LABORATORY MANUAL

Chemistry 1003

Table of Contents

Laboratory Safety, Protocols, and Check-In

Your instructor may make adjustments and/or include additional instructions to the following listed items. If so, record these on the following pages.

Attire

1. Eye protection must be worn at all times. You may wear your regular prescription glasses, but you will need to find a pair of lab goggles/safety glasses that can fit over them. Your first action during any lab session is to put them on. If you wear contacts, you need to be prepared to remove them if your eyes become irritated.
2. Dress appropriately.
   - Closed-toe shoes. No sandals or flip-flops.
   - Legs covered.
   - Torsos covered. No bare midriffs, bare backs or bare shoulders.
   - Avoid wearing bulky clothes and hanging jewelry. If possible, remove these items and stow them until you are through with the lab.
   - Long, loose hair should be pulled back with a ponytail holder or scrunchy.
   - No hats.

Protocols not involving chemical reagents/equipment

3. NO horseplay.
4. Do not place items on the floor where you are working. Place backpacks and other items elsewhere (not where people are walking).
5. No food, beverage, gum, water, makeup application. If you have a bottle or food with you, it needs to be in your backpack or whatever you have before entering the lab. Do not touch the bottle or food until you have left the lab.
6. Only ULM personnel or students enrolled in the lab are allowed to be in the lab. If you must talk to someone not enrolled in the lab, you must talk out in the hall.
7. Students may not use cell phones in the laboratory. Cell phones must be turned off during pre-lab lecture and in the laboratory. If there is a reason you must have your phone on, it needs to be on vibrate mode and you need to exit the laboratory to answer it.
8. Know the location <u>and</u> use of safety equipment (fire extinguishers, eyewashes, fire blanket, and fire alarms). You may be required at any time during the semester to draw location diagrams and to give instructions on the use of this safety equipment.

Protocols involving chemical reagents/equipment

9. Do only what you are instructed to do; do not change any procedure.
10. If ANY piece of glassware is chipped, cracked, or broken, show it to the instructor. The instructor will probably tell you to put it in the broken glass box. ONLY broken glass goes into this box; no paper.
11. Double check chemical labels. Make sure you are using the required chemical.
12. If you knock over a reagent bottle or break it, immediately let your instructor know.

13. The instructor will tell you during the pre-lab lecture how to dispose of chemicals. If in doubt, do not throw chemicals down the drain or into the trash can without first consulting with your laboratory instructor.
14. Before you leave, clean all of your glassware AND wipe off the bench where you worked.
15. When no longer required, equipment should be returned to its original space (e.g. drawer, shelf, hood).

Protocols involving the avoidance of injuries

16. Do not touch any part of your face (especially the mouth or eyes) while in the lab.
17. Never smell any vapor or gas directly by putting your nose over the top of a test tube, bottle, etc. and inhaling.
18. If your skin feels tingly or itchy, DO NOT scratch. Rinse it off with some water.
19. Students must NOT handle broken glass with bare hands. If any glass item breaks, inform the instructor. Broken glass may also be found on the floor, in your desk drawer, and in the sinks. Don't reach carelessly into a desk drawer or a sink. When broken glass is found, inform your instructor who will get the proper equipment for cleaning it up.
20. Wash your hands with soap before leaving the lab.

Protocols involving injures

21. Report ALL injuries to your instructor.
22. If you have a reaction to one of the chemicals, it should be reported.
23. If chemicals come into contact with your eyes, rinse your eyes out with water from the eye wash for several minutes.
24. If chemicals come into contact with other body parts not covered with clothing, and you cannot get these areas under the faucet to rinse, you should use the eye wash to rinse off these areas.
25. If harmful chemicals come into contact with your clothing, it may be necessary for you to remove the clothing and then rinse off your skin that was under the clothing. The eye wash can be used for this rinsing.

Protocols involving fire

26. If a large fire occurs in the lab, follow the Evacuation Plan. Close open gas jets and turn off any running water.
27. Most fires in the lab are small and will stop as soon as the burning material is consumed. For instance, a piece of paper towel catches on fire. For such a fire, remove other things that might catch on fire (like the roll of paper towels), step back and do not scream. Do not pick it up and try running to the sink. Do not get something and try to beat out the fire. If necessary, a small fire can easily be extinguished by smothering. For instance, sometimes fumes over a beaker or test tube will ignite. First, turn off the laboratory burner; then, smother the fire by holding a flat object (such as a book or a larger beaker) over the opening.
28. If a fire occurs in the trash can, using the fire extinguisher is the best course of action.

29. If a fire occurs in a hood, close the hood door. Turn off the gas or water if either is being used in the hood. If you are instructed to follow the Evacuation Plan, then do so.

30. Flaming hair or clothes can be extinguished by smothering or with water. A fire blanket is in the laboratory. If using a fire blanket, place the blanket over the burning area. DO NOT BEAT THE FIRE with the blanket. The eye wash may be used in a similar way as a water hose.

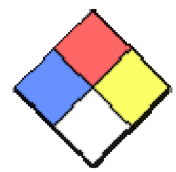
Safety Labels for Reagents

Two types of safety labels may be on a reagent container. The two types are a NFPA (National Fire Protection Association) label and the other is a GHS (Globally Harmonized System) label.

NFPA

The NFPA is a United States association that establishes codes for hazardous material (hazmat) response. The NFPA 704 standard presents a system of markings (commonly referred to as the "NFPA hazard diamond") that provides information as to the hazards of a material.

The four divisions are color-coded with red indicating flammability, blue indicating health hazard, yellow for chemical reactivity, and white containing codes for special hazards. Each of health, flammability and reactivity is rated on a scale from 0 (no hazard) to 4 (severe risk).



Flammability (red)

0: Materials that will not burn under typical fire conditions

1: Materials that require considerable preheating

2: Must be moderately heated

3: Liquids and solids (including fine divided suspended solids) that can be ignited under almost all ambient temperature conditions

4: Will rapidly or completely vaporize at normal atmospheric pressure and temperature, or is readily dispersed in air and will burn readily.

Health (Blue)

0: Poses no health hazard

1: Exposure would cause irritation with only minor residual injury

2: Intense or continued but not chronic exposure could cause temporary incapacitation or possible residual injury

3: Short exposure could cause serious, temporary or moderate residual injury

4: Very short exposure could cause death or major residual injury

Instability/Reactivity (Yellow)

0: Normally stable, even under fire exposure conditions, and is not reactive with water

1: Normally stable, but can become unstable at elevated temperatures and pressures

2: Undergoes violent chemical change at elevated temperatures and pressures, reacts violently with water, or may form explosive mixtures with water

3: Capable of detonation or explosive decomposition but requires a strong initiating source, must be heated under confinement before initiation, reacts explosively with water, or will detonate if severely shocked

4: Readily capable of detonation or explosive decomposition at normal temperatures and pressures

Special Notice (White)

The white "special notice" area can contain several symbols. The following symbols are defined by the NFPA 704 standard.

OX: Oxidizer, allows chemicals to burn without an air supply

W: Reacts with water in an unusual or dangerous manner

SA: Simple asphyxiant gas
* An asphyxiant is a substance that can cause unconsciousness or death by suffocation

Non-standard symbols: These hazard codes are not part of the NFPA 704 standard, but are occasionally used in an unofficial manner. The use of non-standard codes may be permitted, required or disallowed by the authority having jurisdiction (e.g. fire department).

COR: Corrosive

ACID, ALK: Acid or alkaline

BIO or ☣: Biological hazard

POI: Poisonous

RA, RAD or ☢: Radioactive

CRY or CRYO: Cryogenic (freezing)

GHS

The GHS was created by the United Nations. It was designed to use consistent criteria on a global level. Its usage in the United States is governed by the Occupational Safety and Health Administration (OSHA).

Physical, health, and environmental hazards of chemicals along with protective measures are on labels and Safety Data Sheets (SDS).

Hazard Classes
- 16 classes of physical hazards
- 10 classes of health hazards
- 3 classes of environmental hazards

GHS Physical Hazards
- Explosives
- Flammable Gases
- Flammable Aerosols
- Oxidizing Gases
- Gases Under Pressure
- Flammable Liquids
- Flammable Solids
- Self-Reactive Substances
- Pyrophoric Liquids
- Pyrophoric Solids
- Self-Heating Substances
- Substances which, in contact with water emit flammable gases
- Oxidizing Liquids
- Oxidizing Solids
- Organic Peroxides
- Corrosive to Metals

GHS Health Hazards
- Acute Toxicity
- Skin Corrosion/Irritation
- Serious Eye Damage/Eye Irritation
- Respiratory or Skin Sensitization
- Germ Cell Mutagenicity
- Carcinogenicity
- Reproductive Toxicology
- Target Organ Systemic Toxicity-Single Exposure
- Target Organ Systemic Toxicity-Repeated Exposure
- Aspiration Toxicity

GHS Environmental Hazards
- Hazardous to the Aquatic Environment
- Acute aquatic toxicity
- Chronic aquatic toxicity

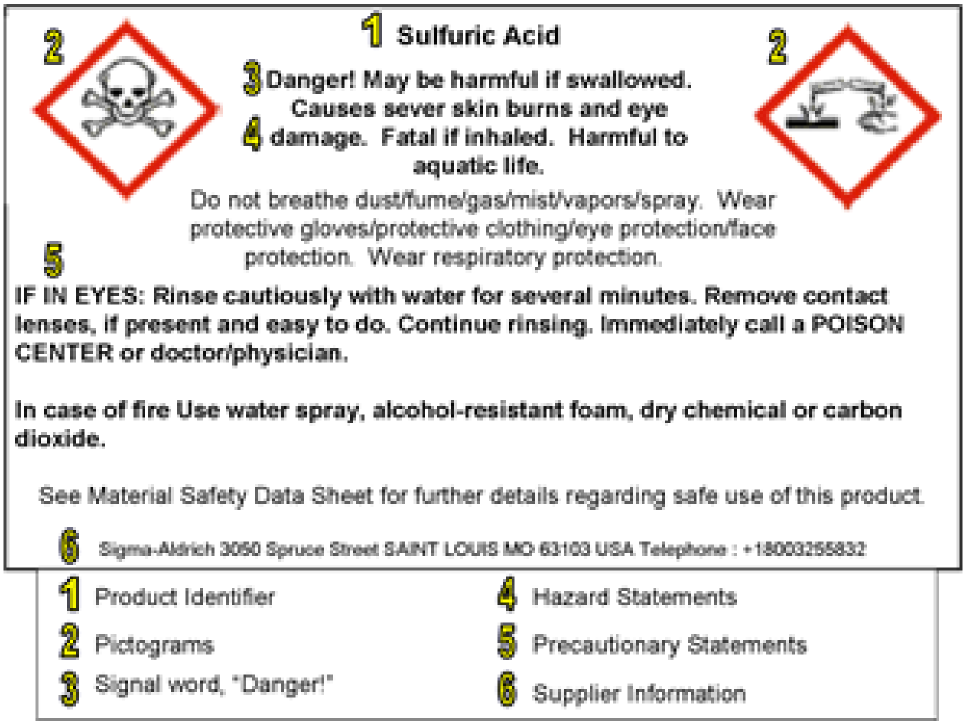
GHS Label Elements

- Product Identifier
- Signal Word
  - ·Danger for more severe hazards
  - ·Warning for less severe hazards
- Hazard Statement
  - Hazard statements are standardized and assigned phrases that describe the hazards.
- Pictogram(s)
- Precautionary Statement
  - Precautionary statements are explanations of the measures to be taken to minimize or prevent adverse effects.
- Name, Address and Telephone Number of chemical manufacturer

Safety Data Sheets (SDS): The following information is found on SDS.

