \frac{5\ \text{sandwich}'\text{s}}{6\ \text{sandwich}'\text{s}} \times 100 = 83\% \tag{32}$$

Mol to Mol ratio

When making sandwich's we know the amount of ingredients to use, so to in chemistry we have a chemical reaction that tells us the amount of each chemical to use. Instead of slices chemists use a mole (or mol for short) which is just a ridiculously large number of slices. 6.02×10^{23} to be exact. To give credit where it is due, Avagadro discovered this and so its generally referred to as Avagadro's number: 1 mol = 6.02×10^{23} anything.

For our sandwich recipe we might write the reactions below. The first reaction we are measuring everything by the "slice". In the second reaction we use the chemistry term "mol" which just means lots and lots and lots of slices. In the third recipe/reaction we omit the word "mol" because we are chemists and we know we measure everything in "mols"

2	slice bread	+	3	slice ham	+	3	slice turkey	⟶	1	sandwich
2	mol bread	+	3	mol ham	+	3	mol turkey	⟶	1	mol sandwich
2	bread	+	3	ham	+	3	turkey	⟶	1	sandwich

Thus, we can have a slice/slice ratio if making sandwich's and a mol/mol ratio if we are in the chemistry lab.

$$\frac{1\ \text{sandwich}}{3\ \text{slices ham}} \text{ or } \frac{1\ \text{mol sandwich}}{3\ \text{mol ham}} \tag{33}$$

Molecular Weight (MW)

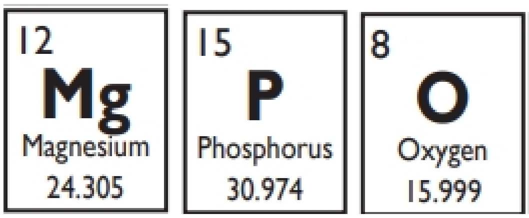
If we went to the Deli to buy sandwich meat to make our sandwich's we would probably buy everything by the ounce (oz). If we did this then we might need to know how many slices of ham are in each ounce of ham we order (same goes for turkey). This would then be a conversion factor, for example we might say that 1 ounce of ham = 6 slices, but turkey is sliced thicker and 1 ounce of turkey = 3 slices. Thus, to order the same number of ham and turkey slices we would need to buy twice as many ounces of turkey as of ham. The number of ham/turkey slices and the mass of the ham/turkey are different.

$$\frac{1\text{ oz ham}}{6\text{ slices ham}} \text{ and } \frac{1\text{ oz turkey}}{3\text{ slices turkey}} \tag{34}$$

If we went to the store and wanted to buy 24 slices of ham we could calculate how many ounces we needed:

$$\frac{24\text{ slices ham}}{} \times \frac{1\text{ oz ham}}{6\text{ slices ham}} = 4\text{ oz ham} \tag{35}$$

In the kitchen we might use cups or ounces to measure things but in the chemistry lab we measure out our chemicals in grams. Just like our ham and turkey each chemical compound has a different mass. In order to get from the grams that we measure in lab to the moles that our balanced equation is written in we will need to calculate the Molecular Weight (MW) or it is sometimes referred to as the Molar Mass. This is the weight of 1 mole (6.02×10^{23} atoms or molecules) of a substance and is calculated using the weight of 1 mole of each element on the periodic table.



For example we can calculate the MW of $Mg_3(PO_4)_2$:

$$\begin{array}{rr} \text{Mg} = 3 \times 24.305 & = 72.915\text{ g/mol} \\ \text{P} = 2 \times 30.974 & = 61.948\text{ g/mol} \\ \text{O} = 8 \times 15.999 & = 127.992\text{ g/mol} \\ \hline \text{Mg}_3(\text{PO}_4)_2 & = 262.853\text{ g/mol} \end{array} \tag{36}$$

General Procedure

Generically we can apply the following procedure to calculate our Limiting Reactant (LR), Excess Reactant left over (ER), Theoretical and Percent Yields.

Quantity	**Calculation**
Limiting Reactant	$\mathrm{g\,R_1 \longrightarrow mol\,R_1 \longrightarrow mol\,P_1 \longrightarrow g\,P_1}$ $\mathrm{g\,R_2 \longrightarrow mol\,R_2 \longrightarrow mol\,P_1 \longrightarrow g\,P_1}$ The Limiting Reactant (LR) is the reactant that produces the least amount of Product.
Excess Reactant	$\mathrm{g\,LR \longrightarrow mol\,LR \longrightarrow mol\,ER \longrightarrow g\,ER}$ The Excess Reactant is found by subtraction: Start g ER - Used g ER = Left Over ER
Theoretical Yield	Already calculated! The minimum amount of product made when you calculated your LR.
Percent Yield	$\frac{\text{Actual}}{\text{Theoretical}} \times 100$

Chemistry Example

We can now graduate from the kitchen to the laboratory with one good example below. Given the reaction below assuming you start with 25.0 g of H_2SO_4 and 50.0 g H_3PO_4.

$$\begin{array}{ccccccc} \underline{\ 3\ }\,\mathrm{Mg(OH)_2(aq)} & + & \underline{\ 2\ }\,\mathrm{H_3PO_4} & \longrightarrow & \underline{\ 1\ }\,\mathrm{Mg_3(PO_4)_2(s)} & + & \underline{\ 6\ }\,\mathrm{H_2O(l)} \\ 58.32\ \mathrm{g/mol} & & 98.00\ \mathrm{g/mol} & & 262.85\ \mathrm{g/mol} & & 18.02\ \mathrm{g/mol} \end{array} \tag{37}$$

Below each reactant and product is the molecular weight which we will need in our calculations. You can double check my calculations to make sure that you remember how to do it!

We will answer the following 5 questions (given that you start with 25.0 g of H_2SO_4 and 50.0 g H_3PO_4:

1. Calculate the Limiting Reactant (LR).
2. Calculate the Theoretical Yield of $Mg_3(PO_4)_2$.
3. Calculate the Theoretical Yield of H_2O.
4. Calculate the amount of Excess Reactant left over.
5. Calculate the Percent Yield if in lab you produced 25.0 g of $Mg_3(PO_4)_2$.

Calculating the Limiting Reactant by determining how much of a product can be produced with each reactant. The reactant that produces the least product is your limiting reactant, and as a bonus, you also get the theoretical yield of that product (a 2 for 1 deal!).

$$\frac{25.0\,\mathrm{g\,Mg(OH)_2}}{} \times \frac{1\,\mathrm{mol\,Mg(OH)_2}}{58.32\,\mathrm{g\,Mg(OH)_2}} \times \frac{1\,\mathrm{mol\,Mg_3(PO_4)_2}}{3\,\mathrm{mol\,Mg(OH)_2}} \times \frac{262.85\,\mathrm{g\,Mg_3(PO_4)_2}}{1\,\mathrm{mol\,Mg_3(PO_4)_2}} = 37.6\,\mathrm{g\,Mg_3(PO_4)_2}$$

$$\frac{50.0\,\mathrm{g\,H_3PO_4}}{} \times \frac{1\,\mathrm{mol\,H_3PO_4}}{98.00\,\mathrm{g\,H_3PO_4}} \times \frac{1\,\mathrm{mol\,Mg_3(PO_4)_2}}{2\,\mathrm{mol\,H_3PO_4}} \times \frac{262.85\,\mathrm{g\,Mg_3(PO_4)_2}}{1\,\mathrm{mol\,Mg_3(PO_4)_2}} = 67.1\,\mathrm{g\,Mg_3(PO_4)_2}$$

From the calculations above the least amount of product (37.56 g $Mg_3(PO_4)_2$ which is our Theoretical

Yield and is produced by $Mg(OH)_2$ which is the Limiting Reactant. We have answered question 1 and 2.

To determine theoretical yield of our other product we will start with our Limiting Reactant and determine the amount of water produced.

$$\frac{25.0\,\mathrm{g\,Mg(OH)_2}}{} \times \frac{1\,\mathrm{mol\,Mg(OH)_2}}{58.32\,\mathrm{g\,Mg(OH)_2}} \times \frac{6\,\mathrm{mol\,H_2O}}{3\,\mathrm{mol\,Mg(OH)_2}} \times \frac{18.02\,\mathrm{g\,H_2O}}{1\,\mathrm{mol\,H_2O}} = 15.5\,\mathrm{g\,H_2O}$$

To calculate the amount of Excess Reactant left over we start with our Limiting Reactant, calculate the amount of the other reactant used and subtract it from our starting amount.

$$\frac{25.0\,\mathrm{g\,Mg(OH)_2}}{} \times \frac{1\,\mathrm{mol\,Mg(OH)_2}}{58.32\,\mathrm{g\,Mg(OH)_2}} \times \frac{2\,\mathrm{mol\,H_3PO_4}}{3\,\mathrm{mol\,Mg(OH)_2}} \times \frac{98.00\,\mathrm{g\,H_3PO_4}}{1\,\mathrm{mol\,H_3PO_4}} = 28.0\,\mathrm{g\,H_3PO_4}$$

Start - Used = Left Over
50.0 - 28.0 = 22.0 g H_3PO_4 Left Over

The last calculation is the Percent Yield.

$$\%\,\mathrm{Yield} = \frac{\mathrm{Actual}}{\mathrm{Theoretical}} \times 100$$

$$\%\,\mathrm{Yield} = \frac{25.0\,\mathrm{g\,Mg_3(PO_4)}}{37.6\,\mathrm{g\,Mg_3(PO_4)_2}} \times 100 = 66.5\%$$

To summarize our answers for the 5 questions we posed:

1. Calculate the Limiting Reactant (LR). **$Mg(OH)_2$**
2. Calculate the Theoretical Yield of $Mg_3(PO_4)_2$. **37.6 g $Mg_3(PO_4)_2$**
3. Calculate the Theoretical Yield of H_2O. **15.5 g H_2O**
4. Calculate the amount of Excess Reactant left over. **28.0 g H_3PO_4**
5. Calculate the Percent Yield if in lab you produced 25.0 g of $Mg_3(PO_4)_2$. **66.5 %**

Procedure

In this experiment, you will prepare copper metal (Cu) from the reaction of aluminum metal (Al) with a solution of copper(II) sulfate ($CuSO_4$). From the amounts of the reactants, you will determine which reactant is the limiting reactant, and from this amount, calculate the theoretical yield of copper metal. From the actual amount of copper obtained, you can then calculate your percent yield of copper.

The reaction we will be studying is given below:

$$\underline{\ 2\ }\ \text{Al(s)} + \underline{\ 3\ }\ \text{CuSO}_4\text{(aq)} \longrightarrow \underline{\ 3\ }\ \text{Cu(s)} + \underline{\ 1\ }\ \text{Al}_2(\text{SO}_4)_3\text{(aq)}$$

1. Weigh a clean, dry 150 mL beaker and record its weight.
2. Weigh out approximately 2.0 g of $CuSO_4$ into the beaker.
3. Add approximately 10. mL of water to the beaker and swirl to dissolve the $CuSO_4$. Record the color of the solution.
4. Measure approximately 2.0 mL of 6 M HCl and add it to the beaker
5. Weigh out approximately 0.25 g of aluminum foil in small pieces and record the mass.
6. Add the pieces of Al foil to the beaker containing the $CuSO_4$ little at a time, stirring the mixture continuously. Use caution as the reaction is very Exothermic. After all the aluminum foil is added to the beaker add an additional 5.0 mL of 6 M HCl to the beaker to facilitate the reaction of any excess Al foil. Record the color of the solution and the solid precipitate.
7. After the reaction is complete allow the solid particles of copper to settle and carefully decant the solution from the solid (leaving the copper behind in the beaker) into a 2nd beaker.
8. Add 20 mL of water to the beaker containing the copper stir well and decant again. Do this step twice.
9. Add 10 mL of methanol to the copper, stir and decant.
10. Dispose of liquids decanted into the waste container labeled "Experiment 14 Waste".
11. Heat the beaker on a hot plate at a medium heat setting (4 out of 10) until the solid and beaker are thoroughly dry. Allow the beaker to cool (so you can touch it) and weight it on the scale. Record the mass.
12. If time permits, reheat the beaker for an additional 10 minutes and record the mass to ensure that the sample was completely dry. This is called heating to a constant mass.
13. Dispose of your copper metal in the container labeled "Solid Copper Metal Waste"

☠ Dispose of any excess materials in the Heavy Metals waste jug (if liquid) or in the provided container (if solid).

.

Results

1. Mass of 150 mL beaker 1. ______________
2. Mass of 150 mL beaker + $CuSO_4$: 2. ______________
3. Mass of $CuSO_4$ used: 3. ______________
4. Color of solution (before adding Al foil) 4. ______________
5. Mass of Al foil: 5. ______________
6. Color of solution (after adding Al foil) 6. ______________
7. Mass of dry Cu + beaker (after 1st heating): 7. ______________
8. Mass of dry Cu + beaker (after 2nd heating): 8. ______________
9. Mass of dry Cu: 9. ______________

Post Lab Questions

1. Write the complete (balanced, and include states) chemical reaction that occurred in the lab.

2. Below each Reactant and Product in the previous reaction write the molecular weight of the compound.

3. Determine the theoretical amount of Copper metal from the starting values for the mass of $CuSO_4$. (Show calculation) 3. ______________

4. Determine the theoretical amount of Copper metal from the starting values for the mass of Al. (Show calculation) 4. __________

5. What is the Theoretical Yield of Copper metal. Explain. 5. __________

6. What is the Limiting Reactant. Explain. 6. __________

7. What is the Percent Yield of Copper? (Show calculation) 7. __________

8. What is the amount of Excess Reactant Left over? (Show calculation) 8. __________

9. What is the theoretical mass of the $Al_2(SO_4)_3$ produced? (Show calculation.) 9. __________

10. Did you obey Lavoisier Conservation of Mass law? Explain. 10. __________

Name: ______________________ Class: ____________ Date: ____________

Prelab Questions

1. Complete the following table.

Name	**Formula**	**Molecular Weight**
Aluminum Sulfate		
Copper (II) Sulfate	$CuSO_4$	
	HCl	
	$AlCl_3$	
Methanol	CH_3OH	

.

Experiment 13
Titrations

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Nomenclature - Titration, End point, Indicator, Standard solution, Meniscus.
2. Solving titration problems.
3. Use of molecular weight and mol-to-mol ratio's.
4. Determine the concentration of an unknown solution.

Discussion

Titrations are a standard technique used to determine the concentration of a solution by chemically reacting it with a known solution. There are a variety of techniques for doing this, we will focus on the most basic technique realizing that we can apply it to more complicated problems if required.