- Identification of the substance or mixture and of the supplier
- Hazards identification
- Composition/information on ingredients
- First aid measures
- Firefighting measures
- Accidental release measures
- Handling and storage
- Exposure controls/personal protection
- Physical and chemical properties
- Stability and reactivity
- Toxicological information
- Ecological information
- Disposal considerations
- Transport information
- Regulatory information
- Other information including information on preparation and revision of the SDS

Example of Label with Elements Numbered



Example of Label

Pictograms

Health Hazard

- Carcinogen
- Mutagenicity
- Reproductive Toxicity
- Respiratory Sensitizer
- Target Organ Toxicity
- Aspiration Toxicity

Flame

- Flammables
- Pyrophorics
- Self-Heating
- Emits Flammable Gas
- Self-Reactives
- Organic Peroxides

Exclamation Mark

- Irritant (skin and eye)
- Skin Sensitizer
- Acute Toxicity (harmful)
- Narcotic Effects
- Respiratory Tract Irritant
- Hazardous to Ozone Layer (Non-Mandatory

Gas Cylinder

- Gases Under Pressure

Corrosion

- Skin Corrosion/Burns
- Eye Damage
- Corrosive to Metals

Exploding Bomb

- Explosives
- Self-Reactives
- Organic Peroxides

Flame Over Circle

- Oxidizers

Environment (Non-Mandatory)

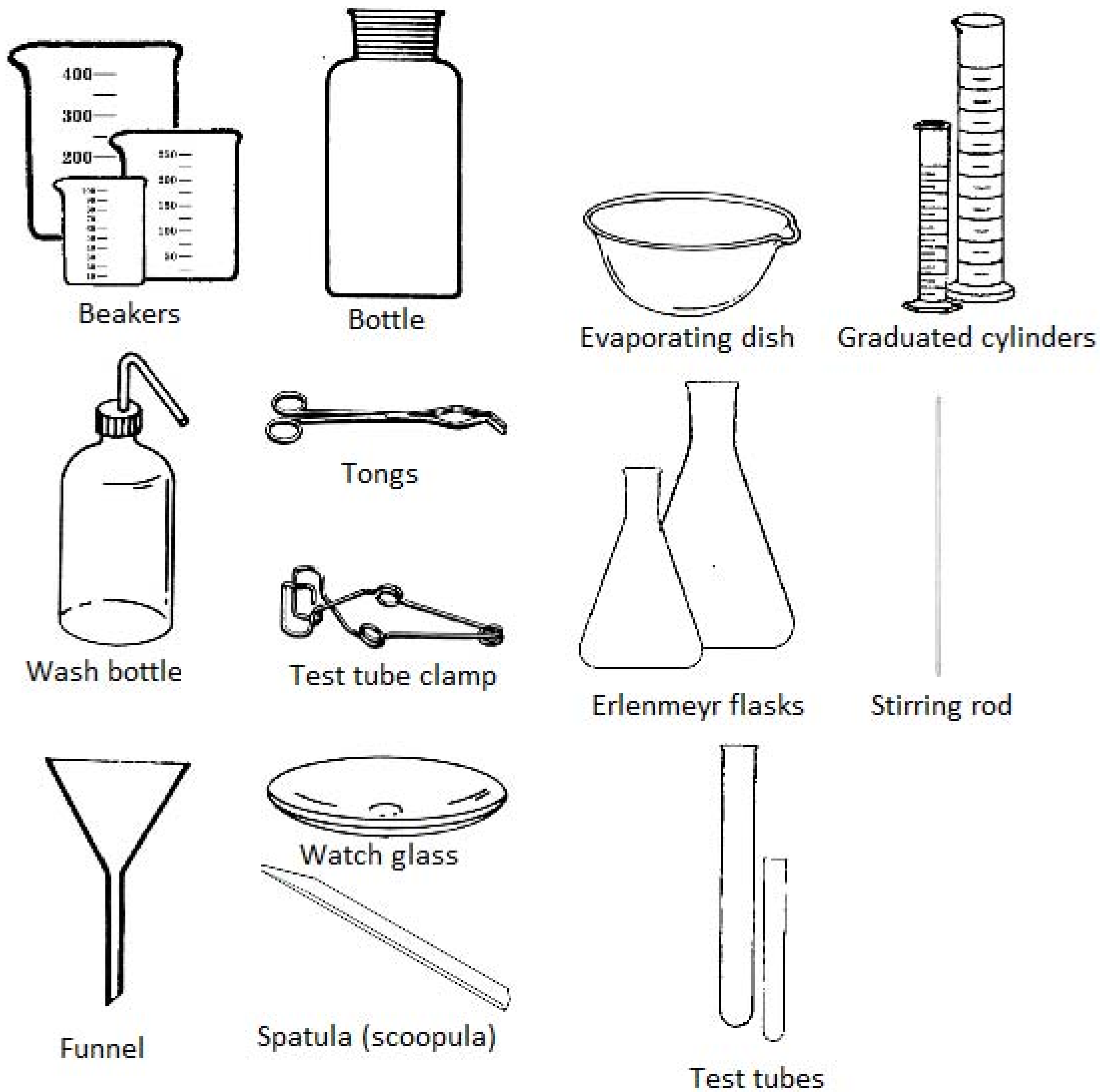
- Aquatic Toxicity

Skull and Crossbones

- Acute Toxicity (fatal or toxic)

Check In Procedure

After being assigned a partner and a lab desk number, check to see if your desk has all of the listed items on the Desk Assignment Sheet, which is on the next page. Below are pictures of this equipment. Once you have all of the items, finish filling in the Desk Assignment Sheet and give it to your instructor, if so instructed to do so.



Desk Assignment Sheet (Chemistry 1003)

Please Print

Name____ULM ID______________________

Name____ULM ID______________________

Section #__________________Room #__________________Desk #___________________

- ☐ 2 Beaker, 100 or 150 ml
- ☐ 1 Beaker, 250 ml
- ☐ 1 Beaker, 400 ml
- ☐ 1 Bottle
- ☐ 1 Wash Bottle
- ☐ 1 Clamp, Test-Tube
- ☐ 1 Cylinder, Graduated 10 ml
- ☐ 1 Cylinder, Graduated, 100 ml
- ☐ 1 Evaporating Dish
- ☐ 2 Flask, Erlenmeyer, 125 ml
- ☐ 1 Funnel, Narrow Stem
- ☐ 1 Watch Glass
- ☐ 1 Spatula
- ☐ 5 Test Tube, 15 × 125
- ☐ 1 Tongs
- ☐ 1 Stirring rod

Check In:

Signature_______________________________________

Signature_______________________________________

Checked in by___________________________________Date_________________________

I MUST WEAR SAFETY GOGGLES DURING EACH LAB SESSION.

Initials_____ Date________ Initials_____ Date________

Check Out:

Checked out by__________________________________Date_________________________

Introductory Exercises

If a property is measured, the measurement is expressed with a number and a unit. For instance, if you place an object on a balance to determine the object's mass and the balance reads 34 grams; 34 is the number and grams is the unit.

Accuracy refers to the closeness of a measurement to the correct value. For example, if in lab you record the mass of an object as 3.2 g, and the true value is 5.7 g, then your measurement is not accurate.

Precision refers to the closeness of two or more measurements to each other. For example, let us say you measure the mass three times of the object mentioned in the accuracy example. You get 3.2 g, 3.3 g, and 3.2 g, then your measurements are precise. But, this precision is not worth much, because none of them are accurate.

When making measurements, the measuring device determines the number of significant figures in the recorded measurement. Having more significant figures in a measurement will result in a measurement that is more accurate. For example, you measure the mass of an object on two different balances. The actual mass of the object is 4.226 grams. One balance only measures to the nearest gram. When you place the object on that balance, the reading is 4 grams. This measurement has one significant figure. The other balance measures to the nearest hundredth of a gram. When you place the object on that balance, the reading is 4.23 grams. This measurement has three significant figures. The second balance gives a more accurate measurement.

Whenever a measurement is made, there is some uncertainty about the measurement. The digit that is written last is the uncertain one. It is the "best call" one.

It is necessary to determine the number of significant figures in measurements whenever doing calculations with these measurements. The answers to the calculations should not be more/less accurate than the measurements. The number of significant figures in the measurements determine the number of significant figures in the answers for the calculations.

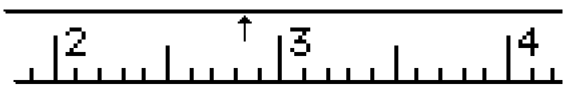
<u>Rules for Determining the Number of Significant Figures in a Number</u>

1. All nonzero digits are significant.
2. Zeros:
   a. Zeros between nonzero digits are significant.
   b. Leading zeros are not significant.
   c. Trailing zeros are significant if there's a decimal.
      - 40. has two significant figures.
      - 40 has the possibility of having one or two significant figures. If you were counting pencils, and there were exactly 40 pencils, this would be two significant figures. If you were counting birds flying around and were not really sure about the count, 40 would have one significant figure.
      - 0.02020 has four significant figures.
3. Exact numbers are not used to determine the number of significant figures in calculation answers. Examples include the following:
   a. counting numbers
   b. conversion factors
   c. numbers obtained from a table or other source of a property's value

Rules for Rounding Calculation Answers

Multiplication/division:	The answer should be rounded to the same number of significant figures as the measurement with the fewest significant figures.
Addition/subtraction:	The result should be rounded to the same decimal place as the measurement with the fewest number of decimal places.

Length

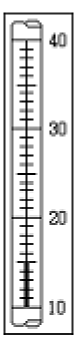


Above is a section of a meter stick. It is NOT drawn to scale. Each numbered division is for centimeters (cm). Between each of the centimeter divisions are 10 smaller divisions. Above the meter stick is a line. There is an arrow pointing to the line. How long is the line at the location at which the arrow is pointing?

You know for sure that your answer will include 2.8 cm. This is because the arrow is between 2 cm and 3 cm and between the .8 and the .9 mark. Now for the "best call" digit. Let us say you think the arrow is at about .83. It does not seem to be exactly half way between the .8 and .9, but a little less than the half way point.

Your answer would be 2.83 cm. This measurement has 3 significant figures. All are known for sure except the last one.

Temperature



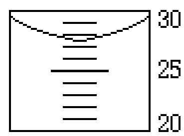
To the left is a section of a thermometer. The divisions are for one degree Celsius (C). When using such a thermometer, you should record the temperature to the nearest tenth of a degree. If, in your judgment, the top of the line is exactly at 15, then you would record the temperature as 15.0 °C. The 0 is the "best call" digit. This measurement has 3 significant figures.

Mass

Mass is a measure of the quantity of matter. The weight of an object is related to the pull of gravity on the object and is not the same as its mass. An astronaut's mass does not change as he or she is in space, but once free of the gravitational pull of the earth, the astronaut becomes weightless. On the moon, an astronaut's weight is much less than that on earth because of the moon's much weaker gravitational pull. A scale is used to measure weight; a balance is used to measure mass.

Volume

Volume is the amount of space occupied by matter. The volume of a liquid can be determined by using a graduated cylinder. To read the volume of a liquid in a graduated cylinder properly, read the bottom of the meniscus. The meniscus is the curved surface of the liquid.

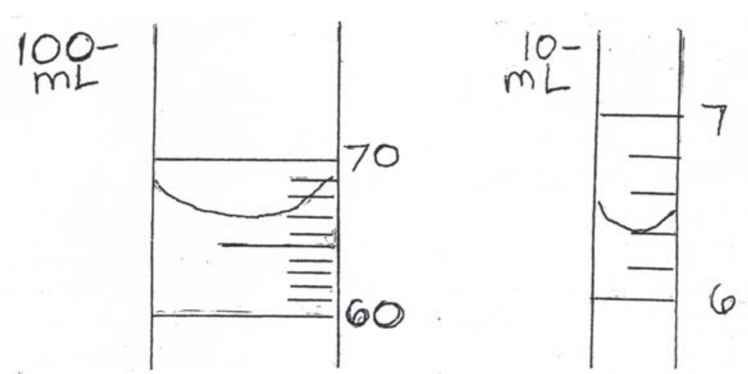


Above is a section of a graduated cyclinder. The divisions are for one milliliter (mL). You should record the volume to the nearest tenth of a mL. Since the bottom of the meniscus is about half way between the 27 and the 28 mark, the recording would be 27.5 mL. The 5 is the "best call" digit. This measurement has 3 significant figures.

How many mL of water are in the following graduated cylinders?

(a) 100 mL graduated cylinder?

(b) 10 mL graduated cylinder?



It is possible to measure the volume of a solid object by placing the solid under water. The object will cause the water level to rise. The difference in the water level before and after the object is submerged is due to the volume of the object. The object has displaced its own volume of water. By measuring the water level before and after the object is added, the volume of the object can be calculated. This procedure to determine an object's volume is called volume displacement.