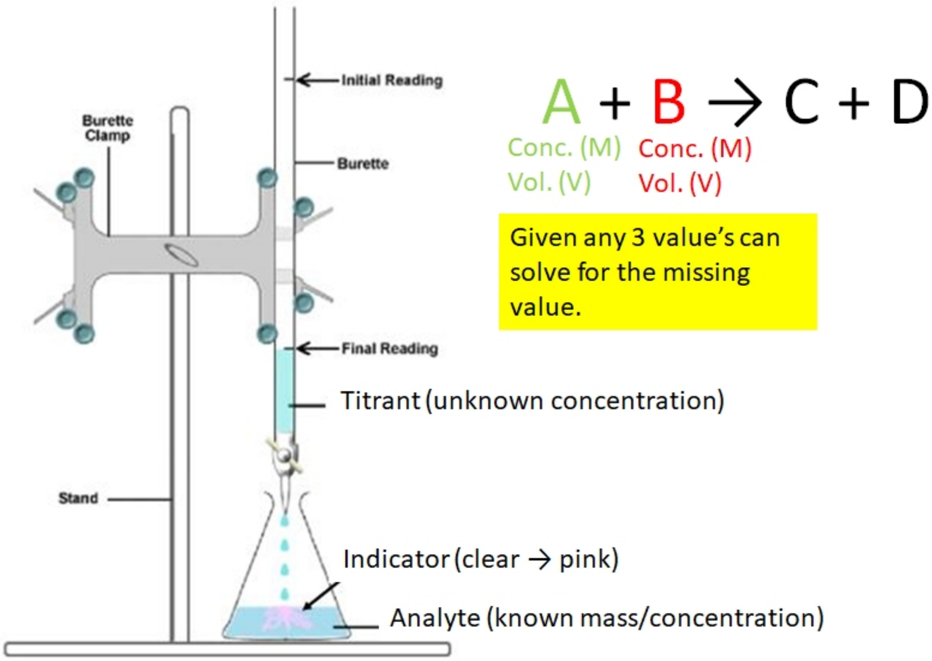


Figure 13.1: Typical setup for performing a titration. credit: author

A discussion of titrations can be found in Hein Ch. 15.10 or in McMurry Ch 3.10. **Titrations** are a method for determining the amount or concentration of an unknown in a solution. Titrations react a solution of known concentration (the **standard**) with a solution of unknown concentration (the unknown). If the reaction goes to completion, and the yield is 100%, then the concentration of the unknown solution can be determined through the stoichiometry of the chemical reaction. The completion of the reaction is determined by a variety of methods including color changes, pH, and electrical conductivity.

Concentration is generally measured in terms of **Molarity** (M), which has units of (moles of solute/L of solution), or (mol/L).

A typical reaction which goes to completion with 100% yield is the reaction of an acid with a base. Such reactions are also commonly referred to as neutralization reactions. Generally we can write:

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water} + \text{Heat} \tag{38}$$

In this lab we will use the reaction of potassium hydrogen phthalate, $KHC_8H_4O_4$, commonly known as “KHP” with Magnesium Hydroxide ($Mg(OH)_2$). The reaction is balanced as shown below.

$$\underline{\ 2\ }\mathrm{KHC_8H_4O_4(aq)} + \underline{\ 1\ }\mathrm{Mg(OH)_2(aq)} \longrightarrow \underline{\ 1\ }\mathrm{Mg(KC_8H_4O_4)_2(aq)} + \underline{\ 2\ }\mathrm{H_2O(l)} \tag{39}$$

The choice of reaction to study is made because KHP is a solid at room temperatures, we can measure its concentration by mass, which is much more accurate and precise then measuring liquid volumes. We will use this to titrate an unknown concentration of $Mg(OH)_2$.

The **end point** of a reaction is the point at which the reaction is neutralized (the concentration of the reactants are equal) or 100% complete and is determined by using an **indicator**. Most indicators signal the end point of a reaction by changing colors, though other changes are possible. The indicator chosen for this reaction is phenophthalein, an organic dye, that is colorless in acid solutions (the KHP solution) and turns pink in basic solution (when enough $Mg(OH)_2$ is added).

The standard instrument used in titrations is the **Buret** which is a piece of calibrated glassware used to accurately measure volumes of liquids. Figure 13.1 shows a typical example. The buret is filled with the solution ($Mg(OH)_2$) to be titrated and is often referred to as the titrant.

Volume measurements are made by reading the point on the graduated scale that coincides with the bottom of the curved surface called the **meniscus** of the liquid as shown in Figure 13.2. A discussion of why a meniscus is formed can be found in Hein Ch. 13.4, Chang Ch. 11.3, or McMurry Ch. 10.4.

Do not waste your time trying to fill the buret to the zero line or some other round number, because The exact volume to which the buret is filled is unimportant because we are measuring the difference in volume (Δv).

The titrant is now added to a solution made from the KHP, water, and indicator. Titrant is added until just one drop changes the solution from colorless to a faint pink color. The exact process will be demonstrated in class by your instructor. The final volume of liquid is then measured.

Using the mass of the KHP, the volume of $Mg(OH)_2$ used and the stoichiometry of the reaction it is then possible to determine the molarity of the $Mg(OH)_2$ solution.

For example, if you start with 1.025 g of KHP and add 22.50 mL of an unknown Molarity of $Mg(OH)_2$ calculate the Molarity of the solution.

$$\frac{1.025\,\mathrm{g}\ \text{KHP}}{22.50\,\mathrm{mL}} \times \frac{1\,\mathrm{mol}\ \text{KHP}}{204.2\,\mathrm{g}\ \text{KHP}} \times \frac{1\,\mathrm{mol}\,\mathrm{Mg(OH)_2}}{2\,\mathrm{mol}\ \text{KHP}} \times \frac{1\,\mathrm{mL}}{0.001\,\mathrm{L}} = 0.1116\,\text{M}\,\mathrm{Mg(OH)_2}$$

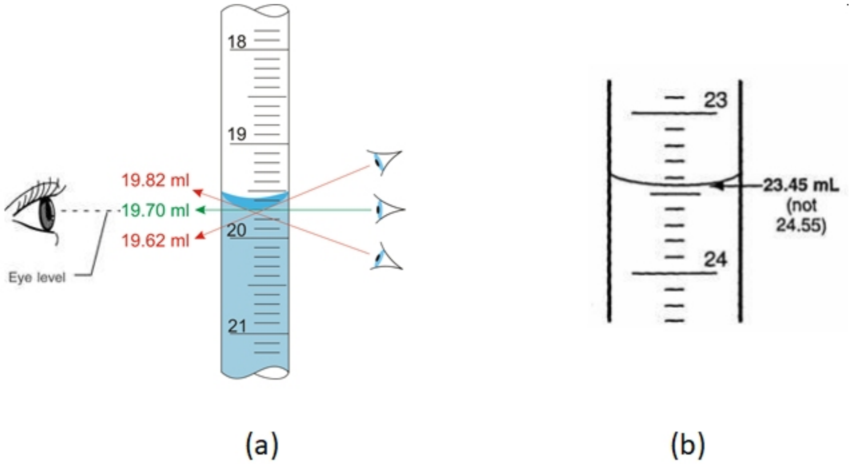


Figure 13.2: (a) Always read a buret with your eyes level with the bottom of the meniscus. (b) Notice the numbering on the scale, it goes from top to bottom (opposite of a beaker). credit: (a) E. Generalic, `https://glossary.periodni.com/glossary.php?en=meniscus` (b) `https://effectiveness.lahc.edu/academic_affairs/sfcs/chemistry/Shared`

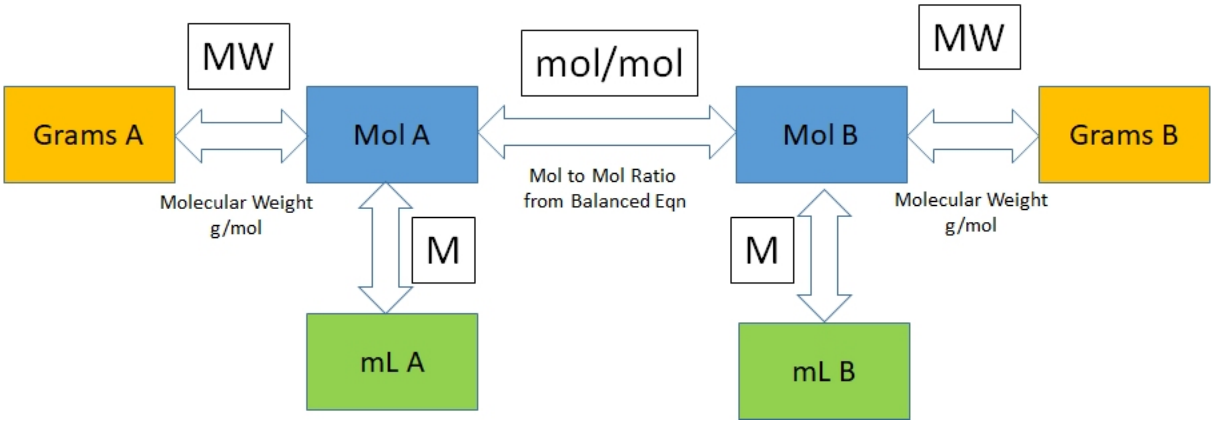


Figure 13.3: Using conversion factors to solve a titration problem. credit: author

Procedure

In this experiment you will be titrating a known mass of an acid (KHP) with an unknown concentration of base ($Mg(OH)_2$). After determining the concentration of the base solution you will use it to determine the concentration of acetic acid in vinegar and compare it to the manufactures claimed concentration.

Read the entire procedure before beginning the experiment.

Part 1 - Practice using the buret

1. When reading the meniscus level in the buret, you may find it helpful to hold a white card (Figure **??**) marked with a dark strip (or a dark colored object) behind and slightly below the meniscus. Remember, your line of sight should be level with the meniscus to obtain the most accurate reading.

2. To what uncertainty can you measure volume using a buret?
3. Add between 40 and 45 mL of distilled water to the buret using a funnel placed on the top of the buret. Be sure to check for air bubbles in the tip of the buret, if found you need to dispense your liquid until they are removed. Record your reading on your data sheet and have your instructor verify the reading.
4. Weigh an empty beaker.
5. Place an empty beaker below your buret and dispense between 10 and 12 mL of distilled water into it. Record your final reading, and have your instructor verify it.
6. Record the mass of water dispensed.
7. Calculate the volume of water dispensed into your beaker. (Show work)
8. Calculate the difference between the volume of water dispensed and the mass of water dispensed. Does the calculated volume of water dispensed into your beaker agree with the mass of the water dispensed into the beaker within 0.1g of the expected value? If not consult your instructor.
9. If you have any troubles with this portion of the lab, or your mass and volume measurements disagree too much, consult your instructor, and do a second trial if needed.
10. Continue practicing dispensing varying amounts of water, at varying rates from the buret into a beaker until you feel comfortable controlling the water flow using the stopcock.
11. Finish by draining all of the water from the buret into the beaker.
12. Any water used in this part of the experiment may be disposed of down the sink.

Part 2 - Obtaining a $Mg(OH)_2$ sample, and filling the buret.

You will need the $Mg(OH)_2$ solution for titrating both the KHP standard Each titration should take between 10-20 mL of the $Mg(OH)_2$ solution.

1. Obtain about 250 mL of the $Mg(OH)_2$ solution from your instructor. Put your name on the Erlenmeyer flask.
2. Close the stopcock on your buret and carefully add through a funnel 5 mL of your $Mg(OH)_2$ solution. Remove the funnel. Remove the buret from the support stand, and hold it almost horizontally. Carefully rotate so that the $Mg(OH)_2$ solution contacts the entire inner surface.
3. Drain the $Mg(OH)_2$ solution through the buret tip into a beaker labeled “Discarded Rinse Solutions” or dispose of it down the sink.
4. Clamp the buret into the support stand. Fill the buret to a point slightly below the 0 mL calibration mark.
5. Dispense the $Mg(OH)_2$ solution though the stopcock and remove any air bubbles present. If you

can not remove the air bubbles consult your instructor.

6. Once all air bubbles are removed, refill the buret to between the 0 and 10 mL mark.

Part 3 - Preparing the KHP sample

1. Obtain a 250 mL Erlenmeyer flask.
2. Calibrate your scale (hold button down until the calibration menu is reached. Then hit the on button, wait until it flashes 100 g. Place the 100 g calibration weight on the scale and push the on button again. Wait until the calibration is complete). Be sure to recalibrate the scale if you move it or spill anything on it.
3. Using weighing paper or boat, carefully weigh out about 1.0 grams (between 0.9 and 1.1 grams) of KHP onto the paper. (Be sure to tare the mass of the weigh boat or paper).
4. Carefully pour the KHP into the 250 mL Erlenmeyer flask. Add about 75 mL of distilled water (measured with graduated cylinder) into the Erlenmeyer flask and gently swirl the flask and its contents until the KHP is fully dissolved.
5. Add 3 drops of phenolphthalein indicator solution to the KHP solution and swirl gently to mix.

Part 4 - Titrating the KHP

1. Place the Erlenmeyer flask containing the KHP sample under the buret as shown in Figure 13.1.
2. Place a white sheet of paper under the solution. This will make it much easier to detect the color change.
3. Record the value of the starting volume on your data sheet.
4. Add 1-2 mL volumes of the $Mg(OH)_2$ solution from the buret to the KHP solution, while swirling gently.
5. A faint pink color will appear with each addition. When the pink persists for a second or two start adding the $Mg(OH)_2$ in smaller volumes. The goal of a titration is to have the solution be clear and after adding a single drop of KHP to the solution it will turn a faint pink color. You will be required to be this accurate in order to complete the lab.
6. You may wash down the sides of your container if you need to with distilled water from a wash bottle.
7. When a single drop of the $Mg(OH)_2$ solution causes a pink color to persist for 30 seconds you are done with the titration.
8. Have your instructor check your result to see if it is a good trial. This will be accomplished by adding several drops of an acid solution to your now pink KHP solution. If the color changes back to clear it is a successful trial. If it does not you will need to repeat the trial.
9. Record the final volume on your data sheet.

10. Repeat Parts 4 and 5 until you have a total of 3 good trials approved by your instructor. If you need another data sheet ask your instructor for one.
11. Finish your calculations for each trial and calculate the average molarity of the unknown $Mg(OH)_2$ solution. You will be graded on both the accuracy of your result and the precision.

Results

Be sure to measure as accurately and precisely as possible. Make sure all measurements have the proper uncertainty, significant figures and units. Show work for all calculations.

Practicing with a buret (Part 1)

	Trial 1	**Trial 2**	**Trial 3**
1. Uncertainty in volume measurements:	______________	__________	__________
2. Mass of empty beaker:	__________	__________	__________
3. Initial buret reading:	__________	__________	__________
4. Verify with your instructor:	__________	__________	__________
5. Final buret reading:	__________	__________	__________
6. Verify with your instructor:	__________	__________	__________
7. Volume of distilled water delivered (show calculation):	__________	__________	__________
8. Mass of beaker + water dispensed:	__________	__________	__________
9. Mass of water dispensed (show calculation):	__________	__________	__________
10. Difference between Volume (7) and Mass (9) (show calculation):	__________	__________	__________

1. Does the volume of water measured agree with the mass of the water measured? Explain.

2. List two sources of error that might exist when making these measurements? Explain.

Titration of KHP (Part 4)

	Trial 1	**Trial 2**	**Trial 3**
1. Mass of KHP::	___	___	___
2. Uncertainty in volume measurements:	___	___	___
3. Initial buret reading:	___	___	___
4. Final buret reading:	___	___	___
5. Verify titration (drops of acid added):	___	___	___
6. Volume of $Mg(OH)_2$ delivered (show calculation):	___	___	___

	Trial 4	**Trial 5**	**Trial 6**
1. Mass of KHP::	___	___	___
2. Uncertainty in volume measurements:	___	___	___
3. Initial buret reading:	___	___	___
4. Final buret reading:	___	___	___
5. Verify titration (drops of acid added):	___	___	___
6. Volume of $Mg(OH)_2$ delivered (show calculation):	___	___	___

	Trial 7	**Trial 8**	**Trial 9**
1. Mass of KHP::	__________	__________	__________
2. Uncertainty in volume measurements:	__________	__________	__________
3. Initial buret reading:	__________	__________	__________
4. Final buret reading:	__________	__________	__________
5. Verify titration (drops of acid added):	__________	__________	__________
6. Volume of $Mg(OH)_2$ delivered (show calculation):	__________	__________	__________

	Trial 10	**Trial 11**	**Trial 12**
1. Mass of KHP::	__________	__________	__________
2. Uncertainty in volume measurements:	__________	__________	__________
3. Initial buret reading:	__________	__________	__________
4. Final buret reading:	__________	__________	__________
5. Verify titration (drops of acid added):	__________	__________	__________
6. Volume of $Mg(OH)_2$ delivered (show calculation):	__________	__________	__________

1. What is the molecular weight of KHP. (Remember KHP is an abbreviation for Potassium Hydrogen Phthalate).

2. Write the complete and balance reaction that occurs when KHP reacts with $Mg(OH)_2$.

3. In the space below show a sample calculation for determining the Molarity of your solution for your first "Good" trial.

4. List the Trial Number and Molarity of your 3 best trials. Calculate the Average Molarity (show calculation).

Trial Number: _____ __________

Trial Number: _____ __________

Trial Number: _____ __________

Average Molarity: __________

Post Lab Questions

5. In choosing an acid for the experiment, a solid (KHP) was chosen instead of a liquid (HCl for example). Why would a solid acid be chosen over a liquid acid. Explain?

6. Why is buret considered dirty (and must be cleaned) if water droplets adhere to the inner surface? (or why don't we want to use dirty burets?)

7. Why do we want to eliminate air bubbles in the buret, especially in the area below the stopcock?

8. Why are you allowed to wash down the sides of your Erlenmeyer flask during the titration? (Hint: Why is having no water in the buret important, but adding water to the acid in the Erlenmeyer unimportant.) Explain.

9. A student determined the percent KHP in an impure sample of KHP. A 2.750 g sample of impure KHP required 39.55 mL of 0.1215 M $Mg(OH)_2$ solution for titration. What is the percentage of KHP in the impure sample? (Hint: You can not start with the impure sample of KHP because you can't calculate the molecular weight of the mixture, since you don't know what it is contaminated with. Instead, determine the amount of pure KHP in the sample by using the titration information.) The percent of KHP in the impure sample is given by:

$$\text{Percent of KHP} = \frac{\text{mass pure KHP in sample}}{\text{mass of impure sample}} \times 100\%$$

10. What assumption must be made about the impurity in the previous question.

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Name: ____________________ Class: __________ Date: __________

Prelab Questions

1. Define each of the following terms as they relate to the experiment:
   (a) Titration
   (b) End point
   (c) Indicator
   (d) Standard Solution
   (e) Meniscus
2. What is the molecular weight of KHP. (Remember KHP is an abbreviation for Potassium Hydrogen Phthalate).
3. Write the complete and balance reaction that occurs when KHP reacts with Magnesium Hydroxide.
4. A student standardized a $Mg(OH)_2$ solution, following the procedures outlined in this experiment. The student weighed a samples of KHP (1.350 g) and transferred it to a 250 mL Erlenmeyer flasks. She added distilled water, the indicator, and then titrated the sample with 18.75 mL an unknown concentration of $Mg(OH)_2$ solution. Calculate the Molarity of the $Mg(OH)_2$ Be sure to include the correct units and SF for all answers.

Experiment 14
Water in Hydrates

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Calculate molecular weight of compounds.
2. Calculate percent composition of compounds.

Discussion

Certain ionic compounds form with large spaces in their crystalline structures as shown in Figure 14.1. These spaces are a perfect fit for many small molecules which are trapped there. When water is trapped in the lattice these compounds are often referred to as **hydrates**. The water which is chemically bonded to the hydrate is often called the **water of crystallization** or water of hydration. In a hydrate the water molecules are a distinct part of the compound and are joined to it by bonds that are significantly weaker than those forming the ionic lattice. To indicate this, the formula for hydrates is written with a dot used to separate the formula of the ionic salt from that of the trapped water molecules. The name of the compounds are slightly different also, the ionic part is named as usual, and we use indicate the additional water by giving the number of water molecules (using di, tri, etc.) and the word hydrate. For example $CaSO_4 \cdot 2\,H_2O$ would be called calcium sulfate dihydrate.

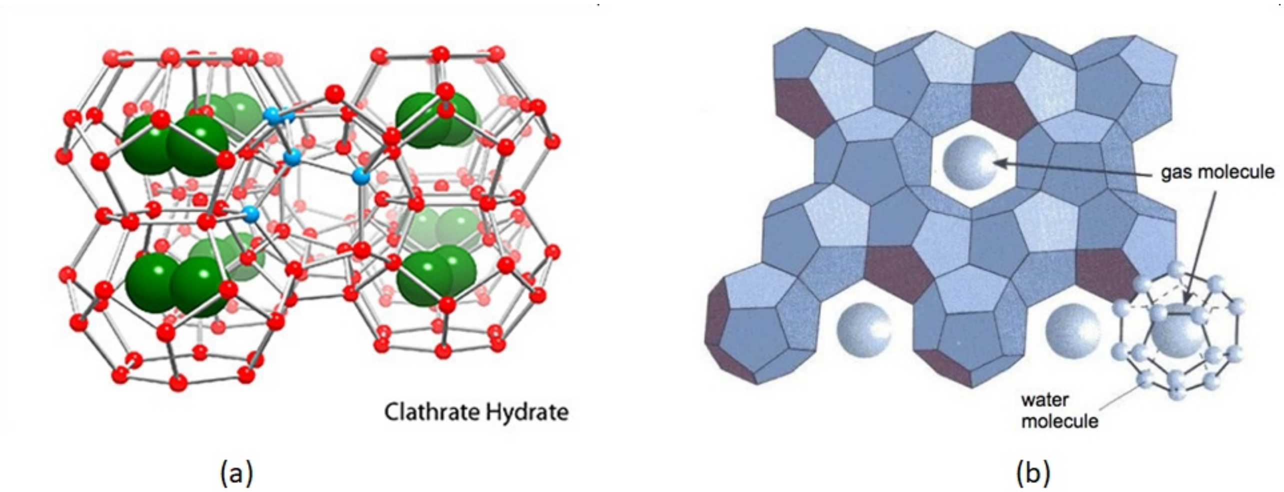

Figure 14.1: (a) and (b) Hydrates are crystalline compounds that have water trapped in the empty space within the crystalline lattice. Similar compounds shown here are chlathrate hydrate structures which are crystalline compounds in gas molecules are trapped in the spaces inside the crystal structure of frozen water. credit: (a) `https://ps.uci.edu/group/kcjanda/research/gas-hydrate-structure` (b) `http://dusk.geo.orst.edu/oceans/deep_currents.html`

These compounds have many industrial uses, as well possibly playing an important role in environmental issues such as global warming. We will briefly discuss two of the many uses for these types of compounds.

Natural gas hydrates (also commonly referred to as: clathrates, gas hydrates, methane hydrates) are solid compounds formed from small gas molecules (natural gas/methane), and water that are formed at low temperatures and high pressures. An example of a clathrate is shown in Figure 14.1. Some estimates predict that the total amount of gas hydrates could be as large as all the known reserves of coal, oil and gas combined, and could last for 1,000 years at current rates of use. Clathrates are crystalline water based solids physically resembling ice, in which small non polar molecules (typically gases) are trapped inside "cages" of hydrogen bonded water molecules. Without the support of the trapped molecules, the lattice structure of hydrate clathrates would collapse into conventional ice crystal structure or liquid water. Most low molecular weight gases (including O_2, H_2, N_2, CO_2, CH_4, H_2S, Ar, Kr, and Xe), as well as some higher hydrocarbons and freons will form hydrates at suitable temperatures and pressures. Clathrate hydrates are not chemical compounds as the sequestered molecules are never bonded to the lattice.

Hand warmers also make use of crystalline compounds with trapped molecules in them. Solid sodium acetate trihydrate ($NaC_2H_3O_2 \cdot 3\,H_2O$) when heated forms a solution of sodium acetate dissolved in the water which was trapped in the solids lattice. This solution is super-saturated and when disturbed will reform the solid compound liberating large quantities of heat. The compound makes excellent, reusable hand warmers because the reaction can be cycled many times. We can write a chemical equation for the reaction as:

$$\mathrm{NaC_2H_3O_2 \cdot 3\,H_2O(s) + Energy \xrightarrow[\text{Boiling Water}]{\text{Heator}} NaC_2H_3O_2 + 3H_2O(l)} \qquad (40)$$

The reaction is of course reversible if water is added to the anhydrous sodium acetate, crystalization occurs as the water is absorbed into the solid lattice resulting in an exothermic reaction occurring which releasing energy/heat (i.e. a hand warmer).

$$\mathrm{NaC_2H_3O_2 + 3H_2O(l) \xrightarrow{\text{Crystallization}} NaC_2H_3O_2 \cdot 3\,H_2O(s) + Energy} \qquad (41)$$

(a) (b)

Figure 14.2: (a) Cold hands? (b) Get some mittens and a reusable hot pack! credit: (a) `https://www.pxfuel.com/en/free-photo-qtjlg` (b) `https://commons.wikimedia.org/wiki/File:Handwaermer12.jpg`

Calculating Molecular Weight and Percent Composition

In the quantitative section of the lab you will measure the mass of water lost by a hyrdated sample upon heating. Using this information you will be able to determine the percentage of water in the sample, and compare it to several known hydrates to determine the identity of an unknown hydrate. In order to do so we will need to calculate the molar mass or molecular weight of each compound and the percentage of water in each compound.

The molar mass/molecular weight of a compound is the atomic mass of all the elements in the compound expressed in grams/mol. For example, the molecular weight of NaCl is the sum of the molar masses of Na and Cl which is 22.99 g/mol and 35.45 g/mol for a total of 58.44 g/mol. For a more complicated example we can calculate the molecular weight of $CaSO_4 \cdot 2\,H_2O$.

Element	# Atoms	Molar Mass (g/mol)	Total Mass (g/mol)
Ca	1	40.08	1(40.08) = 40.08
S	1	32.07	1(32.07) = 32.07
O	6	16.00	6(16.00) = 96.00
H	4	1.008	4(1.008) = 4.032
$CaSO_4 \cdot 2\,H_2O$	-	-	172.2

Figure 14.3: Example of calculating the Molecular Weight (MW) of a compound.

The percent composition of a compound is found by dividing the total mass of each element by the total mass of the compound and multiplying by 100. An example calculation for $CaSO_4 \cdot 2\,H_2O$ in Figure 14.4.

$$\frac{\text{MW of element}}{\text{MW compound}} \times 100$$

Element	# Atoms	Molar Mass (g/mol)	Percent Composition
Ca	1	40.08	$\frac{1(40.08)}{172.2} \times 100 = 23.28\%$
S	1	32.07	$\frac{1(32.07}{172.2} \times 100 = 18.63\%$
O	6	16.00	$\frac{6(16.00)}{172.2} \times 100 = 55.76\%$
H	4	1.008	$\frac{4(1.008)}{172.2} \times 100 = 2.341\%$
$CaSO_4 \cdot 2\,H_2O$	-	-	100.0%

Figure 14.4: (a) Formula for calculating percent composition of an element in a compound, (b) Table showing example calculation for $CaSO_4 \cdot 2\,H_2O$.

If we want to know the percentage of water in the compound than we need to take the molar mass of the water in the compound divided by the total molar mass as shown in the equation below.

$$\frac{\text{total mass of the } H_2O}{\text{total molar mass}} \times 100$$

For $CaSO_4 \cdot 2\,H_2O$ the equation would be:

$$\frac{36.03 \text{ g/mol } H_2O}{172.2 \text{ g/mol } CaSO_4 \cdot 2\,H_2O} \times 100 = 20.92\ \%\ H_2O$$

Procedure

This experiment will consist of two parts, heating a hydrated compound to remove the trapped molecules inside from the lattice leaving us with an anhydrous salt (the compound without water) and an unknown liquid. We will then attempt to characterize the liquid component of the hydrate and prove that it is indeed water. In the second part of the experiment we will quantitatively determine the percentage of water trapped in an unknown hydrate.

Part 1 - Qualitative Determination of Water

In this part of the experiment our goal is to place a finely ground sample of copper (II) sulfate pentahydrate in the bottom of a test tube without having any stick to the sides of the test tube. We will then heat the bottom of the test tube slowly in order to free the molecules trapped in the lattice of the copper (II) pentahydrate. The resulting gas will condense on the sides of the test tube and drip into the receiving test tube. We will then perform a series of tests on the collected liquid to prove it is water.

1. Obtain a medium sized test tube, a ring stand, clamp, erlenmeyer flask, and a bunsen burner.

2. Fold a 2.5 by 20 cm strip of paper lengthwise to form a V-shaped trough or chute.

3. Weigh about 3 grams of copper (II) sulfate pentahydrate and spread it evenly along the length of the trough. Assemble the apparatus as shown below in Figure 14.7. Clamp the test tube so that its mouth is 15-20 degrees above the horizontal. Carefully place the loaded trough containing the copper (II) sulfate pentahydrate into the bottom of the test tube. Tap the paper chute gently if needed, being careful not allow any copper (II) sulfate pentahydrate to adhere to the sides of the test tube.

4. Remove the chute and turn the test tube until it slants downward at an angle of 15-20 degrees below the horizontal (Figure 14.7c). Make sure the copper (II) sulfate pentahydrate does not slide downward, and remains at the bottom of the tube. Place a test tube in an erlenmeyer flask t0 catch the liquid that will condense in the test tube as it is heated.

5. Heat the hydrate gently at first, to avoid excessive splattering, by waving the bunsen burner underneath the solid sample. Gradually increase the rate of heating. Note any changes that occur in the results section.

6. Continue heating and collect the liquid which will condense in the cooler part of the test tube.

7. Heat the sample until the original blue color has disappeared, but **do not** overheat the sample causing the residue to turn black. (At excessively high temperatures (above 600℃) copper (II) sulfate decomposes into sulfur trioxide which is driven off as a gas, and copper (II) oxide which is a black solid.)

8. It may be necessary to heat the entire length of the test tube for a minute or two to drive most of the condensing liquid down the test tube and into the collecting test tube. Be sure to not heat the clamp holding the test tube in place.

9. Allow the test tube and its contents to cool.

(a) (b)

Figure 14.5: (a) Use a folded sheet of paper to place the sample in the test tube being careful not to spill. (b) Heating the test tube to drive off the trapped water, be sure to not heat by the clamp. credit: author.