Density

The density is the mass of one unit volume of a substance (d = mass/volume). For instance the density of sulfuric acid (a liquid) is 8.4 g/mL. This means that if you measure out 1 mL of sulfuric acid, the mass would be 8.4 g. The density of cork (a solid) is 0.3 g/cm^3. This means that if you have a cube that is 1 cm by 1 cm by 1 cm, the mass would be 0.3 g. Density is always expressed as the mass per unit of volume. So the number in front of the volume unit has to be 1.

To determine the density of a substance, you need to know the mass of a certain volume. Once you know this, you can calculate the mass of just one unit of the volume. For instance, if 36 mL of a liquid has a mass of 24 g, then to calculate the mass of 1 mL, you divide both numbers by 36. You are only dividing the numbers; you are not doing anything with the units. The units stay.

$$\frac{24\ \text{g}}{36\ \text{mL}} \qquad\qquad \text{Density} = \frac{0.67\ \text{g}}{1\ \text{mL}}$$

You can also write density this way: 0.67 g/mL

If you write it the second way (all on one line), you do not need to write the number 1 in front of the volume unit. If you write it the first way, you have to include the number 1 in the denominator.

Items Needed for Each Section of Lab (Stockroom)

- 60 pennies
- 12 rubber stoppers, size #2

Items Needed for Each Pair of Students (Already on shelves or in drawer)

- 10-mL graduated cylinder
- 100-mL graduated cylinder
- Wash bottle
- 100 or 150-mL beaker
- Thermometer
- Plastic dropper

Procedure

Balance and Mass

1. Select five pennies.
2. Place a sheet of weighing paper on the pan. Zero the balance. When possible, use the cover to protect the pan from air drafts.
3. Place a penny on the balance. Record its mass below. DO NOT ROUND ANY BALANCE READINGS. Remove the penny.
4. Zero the balance. Place the second penny on the balance. Record its mass below. Remove the penny.
5. Do step #4 for the other three pennies.

	Mass (g)
#1	
#2	
#3	
#4	
#5	

6. Calculate the total sum of the masses of the five pennies (Σ mass). Record answer here.

   Σ mass =

   If all of the mass recordings in #5 do not have the same number of decimal places, how many decimal places should your sum have?

7. Calculate the mean or average mass of the pennies. Use the equation below to do this. The letter n stands for the number of pennies weighed. Round your answer to have the correct number of significant figures.

$$\text{mean} = \frac{\Sigma \text{ mass}}{n}$$

8. Calculate the absolute deviation (d) from the mean of each penny's mass using the equation below. The lines on the left and right of the words 'mean-mass' stands for absolute. Absolute means positive. So, if you get a negative answer, reverse the sign. For each d, remember the rule to determine the correct number of decimal places when subtracting.

$$d = |\text{mean-mass}|$$

#1
#2
#3
#4
#5

9. Calculate the total sum of the means. Determine the average absolute deviation by using the equation below. Σd stands for the sum of all of the absolute deviations. The letter n stands for the number of absolute deviations calculated. Round your answer to have the correct number of significant figures.

$$\frac{\Sigma d}{n}$$

<u>Density of an Object Determined by Using Volume Displacement</u>

1. Get a rubber stopper.

2. Place a sheet of weighing paper on the balance. Zero the balance. When possible, use the cover to protect the pan from air drafts.

3. Place the rubber stopper on the balance. Record its mass below. DO NOT ROUND THE BALANCE'S READING.

______________ g

4. Get a 100-mL graduated cylinder. Add almost, but less than, 50.0 mL of water to the 100-mL graduated cylinder. Use water from a wash bottle to adjust the level to the 50.0 mL mark. To read the volume of water in a graduated cylinder properly, set the cylinder on a level surface and bring your eyes even with the liquid level. Putting a piece of paper behind the graduated cylinder makes it easier to see the reading. Record this volume (50.0 mL) on line (2) below.

5. Carefully put the small rubber stopper into the 100-mL graduated cylinder (do not splash out any of the water). Record the level of the water below. Do your best to read the level to the nearest tenth of a mL. Record the volume occupied by the water and stopper on line (1) below.

6. Calculate the volume of the stopper and record the stopper's volume on line (3) below. How many decimal places should your answer have for line (3)?

(1) ___________ mL (water and stopper)
(2) ___________ mL (water)
(3) ___________ mL (stopper)

7. Calculate the density of the stopper. Do not forget to include the units in your answer. Round your answer to have the correct number of significant figures.

<u>Density of Water</u>

1. Put about 50 mL of water into a small beaker and determine its temperature by holding a thermometer in it for one minute. Record the temperature here.

2. Place a 10-mL graduated cylinder on the balance and zero the balance.

3. With a long, plastic dropper transfer 10.0 mL of water into the cylinder. Read the bottom of the meniscus for the most accurate reading. DO NOT GET ANY DROPS OF WATER ON THE BALANCE. Record the mass of the water. Do not round the balance's reading.

   Mass of H_2O = _________________ g

4. Calculate the density of the water. Do not forget to include the units in your answer. Round your answer to have the correct number of significant figures.

5. Consult the Table of Water Densities (below) and calculate the percent error in your determination according to the following formula. In the equation below, the number 100 is a counting number. The true value is obtained from the table. Which number is going to be used to determine the number of significant figures in your answer? Round your answer to have the correct number of significant figures.

$$\frac{|\,\textbf{True - Actual}\,|}{\textbf{True}} \quad \textbf{X 100 =} \quad \textbf{\% Error}$$

Table of Water Density at Various Temperatures

T,°C	D, g/mL	T,°C	D, g/mL	T,°C	D, g/mL
15	0.9991	20	0.9982	25	0.9970
16	0.9989	21	0.9980	26	0.9967
17	0.9987	22	0.9978	27	0.9965
18	0.9986	23	0.9975	28	0.9962
19	0.9984	24	0.9973	29	0.9960

Pre-Laboratory Questions
Introductory Exercises

1. Indicate how many significant figures there are in each of the following measured values. Write answer on blank.

246.32	____	1.008	____	700000.	____
107.854	____	0.00340	____	350.670	____
100.3	____	14.600	____	1.0000	____
0.678	____	0.0001	____	320001	____

2. Calculate the answers to the appropriate number of significant figures.

$$\begin{array}{r} 32.567 \\ 135 \\ \underline{+\quad 1.4567} \end{array} \qquad \begin{array}{r} 246.24 \\ 238.278 \\ \underline{+\ \ 98.3} \end{array} \qquad \begin{array}{r} 658.0 \\ \underline{-\ 23.5678} \end{array}$$

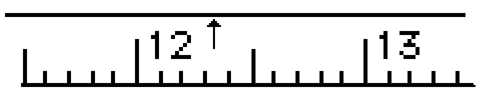
3. Calculate the answers to the appropriate number of significant figures.

a) 23.7 X 3.8 =

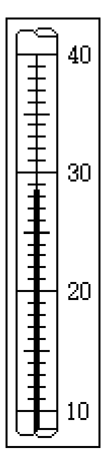
b) 28.367 / 3.74 =

c) 1.678 / 0.42 =

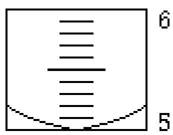
4. In the sketch below, how long is the line at the point indicated by the arrow? The numbered divisions are for cm. Report your answer with the correct number of significant figures.



5. What is the temperature indicated on the thermometer below? The thermometer measures temperature in degrees Celsius. Report your answer with the correct number of significant figures.



6. How many mL of water are in the graduated cylinder below? Report your answer with the correct number of significant figures.



7. Fill in the blanks.

   (a) If a property is measured, the measurement is expressed with a _________________ and a ___________.

   (b) Accuracy refers to the ________________ of a measurement to the _________________ value.

   (c) Precision refers to the _______________ of two or more measurements to _______________________.

   (d) When making measurements, the ________________ device determines the _______________________________________ in the recorded measurement.

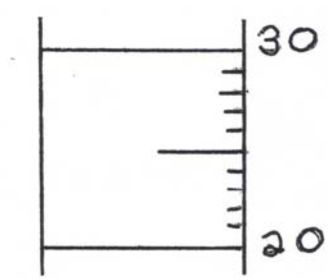
   (e) Having more significant figures in a measurement will result in a measurement that is more ____________________.

   (f) The digit that is written last is the _____________________ one.

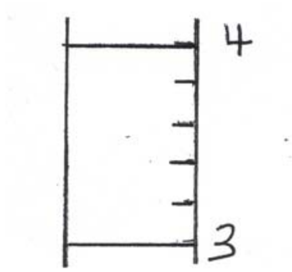
   (g) The answers to the calculations should not be more/less ____________ than the measurements.

8. A piece of lead has a mass of 226 g. When placed in 222 mL of water, the level rose to 242 mL. What is the density of the lead?

9. Draw a meniscus on the drawing below to indicate 21.0 mL.



10. Draw a meniscus on the drawing below to indicate 3.70 mL.



11. A student weighed 3 rocks. The masses were 2.5 g, 2.6 g, and 2.7 g. What is the average absolute deviation of these masses? Report your answer with the correct number of significant figures. Show work.

12. A student calculated the density of a liquid to be 1.555 g/mL. According to a table, it should have been 1.888 g/mL. What was his percent error? Show work.

13. •If you have a cube that is 1 cm x 1 cm x 1 cm, the volume is 1 cm^3.
•If we fill this cube with water and transfer the water to a graduated cylinder, the volume would be 1 mL.
•So, the volumes 1 mL and 1 cm^3 are the same.
•cm^3 is read cubic centimeter (sometimes abbreviated cc).

(a) 1 mL = ___ cm^3 = ___ cc

(b) 2 mL = ___ cm^3 = ___ cc

(c) What is the volume of a box that is 2 cm x 3 cm x 1 cm?

(d) Express your answer for question (c) with mL as the unit.

14. If a line is 12.55 cm long, what is the unit for this measurement?

15. Know the definitions of the following in order to answer questions.

(a) meniscus
(b) volume displacement
(c) density
(d) mass
(e) weight
(f) volume
(g) accuracy
(h) precision

Report Sheet: Introductory Exercises

Name ________________________________ Lab Partner's Name ________________________________

Balance and Mass

1. What were the masses of each of the five pennies?

   #1
   #2
   #3
   #4
   #5

2. What was the sum of the masses (Σ mass) of the five pennies?

3. What was the average mass of the five pennies?

4. What was the absolute deviation for each of the five pennies?

   #1
   #2
   #3
   #4
   #5

5. What was the sum of the deviations (Σd) of the five pennies?

6. What was the average absolute deviation of the five pennies?

Density of an Object Determined by Using Volume Displacement

1. What was the mass of the stopper?

2. What were your measurements for (1)-(3)?

   (1) __________ mL (water and stopper)

   (2) __________ mL (water)

   (3) __________ mL (stopper)

3. What was the density of the stopper?

Density of Water

1. What was the mass of the 10.0 mL of water?

2. What was the calculated density of the water?

3. What was your percent error?

Writing Formulas and Nomenclature of Compounds
Balancing Chemical Equations

Ions and Ionic Compounds

When an atom reacts in a chemical reaction, it is doing so to achieve stability. Some elements are more reactive than others. This means that the atoms of these elements are less stable. The noble gases undergo very few reactions so these elements are very stable. The stability of the noble gases is due to the electron placement in these elements.

Some of the other elements can achieve stability by losing or gaining electrons. This process produces a structure that has the electron placement of a noble gas. The number of protons and neutrons in the nucleus does not change. Before the change, the number of protons and the number of electrons were equal. The resulting structure does not have an equal number of protons and electrons. It is no longer an atom; it is no longer electrically neutral. The structure is an ion, which is a charged structure.

The metals in Group 1 lose one electron. By doing so, each will have the electron placement of the noble gas preceding it on the Periodic Table. Any atom losing one electron produces an ion with a 1+ charge. To write the formula of the ion

1) write the symbol of the element that formed the ion
2) write the charge as a superscript in the upper right-hand corner.

Example: Na^{1+}

The metals in Group 2 lose two electrons. By doing so, each will have the electron placement of the noble gas preceding it on the Periodic Table. Any atom losing two electrons produces an ion with a 2+ charge.