10. Observe and record the odor of the liquid that has been collected.
11. While the tube is cooling, dry a piece of cobalt chloride test paper by holding it about 25 cm above the bunsen burner. Heat the paper but do not char or ignite it. When properly dried the test paper should be a blue color. Using a clean stirring rod, place a drop of the collected liquid on the dried cobalt chloride test paper. Record your observations
12. For comparison, place a drop of distilled water on a second piece of dried cobalt chloride test paper. Record your observations.
13. Empty the anhydrous salt residue onto a watch glass, and divide it into two portions.
14. Place one portion in a test tube, and carefully add 3-4 drops of the liquid collected earlier to the anhydrous salt. Record your observations.
15. Place the second portion into a test tube, and carefully add 3-4 drops distilled water to the anhydrous salt. Record your observations.
16. Dispose of the solid residue in the waste container labeled “Copper Waste” or “Cu Waste”.

Part 2 - Quantitative Determination of Water in a Hydrate

To insure that all of the water in the sample has been driven off, we will use a technique known as **heating to constant weight**. Essentially, we will heat and weight the sample until the mass of the sample does not change between heatings indicating that all of the molecules trapped in the lattice have been driven off. Due to limited time, constant weight in this experiment will be achieved when the sample is heated and weighed in successive trials until the weight differs by no more than 0.05 grams between two weighings. This may take 2-3 trials to occur.

An example would be if a student made several heatings as shown in the table below. He would only be able to stop at the last heating because the difference between the 3rd and 4th trials was less than 0.05 grams.

Trial	Initial Mass of Sample	Difference in Mass	Conclusion
Initial Weight	4.000 g	-	Start heating
Mass after 1st Heating	3.500 g	-	Always do a 2nd heating
Mass after 2nd Heating	3.250 g	0.250 g	$0.250 > 0.05$ therefore keep heating
Mass after 3rd Heating	3.150 g	0.100 g	$0.100 > 0.05$ therefore keep heating
Mass after 4th Heating	3.125 g	0.025 g	$0.025 < 0.05$ therefore we are done!

Figure 14.6: An example of heating to a constant mass. Only when the difference in mass between two trials is less than 0.05 is the sample considered dry.

1. Obtain a sample of an unknown hydrate from your instructor. Record its identity in the results section.

2. Obtain a crucible and cover from your instructor. Clean and dry it. Record the mass of the crucible and cover.

3. Place between 2 and 3 grams of your unknown into the crucible. Place the cover on top and weigh the crucible + sample.

4. Obtain a bunsen burner, ring stand, ring clamp, clay triangle, tongs, and construct the apparatus shown in Figure **??**.

5. Place the covered crucible on the clay triangle and adjust the cover so that it is slightly ajar to allow the water vapor to escape.

6. Heat the crucible very gently for around 5 minutes by waving the flame gently underneath. Be careful not to splatter the sample.

7. After 5 minutes readjust the flame so that a sharp inner-blue cone is formed. Place the bunsen burner underneath the sample and heat for 12 minutes with the tip of the inner-blue cone just touching the bottom of the crucible. The crucible bottom should be come a dull red color during this period.

8. Recall “Hot things are hot!” Handle the crucible with the crucible tongs. It will prove necessary to remove the cover first and place it on a ceramic pad. Then place the crucible on the pad, and place the cover on it again.

9. After the first heating is completed, close the cover and allow the crucible to cool for 10 minutes (or longer if required).

10. **The crucible should be cool to the touch before placing it on the scale to weigh it.**

11. Weigh the cooled crucible.

12. To determine if all of the water in the sample was removed during the initial heating, reheat the covered crucible and sample for an additional 5 minutes at maximum heat. Allow the sample to cool for 10 minutes (or longer if required) and weight the sample. The results of the last two

weighing should agree within 0.05 grams. If the decrease in mass is larger than this, repeat the heating and weighing until the results of two successive weighings are within 0.05 grams of each other.

13. Calculate the percentage of water in your sample on the basis of the final weight measured.

14. Perform 2 trials if time permits (it should) and average the results.

15. Dispose of the solid residue in the waste container labeled “Unknown Heavy Metal Waste” or ‘’Unknown Hydrate Waste”.

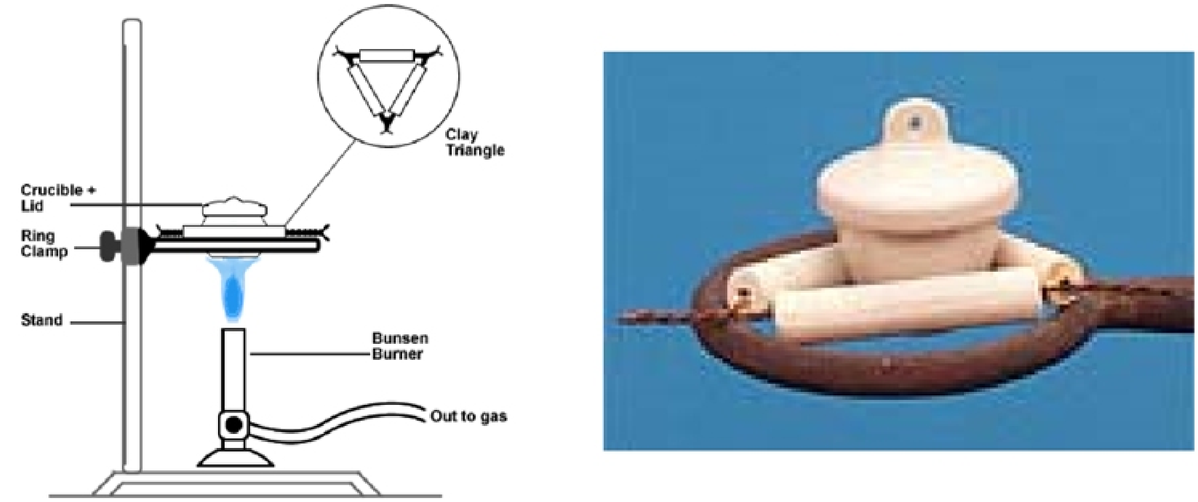

Figure 14.7: Setup for heating a crucible. Leave the cover ajar, to allow gas to escape but larger particles can not. credit: " The Composition of Potassium Chlorate (Experiment)" by Santa Monica College, LibreTexts is licensed under CC BY-NC .

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Name: ______________________ Class: ________________ Date: ________________

Results

A. Qualitative Determination of Water in a Hydrate

Property Observed	**Before Heating**	**After Heating**
Describe the appearance of the solid:		
Property Observed	**Unknown (Collected) Liquid**	**Water**
Describe the appearance of the liquid:		
Describe the odor of the liquid:		
Property Observed	**Unknown (Collected) Liquid**	**Water**
Original Color of Cobalt Chloride Paper:		
Color after adding liquid:		
Temperature change of anhydrous solid:		
Color change in anhydrous solid after adding liquid:		

1. What evidence did you see that indicated the liquid obtained above from the copper (II) sulfate pentahydrate was water. (2 answers)

2. What evidence of a chemical reaction occurring was observed when the unknown liquid, and distilled water were added to the anhydrous salt. (2 answers)

B. Quantitative Determination of Water in a Hydrate

	Trial 1	Trial 2
1. Unknown Identity:	________	________
2. Mass of crucible and cover:	________	________
3. Mass of crucible, cover and sample:	________	________
4. Mass of crucible, cover, and sample after 1st heating:	________	________
5. Mass of crucible, cover, and sample after 2nd heating:	________	________
6. Mass of crucible, cover, and sample after 3rd heating:	________	________
7. Mass of the original sample (show calc.):	________	________
8. Total mass lost by the sample during heating (show calc.):	________	________
9. Percentage of water in the sample (show calc.):	________	________

10. Average percentage of water in the sample (show calc.): ________

11. What is the chemical formula of your unknown: ________

Explain how you identified your unknown:

Post Lab Questions

1. Why do we want to avoid getting any copper (II) sulfate pentahydrate onto the sides of the test tube when we heated the sample to collect the molecules trapped inside the lattice?

2. Write a complete (balanced, states etc) chemical equation for the decomposition of copper (II) sulfate pentahydrate assuming it decomposes into copper (II) sulfate and dihydrogen monoxide.

3. Write the complete (balanced, states etc) chemical reaction (including states) for the decomposition of copper (II) sulfate at high temperatures.

4. When the unknown was heated, could the decrease in mass have been partly due to the loss of some other substance than water? Explain.

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Name: ____________________ Class: ________________ Date: ________________

Prelab Questions

1. Complete the following table by filling in the missing name, formula, and molecular weight of the hydrates. Show work for one of the compounds on the back of the page.

Name	**Formula**	**Molecular Weight (g/mol)**	**% Water**
Calcium sulfate dihydrate		172.1816	
	$CoCl_2 \cdot 6\,H_2O$		
	$MgSO_4 \cdot 7\,H_2O$		
Sodium carbonate decahydrate		286.148	
Barium chloride dihydrate		244.2616	
	$CuSO_4 \cdot 5\,H_2O$		
	$SrCl_2 \cdot 6\,H_2O$		

2. A student heated a hydrated salt sample with an initial mass of 4.8702 grams. After the first heating the mass had decreased to 3.0662 grams. If the student is using the method of **heating to a constant weigh** described in this experiment, what is the minimum mass that the sample can have after the second heating to avoid having to heat it a third time? Explain.

2. ____________

3. The student in the previous problem determined that the mass lost by the sample was 1.8053 grams. What was the percent water in the original hydrated sample? Explain.

3. ____________

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Experiment 15
Molecular Compounds and Lewis Structures

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Draw Lewis structures.
2. Count valence electrons.
3. Predict shape and bond angles around atoms.
4. Determine if a molecule is polar or non-polar.
5. Memorize different shapes/bond angle combinations.

Discussion

All atoms are formed of a nucleus containing protons and neutrons surrounded in space by electrons which are held within specific regions of space by the attractive force of the protons. An early theory for predicting the formation, and structure of molecular compounds was formulated by Lewis. The theory says that the outermost electrons in an atom, often referred to as **valence electrons** are involved in bonding atoms together to form compounds. The valence electrons for the representative elements are the sum of the s and p electrons in the outermost shell (largest principle quantum number), and is also the same as the group number on most periodic tables.

Lewis Dot Structures

Lewis electron dot structures or simply **Lewis structures** are a useful construct to keep track of valence electrons in representative elements. In this notation the valence electrons are represented by dots surrounding the atomic symbol of an element. Several examples are shown below in Figure 15.1. Ions are shown in brackets with the corresponding charge. The formation of ionic compounds will be discussed in lecture.

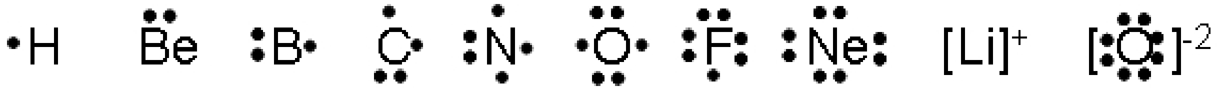


Figure 15.1: Example Lewis dot structures for several atoms, and ions. Note that cations and anions are drawn in brackets and show the charge of the ion. credit: author

The formation of molecular bonds, and the sharing of electrons is driven by the desire for atoms to achieve the **noble gas configuration**. From quantum mechanics this is represented as having filled orbitals or a s^2p^6 configuration like that of the noble gases. This is often referred to as the **octet rule**, meaning that all atoms (except H and He) want to have 8 electrons in the outermost orbital. Ionic compounds achieve this by cations losing electrons and anions gaining electrons, but molecular compounds are forced to share the electrons in order to achieve octets. The octet rule only a general guideline, and breaks down when considering d-orbitals, and in several other cases as discussed in lecture and in your book. The most important exception is for hydrogen, which only requires two electrons to achieve a noble gas configuration (He).

A Lewis structure for molecular compounds is a 2D representation in which electrons that are shared between two atoms are represented as a single line connecting the atoms. If multiple pairs of electrons are shared they are represented by multiple lines between the atoms. Unshared or **lone-pair electrons** are represented by dots located around the atom. For polyatomic ion, the rules are the same except that the group of atoms is enclosed in brackets and the overall charge of the ion is shown. Figure 15.2 shows several examples.

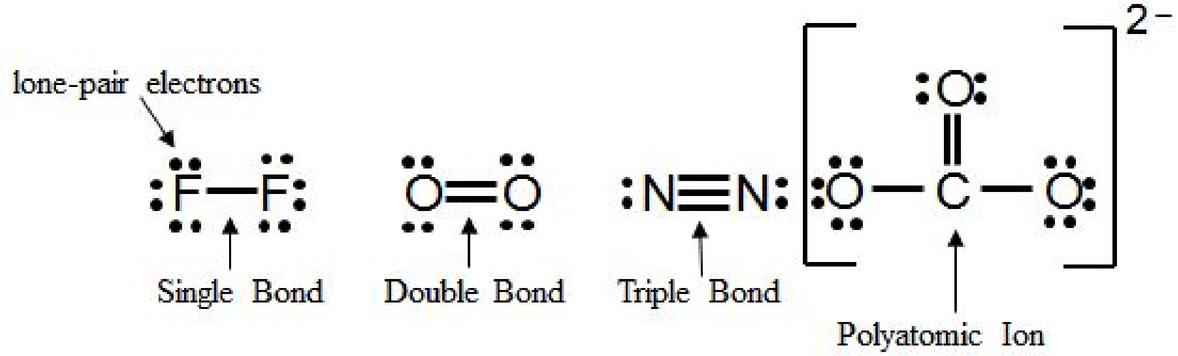


Figure 15.2: Example Lewis dot structures for several molecules illustrating single, double, triple bonds and polyatomic ions. credit: author

Molecular Model Building (3D Models)

The 3D structure of molecules is often difficult to visualize from a 2D Lewis structure. In order to understand the true 3D shape of molecules molecular model kits will be used to create 3D models. This will make it easier to see the common geometric patterns which Lewis theory predicts molecules will form.

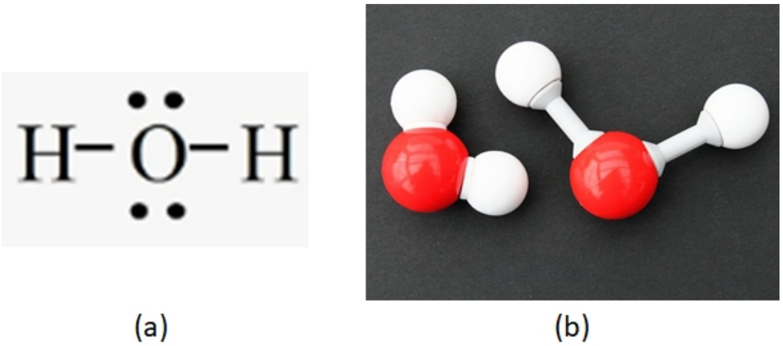


Figure 15.3: (a) Lewis structure of water (H_2O). (b) 3D model of water. Note that the bond angles in the drawn lewis structure do not match the bond angle in the 3D model. credit: (a) author (b) `https://commons.wikimedia.org/wiki/File:H2O_Kalottenmodell_und_St%C3%A4bchenmodell_8127.JPG`

Atoms in molecules or polyatomic ions are arranged into geometric shapes which allow the electron pairs to remain as far apart as possible in order to minimize the repulsive forces between them. The underlying theory is call **valence shell electron pair repulsion (VSEPR)** theory. For well behaved molecules that obey the octet rule there are six basic shapes molecules can assume as shown in Figure 15.4.