Example: Mg^{2+}

Aluminum in Group 13 will lose three electrons to have the same number of electrons as neon. The charge on an aluminum ion is 3+.

These are the only metals (Group 1, 2, and Al) that you can look at the Periodic Table to determine the charge on the ions that form.

The nonmetals in Group 17 will gain one electron to have the electron placement of the noble gas directly after them on the Periodic Table. These nonmetals form ions with a 1- charge.

Example: Cl^{1-}

The nonmetals in Group 16 will gain two electrons to have the electron placement of the noble gas directly after them on the Periodic Table. These nonmetals form ions with a 2- charge.

Example: O^{2-}

The nonmetals in Group 15 will gain three electrons to have the electron placement of the noble gas directly after them on the Periodic Table. These nonmetals form ions with a 3- charge.

Example: N^{3-}

Positive and negative ions always form together. It is a complementary process. For ions to form, at least two atoms must be present: one to lose an electron and another to gain the donated electron. When table salt (NaCl) is formed, a sodium atom transfers a single electron to a chlorine atom. The sodium atom, in losing an electron, becomes positively charged. The chlorine atom, in gaining an electron, becomes negatively charged. Because of their opposite charges, sodium ions (Na^{1+}) and chloride ions (Cl^{1-}) are attracted to each other. This attraction that holds the ions together is called an ionic bond, and the compounds that are formed, such as sodium chloride, are called ionic compounds.

An ion of an element is not the same thing as an atom of an element. You cannot put only sodium ions in a container. If you have sodium ions, there will also be negative ions present. However, you can put a sample of sodium (which is composed of sodium atoms) in a container. Another example of the difference in an atom and an ion is that if you put sodium (atoms) into your body, you will regret it. You will be severely injured. However, to stay alive your body has to have sodium ions.

Compounds are substances that are made from two or more elements. The composition of a compound can be represented by a chemical formula. The formula indicates the number of atoms of each element that were used to make the smallest unit of the compound. The numerical subscripts to the right of the symbols for the elements indicate the number of atoms of each element used. If there is no subscript following the element, it means that only one atom was used. Here are some examples of formulas of compounds and an explanation of each formula.

NaCl: one sodium atom and one chlorine atom were used to make a formula unit of the compound sodium chloride

$MgCl_2$: one magnesium atom and two chlorine atoms were used to make a formula unit of the compound magnesium chloride

Ionic compounds are formed primarily when metals on the left-hand side of the periodic table donate electrons to nonmetals (excluding the noble gases) on the right-hand side of the table.

Ionic compounds containing just two different elements are named as follows: The name of the metallic element is given and then, as a separate word, the name stem of the nonmetallic element with the suffix, -ide. The name of Mg_3N_2 is magnesium nitride.

Names of Common Nonmetallic Ions

Fluoride	Chloride	Bromide	Iodide
Oxide	Sulfide	Nitride	Phosphide

In the previous material, the charges on the ions made from metals were based on the losing of electrons to attain the electronic placement of a noble gas. Not all metals do this. They do lose electrons, but will not have the same number of electrons as any noble gas. For example, zinc forms an ion with a 2+ charge. This ion has the number of electrons that are found in an atom of nickel, which is not a noble gas. Some of these metals can form more than one ion. The number of electrons lost depends on reaction conditions. Below are examples. When writing the name of an ionic compound that has such a metal in it, a Roman numeral is required to let you know which ion is being made or used.

Cu^{1+}	copper (I)
Cu^{2+}	copper (II)
Fe^{2+}	iron (II)
Fe^{3+}	iron (III)
Sn^{2+}	tin (II)
Sn^{4+}	tin (IV)
Pb^{2+}	lead (II)
Pb^{4+}	lead (IV)

When writing formulas of ionic compounds, the formulas are written as follows:

$$\text{Positive ion}_x\ \text{Negative ion}_y$$

where x and y are subscripts indicating the combining ratio of positive and negative ions. If a subscript is one, it is not written. The combining ratio is such that the number of electrons lost by the atoms forming positive ions is exactly the same as the number of electrons gained by the atoms forming negative ions.

The formula of the ionic compound, sodium chloride, is NaCl. This formula indicates a combining ratio of one sodium ion with one chloride ion. An atom of sodium loses one electron when forming an ion. An atom of chlorine gains one electron when forming an ion. Therefore, the combining ratio is one to one. To determine the combining ratio, you need to

(1) know the charge of each ion

Na^{1+} Cl^{1-}

(2) determine how many of each of the positive and negative ions are needed to get a sum of zero.

1+ 1- = 0

*Notice: When writing the formula of an ion, the charge is included. When writing the formula of an ionic compound, the charges of the ions are not written. Also the positive ion is written first.

Formula of Ions: Na^{1+}, Cl^{1-}

Formula of Compound: NaCl

The formula of the ionic compound, magnesium chloride, is $MgCl_2$. Magnesium loses two electrons when forming an ion; chlorine gains only one. The combining ratio cannot be one to one. Two chlorine atoms will be needed so that the magnesium ion will form. Each of the chlorines will accept one of magnesium's lost electrons.

(1) Mg^{2+} Cl^{1-}

(2) 2+ 1- 1- = 0

Mg Cl Cl $MgCl_2$

The formula of the ionic compound, aluminum oxide, is Al_2O_3.

(1) Al^{3+} O^{2-}

(2) 3+ 3+ 2- 2- 2- = 0

Al Al O O O Al_2O_3

The compound's formula represents the smallest combining ratio, so if you use the crossover short cut, be careful. Some instructors will mark an answer such as Mg_2O_2 as an incorrect answer. The person writing this document will mark it as an incorrect answer.

When naming the compounds of metallic elements that form more than one ion, the Roman numeral needs to be included. $FeCl_2$ is named iron (II) chloride. $FeCl_3$ is named iron (III) chloride. If given the formula of such a compound and asked to name it, you will need to determine the charge on the metallic ion. For $FeCl_2$, here is the process for doing so. The formula of the compound tells you that one iron ion combines with two chloride ions. The 1- charge on the chloride ion is known because chlorine is one of the elements that you can determine its charge by looking at the Periodic Table. The sum of all the charges in the compound's formulas should be zero. This can be only if the charge on the iron ion is 2+. You might do the following scratch work to determine the charge on the iron ion.

Fe Cl^{1-}
Cl^{1-}

The charge on Fe has to be 2+ to get a sum of zero.

A polyatomic ion is made from more than one atom. These atoms are bonded together by using covalent bonds. There is an unequal number of protons and electrons. An example of a positive polyatomic ion is the ammonium ion, NH_4^{1+}. This ion is made from one nitrogen atom and four hydrogen atoms. It has a one positive charge because there is one more proton than electrons. An example of a negative polyatomic ion is the nitrate ion, NO_3^{1-}. This ion is made from one nitrogen atom and three oxygen atoms. It has a one negative charge because there is one more electron than protons. Below is a list of the formulas and names of some polyatomic ions.

Ion Name	Formula
Ammonium	NH_4^{1+}
Hydronium	H_3O^{1+}
Acetate	$C_2H_3O_2^{1-}$
Bicarbonate	HCO_3^{1-}
Hydroxide	OH^{1-}
Nitrate	NO_3^{1-}
Permanganate	MnO_4^{1-}
Carbonate	CO_3^{2-}
Dichromate	$Cr_2O_7^{2-}$
Sulfate	SO_4^{2-}
Phosphate	PO_4^{3-}

If an ionic compound contains a polyatomic ion, then an extra rule for writing the compound's formula is required. If more than one of the polyatomic ions is used, then parentheses are to be placed around the formula of the polyatomic ion before writing the subscript that indicates how many of the polyatomic ions are used.

The formula of the ammonium ion is NH_4^{1+}. The formula of the chloride ion is Cl^{1-}. The formula of ammonium chloride is NH_4Cl. The ratio of positive ion to negative ion is one to one. No parentheses are used since only one ammonium ion is used.

The formula of ammonium sulfide is $(NH_4)_2S$. The sulfide ion has a 2- charge, so two ammonium ions are needed for a sum of zero (1+ 1+ 2- = 0). The formula of the ammonium ion is NH_4^{1+}. The 1+ charge is not written in the formula of the compound. The 4 after the H is part of the polyatomic ion's formula. It is not a subscript for indicating the ratio of positive to negative ions in the formula of the compound. The 4 needs to be inside the parentheses.

The formula of the sulfate ion is SO_4^{2-}. The formula of the sodium ion is Na^{1+}. The formula of sodium sulfate is Na_2SO_4. Two sodium ions are needed to give a sum of zero (1+ 1+ 2- = 0). No parentheses are placed around the Na because the sodium ion is not a polyatomic ion.

The formula of ammonium sulfate is $(NH_4)_2SO_4$. Two ammonium ions are needed for a sum of zero (1+ 1+ 2- = 0). Parentheses are not required for the sulfate ion, since only one was used.

Binary Covalent Compounds

Compounds composed of two nonmetals (binary) usually are not ionic, but covalent. Covalent compounds are composed of atoms bonding together with covalent bonds. There are no ions, only atoms bonding together. Formulas of compounds such as these are written so that the subscripts (x and y) are whole numbers. If not specified, a subscript of 1 is implied. Use the following rules for naming binary covalent compounds:

1. If MORE than one atom of the first element is present in the formula, state the prefix corresponding to that number. Do not use a prefix if there is only one atom present.

2. State the name of the first element.

3. ALWAYS state the prefix corresponding to the subscript of the second element (mono for one).

4. State the name of the root of the second element followed by the suffix –ide. This will be the ionic name of the second element even though it is not an ion.

The following list includes the prefixes corresponding to the number of an element's atoms present in the formula.

mono = 1	hexa = 6
di = 2	hepta = 7
tri = 3	octa = 8
tetra = 4	nona = 9
penta = 5	deca = 10

Example: Formula: OF_2
Name: Oxygen difluoride

Example: Formula: P_2O_5
Name: diphoshorous pentoxide (or pentaoxide)

Example: Formula: CO
Name: carbon monoxide (or monooxide)

Chemical Equation

•A chemical equation is written for a chemical reaction. The reactant(s) is/are written first, then draw an arrow pointing toward the product(s).

$$\text{reactant(s)} \xrightarrow{\text{form}} \text{product(s)}$$

•Formulas of reactants and products are separated with a + sign. The chemical equation below would be read: A reacts with B to form C ***and*** D.

$$A + B \rightarrow C + D$$

•Chemical reactions do not form other elements. Any element appearing in a reactant will be in a product.

•Law of Conservation of Mass: the mass of the reactants is equal to the mass of the products. No new mass is made, nor is mass lost. The equation, therefore, needs to be balanced.

Balancing the equation

FIRST: Correct formulas of reactants and products are written first.

SECOND: A coefficient is the number in front of the formula of each reactant and each product. If you see no number, it is understood to be one. Balance the equation by altering coefficients, NOT a subscript. Formulas of reactants and products use subscripts. If you change a subscript, you are changing a correct formula to an incorrect one.