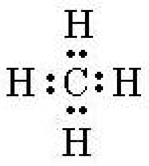
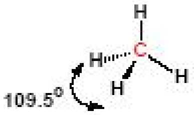
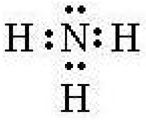
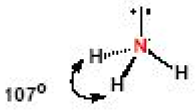
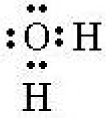
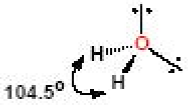
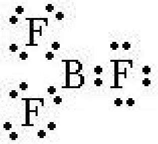
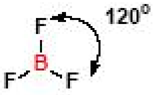
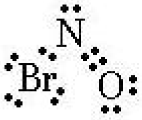
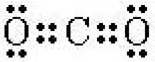
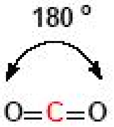
Lewis Structure	# atoms bonded	# lone pairs	Molecular Shape	Bond Angle	Molecular Polarity	3D Structure
	4	0	Tetrahedral	109.5	Non-polar or Dipolar	
	3	1	Trigonal Pyramidal	109.5	Dipolar	
	2	2	Bent - 109.5	109.5	Dipolar	
	3	0	Trigonal Planar	120	Non-polar or Dipolar	
	2	1	Bent - 120	120	Dipolar	
	2	0	Linear	180	Non-polar or Dipolar	

Figure 15.4: Six basic shapes for Lewis Structures using s and p electrons only and obeying the octet rule. credit: author

Bond angles (Figure 15.5) always refer to the angle formed between two end atoms with respect to a central atom. If there is no central atom there is no bond angle. The size of the angle depends mainly on the repulsive forces between electron pairs around the central atom. According to VSEPR theory the atoms and electrons around the central atom try to remain as far apart as possible. The bond angles determined are estimates only, and the real bond angles can differ by several degree's depending on the molecule studied.

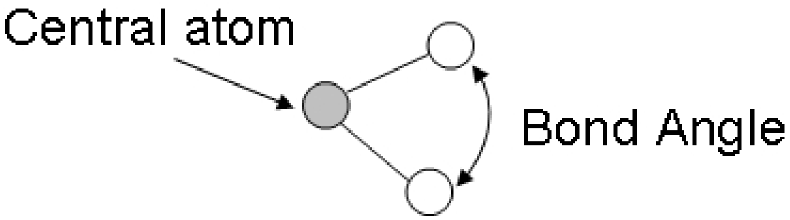
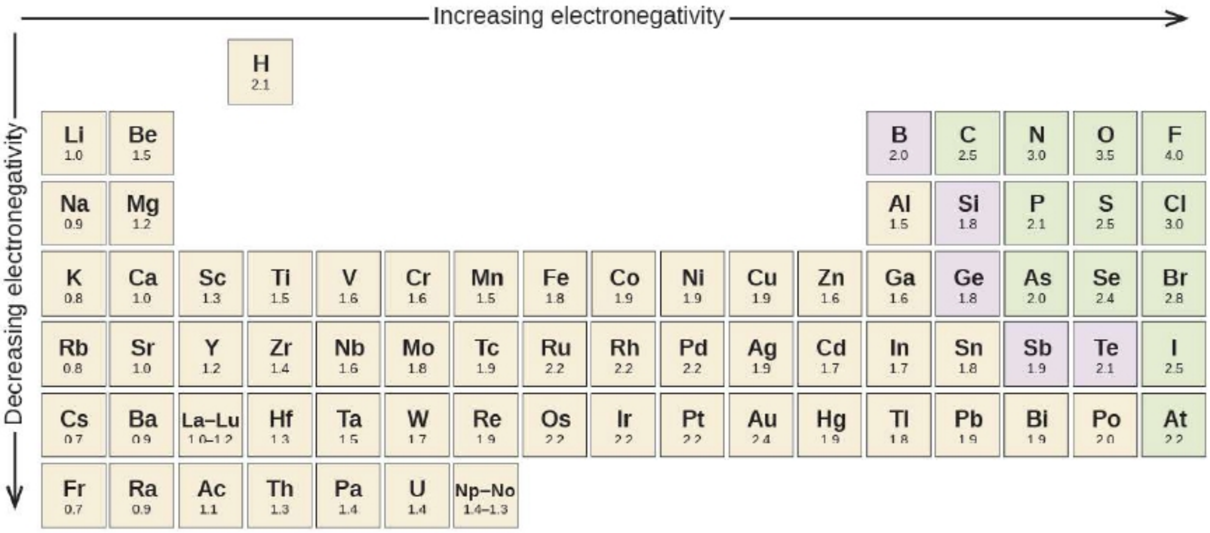


Figure 15.5: Bond angles occur when three or more atoms are bonded together to form a molecule. credit: author

Bond Polarity

Electrons in molecular compounds are shared between two atoms to form bonds. For atoms that are alike (diatomics) the sharing is equal, but for most other molecules the sharing is unequal. **Electronegativity** is the attractive force an atom has for the shared electrons in a bond. Electronegativity values are assigned to elements, and can be found in Figure 15.6. In general electronegativity increases as we move across a row, and decreases as we move down a column.



			H 2.1													
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La–Lu 1.0–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np–No 1.4–1.3										

Figure 15.6: Electronegativity values for different elements. The larger the value the more electronegative the element is and the stronger the attractive force for the shared electrons in a molecular bond. credit: Rice University Openstax CC BY NC SA `https://openstax.org/books/chemistry-2e/pages/7-2-covalent-bonding`

Atoms that share the electrons in a bond equally are called **non-polar covalent bonds or covalent** bonds, while those that are shared unequally are called **polar covalent** bonds. Bond polarity in atoms can be indicated by using the greek symbols $\delta+$ to indicate a small excess of positive charge, or $\delta-$ to indicate a small excess of negative charge from the unequal sharing. Another common method is to use a modified line for the bond with an arrow pointed toward s the more electronegative atom, and a small cross toward s the more electropositive atom. Figure 15.7 shows an example of each method.

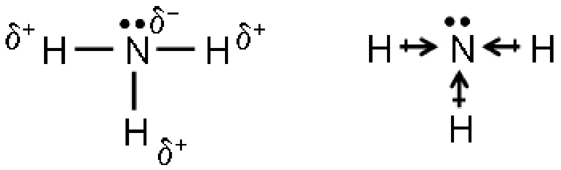


Figure 15.7: Bond polarity in an ammonium molecule. Nitrogen is more electronegative than hydrogen therefore electrons spend more time closer to the nitrogen than hydrogen resulting in the formation of partial charges. credit: author.

Molecular Dipoles

Just as individual bonds in molecules can be polar and non-polar, molecules as a whole are often polar because of the net sum of individual bond polarities and lone-pair contributions in the molecule. The resulting **molecular dipoles** can be thought of as the center mass of all positive charges being different than the center of mass for all negative charges. Another way of looking at it is a “tug of war” between the positive and negative ends of the polar bonds, if polar bonds tug in opposite directions as shown in Figure 15.8 then the molecule is considered **nonpolar**, but if the polar bonds align, or do not cancel out then there is a net dipole and we consider the molecule to be **dipolar** as shown in Figure 15.8.

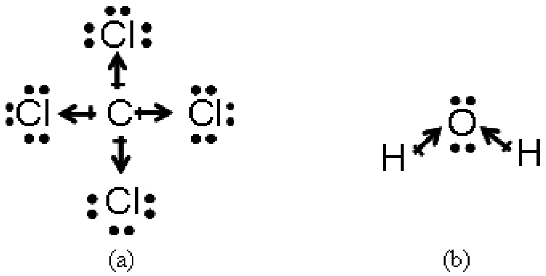


Figure 15.8: Molecular Dipole occur when there is an unsymmetrical distribution of electrons around an atom. (a) Nonpolar molecule due to symmetry (b) Dipolar molecule. credit: author

Drawing Lewis Structures

Drawing Lewis structures takes time and practice, and there is no single set of steps that will always yield the correct answer. Expect to occasionally draw several incorrect models before you find the correct one. Learn from each incorrect model what does and does not work, and apply it to drawing future Lewis structures. The general rules below will generally lead to the correct structure in one or two iterations.

1. Add up the valence electrons
   - (a) Add up the valence electrons for all regular atoms (s and p orbitals with the highest quantum number)
   - (b) Add electrons for molecules with a negative charge (ex: CO_3^{-2})
   - (c) Subtract electrons for molecules with a positive charge (ex: NH_4^+)
2. Write a trial structure
   - (a) Place the least electronegative atom in the center
   - (b) Carbon is generally a central atom and forms bonds with itself frequently
   - (c) Make molecules as symmetrical as possible
   - (d) Hydrogen has only one valence electron, and can only form one bond and is therefore never the central atom
   - (e) Draw one bond between all atoms
   - (f) Typical bond numbers formed (H = 1, O = 1 or 2, C = 4, N = 1, 2, 3 or 4)
   - (g) Oxygen rarely bonds to another oxygen (except for peroxides), instead forming single or double bonds to other atoms
   - (h) F, Cl, Br, and I generally form 1 bond (but not always)
3. Count electrons - Subtract 2 electrons for every bond formed
4. Distribute the remaining electrons to give noble gas configurations (octet rule)
   - (a) Surround each atom with 8 electrons (except H)
   - (b) Start with the most electronegative atoms first
   - (c) If all atoms have 8 electrons around them you are done, if not remove unshared electron pairs from outer atoms and form double and triple bonds

Building 3D Models

Use the ball and stick kits provided in class to build 3D models of the molecules after you have drawn the Lewis structures. The balls are color coded as shown in Figure 15.9.

Ball/Stick	**Use**
Black (4 holes)	Carbon - tetrahedral
Black (3 holes)	Carbon - trigonal planar
Red (2 holes)	Oxygen
Green (1 hole)	Halogens
White (1 hole)	Hydrogen
Light Blue (4 holes)	Nitrogen
Inflexible bonds	Single bonds, and lone pair electrons
Flexible bonds	Double and Triple Bonds

Figure 15.9: Ball and stick model parts used in class.

Procedure

For each molecule or polyatomic ion complete the following:

1. Calculate the number of valence electrons.
2. Draw a Lewis Structure.
3. Build a 3D model of the structure. Have your model checked by the instructor.
4. Draw a line pointing to each atom (there may be cases where different atoms have different geometry) that has a molecular geometry and label it with both the geometry (Tetrahedral, Trigonal Pyramidal, Bent-109, Trigonal Planar, Bent-120, and Linear) and bond angle.
5. Using your model, and the information on bond polarity, determine if the molecule as a whole is nonpolar (NP) or dipolar (DP).
6. Answer all questions at the end of the lab.

Results

CH_4		Valence e-:	**CS_2**		Valence e-:
		Polarity: NP or DP			Polarity: NP or DP
		Instructor OK:			Instructor OK:
H_2S		Valence e-:	**N_2**		Valence e-:
		Polarity: NP or DP			Polarity: NP or DP
		Instructor OK:			Instructor OK:
SO_4^{-2}		Valence e-:	**SCl_3^{+1}**		Valence e-:
		Polarity: NP or DP			Polarity: NP or DP
		Instructor OK:			Instructor OK:
H_2CO_3		Valence e-:	**CH_3Cl**		Valence e-:
		Polarity: NP or DP			Polarity: NP or DP
		Instructor OK:			Instructor OK:

C_2H_6	Valence e-:	**SO_3^{-2}**	Valence e-:
	Polarity: NP or DP		Polarity: NP or DP
	Instructor OK:		Instructor OK:
C_2H_4	Valence e-:	**CH_2O**	Valence e-:
	Polarity: NP or DP		Polarity: NP or DP
	Instructor OK:		Instructor OK:
OF_2	Valence e-:	**NO_2^-**	Valence e-:
	Polarity: NP or DP		Polarity: NP or DP
	Instructor OK:		Instructor OK:
O_2	Valence e-:	**C_3H_4**	Valence e-:
	Polarity: NP or DP		Polarity: NP or DP
	Instructor OK:		Instructor OK:

HPO$_4^{-2}$		Valence e-:	**HNO$_3$**		Valence e-:
		Polarity: NP or DP			Polarity: NP or DP
		Instructor OK:			Instructor OK:

1. There are three acceptable Lewis structures for $C_2H_2Cl_2$. Draw all three structures below. What geometry does each molecule have? Label each as being nonpolar or dipolar.

2. For the molecules drawn above, one is nonpolar and the other two are dipolar. Explain how this occurs.

3. There are also three Lewis structures for $[NO_3]^-$. Draw all three structures below. What geometry does each molecule have? Label each as being nonpolar or dipolar. What property of Lewis Structures does this illustrate?

.

Name: ______________________ Class: ______________ Date: ______________

Prelab Questions

1. Complete the following table by filling in the missing blanks.

Atom	**# Valence e^-**	**Electronegativity**	**Model Color**
Hydrogen			
Carbon			
Nitrogen			
Oxygen			
Fluorine			

Table 15.1: Molecular Model Basics

2. What is the octet rule? What common atom is allowed to violate the octet rule? Explain.

3. What is the difference between Polar Covalent and Non-polar Covalent bonds?

,

Experiment 16
Properties of Solutions

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Terminology - solution, solute, solvent, precipitate, miscible, immiscible, soluble, insoluble.
2. Effect of temperature on solubility.
3. Effect of particle size on solubility.
4. Use of ratios/proportions to solve solubility problems.

Discussion

Solutions are an important class of materials we each encounter daily. This experiment will probe the physical properties of solutions. A complete discussion of solutions can be found in Hein Chapter 14, and review of the properties of mixtures can be found in Hein Chapter 3. In addition we will review our graphing skills and how to use information in graphs and equations to perform calculations.

Solutions are mixtures of two or more substances. The simplest solution is composed of a **solute** dissolved in a **solvent** as shown in Figure 16.3. The solute is the substance that is least abundant in the solution, while the solvent is the substance that is most abundant. Solutions are commonly named after the solute, for instance when 10.0 g of potassium chloride is dissolved in 100 mL of water, we would refer to the solution as a potassium chloride solution.

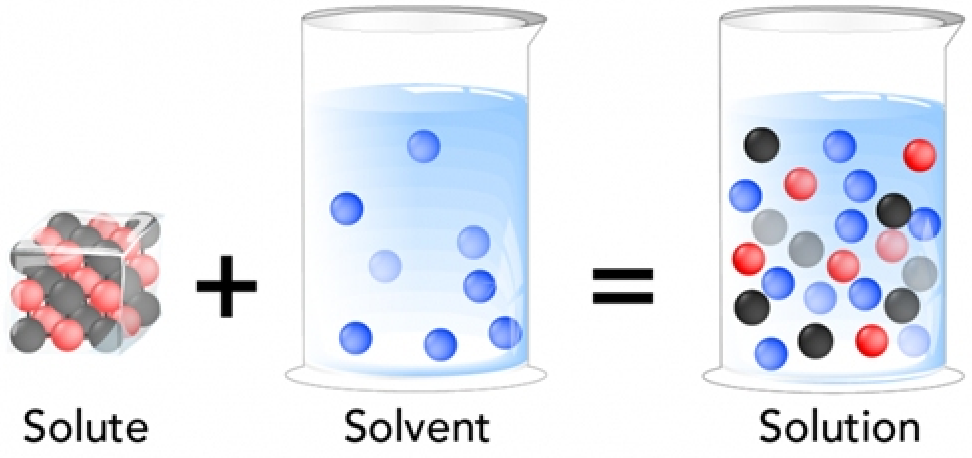
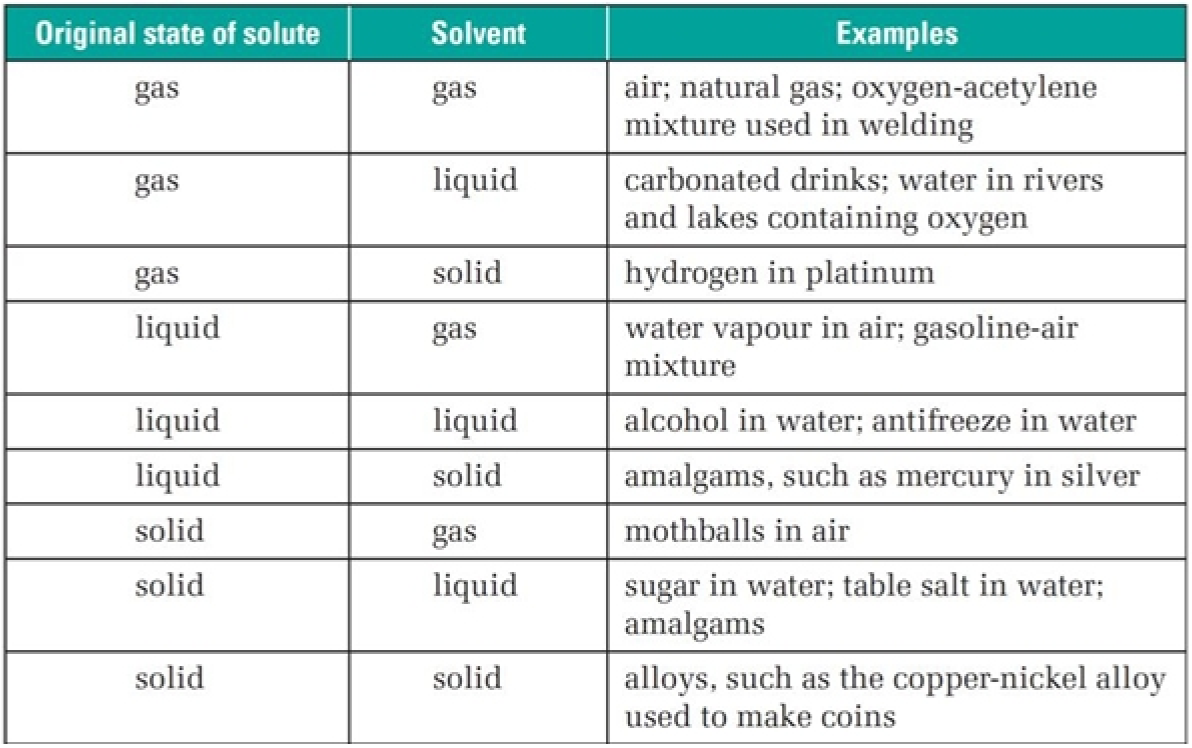


Figure 16.1: A solute is dissolved in a solvent to create a solution. credit: unknown (modified by author)

Mixtures are homogeneous, the solute is uniformly distributed throughout the solvent and will not form separate phases. Mixtures may have variable compositions, meaning the amount of solute dissolved in the solvent can vary. The particles of a solution are generally small molecules or ions, larger molecules tend to make colloids.

Solutions can be formed from any combination of the normal phases of matter (solid, liquid, and gas) resulting in nine possible combinations. The most commonly encountered types of solution in chemistry lab being solids in liquids or liquids in liquids.



Original state of solute	Solvent	Examples
gas	gas	air; natural gas; oxygen-acetylene mixture used in welding
gas	liquid	carbonated drinks; water in rivers and lakes containing oxygen
gas	solid	hydrogen in platinum
liquid	gas	water vapour in air; gasoline-air mixture
liquid	liquid	alcohol in water; antifreeze in water
liquid	solid	amalgams, such as mercury in silver
solid	gas	mothballs in air
solid	liquid	sugar in water; table salt in water; amalgams
solid	solid	alloys, such as the copper-nickel alloy used to make coins

Figure 16.2: Nine different types of solutions. credit: unknown

Solutions are important in chemistry because many reactions will only occur at a meaningful rate in solutions. For instance solid sodium chloride and silver nitrate will not react, while a solution of sodium chloride and silver nitrate will react instantly to produce a precipitate.

$$\mathrm{NaCl(s)} + \mathrm{AgNO_3(s)} \longrightarrow \mathrm{No\,Reaction} \tag{42}$$

$$\mathrm{NaCl(aq)} + \mathrm{AgNO_3(aq)} \longrightarrow \mathrm{NaNO_3(aq)} + \mathrm{AgCl(s)} \tag{43}$$

The rate of a reaction is directly proportional to the number of molecules that can interact, and how much kinetic energy each molecule has, thus reactions between solids are slow (very few molecules can interact, and low energy) compared to those between solutions (large numbers of molecules can interact, higher energy).

Solids in Liquids

The rate at which a solid solute dissolves in a liquid solution depends on a number of factors. Solids can only dissolve at the surface that is in contact with the solution, thus factors which effect this can cause a solid to dissolve at different rates. We will examine each of these factors in turn:

- Particle size

- Temperature
- Concentration of solute
- Agitation or stirring.

Solubility is defined as the maximum amount of solute that will dissolve in a specific amount of solvent under stated conditions (normally temperature). Solubility is often expressed qualitatively as being **soluble** (will form an aqueous solution), slightly soluble (small amounts will form an aqueous solution), or **insoluble** (the solute will not dissolve in the solution). Solubility can also be expressed quantitatively as the grams of solute that will dissolve in 100 grams of water. Table 16.1 shows the solubility data for four different ionic salts in water.

Temperature (℃)	KCl (g)	NaCl (g)	KBr (g)	$BaCl_2$ (g)
0.0	27.6	35.7	53.5	31.6
10.	31	35.8	59.5	33.3
20.	34	36	65.2	35.7
30.	37	36.3	70.6	38.2
40.	40.	36.6	75.5	40.7
50.	42.6	37	80.2	43.6
60.	45.5	37.3	85.5	46.6
70.	48.3	37.8	90	49.4
80.	51.1	38.4	95	52.6
90.	54	39	99.2	55.7
100.	55.6	39.8	104	58.8

Table 16.1: Solubility Data for Four Salts (g solute/100 g water).

Liquids in Liquids

To qualitatively describe the solubility of two liquids we use the term **miscibility**. A **miscible** solution is one in which the liquids are soluble in each other and will mix completely, while the term **immiscible** is used to describe two liquids that are insoluble in each other, thus will not mix, and will show a visible line separating the two liquids (or phases). An example of liquids that are miscible would be ethanol and water (beer), while the phrase "oil and water don't mix" describes two substances that are immiscible.

Concentration - Qualitative Descriptions

The **concentration** of a solution is a measure of how much solute is dissolved in a solution. It can be given both qualitatively and quantitatively.

Qualitatively a **dilute** solution contains relatively small amounts of solute in a given volume, while a **concentrated** solution contains relatively large amounts of solute in a given volume of solution. The terms are not well defined, and are normally only used to distinguish between two solutions.

A solution can also be described as saturated, unsaturated or supersaturated. A **saturated** solution contains the maximum possible amount of solute dissolved in a solution at a given temperature and pressure. For example a solution consisting of 42.6 grams of KCl dissolved in 100 grams of water at

50℃ would be a saturated solution. A saturated solution is an **equilibrium** solution in which the rate at which solute dissolves is equal to the rate at which solute precipitates from the solution.

$$\text{Solute(solid)} \rightleftharpoons \text{Solute(aq)} \tag{44}$$

A **unsaturated** solution is one in which less than the maximum amount of solute is dissolved in the solution. For example at 50℃ any amount of KCl less than 42.6 grams dissolved in 100 grams of water would be unsaturated.

A **supersaturated** solution contains more than the maximum amount of solid dissolved in the solution. This may seem like a contradiction of the term “maximum amount”, but a supersaturated solution is an inherently unstable state, and will form a saturated solution (with excess solid precipitating out) if disturbed. An example of this would be making Jello, a solution of Jello in water is made at a high temperature and then allowed to cool, as the solution cools, the solubility of the Jello decreases, a supersaturated solution forms, and if disturbed will form a solid.

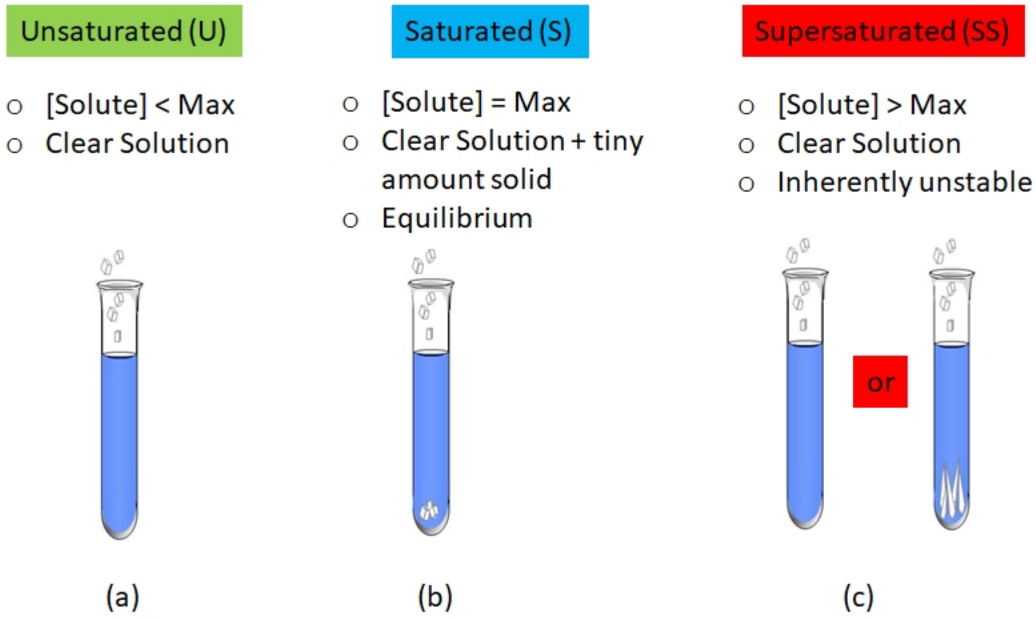


Figure 16.3: The difference between unsaturated, saturated and supersaturated solutions. credit: author

Concentration - Quantitative Descriptions

Quantitatively the amount of solute dissolved in solution can be expressed in a number of different ways depending on what is being described. The most common expressions are given below.

Mass-percent solution is defined as the mass of solute divided by the mass of the solution. For example, a 10% solution of NaCl in water contains 10 grams of NaCl in 90 grams of water, or 10 grams of NaCl in 100 grams of solution.

$$\text{Mass Percent} = \left(\frac{\text{grams solute}}{\text{grams solute} + \text{grams solvent}}\right) * (100) = \left(\frac{\text{grams solute}}{\text{grams solution}}\right) * (100) \tag{45}$$

Mass per 100 grams solvent is similar to Mass-percent, but is expressed in terms of 100 grams of solvent not solution. Thus, the 10 mass-percent solution described above, the Mass per 100 grams solvent would be 11 grams NaCl per 100 grams of H_2O.

$$\left(\frac{\text{10 grams NaCl}}{\text{90 grams H}_2\text{O}}\right) * (100) = 11 grams = \frac{\text{11 grams NaCl}}{\text{100 grams H}_2\text{O}} \tag{46}$$

Molarity is the most common quantitative expression of concentration used in chemistry. Molarity is defined as the moles of solute per liter of solution. The concentration of 10 grams of NaCl in 90 grams of water discussed above would produce a 1.90 M solution of NaCl.

$$\text{Molarity} = \text{M} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}} \tag{47}$$

Procedure

The experiment is broken apart into several sections. The sections may be completed in any order.

A. Determining the Concentration of a Saturated Solution

In this section you will determine the concentration of KCl in a saturated solution and compare it to the theoretical value. Make all measurements to the highest precision possible. Perform only one trial, use the second trial only if you make a mistake in trial 1. The method used in this section is refereed to as **"heating to a constant mass"**, and involves heating the sample multiple times to ensure that all of the water is removed.

1. Prepare a water bath using a 400 mL beaker 3/4 full of water on which you will place an evaporating dish. Before you heat the water, make sure the evaporating dish will fit on top of the beaker (and won't fall in).
2. Record the temperature of the room.
3. Weigh an empty evaporating dish. Obtain 5-6 mL of saturated potassium chloride solution and pour it into the dish. Weigh the dish + potassium chloride solution. Record all masses on your data table.
4. Place the evaporating dish on the beaker of water. Boil the solution until all of the potassium chloride solution has evaporated to dryness (it should take 30-45 minutes). You may need to add water to the beaker (**DO NOT** let it boil dry, you will shatter your beaker and receive a zero for this lab). You may want to periodically break the crust formed over the solution of potassium chloride to speed evaporation up.
5. While evaporating perform the other sections of the lab.
6. When all of the liquid in the potassium chloride solution in the evaporating dish has evaporated cool the dish.
7. When cool, wipe any excess water off the bottom of the dish with a paper towel. Place the dish on the hot plate and heat gently for an additional 1-2 minutes to evaporate any last traces of water. Do not heat strongly because your sample might splatter.
8. Allow your sample to cool and weigh it.
9. Reheat your sample on the hot plate for an additional 1-2 minutes, and allow the sample to cool and reweigh it.
10. When finished allow the dish to cool, dissolve the remaining potassium chloride in water, and dispose of it in the sink.
11. Show calculations for lines 5-12.

B. Solubility versus Temperature (Comparing Saturated and Unsaturated Solutions)

1. Label 4 weighing boats or papers and weigh out the following: 1.0 g NaCl, 1.4 g NaCl, 1.0 g NH_4Cl and 1.4 g NH_4Cl.
2. Add 1.0 grams of NaCl and NH_4Cl to **separate** test tubes labeled A and B. Add 5.0 mL of distilled water to each. Mix each test tube for 2 minutes. Record your observations by circling the appropriate answers.
3. Add 1.4 grams of NaCl to test tube A, and NH_4Cl to test tube B. Mix each test tube for 4 minutes. Record your observations.
4. Place test tube A and B into a beaker of boiling water. Stir occasionally (caution, hot things are hot). Note the results after 5 minutes.
5. Remove the tubes and run under tap water for 2 minutes. Let stand for 5 minutes. Record your observations.
6. Dispose of the solutions in the sink.

C. Relative Solubility of a Solute in Two Solvents

1. Add 2 mL of heptane and 5 mL of water to a "large" test tube. Mix vigorously for 30 seconds. Record your results. Wait 2 minutes, and record your results.
2. To the previous test tube add 5 drop of saturated iodine-water solution. Note the color of each layer. Mix the contents for 30 seconds. Record your results. Wait 2 minutes, and record your results.
3. Dispose of the contents of the test tube in the waste bottle marked “Volatile Wastes”.