If there is a polyatomic ion on the left side of the arrow and it is also on the right side of the arrow, do not break it up into its individual elements. In the chemical equation below, the sulfate ion (SO_4) and the nitrate ion (NO_3) are on both sides of the arrow. So, at the start of balancing this equation, you would count how many sulfates and how many nitrates are on the left and right. You should not say how many sulfurs or how many nitrogens are on the left and right. After balancing, there are 3 sulfates on each side and 6 nitrates on each side.

$$Al_2(SO_4)_3 + 3\ Pb(NO_3)_2 \longrightarrow 2\ Al(NO_3)_3 + 3\ PbSO_4$$

Pre-Laboratory Questions
Ionic Compounds and Covalent Compounds
Writing Formulas and Nomenclature

1. Which elements are the most stable?

2. By being stable, do these elements undergo a lot or very few chemical reactions?

3. In an atom, the number of protons is the same as the number of __________.

4. How do some elements achieve stability?

5. When an atom forms an ion, does the nucleus change?

6. Which subatomic particles are in the nucleus?

7. What is an ion?

8. Is the number of protons equal to the number of electrons in an ion?

9. Is an atom electrically neutral?

10. Is an ion electrically neutral?

11. List the METALS in Group 1.

12. Lithium is a metal that loses one electron when forming ions. Write the symbol of a lithium ion.

13. List the metals in Group 2.

14. Calcium is a metal that loses two electrons when forming ions. Write the symbol of a calcium ion.

15. Aluminum is a metal that loses three electron when forming ions. Write the symbol of the aluminum ion.

16. When metals form ions, a (gain, loss) of electrons results in (negative, positive) charges.

17. Bromine is a nonmetal that gains one electron when forming ions. Write the symbol of a bromine ion.

18. Sulfur is a nonmetal that gains two electrons when forming ions. Write the symbol of a sulfur ion.

19. Phosphorus is a nonmetal that gains three electrons when forming ions. Write the symbol of a phosphorus ion.

20. When nonmetals form ions, a (gain, loss) of electrons results in (negative, positive) charges.

21. Define compound.

22. What is a chemical formula used for?

23. Is a subscript written above a symbol or below a symbol?

24. Is a charge written above a symbol or below a symbol?

25. For the following formulas of ionic compounds, give the indicated information.

    On the first blank, write the symbol of the metal. In the set of (), write how many atoms of the metal were used. On the second blank, write the symbol of the nonmetal. In the second set of (), write how many atoms of the nonmetal were used. On the long blank, write the name of the ionic compound.

    (a) Al_2O_3 _____ ( ), _____ ( ), ___________________________

    (b) LiBr _____ ( ), _____ ( ), ___________________________

    (c) Ba_3N_2 _____ ( ), _____ ( ), ___________________________

    (d) K_2S _____ ( ), _____ ( ), ___________________________

26. Do oppositely charged ions attract or repel each other?

27. The bond that holds ions together is called an __________ bond and a compound that is made from ions is called an _________ compound.

28. When writing the formula of an ionic compound, the (negative, positive) ion is written first.

29. When writing the formula of an ionic compound, subscripts (are, are not) written and charges (are, are not) written.

30. Ionic compounds are formed primarily when metals __________ electrons to nonmetals, excluding the ____________________.

31. (a) Gaining electrons results in (negative, positive) ions.
    (b) Losing electrons results in (negative, positive) ions

32. Write the symbols of the ion with the following number of protons and electrons. Do not forget number and sign.

    (a) 3 protons, 2 electrons

    (b) 12 protons, 10 electrons

    (c) 9 protons, 10 electrons

    (d) 26 protons, 23 electrons

    (e) 30 protons, 28 electrons

33. How many protons and electrons are in the following ions?

		Protons	Electrons
(a)	O^{2-}		
(b)	K^{1+}		
(c)	S^{2-}		
(d)	Fe^{3+}		

34. Which of the following pairs of elements are likely to form an ionic compound? Look at your answer for #30 to help answer this question. If your answer is no, explain why.

    (a) lithium and chlorine

    (b) potassium and oxygen

    (c) sodium and neon

    (d) oxygen and fluorine

    (e) strontium and copper

35. Do metals gain or lose electrons when forming ions?

36. Can some metals form more than one ion?

37. Why is a Roman numeral placed after the names of some metallic ions?

38. Why is a Roman numeral not placed after the names of some metallic ions?

39. Write the symbol of the following ions:

copper (II)	_______	iron (III)	_______
silver(I)	_______	gold (I)	_______
lead(II)	_______	lead(IV)	_______
zinc(II)	_______	nickel(II)	_______
iron (II)	_______	iron(III)	_______

40. Write the formulas of these ions.

sodium	_______	fluoride	_______
lithium	_______	iodide	_______
potassium	_______	chloride	_______
magnesium	_______	bromide	_______
calcium	_______	oxide	_______
barium	_______	sulfide	_______
strontium	_______	selenide	_______
aluminum	_______	nitride	_______

41. What is the extra rule for writing the formula of an ionic compound's formula if a polyatomic ion is used?

42. The formula of the nitrate ion is NO_3^{1-}.

(a) Will the 1- appear in the compound's formula?

(b) Will the 3 appear in the compound's formula?

(c) If parentheses are needed for the nitrate ion, will the 3 be inside or outside of the parentheses?

43. The formula of the ionic compound with the name puppies kitties shows a 1:2 ratio of positive to negative ions. Pp is the symbol for puppies. Ki is the symbol for kitties. Write the compound's formula.

44. The formula of the ionic compound dogs cats shows a 3:2 ratio of positive to negative ions. Do is the symbol for dogs. Ct is the symbol for cats. Write the compound's formula.

45. Do all metals achieve the electronic placement of a noble gas when forming ions?

46. Silver loses one electron when forming an ion. The ion will have _______ for its charge and will have the same number of electrons as an atom of _________ .

47. Can some metals form more than one ion?

48. What needs to be around the Roman numeral in the name for an ion that uses a Roman numeral?

49. How many elements are in a binary compound?

50. If a binary compound only is composed of nonmetals, then it is probably (covalent, ionic).

51. Covalent compounds are made up of (atoms, ions).

52. Are binary covalent compounds named in the same manner as ionic compounds?

53. Do ionic compounds use prefixes in the names?

54. Do binary covalent compounds use Roman numerals in the names?

55. Use the equation below to answer questions a-h .

$$2\ HCl + Mg(OH)_2 \rightarrow 2\ H_2O + MgCl_2$$

a) Which compounds are the reactants?

b) Which compounds are the products?

c) The + on the left is read as (reacts with, and).

d) The $\rightarrow$ is read as _________.

e) The + on the right side is read as (reacts with, and).

f) The number in front of each formula is called a (coefficient, subscript).

g) The small numbers contained within the formulas are called (coefficients, subscripts).

56. Balance the following equations. Do not change any formulas; only the coefficients.

a) $H_2 + Cl_2 \rightarrow HCl$

b) For this equation, notice that C is in both products.

$C + SO_2 \rightarrow CS_2 + CO$

c) $Mg + O_2 \rightarrow MgO$

d) $PCl_5 \rightarrow PCl_3 + Cl_2$

e) $Mg + HNO_3 \rightarrow Mg(NO_3)_2 + H_2$

f) $Fe + O_2 \rightarrow Fe_2O_3$

g) $CH_4 + O_2 \rightarrow CO_2 + H_2O$

h) $Sb_2S_3 + HCl \rightarrow SbCl_3 + H_2S$

Lab Report
Writing Formulas and Nomenclature of Compounds
Balancing Chemical Equations

1. Write the formula of the ionic compound formed between ions of the following elements. Then name the compound.

a. sodium and chlorine

b. magnesium and fluorine

c. aluminum and iodine

d. potassium and oxygen

e. calcium and sulfur

f. aluminum and sulfur

g. lithium and nitrogen

h. aluminum and nitrogen

i. barium and nitrogen

j. mercury (II) and oxygen

k. lead (IV) and sulfur

l. lead (II) and sulfur

m. copper (I) and oxygen

n. copper (II) and oxygen

o. silver (I) and selenium

p. sodium and oxygen

2. Write the formulas of the following ionic compounds that contain polyatomic ions.

a. potassium sulfate

b. beryllium hydroxide

c. lead (II) acetate

d. tin (IV) carbonate

e. copper (II) sulfate

f. ammonium sulfate

g. calcium carbonate

h. magnesium phosphate

i. silver (I) phosphate

j. aluminum nitrate

k. tin (IV) hydroxide

l. lithium sulfate

m. barium carbonate

n. magnesium carbonate

o. ammonium phosphate

p. magnesium hydroxide

q. lithium hydroxide

3. Name the following ionic compounds. None of these need a Roman numeral.

a. K_2S

b. $CaCl_2$

c. Na_2CO_3

d. AlP

e. $BaBr_2$

f. LiCl

g. $CaSO_4$

h. Mg_3N_2

i. NaBr

j. K_ePO_4

k. K_3P

4. Name these ionic compounds. These do need Roman numerals in the names.

a. SnO_2

b. HgO

5. Name these binary covalent compounds.

a. CF_4

b. P_4O_{10}

c. SF_6

d. S_2F_2

e. SO_3

f. PCl_3

g. NO

6. Write the formula for these binary covalent compounds.

a. Phosphorus pentachloride

b. Carbon dioxide

c. Dihydrogen monoxide

d. Carbon tetrachloride

e. Dinitrogen trioxide

7. Balance the following equations. Do not change any formulas.

1. ___ H_2 + ___ O_2 $\rightarrow$ ___ H_2O
2. ___ N_2 + ___ H_2 $\rightarrow$ ___ NH_3
3. ___ S_8 + ___ O_2 $\rightarrow$ ___ SO_3
4. ___ HgO $\rightarrow$ ___ Hg + ___ O_2
5. ___ Zn + ___ HCl $\rightarrow$ ___ $ZnCl_2$ + ___ H_2
6. ___ $C_{10}H_{16}$ + ___ Cl_2 $\rightarrow$ ___ C + ___ HCl
7. ___ Fe + ___ O_2 $\rightarrow$ ___ Fe_2O_3
8. ___ $Fe_2(SO_4)_3$ + ___ KOH $\rightarrow$ ___ K_2SO_4 + ___ $Fe(OH)_3$
9. ___ Al + ___ FeO $\rightarrow$ ___ Al_2O_3 + ___ Fe
10. ___ K + ___ Br_2 $\rightarrow$ ___ KBr
11. ___ P_4 + ___ O_2 $\rightarrow$ ___ P_2O_5
12. ___ $MgBr_2$ $\rightarrow$ ___ Mg + ___ Br_2
13. ___ Al + ___ O_2 $\rightarrow$ ___ Al_2O_3
14. ___ SiO_2 + ___ HF $\rightarrow$ ___ SiF_4 + ___ H_2O
15. ___ C + ___ H_2O $\rightarrow$ ___ CO + ___ H_2
16. ___ $KClO_3$ $\rightarrow$ ___ KCl + ___ O_2
17. ___ $Al_2(SO_4)_3$ + ___ $Ca(OH)_2$ $\rightarrow$ ___ $Al(OH)_3$ + ___ $CaSO_4$
18. ___ $FeCl_3$ + ___ NH_4OH $\rightarrow$ ___ $Fe(OH)_3$ + ___ NH_4Cl
19. ___ Sb + ___ O_2 $\rightarrow$ ___ Sb_4O_6
20. ___ Al + ___ HCl $\rightarrow$ ___ $AlCl_3$ + ___ H_2
21. ___ H_2S + ___ Cl_2 $\rightarrow$ ___ S_8 + ___ HCl
22. ___ Mg + ___ N_2 ---> ___ Mg_3N_2

Double Replacement Reactions

In a double replacement reaction, two compounds react. If the two compounds are ionic, the positive ion of one of the compounds swaps places with the positive ion of the other compound. We can use letters to illustrate this type of reaction.

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

The compounds that are going to be used in today's double replacement reactions are ionic compounds and are solids. They are soluble in water. In today's experiment, you will work with aqueous solutions of these compounds.

A precipitate is a solid product that forms and is not water soluble. Reactions between two ionic compounds can produce two new ionic compounds. If solutions of two ionic compounds are mixed and a precipitate forms, then you know a reaction has occurred. This is known because the original compounds were soluble in water; the precipitate is not.

An example of such a reaction is below.

$$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow NaNO_{3(aq)} + AgCl_{(s)}$$

You will notice in the above equation that the ions swapped partners during the reaction. The positive ion in each of the reactants has the other negative ion associated with it in the products. The (s) after AgCl means it is a solid and not soluble in water. It is the precipitate. The (aq) and the (s) are not part of the chemical's formula. These are written to provide information.

The following are the names and formulas of the compounds used in today's work.

Potassium chloride, KCl
Lead (II) nitrate, $Pb(NO_3)_2$
Aluminum sulfate, $Al_2(SO_4)_3$
Sodium oxalate, $Na_2C_2O_4$
Copper (II) chloride, $CuCl_2$

In today's work, you will mix solutions together to see if a precipitate forms. If one does form, you will write the formulas of the ionic products. Using these formulas, you will then write a balanced equation for the reaction. To determine which product is the precipitate, you will use the following solubility rules.

Solubility Rules for This Laboratory

NO_3^- All nitrates are soluble.