D. Miscibility of Liquids

1. Obtain 3 dry test tubes.
2. In test tube 1 mix 1 mL of kerosene and 1 mL of isopropyl alcohol.
3. In test tube 2 mix 1 mL of kerosene and 1 mL of water.
4. In test tube 3 mix 1 mL of water and 1 mL of isopropyl alcohol.
5. Mix each of the three test tubes for 15 seconds. Note which pairs are miscible and which are not.
6. Dispose of the contents of the test tubes in the waste bottle marked “Volatile Wastes”.

E. Rates of Dissolving - Particle Size

1. Obtain two dry test tubes. Fill the first test tube with 0.5 grams of fine NaCl. Fill the second tube with 0.5 g of coarse (Rock Salt or Tablets) NaCl.
2. Add 10 mL of water to each test tube.
3. Stir both test tubes at the same time, noting the number of seconds each takes to fully dissolve. (You should not have to shake the tubes for more than 2 minutes, if you do inform your instructor).
4. Dispose of both solutions in the sink.

F. Rates of Dissolving - Temperature

1. Weigh out two 0.5 gram samples of fine NaCl.
2. Fill two beakers with 50 mL of water. Heat the water in one beaker to boiling, and allow it to cool for one minute.
3. Add the 0.5 gram sample of NaCl to each beaker and record the time required for each to dissolve completely. Stir gently. Record your results.
4. As soon as the crystals are dissolved in the hot water, take the beaker containing the hot solution, and slowly tilt it back and forth. Observe the layer of denser salt solution on the bottom. Repeat this step with the cold solution.
5. Dispose of both solutions in the sink.

G. Ionic Reactions in Solution

1. Into four labelled test tubes place a pea-sized quantity of the following salts (one per tube): barium chloride, sodium sulfate, sodium chloride, and barium sulfate.
2. Add 5 mL of water to each tube, and stir until three of the salts are dissolved. (One of the four salts will not dissolve).
3. Mix the barium chloride and sodium sulfate solutions together. Record your results.
4. Dispose of all solutions in the "Ba/Barium Waste Container".

Results

A. Determining the Concentration of a Saturated Solution

		Trial 1	**Trial 2**
1.	Room Temperature:		
2.	Mass of evaporating dish:		
3.	Mass of evap. dish + KCl solution:		
4.	Mass of evap. dish + dry KCl after 1st heating:		
5.	Mass of evap. dish + dry KCl after 2nd heating:		
6.	Mass of solution. (show calc.)		
7.	Mass of the KCl (solute) (show calc.):		
8.	Mass of H_2O (solvent) (show calc.):		
9.	Mass percent of KCl in solution (show calc.):		
10.	Grams of KCl per 100 grams water (experimental) (show calc.):		
11.	Grams of KCl per 100 grams water (theoretical) (Explain):		
12.	Percent Error in grams KCl per 100 grams water (show calc.):		

1. In determining the concentration of potassium chloride in a saturated solution you recorded the temperature of the room. Why?

2. In determining the concentration of potassium chloride in a saturated solution you reheated the sample several times. This procedure is called heating to a (two words). Why do we use this method?

B. Solubility versus Temperature (Comparing Saturated and Unsaturated Solutions

Experiment	**NaCl**	**NH_4Cl**
1.0 g X + 5 mL water	Dissolved: (yes / no) Saturated / Unsaturated	Dissolved: (yes / no) Saturated / Unsaturated
1.0 g X + 5 mL water + 1.4 g X	Dissolved: (yes / no) Saturated / Unsaturated	Dissolved: (yes / no) Saturated / Unsaturated
2.4 g X + 5 mL water + heat	Dissolved: (yes / no) Saturated / Unsaturated	Dissolved: (yes / no) Saturated / Unsaturated
2.4 g X + 5 mL water + cooled	Dissolved: (yes / no) Saturated / Unsaturated	Dissolved: (yes / no) Saturated / Unsaturated

Answer the following questions both **qualitatively and quantitatively.**

1. At room temperature what can you say about the solubility of NaCl?

2. At room temperature what can you say about the solubility of NH_4Cl?

3. At high temperature (boiling) what can you say about the solubility of NaCl?

4. At high temperature (boiling) what can you say about the solubility of NH_4Cl?

C. Relative Solubility of a Solute in Two Solvents

1. Sketch a picture of the test tube containing heptane + water after 30 seconds and after two minutes (label your pictures).

2. Which liquid is denser? Explain (based on your experimental observations).

3. Sketch a picture of the test tube containing heptane + water + iodine after 30 seconds and after two minutes (label your pictures, including colors).

4. In which solvent is iodine more soluble? Explain (based on your experimental observations).

5. Explain (in terms of IMF's) why Iodine was more soluble in the solvent above.

D. Miscibility of Liquids

1. Which liquid pairs tested were miscible?

2. Which liquid pairs tested were immiscible?

3. From the previous section how would you classify the mixture of heptane and water? Explain.

E. Rates of Dissolving - Particle Size

1. Time required for fine salt crystals to dissolve: ______________

2. Time required for coarse salt crystals to dissolve: ______________

3. The amount of salt, the volume of water, and the temperature of the systems were kept constant between the two test tubes, only particle size was varied. What trend was observed? Explain why the trend occurs.

F. Rates of Dissolving - Temperature

1. Time required for fine salt crystals to dissolve in the hot solution: ______________

2. Time required for fine salt crystals to dissolve in the cold solution: ______________

3. The amount of salt, the volume of water, and the particle size of the salt particles were kept constant between the two test tubes, only temperature was varied. What trend was observed? Explain why the trend occurs.

G. Ionic Reactions in Solution

Test Tube	**Observation**
TT 1: $BaCl_2$	
TT 2: Na_2SO_4	
TT 3: NaCl	
TT 4: $BaSO_4$	
TT 5: $BaCl_2 + Na_2SO_4$	

1. Write complete word **AND** formula equations representing the chemical reaction that occurred when the first two test tubes were mixed together.

2. Based on your observations, which of the products produced is the white precipitate: ____________

3. What experimentally observed evidence leads you to this conclusion?

Post Lab Questions

Using the solubility data in Table 16.1 answer the following questions.

1. What is the percentage by mass of NaCl in a saturated solution of sodium chloride at 50.0℃. Explain. 1. ____________

2. Calculate the solubility of potassium bromide at 23℃. Hint: Assume that the solubility increases by an equal amount for each degree between 20.0℃, and 30.0℃. Explain. 2. ____________

3. A saturated solution of barium chloride at 40.0℃ contains 100. grams of water. How much additional barium chloride can be dissolved by heating the solution to 70.0℃? Explain. 3. ____________

4. A saturated solution of barium chloride at 30.0℃ contains 25.0 grams of water. How many grams of barium chloride is dissolved in the water? Explain. 4. ____________

5. A saturated solution of barium chloride at 30.0℃ contains 150. grams of water. How many additional grams of barium chloride can be dissolved by heating the solution to 60.0℃? Explain. 5. ____________

6. A solution of KCl is saturated at 50°C.

(a) How many grams of solute are dissolved in 100. grams of water? 6(a) ____________

(b) What is the total mass of the solution? Explain. 6(b) ____________

(c) What is the mass percent of this solution? Explain. 6(c) ____________

(d) If the solution is heated to 100°C, how much more KCl can be dissolved in the solution without adding more water? Explain. 6(d) ____________

(e) If the solution is saturated at 100°C, and then cooled to 30°C, how many grams of solute will precipitate out? Explain. 6(e) ____________

Blah blah blah.

Name: ____________________________ Class: _____________________ Date: _____________________

Prelab Questions

1. In the discussion section, what class of reaction occurred in equation (2)?

2. What does the term homogeneous mean?

3. Using Excel, graph the data in Table 16.1 showing the solubility of the four salts as a function of temperature. Attach the graph to the back of the pre-lab. Include a curve fit for the $BaCl_2$ data.

4. Using the graph **qualitatively** what amount of KCl that will dissolve in 100 grams of water at 95℃. Explain.

5. Using the graph **quantitatively** determine the amount of $BaCl_2$ that will dissolve in 100 grams of water at 43℃. Explain.

6. Prove that a 10-mass percent solution of NaCl is the same as a 1.90 M solution of NaCl.

Experiment 17
Acids and Bases

Jay C. McLaughlin
Colorado Northwestern Community College

Name:

Date:

Key Objectives

1. Definitions of Acid and Base.
2. Use of litmus paper to determine if a solution is acidic or basic.
3. Use of pH paper and pH meter to determine pH.
4. Complete Acid/Base reactions.
5. Recognize acid anhydrides, base anhydrides, and neutral salt.

Discussion

Acid-Base reactions are an important, fundamental class of reactions. Equally important is learning how to make good pH measurements using a variety of tools including a pH meter, universal indicator, red and blue litmus paper and pH paper.

Acid and Bases

A complete discussion of Acid and Bases can be found in your textbook, a brief summary is provided below.

There are many definitions of Acids and Bases, the most simple was from Arrhenius which suffices for most discussions, though the Bronsted-Lowry defintion is useful in that it allows one to include the common base NH_3. The Lewis definition is useful in that it extends acid/base chemistry to a whole new series of compounds, though that is beyond the scope of this course.

	Arrhenius	**Bronsted-Lowry**	**Lewis**	**Jay's Dfn.**
Acid:	excess H^+	proton (H^+) donor	electron pair acceptor	H___
Base:	excess OH^-	proton (H^+) acceptor	electron pair donor	___OH (+NH_3)

Many types of compounds when added to water affect the pH of the solution. A not so short list includes:

1. Strong Acids or Strong Bases when dissolved in water, dissociate completely (100%) to produce ions and make an acidic or basic solutions.

$$\mathrm{HCl(aq) + H_2O(l) \longrightarrow Cl^-(aq) + H_3O^+(aq)} \quad (48)$$

$$\mathrm{NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)} \quad (49)$$

2. Weak Acids and Weak bases when dissolved in water, dissociate only partially (normally $<$ 10%) to produce ion and make acidic and basic solutions.

$$\mathrm{HF(aq) + H_2O(l) \rightleftharpoons F^-(aq) + H_3O^+(aq)} \tag{50}$$

$$\mathrm{NH_3(aq) + H_2O \rightleftharpoons {NH_4}^+(aq) + OH^-(aq)} \tag{51}$$

3. Salts (Ionic Compounds) typically dissociate (100%) in water and the resulting cations and anions can react with water to form acidic, basic, or if they fail to react a neutral solution.

4. Metal cations will form a hydrated cation (sometimes refered to as a complex, hydration shell, or coordination compound) with water that can then undergo hydrolysis resulting in an acidic solution.

5. Acid Anhydrides are non-metal oxides that produce acids when reacting with water. This is a new type of reaction called a **Combination** reaction in which the reactants combine together to for a single product. In the case of a non-metal oxide it will combine to produce a acid.

$$\mathrm{SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq)} \tag{52}$$

6. Base Anydrides are metal oxides that produce bases when reacting with water. This is a new type of reaction called a **Combination** reaction in which the reactants combine together to for a single product. In the case of a non-metal oxide it will combine to produce a base.

$$\mathrm{Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)} \tag{53}$$

pH and Concentration of [H$^+$] ions

The strength of an acid or base is generally measured one of two ways. First by calculating the concentration (Molarity) of hydrogen ions [H$^+$] or hydroxide ions [OH$^-$] in solution. Second by using a log scale, pH or pOH which is defined as follows. The equations do not need to be memorized and can also be found on your Cheat Sheet.

$$\mathrm{pH} = -\log[\mathrm{H^+}] \qquad [\mathrm{H^+}] = 10^{-\mathrm{pH}}$$

$$\mathrm{pOH} = -\log[\mathrm{OH^-}] \qquad [\mathrm{OH^-}] = 10^{-\mathrm{pOH}}$$

$$\mathrm{pH} + \mathrm{pOH} = 14 \qquad [\mathrm{H^+}]\,[\mathrm{OH^-}] = 1 \times 10^{-14}$$

Figure 17.1: Equations relating pH, pOH, [H$^+$] and [OH$^-$]. The most frequently used scales are the pH and [H$^+$] scales.

Figure 17.2 the relationship between concentration of hydrogen ions [H$^+$] and pH.

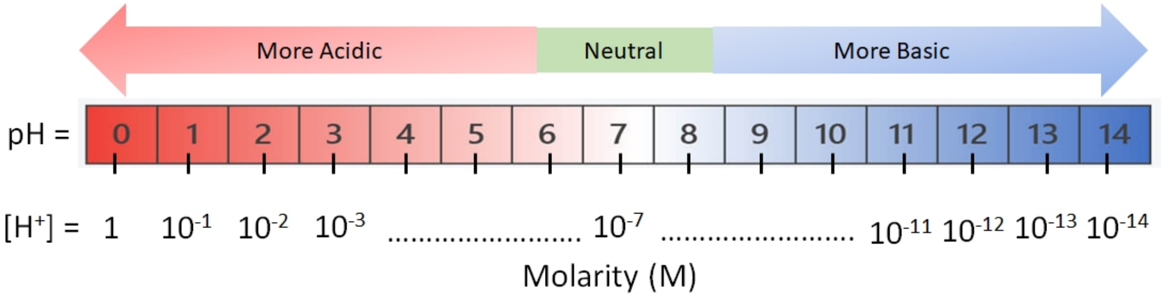


Figure 17.2: Hydrogen ion concentration scales [H^+] in Molarity (M) and pH scale. credit: author

Measuring pH - Test Paper and Indicators

Two methods will be used to measure the pH of the solutions used in lab.

Indicators are organic compounds whose structure is sensitive to the H^+ concentration, and the resulting change in structure is accompanied by a change in the color of the solution.

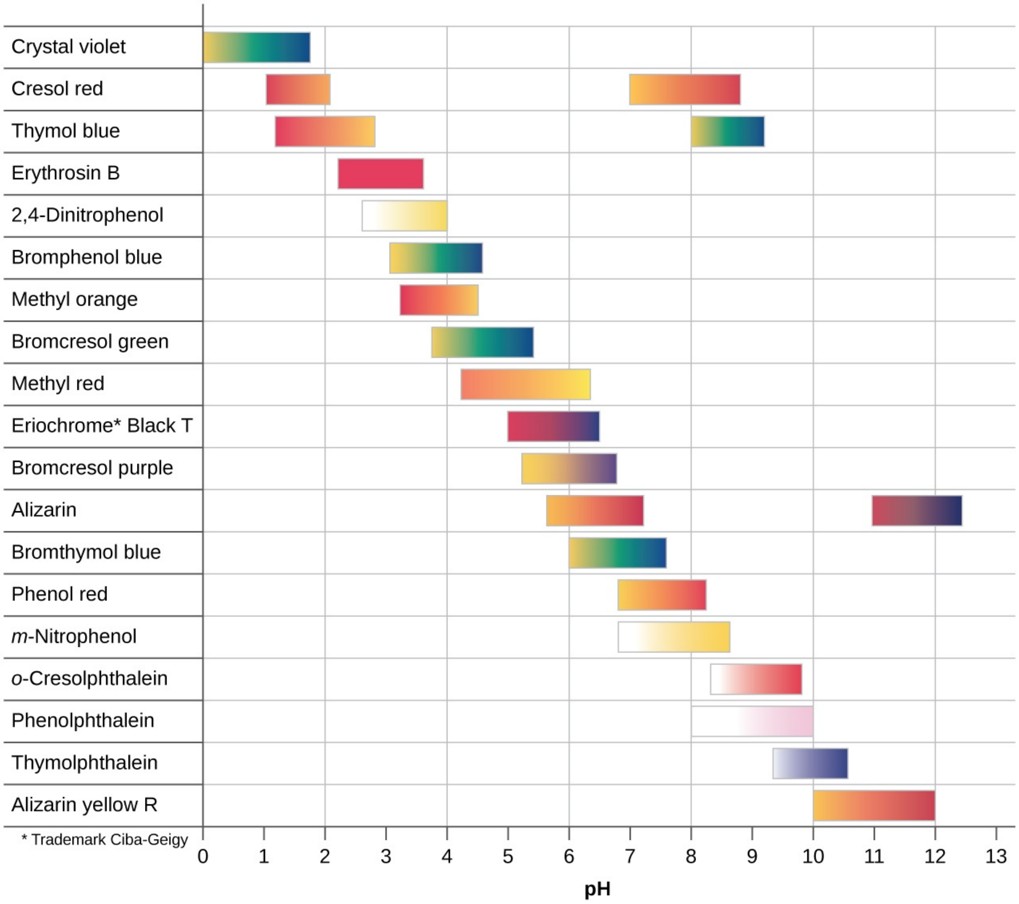


Figure 17.3: pH Range of Different Indicators. credit: `https://commons.wikimedia.org/wiki/File:CNX_Chem_14_07_indicators.png`

One of the most common indicators, Phenolphthalein, you may recall was used as the indicator in the

Titrations Lab.

pH paper is generally made with one or more acid-base indicators impregnated on paper, which will turn different colors depending on the pH. There are a wide variety of pH papers, some with broad ranges (measure pH 1-14), some with very narrow ranges (measure pH ± 1.0 units). We will use a variety of pH papers in lab, so be careful to note the type used and the observed color and pH.

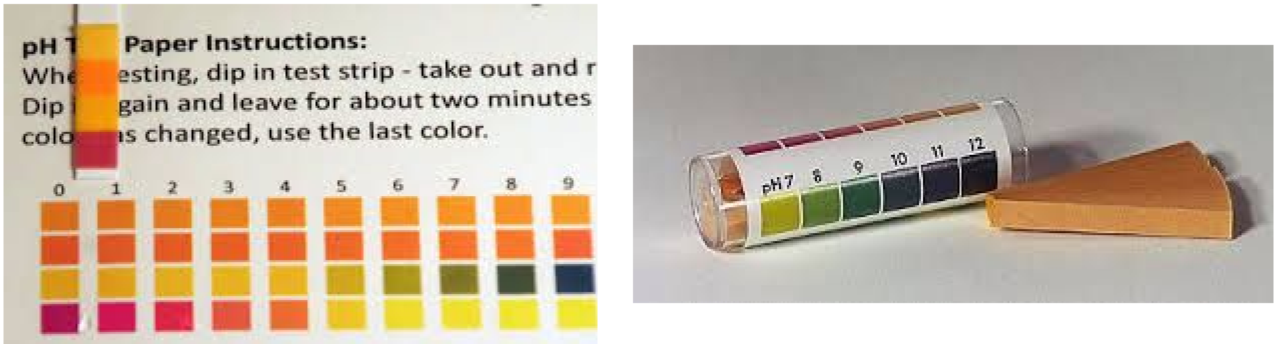


Figure 17.4: Different Brands of pH paper. Those used in lab may be different than those shown here. credit: unknown

Measuring pH - pH Meters

A pH meters is a device that measures the H^+ concentration and converts it into an electrical potential which is then read by a volt meter. The electrical conductivity of a solution is measured using an electrochemical half-cell consisting of twi glass electrodes, a reference electrode and the measuring electrode. For this laboratory we will just focus on learning to calibrate and use a pH meter.

A pH meter (like all electronic instruments) needs to be calibrated. The Vernier pH meter needs to know to pH values so it can extrapolate a line between them to measure the pH. We will use two buffer solutions pH = 4.00 and pH = 10.00 to calibrate the meter. Once calibrated it is important to occasionally check the calibration with any one of the buffer (pH = 4,5,7, or 10) solutions available.

Figure 17.5: pH Probe and Lab Quest 2 from Vernier. credit: (a) https://www.vernier.com/files/manuals/gdx-ph-bnc/gdx-ph-bnc.pdf (b) `https://www.vernier.com/product/labquest-2/`

"Pure" Water, Dissolved Carbon Dioxide, and Distilled Water

A solution of pure water (tap water) should be neutral and have a pH = 7.0. However, various substances can affect the pH of the aqueous solutions used in lab. Carbon Dioxide in the atmosphere will dissolve in water and produce a slightly acidic solution due to the formation of hydrogen carbonate (bicarbonate). This will cause tap water in the laboratory to have a pH in the range of 5.5-6.5.

$$\mathrm{CO_2(aq) + H_2O(l) \rightleftharpoons HCO_3{}^-(aq) + H^+(aq)} \qquad (54)$$

The distilled water produced in lab is very pure and should have a pH of 7.0. However, the water is "too pure" and most pH meters have a hard time reading the pH level, and so can often give "wrong" reading. In this instance pH paper is probably the more accurate measure of pH.

Procedure

A. Qualitative Measurements of pH

1. Obtain 7 Test Tubes, fill each with 1.0 mL of the solution to be tested.
2. Using a glass rod place a small drop of each solution onto the pH test paper. Record the color observed and the value of the pH according to the chart.
3. Using a glass rod place a small drop of each solution onto the Blue Litmus paper and note the results.
4. Using a glass rod place a small drop of each solution onto the Red Litmus paper and note the results.
5. Add 1 drop of Universal Indicator. Record your results.
6. In the last column indicate if the solution was Acidic (A), Basic (B) or Neutral (N). If there are contradictory results consult with your instructor.
7. Dispose of pH paper in the trash can.
8. Dispose of the Test Tube containing the Universal Indicatior in the ACID/WASTE jar.

B. Quantitative pH Measurements of Aqueous Solutions

1. Calibrate the pH meter.
   (a) On the Lab Quest, choose the Sensors ⟶ Calibrate ⟶ pH ⟶ Calibrate Now.
   (b) Place the pH probe in the pH 4.0 buffer and while swirling input the value of the buffer on the Lab Quest and click "Keep".
   (c) Click Calibrate Now.

   (d) Place the pH probe in the pH 10.0 buffer and while swirling input the value of the buffer on the Lab Quest and click "Keep".

   (e) Click "OK".

   (f) Test your calibration using the 7.0 buffer, if it does not read between 6.9 and 7.1 consult your instructor.

2. Fill a small beaker/Erlenmeyer flask with 15-20 mL of the solution to be tested.

3. Measure the pH of each solution, record your results. Be sure to rinse the electrode between measurements with de-ionized water.

4. When you are done making he pH measurements, pour the solution back into the bottle.

C. Reactions of Acids

1. Perform the following reactions in test tubes 1-7.

2. For each test tube write the complete reaction (balanced, states etc).

3. Perform reactions 1-4 by filling a test tube with 2.0 mL of the acid listed on the table and place a few strips of magnesium ribbon into each tube. Place a cork in the test tube and allow the reaction to proceed for 1 minute. In the hood, test the gas generated by placing a burning split into the mouth of the tube. **EXTINGUISH THE SPLINT IN THE WATER.** Record your observations, was the gas flammable or non-flammable

4. Perform reactions 5-6 by placing a pea sized (or tip of spatula sized) portion of the solid in the bottom of a test tube. Place 4-5 mL of the acid listed in the table in the test tube. Place a cork in the test tube and allow the reaction to proceed for 1 minute. In the hood test the gas generated by placing a burning split into the mouth of the tube. **EXTINGUISH THE SPLINT IN THE WATER.** Record your observations, was the gas flammable or non-flammable

5. Perform reaction 7 by placing 25 mL of water into a beaker, add 3 drops of phenolphthalein and 5 drops of HCl to the solution. Note the color. Drop by drop slowly add the 6.0 M NaOH solution to the beaker. Stir after each drop. Stop adding the base when the phenolphthalein changes color.

6. Dispose of all solutions in the waste container labeled "ACID/BASE WASTE".

D. Reactions of Metal Oxides

1. Perform the following reactions in test tubes 1-2.

2. For each test tube write the complete reaction (balanced, states etc).

3. Perform reactions 1-2 by placing 10 mL of water, 2 drops of phenolphthalein into each test tube. Then add about a peas sized amount of solid, swirl the solution and record the results.

4. Test the solution with both blue and red litmus paper, and pH paper and note the results.

5. Dispose of all solutions in the waste container labeled "ACID/BASE WASTE".

E. Reactions of Nonmetal Oxides and Water

1. Do the following reaction in the hood.
2. For each test tube write the complete reaction (balanced, states etc).
3. Prepare a wide mouth bottle by placing 15 mL of water in it and obtaining a stopper for it.
4. Place a small amount (marble sized) of sulfur powder in a deflagrating spoon and start it burning by heating in a bunsen burner.
5. Holding the wide mouth bottle near horizontal, place the burning sulfur in the deflagration spoon in the wide mouth bottle and collect the gas given off.
6. **EXTINGUISH THE BURNING SULFUR IN THE DEFLAGRATION SPOON IN THE WATER.**
7. Place a stopper in the bottle and shake to mix the water and gas given off.
8. Remove the deflagration spoon and place it in the distilled water nearby to quench it.
9. Test the solution with both blue and red litmus paper, and pH paper and note the results.
10. Dispose of all solutions in the waste container labeled "ACID/BASE WASTE".

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Results

Data Table A: Qualitative Measurements of pH

TT #	Compound	Color pH Paper	Value pH Paper	Red Litmus	Blue Litmus	Universal Indicator	A/B/N
1	0.1 M KNO_3						
2	0.1 M $HC_2H_3O_2$						
3	0.1 M HCl						
4	0.1 M NaOH						
5	0.1 M NH_3						
6	Tap Water						
7	Distilled Water						

1. Blue Litmus paper turns Red when it reacts with? 1. ______________

2. Red Litmus paper turns Blue when it reacts with? 2. ______________

3. Did all the methods used (pH paper, Litmus paper and Universal Indicator) agree on whether the tested solution was an Acid, Base or Neutral. If not which solutions gave contradictory results? Explain.

Data Table B: Quantitative Measurements of pH

TT #	Compound	pH meter	A/B/N	Calc. [H^+] conc.
1	0.1 M KNO_3			
2	0.1 M $HC_2H_3O_2$			
3	0.1 M HCl			
4	0.1 M NaOH			
5	0.1 M NH_3			
6	Tap Water			
7	Distilled Water			

4. Which Test Tubes (#) were:
   (a) Acidic: 4(a) ____________
   (b) Basic: 4(b) ____________
   (c) Neutral: 4(c) ____________

5. Compare the results from Table A (ph Paper) and the values in Table B (ph Meter). Did any compounds tested produce contradictory results (ie one says a solution is Acidic, one says its Basic, or if the values are greater then 2 pH units apart). Explain.

6. For each of the solutions based on the pH value measured by the meter calculate the concentration of hydrogen ions in solution. Show work for Test Tube # 2.

Data Table C: Reactions of Acids

TT #	Reaction	Observation
1	___$HCl(aq) + $___$Mg(s) \longrightarrow$	
2	___$H_2SO_4(aq) + $___$Mg(s) \longrightarrow$	
3	___$HNO_3(aq) + $___$Mg(s) \longrightarrow$	
4	___$HC_2H_3O_2(aq) + $___$Mg(s) \longrightarrow$	
5	___$HCl(aq) + $___$NaHCO_3(s) \longrightarrow$	
6	___$HCl(aq) + $___$CaCO_3(s) \longrightarrow$	
7	___$HCl(aq) + $___$NaOH(aq) \longrightarrow$	

1. What class of reaction (Double Displacement, Single Displacement, Combustion, Acid/Base) occurred in each of the test tubes.
   (a) Test Tubes 1-4: 1(a) ______________
   (b) Test Tubes 5-6: 1(b) ______________
   (c) Test Tubes 7: 1(c) ______________
2. Answer the following questions about the flame tests for Test Tubes 1-6:
   (a) Based on the products for each reaction, which reactions produced a flammable gas. 2(a) ______________
   (b) Based on the products for each reaction, which reactions produced a non–flammable gas. 2(b) ______________
   (c) Based on the products for each reaction, what was the identity of the flammable gas? 2(c) ______________

(d) Based on the products for each reaction, what was the identity of the non-flammable gas?
2(d) ______________

3. Based on your observations of the flame test, did any of the acids tested in Reactions 1-4 **NOT** react as one would expect? Explain.

4. Based on your observations in reaction # 7 what color is phenolphthalein in:

(a) Basic Solution: 4(a) ______________

(b) Acidic Solution: 4(b) ______________

Data Table D: Reactions of Metal Oxides

TT #	Reaction	Red Litmus	Blue Litmus	pH Paper
1	___$H_2O(l) + $___$CaO(s) \longrightarrow$			
2	___$H_2O(l) + $___$MgO(s) \longrightarrow$			

Data Table E: Reactions of Nonmetal Oxides

TT #	Reaction	Red Litmus	Blue Litmus	pH Paper
1	___ $H_2O(l) + $___$SO_2(g) \longrightarrow$			

1. Are the solutions produced in Table D (Acidic/Basic/Neutral)? 1. _______________

2. Are the solutions produced in Table E (Acidic/Basic/Neutral)? 2. _______________

Post Lab Questions

1. Are the following compounds Acids, Bases, Neutral Salt, an Acid Anhydride, Base Anhydride or none.

Compound	**Category**	**Compound**	**Category**
CuF_2		$CaSO_4$	
$Ba(OH)_2$		C_2H_4	
LiOH		$C_{12}H_{22}O_{11}$	
$HBrO_3$		HI	
$RaCO_3$		P_2O_5	
KNO_3		HCN	
$H_2C_2O_4$		MgO	

2. Complete the following Combination reactions and name the product formed.

(a) $\_\_\_K_2O(s) + \_\_\_H_2O(l) \longrightarrow$

(b) $\_\_\_SrO(s) + \_\_\_H_2O(l) \longrightarrow$

(c) $\_\_\_SO_3(g) + \_\_\_H_2O(l) \longrightarrow$

(d) $\_\_\_N_2O_5(g) + \_\_\_H_2O(l) \longrightarrow$

Name: ________________________ Class: _________________ Date: _________________

Prelab Questions

1. Give the name or formula for the 10 acids you were required to memorize in class.

2. Give the name or formula for 4 bases.

3. We covered many reactions in class that included acids and bases. There are also two new reactions to learn. Complete the following reactions and in the answer blank include the class of reaction.

   (a) ___$HCl(aq) + $___$NaOH(aq) \longrightarrow$

   (b) ___$Na_2CO_3(aq) + $___$HNO_3(aq) \longrightarrow$

   (c) ___$H_2SO_4(aq) + $___$Na(s) \longrightarrow$

   (d) ___$K_2O(aq) + $___$HOH(l) \longrightarrow$

   (e) ___$HOH(l) + $___$CO_2(g) \longrightarrow$