Cl^- All chlorides are soluble except those of silver, mercury (I), and lead (II)

Items Needed for Each Section of Lab (Stockroom)

Each of the solutions below should be in bottles with droppers

- KCl, 0.1 M, 100 mL
- $Pb(NO_3)_2$, 0.1 M, 100 mL
- $Al_2(SO_4)_3$, 0.1 M, 100 mL
- $Na_2C_2O_4$, 0.1 M, 100 mL
- $CuCl_2$, 0.1 M, 100 mL

Items Needed for Each Pair of Students (Already on shelf or in drawer)

- 3 test tubes that fit into the centrifuge
- Stirring rod

Procedure

1. Get 3 test tubes and a test tube rack. Do not empty the content of the test tubes until told to do so.

2. Add 5 drops of 0.1 M sodium oxalate to one of the test tubes.
   Add 5 drops of 0.1 M potassium chloride to the same test tube.
   Did a precipitate form?

3. (a) Add 5 drops of 0.1 M aluminum sulfate to another test tube.
   Add 5 drops of 0.1 M lead (II) nitrate to the same test tube.
   Did a precipitate form?

   (b) Color, state (solid, liquid, gas), and solubility are examples of physical properties. Fill in the information of these 3 physical properties for the precipitate.

   Color:
   State:
   Is it water soluble?

   (c) Name the 2 reactants.

   (d) Name the 2 products. Do not forget the Roman numeral.

   (e) Write the formulas of the reactants.

   (f) Write the formulas of the ions in the reactants. Separate the formulas with commas. Do not forget to include the charges.

   (g) Write the formulas of the products.

   (h) Write a balanced equation for this reaction.

(i) Look at the water solubility chart. Which product is water-soluble?

(j) Which product is the precipitate?

(k) Indicate the water solubility of all the compounds in the equation by adding (aq) or (s) to your answer for question (h).

AT THIS POINT CHECK YOUR WORK WITH YOUR INSTRUCTOR.

4. (a) Add 5 drops of 0.1 M copper (II) chloride to another test tube.
Add 5 drops of 0.1 M sodium oxalate to the same test tube.
Did a precipitate form?

(b) Color, state (solid, liquid, gas), and solubility are examples of physical properties. Fill in the information of these 3 physical properties for the precipitate.

Color:
State:
Is it water soluble?

(c) Name the 2 reactants.

(d) Name the 2 products. Do not forget the Roman numeral.

(e) Write the formulas of the reactants.

(f) Write the formulas of the ions in the reactants. Separate the formulas with commas. Do not forget to include the charges.

(g) Write the formulas of the products.

(h) Write a balanced equation for this reaction.

(i) Look at the water solubility chart. Which product is water-soluble?

(j) Which product is the precipitate?

(k) Indicate the water solubility of all the compounds in the equation by adding (aq) or (s) to your answer for question (h).

AT THIS POINT CHECK YOUR WORK WITH YOUR INSTRUCTOR.

5. (a) Take your 2 test tubes containing precipitates to a centrifuge. Place them in the centrifuge to evenly distribute the weight of the test tubes and prevent the centrifuge 'walking'. Other students can also place their test tubes in the centrifuge at the same time.

   (b) Turn on the centrifuge and let run for about 60 seconds.

   (c) Turn off the centrifuge. Do not try to make it stop spinning with your hands.

   (d) CAREFULLY remove the test tubes. Do not jiggle them or the contents will get mixed together.

   (e) Read all of this before doing it. Decant means to pour off the aqueous solution as much as possible to separate it from the precipitate. Decant both test tubes. While doing this, try not to jiggle the test tubes, which would result in remixing the contents. You don't want this to happen. There will still be some water in the test tubes, but that is okay.

6. (a) Tear off a sheet of paper towel. Lay it down on the bench.

   (b) With a stirring rod, remove as much as possible the precipitate from a test tube and place it on the paper towel.

   (c) Wipe off the stirring rod and do the same for the other precipitate, but place it on a different area of the paper towel.

   (d) Are these looking more like a solid after getting them out of the water?

When finished, wash out all test tubes with soap and water **USING** a test tube brush. Rinse well. Store upside down.

Pre-Laboratory Questions
Double Replacement Reactions

1. In a double replacement reaction, how many reactants react?

2. In a double replacement reaction, how many products are formed?

3. If the reactants are ionic compounds, the positive ion of one reactant will displace the _______________ ion of the other reactant.

4. What does the word aqueous mean?

5. What is the definition of precipitate?

6. Which letter inside parentheses do you use to indicate a precipitate in an equation?

7. What does decant mean?

8. The following list is for 5 ionic compounds. For the metals without Roman numerals, you can determined the ion's charge by looking at the Periodic Table. For the negative ions, you can determine the charges by looking at the charge of the metal AND by looking at the formula of the compound.

   Potassium chloride, KCl
   Lead (II) nitrate, $Pb(NO_3)_2$
   Aluminum sulfate, $Al_2(SO_4)_3$
   Sodium oxalate, $Na_2C_2O_4$
   Copper (II) chloride, $CuCl_2$

   What is the formula for the following ions? Do not forget to include charges.

   (a) potassium

   (b) chloride

   (c) lead (II)

   (d) nitrate

   (e) aluminum

   (f) sulfate

   (g) sodium

   (h) oxalate

   (i) copper (II)

   (j) chloride

9. For potassium chloride, what would be present for the following elements on a GHS label?

   Pictogram(s):

   Signal Word:

   Hazard Statement:

   Precautionary Statement:

10. For lead (II) nitrate, What would be present for the following elements on a GHS label?

   Pictogram(s):

   Signal Word:

   Hazard Statement:

   Precautionary Statement:

11. For aluminum sulfate, what would be present for the following elements on a GHS label?

   Pictogram(s):

   Signal Word:

   Hazard Statement:

   Precautionary Statement:

12. For sodium oxalate, what would be present for the following elements on a GHS label?

   Pictogram(s):

   Signal Word:

   Hazard Statement:

   Precautionary Statement:

13. For copper (II) chloride, what would be present for the following elements on a GHS label?

   Pictogram(s):

   Signal Word:

   Hazard Statement:

   Precautionary Statement:

Report Sheet: Double Replacement

Name ___________________________________

Lab Partner’s Name _________________________________

1. Write the balanced equations for the 2 double replacement reactions done in lab today. Include (aq) and (s).

2. Name the four products. Don’t forget any Roman numeral needed.

3. (a) All precipitates were (solids, gases, liquids).

   (b) All precipitates (were, were not) water soluble.

   (c) All precipitates were (reactants, products).

Combustion

Usually when someone says oxygen gas, it is a reference to O_2, which makes up about 21% of the air. O_3 is called ozone and is not part of today's discussion. Common reactions with oxygen gas, O_2, are combustion and rusting. Combustion is a fast, observable reaction with oxygen gas. You can usually observe a flame being produced. For combustion to occur not only do you need oxygen gas, you need a combustible material such as a fuel. A fuel is any material that stores energy that can later be extracted to perform work in a controlled manner. Most fuels used by humans undergo combustion. An example of a fuel that does not undergo combustion is nuclear fuel. Hydrocarbons are by far the most common source of fuels. The predominant sources of the hydrocarbons are fossil fuels which are mostly hydrocarbons. A solid fossil fuel is coal; a liquid fossil fuel is petroleum; and a gaseous fossil fuel is natural gas. Petroleum is a mixture of liquid hydrocarbons. Some of these hydrocarbons in petroleum are separated and sold as gasoline. Some others are separated and sold as diesel. Natural gas is a mixture of gaseous hydrocarbons. The main gaseous hydrocarbon in natural gas is methane, CH_4.

Methane (or any fuel) will not react with the oxygen gas if the methane and oxygen molecules do not have enough energy. If the molecules do not have enough energy, the collision of the methane molecules with the oxygen molecules is too weak. They more or less just bounce off of each other. When you want a fuel to burn, you need to supply a little energy to increase the energy content of the fuel and/or the oxygen. This usually involves friction (striking a match) or a spark (spark plugs in a car's engine). Once the fuel starts burning, the reaction releases heat energy. This heat energy supplies the energy to keep the fuel and the oxygen molecules moving fast enough to hit and react. Combustion will stop if the fuel is used up, if there is not enough oxygen gas, or if there is not enough heat available to keep the reactant molecules hitting hard enough.

Complete Combustion of a Hydrocarbon

Complete combustion of a hydrocarbon produces carbon dioxide and water and releases large amounts of energy.

Equation (1) $$\text{Hydrocarbon} + O_{2(g)} \rightarrow \text{Carbon dioxide}_{(g)} + \text{Water} + \text{Energy}$$

Below is the equation for the combustion of methane. The (g) written after the formulas of the reactants and products lets you know that all are gases when you do this reaction at normal atmospheric pressure and room temperature. The big twos in front of the formulas for oxygen gas and water are needed to balance the equation. The big twos are not part of the formulas.

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$$

The energy that is released is in the form of heat energy, light energy, and sound energy. You know this because when you are around a fire, you can feel the heat, see the light, and hear it burning. The heat energy released, along with their availability and relatively low cost, makes hydrocarbons very useful as fuels.

Incomplete combustion of a Hydrocarbon

100% complete combustion of a fuel is not very likely. Here are two equations for other reactions that occur when a hydrocarbon burns.

Equation (2) Hydrocarbon + Oxygen gas $\longrightarrow$ $CO_{(g)}$ + Water

Equation (3) Hydrocarbon + Oxygen gas $\longrightarrow$ $C_{(s)}$ + Water

These carbon products (in Equations 2 and 3) are desired for certain industrial purposes. One industrial use of CO is to convert coal or biomass to diesel fuel. An industrial use of C (carbon black) is in printer toner. To obtain a higher percentage of these carbon products, industry will use less oxygen.

However, when burning hydrocarbons as a fuel the production of these two carbon products is not considered a good thing. There are several reasons that incomplete combustion is not desirable.

For vehicles more energy is released when complete combustion occurs. The more energy released results in your spending less money on fuel and less stress on the vehicle. If less energy is being released, the vehicle is going to have a hard time getting up the hill and will burn more fuel trying to do so. Another reason that incomplete combustion is bad for vehicles is that C (think soot) is not good for engines. It gunks up the engine.

For us, CO and CO_2 are dangerous gases to inhale. Hemoglobin is a protein in red blood cells. One of its functions is to carry O_2 to the other cells. CO bonds with hemoglobin and is hard to dislodge. In fact, the hemoglobin accepts CO more readily than oxygen and hangs on to it longer. A person with an overdose of CO will die sometimes even if they are given pure oxygen because the pure oxygen cannot dislodge the carbon monoxide.

CO_2 is also dangerous, but in a different way. CO_2 does not react with the body as does CO, but if the concentration of CO_2 is too high, then that means that not enough oxygen is available. This can also kill you -- but the effect is more like holding your breath than breathing a toxic chemical.

CO poisoning occurs quite often. Many people do not realize that they are doing something dangerous. Any process that burns a hydrocarbon fuel requires a good supply of oxygen gas. If the combustion is occurring in a closed area or there is not enough fresh air, then the amount of CO will increase. Here are some examples of CO poisoning.

- Early Monday morning (May 20, 2002), security officers found the bodies of two N.C. men in their tent at Lowe's Motor Speedway, apparently victims of carbon monoxide poisoning from a charcoal grill they were using to stay warm.
- Five teenagers celebrating a friend's birthday died from carbon monoxide poisoning after they left their car running in the garage beneath their Florida motel room.
- A couple thought they were playing it safe with the generator set up outside their home. The problem is, the exhaust was spewing into their crawlspace, letting carbon monoxide seep into their home.
- An Oacoma couple were found dead due to apparent carbon monoxide poisoning on their cabin boat at Dock 44 Marina near Snake Creek Recreation area west of Platte.

•ST. LOUIS: A 47-year-old man hooked up a fuel-burning generator in the basement of the house. His wife died.
•An 11-year-old Massachusetts boy died of carbon monoxide poisoning after being overcome as he sat in a running car to keep warm, while his father was shoveling snow to get the car out of a snow bank.
•April 2013: "Buckwild" star Shain Gandee, 21, died from carbon monoxide poisoning from toxic exhaust fumes, according to news reports. People magazine reported that Gandee and his companions were found in his car on Monday, which was stuck in the mud. People explains: With the vehicle knee-deep in mud, toxic fumes from the car were likely unable to escape through the tailpipe, authorities say.

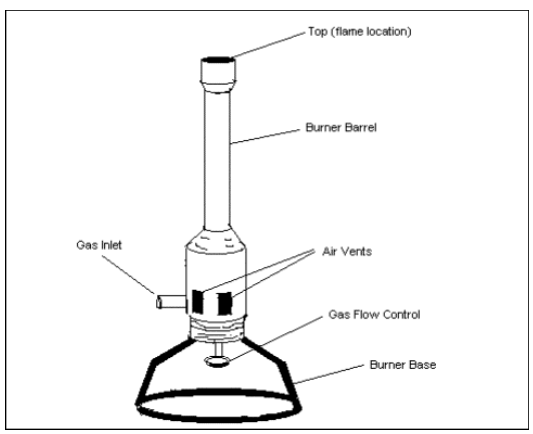
Combustion of Magnesium

Magnesium is silvery-gray metal. It is very flammable. When it burns, it releases a very bright light. Combustion of magnesium can also be classified as a combination reaction. A combination reaction is one in which two or more reactants combine to make one product.

$$2\ Mg_{(s)} + O_{2(g)} \longrightarrow 2\ MgO_{(s)}$$

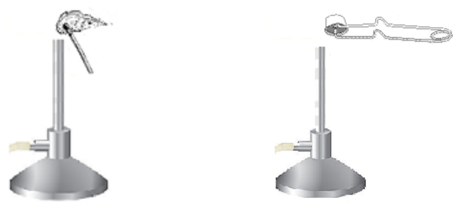
How to Use a Laboratory burner

The laboratory burner most commonly used in an introductory laboratory class is a Tirril burner. It has air and gas adjustments while the Bunsen burner has only an air adjustment. However (even though it is incorrect) many call a Tirril burner a Bunsen burner. In this document, you are just going to see the word 'burner'.



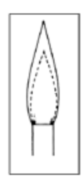
1. Connect the gas inlet of your burner to the bench gas valve by means of a short piece of thin-walled rubber tubing.
2. Close the burner gas control valve by turning the burner upside down and turning the knob clockwise until it just stops (do not overly tighten).
3. Place the burner back down on the bench top. Close the air vents by rotating the barrel of the burner in a clockwise direction until it just stops (do not overly tighten).
4. Turn the bench gas valve to the fully opened position. This will have the movable handle directly over the rubber tubing.

5. Use a match or lighter or striker to ignite the gas. A striker produces sparks. For the match or lighter, bring the flame of the match/lighter from the side ***up*** to the edge of the top of the burner. ***Do not*** put any part of your hand above the top of the burner. The part of the striker that makes sparks has to be over the top of the burner.



6. Open the burner gas valve ***slowly***. If using a striker, make sparks while opening the gas valve ***slowly***.

   *An alternative way is to open the gas valve slowly until you hear hissing. You should have your ear close to the top of the burner to be able to hear as soon as the gas is released. Then use a lighted match, lighter, or striker to ignite the gas.
7. If the flame is too high or too low, adjust the height with the burner gas valve.
8. Open the air control until an all blue flame is produced. Do not open the air too much, or the flame will be blown out. If this happens, close both of the gas valves, close the air control, then start over. Below is a picture of what is desired. The flame will have an outer part and an inner part. This, however, is not always possible with some of the burners in the lab. The lower inner part is the coldest part of the flame. The outer part is hotter with the top of the flame the hottest area.



9. When through with the burner, turn off the bench gas valve; close the burner gas valve and the burner air control.

CAUTION

- Sometimes the whole flame is not readily visible. Do not put any part of your body too near it (especially over the top; the top of the flame can be higher than you realize).
- If you have long hair, keep it away from the flame by pulling it back with a ponytail holder.
- Push back any long sleeves.
- Put the burner near the trough side of the bench, away from the edge nearest you. This will decrease your chances of accidentally reaching over the flame.
- Place other items well away from the burner to avoid catching them on fire.

Explanation of what is happening as you adjust the flame

When a hydrocarbon is burning, the presence of the color yellow in the flame is an indication of insufficient oxygen. Small particles of carbon, produced by incomplete combustion, get heated in the flame and start glowing yellow. Since the air control was closed (Step #3), there is insufficient air (O_2) and a yellow/blue flame is produced.

As you open the air control, the flame starts losing the yellow color and becomes more blue. At this time, less carbon is being made and more CO and CO_2 are being made. By increasing the oxygen supply (air), the goal is to get as much CO_2 as possible. This will produce the hottest flame.

Items Needed for Each Section of Lab (Stockroom)

- 12 wood splints
- 12 strips of Mg (1.5 in. each)

Items Needed for Each Pair of Students (Already on shelf or in drawer)

- Laboratory burner
- Striker, match, or lighter
- Test tube
- Test tube holder
- Tongs

Procedure

1. Get a test tube and a test tube holder.

2. Do steps 1-7 in the instructions for using a burner.

3. With the air vents closed, the flame should be yellow and blue.

4. Do not point the opening of the test tube at you, your partner, or the people on the other side of the bench. The test tube is empty, but this is a good time to learn the correct way for heating a test tube.

   Hold the test tube with a test tube holder (not a pair of tongs) by placing the test tube holder at the very top of the test tube. Place the bottom part of the test tube in the TOP of the flame (this is the hottest part). HOLD the test tube at a 45-degree angle, not straight up and down.

   You should observe a deposit of carbon on the tube. Holding a piece of white paper behind the test tube AFTER you have heated it will enable you to easily see the carbon. Place the HOT test tube on the back of the bench to avoid accidently touching the HOT test tube. Record your observations on the Report Sheet.

5. Do step 8 in the instructions for using a burner. With more oxygen available for the combustion of methane, there is less C produced and more carbon dioxide produced. Try to get a blue flame with an inner cone and an outer cone.

6. (a) Get a wooden splint. Hold it with its edge resting on the top of the burner. Look at the picture below to see the placement of the splint.

(b) Remove it and look at it. Notice how the splint is burned only on the edges of the flame. Record your observations by drawing a picture on the Report Sheet.

(c) After drawing the picture, place the wood splint back on the burner as you did in step (a). Now, start bringing it up toward the top of the flame. Does the splint catch on fire?

7. Get a strip of Mg and a pair of tongs. Hold the strip of Mg with the pair of tongs (not a test tube holder). Place the Mg into the top part of the flame. When it starts burning, LOOK AWAY.

8. Look at the product of the combustion of Mg. It is probably on the bench because it fell down as it was made. Record your observations on the Report Sheet.

9. DO NOT reach over the burner when turning it off. Turn off the burner by turning off the bench gas control, then close the gas valve at the base of the burner and the burner's air control.

10. When the test tube is cool, clean it off with soap and water. All of the carbon will probably not come off. Use a little alcohol (ethyl or isopropyl) poured onto a paper towel to remove the rest of the carbon. Throw away the splint. Wipe off your bench and be sure to clean off all of the magnesium oxide you made.

Pre-Laboratory Questions
Combustion

1. (a) What is the definition of combustion?

   (b) If something is undergoing combustion, you say that it is _______.

2. (a) What is the formula of the oxygen gas used during combustion?

   (b) What is usually the source of the oxygen gas?

3. What are the two things needed for combustion to occur?

4. What is the definition of a fuel?

5. Give an example of a fuel that does not undergo combustion.

6. (a) _________________________ are the most common source of fuels.

   (b) What are the elements in a hydrocarbon?

   (c) What is the predominant source of hydrocarbons?

   (d) Name a solid fossil fuel, a liquid fossil fuel, and a gaseous fossil fuel.

7. (a) The natural gas used in our laboratory is mostly ___________.

   (b) What is the formula of this compound?

8. (a) If you open a valve on a natural gas stove, let the room fill up with gas and there is no source of heat will the gas catch on fire?

   (b) Now, flip a light switch. Could this provide enough energy to make the gas catch on fire? (Sometimes, if a room is dark enough, you can see a spark made in the light switch when you flip it to an on position.)

9. Say you are burning a piece of paper. You use a lighter to start the paper burning. Once the paper ignites, you remove the lighter and the paper keeps burning. Why?

10. A substance will keep burning until one of three things happen. List them.

11. (a) When complete combustion of a hydrocarbon occurs, the chemical products of this reaction are ______ and __________.

(b) In addition to these chemical products, ____________ is released.

(c) List 3 types of energy released that you can easily detect.

(d) Write the equation for the complete combustion of methane.

12. (a) What are the names and formulas of two carbon-containing gases produced when a hydrocarbon fuel is burned?

(b) What is the formula and name of the solid that can result?

(c) Why would industry want to make CO?

(d) In an industrial setting, the C made is called carbon _________ and could be used to make ______________________.

(e) In an industrial setting, to get more CO or more C, less ____________ is used.

13. List 2 reasons that explain why incomplete combustion is bad for a car.

14. If a fuel is burned in an area that does not have a source of fresh oxygen, which gas begins to have an increase in concentration?

15. How does carbon monoxide kill you?

16. How does carbon dioxide kill you?

17. Moving an unconscious person into the open air will be more effective for exposure to excessive levels of (CO, CO_2).

18. Combustion of magnesium can also be classified as a __________ reaction.

19. Define combination reaction.

20. Write the reaction for the combustion of magnesium.

21. Where is the air control on the burner? (top or bottom)

22. Where is the gas control on the burner? (top or bottom)

23. If the flame is yellow and blue, you need more (air, gas).

24. The yellow color in the flame was made by particles of __________ getting hot and glowing.

25. Where is the coldest part of the burner’s flame? Hottest?

26. For magnesium, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

27. For magnesium oxide, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

28. For methane, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

Report Sheet: Combustion

Name ___________________________________

Lab Partner’s Name _________________________________

1. (a) Describe the test tube heated in the oxygen-deprived flame.

   (b) What is the identity of the substance deposited on the test tube?

   (c) What did you do to decrease the amount of this substance in the flame?

   (d) Write the chemical equation for the primary reaction occurring when the flame was very yellow.

   (e) Write the chemical equation for the primary reaction occurring when the flame was blue.

2. (a) Draw a picture of the wood splint after placing it in the lower part of the flame.

   (b) Explain why the splint did not scorch in the middle.

   (c) Why didn’t the splint catch on fire in the lower part of the flame?

   (d) Why did the splint catch on fire after you moved it to the top part of the flame?

3. (a) Write the name and formula of the product formed when Mg burns.

   (b) Describe the product.

   (c) Write the equation for the combustion of magnesium.

Oxygen Gas Made by a Decomposition Reaction Catalyzed by Catalase

The oxygen gas being made in today's class is O_2. Most people say oxygen gas or just oxygen. There is another oxygen gas, but its formula is O_3. It is usually called ozone. The air we breathe contains about 21% O_2 and a very small percent of ozone.

A simple definition for a catalyst is an agent that increases the speed of a reaction. There are some reactions that would really not occur at all if the catalyst is not present. So, the catalyst definitely speeds those reactions up. In living organisms, most reactions are catalyzed by enzymes. Enzymes are proteins.

Unfortunately, some of the reactions in living organisms produce compounds which are harmful. In order to lessen the danger of these harmful compounds, processes have developed to either get rid of them or convert them to something that the organism can use. One harmful compound produced is hydrogen peroxide, H_2O_2. Catalase is an enzyme found in almost all living organisms. It catalyzes the decomposition of hydrogen peroxide to water and oxygen gas, O_2. By doing so, the catalase protects the cells from being damaged by the hydrogen peroxide. The water and oxygen gas can be used by the organism. Catalase is one of the fastest enzymes. One molecule of catalase can decompose millions of molecules of hydrogen peroxide each second.

The reaction we are using to make the oxygen gas is a decomposition reaction. A decomposition reaction starts off with one reactant and produces two or more products. Below is the reaction.

$$2\ H_2O_{2(aq)} \xrightarrow{\text{catalase}} 2\ H_2O_{(l)} + O_{2(g)}$$

The reactant is hydrogen peroxide. The products are water and oxygen gas. The name of the catalyst is written above the arrow. NONE of the elements in the catalyst are in the products. The oxygen in the oxygen gas come from the reactant. The (aq), (l) and (g) indicate the state of a substance for the conditions observations are being made. For today, the conditions are normal room temperature and normal atmospheric pressure. (aq) stands for aqueous; (l) stands for liquid: (g) stands for gas. Aqueous means that the chemical is dissolved in water. The big 2 on the left and the right are not part of the formulas of the hydrogen peroxide and water. They are needed to balance the equation.

The oxygen gas will be collected by displacing water in a bottle. This is possible because oxygen gas is not very soluble in water. After collecting the oxygen gas, a combustion reaction will be done. A combustion reaction is a fast, observable reactions with O_2. A common word for combustion is burning. The combustion reaction will be done to show how an oxygen-rich environment affects combustion. Combustion needs O_2. Things burn easier if there is more O_2.

Items Needed for Each Section of Lab (Stockroom)

- 12 rubber stoppers that have been fitted with a short glass tube and a long section of rubber tubing. The rubber stopper should fit snugly in the top of a 125-mL Erlenmeyer flask.
- 12 wood splints
- 500 mL 3% H_2O_2
- 15 g yeast (Active Dry seems to work better than Rapid Rise)
- Matches/Lighter

Items Needed for Each Pair of Students (Already on shelf or in drawer)

- 125-mL Erlenmeyer flask
- 8 ounce bottle
- Plastic pan
- 100-mL graduated cylinder
- Watch glass
- Scoopula

Procedure

1. Get the following: a 125-mL Erlenmeyer flask, a rubber stopper with a glass tube and rubber tubing attached, a bottle, a plastic pan, a wooden splint, a 100-mL graduated cylinder, and a watch glass.

2. (a) Place the weighing paper on a balance; zero the balance. Weigh out 1 gram of yeast. Yeast is a fungus and will be the catalase source. It was bought at a grocery store.

   (b) Measure out 30 mL of 3% hydrogen peroxide with the graduated cylinder. It the bottle looks familiar, it was also bought at a grocery store.

3. Check to make sure that the glass tube through the rubber topper fits tightly, that the rubber stopper fits tightly in the flask, and that the rubber tube for deliver to the plastic pan fits tightly on the glass tube through the rubber stopper.

4. Fill a plastic pan with water about half full.

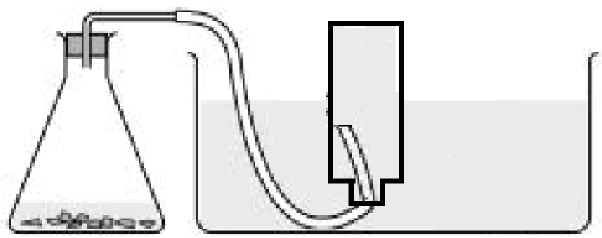
5. Fill the bottle COMPLETELY with water. Place a watch glass on the top of the bottle. Invert the bottle. Position the inverted bottle into the plastic pan. Remove the watch glass and stand the bottle, upside down, in the plastic pan.

6. Tilt the bottle a little and place the end of the delivery hose into the bottle. DO NOT PUT PRESSURE ON THE DELIVERY HOSE WITH THE BOTTLE WHILE COLLECTING THE OXYGEN GAS. The gas needs to get into the bottle. If you close the delivery hose by pressing down on it, the gas will not get into the bottle. Make sure the delivery hose has no bends in it which would prevent the gas from going to the bottle.

7. Remove the stopper from the flask. Add the hydrogen peroxide to the flask.

8. Now, as quickly as possible, pour the yeast into the flask and replace the rubber stopper, making sure it fits snugly.

9. Swirl the flask while the reaction proceeds.



10. After the oxygen gas has displaced all of the water in the bottle, remove the hose from the pan (not just the bottle). The hose should be 100% out of the water.

11. The inverted bottle is still in the pan of water. Place the watch glass over the opening of the upside down bottle. Hold the watch glass in place and take the bottle out of the water. Turn it right side up, KEEPING THE WATCH GLASS ON THE BOTTLE. Set the covered bottle on the lab bench. There may be a foggy appearance inside the bottle. This is from water vapor, not oxygen gas.

12. GET READY, BOTH OF YOU. One of you is going to be moving the watch glass. The other is going to be placing a glowing splint into the bottle and taking it out. Read all of steps 13 and 14 before proceeding.

13. Glowing splint person: Catch the splint on fire with a lighter. It should be on fire. Now, blow on it to put out the flame. It needs to be glowing. Not on fire and not totally out. GLOWING. Do not stand there and look at the glowing splint. It will stop glowing. Also, hold the splint at the far end so as to be able to put the splint further into the bottle. As the oxygen gas at the top of the bottle is depleted, you will have to place the splint further down to reach unreacted oxygen gas.

14. Work together. Person holding the watch glass, move it just enough so the glowing splint person can place the GLOWING splint into the bottle. The GLOWING splint person: Do not drop the splint; hold onto it. After the glowing splint re-ignites, move it out of the bottle QUICKLY. Watch glass person: move the glass back over the bottle. Glowing (now burning) splint person, QUICKLY blow out the flame. Then, both of you QUICKLY repeat this process. Count how many times the splint re-ignites.

15. Throw the contents of the Erlenmeyer flask into a trash can. DO NOT throw them into the sink. Wash out the Erlenmeyer with soap and water. Also, wash out the bottle and the hose.

Pre-Laboratory Questions

Oxygen Gas Made by a Decomposition Reaction Catalyzed by Catalase

1. (a) What are the formulas and names of two oxygen gases?

   (b) Which one will you be making in the laboratory?

2. What are the definitions of

   (a) catalyst

   (b) decomposition reaction

   (c) aqueous

   (d) combustion

3. (a) Write the equation for the decomposition reaction of hydrogen peroxide.

   (b) What is the reactant in this reaction? Give the name and formula.

   (c) What are the products? Give the names and formulas.

   (d) Give the name of the catalyst.

   (e) What is the source of the catalyst in this experiment?

   (f) Does any part of the catalyst show up in a product?

   (g) Why are the (aq), (l) and (g) used in the equation?

   (h) Are the big numbers (coefficients) in the equation part of the formulas of any reactant or the product?

   (i) Why are the coefficients there?

4. Why are you able to collect the oxygen gas by displacing water?

5. Why are you also going to do a combustion reaction?

6. Things burn easier if there is (less, more) O_2. Circle choice.

7. Fill in the blanks.

In living organisms, most reactions are catalyzed by _______________.

Enzymes are __________________.

In order to lessen the danger of harmful compounds in the cells of living organisms, processes have developed to either get ___________ of them or __________________ them to something that the organism can __________.

One molecule of catalase can decompose _______________ of molecules of hydrogen peroxide each _______________.

8. Hydrogen peroxide (is, is not) harmful to living cells.

9. For hydrogen peroxide, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

Report Sheet: Oxygen Gas Made by a Decomposition Reaction
Catalyzed by Catalase

Name ___________________________________

Lab Partner’s Name _________________________________

1. How many times did your stick re-ignite after placing it in the bottle?

2. Write the equation for the decomposition reaction of hydrogen peroxide.

3. Name the reactants.

4. Name the products.

5. Name the catalyst.

6. Why are you able to collect the oxygen gas by displacing water?

7. Why did you do the combustion reaction?

Two Chemical Reactions: Single Replacement and Combustion

Single Replacement

When an element is in an elemental form, it is not combined with another element. It is a single atom or a group of atoms. Some examples of formulas of elemental forms are Na, C, O_2, O_3, H_2 and S_8. A compound is made from two or more elements. The formula of a compound lets you know which elements were used.

In a single replacement reaction, an element will replace another one that is in a compound. The one that is replaced will then be in elemental form. We can use letters to illustrate this type of reaction. If the element that is the reactant forms a positive ion, it will replace the first member of the reactant that is the compound.

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

$$Zn_{(s)} + 2\ HCl_{(aq)} \longrightarrow H_{2(g)} + ZnCl_{2(aq)}$$

If the element that is the reactant forms a negative ion, it will replace the second member of the reactant that is the compound.

$$D + EF \rightarrow F + ED$$

$$Cl_{2(g)} + 2\ NaI_{(aq)} \longrightarrow 2\ NaCl_{(aq)} + I_{2(s)}$$

Combustion

Combustion is a fast, observable reaction with O_2. Hydrogen gas is very flammable and can be used as a fuel. When hydrogen gas burns, there is a lot of heat energy released. A reaction that releases heat energy is an exothermic reaction. The water that is formed will be in the gaseous state due to the heat released.

$$2\ H_{2(g)} + O_{2(g)} \longrightarrow 2\ H_2O_{(g)} + \text{Energy}$$

Normally, when a fuel is burned, it is done so at a controlled rate. When burning natural gas in the laboratory, the laboratory burner is used to allow the natural gas to be released at a rate that produces a usable flame. However, explosions can occur when a fuel ignites in an uncontrolled manner. An explosion is the sudden, violent release of energy. The energy released causes the surrounding air to expand rapidly. This results in a wave of compressed air. This wave of compressed air reaches our eardrums which moves them and we register it as a sound of something exploding.

Combining These Two Reactions

In today’s laboratory, you are going to make hydrogen gas with a single replacement reaction and then you are going to blow it up.

Chemicals Needed for 24 Students (12 Working Pairs)

- 6 M HCl: 75 mL
- Small amount of Mg: 12 strips, each about 0.5 in.
- Large amount of Mg: 12 strips, each about 1.5-2.0 in.

Other Items Needed for Each Pair of Students

- Two test tubes, medium size, 5-6 inches long
- Test tube rack
- Test tube holder
- Wood splint
- Match or lighter

Procedure A-Single Replacement

1. (a) Go to the hood with one of the test tubes. Add 2 droppers (not drops) of 6 M $HCl_{(aq)}$ to it. Go back to your work space before doing step (b).

   (b) While your partner is watching, place the small piece (not the large piece) of magnesium into the $HCl_{(aq)}$ in the test tube.

   (c) Feel the test tube. Did it get warm?

   (d) Record any other sensoria (at least 2). Note: If you write smoke as an sensorium, this means that there is a fire. If you do not see something burning, do not write smoke.

2. (a) Below are the formulas of the reactants of the reaction between magnesium and hydrogen chloride. Read these tips and then write the formulas of the products and balance the equation.

   - Metals form positive ions.
   - The metal will displace the first member of the reacting compound.
   - When writing formulas of ionic compounds, the positive ion is written first.
   - The charges of the ions determines the subscripts in the ionic compound's formula.
   - The elemental form of hydrogen in nature is a diatomic molecule.

   $$Mg_{(s)} + HCl_{(aq)} \rightarrow$$

   (b) One product was the ionic compound magnesium chloride, a solid. Look up its water solubility on the solubility chart on the page after the instructions for Procedure B.

   As it forms in water, will it be dissolved?

   (c) Look at the test tube. Do you see any solid magnesium chloride?

(d) Does your answer for (b) explain your answer for (c)?

(e) You observed bubbles (fizzing) when the Mg was placed in the aqueous HCl. Bubbles indicate the formation of a gas. Was this gas dissolving in the water or was it getting out of the water?

(f) What is the name of the gas that formed? Read the three bullets for the note below to help answer this question.

Note: •Look back at the information in (b).
•Look at the equation in (a) to see what the products are.
•Look at your answer in (e).

(g) By the formulas of the products in (a), write (aq), (s), (l), or (g).

(h) Is this reaction endothermic or exothermic?

(i) Add the additional information, + heat, to the reaction in (a).

Procedure B-Single Replacement and Combustion

Explanation: In the first test tube, you did only one reaction: Mg + HCl making hydrogen gas and magnesium chloride. You recorded your sensoria for this test tube. Two sensoria were that heat was released and that there were bubbles (fizzing, however you described this). For the second test tube, you will do two reactions. The first will be exactly the same as that done in the first test tube: Mg + HCl. But, you will put a bigger piece of Mg in the test tube to make the reaction last longer. This extra time will enable you to light a wooden splint and put the burning splint into the test tube while the Mg is still reacting with the HCl. At the instant you place the splint into the test tube, the second reaction will occur. It is the combustion (burning) of the hydrogen gas being made by the first reaction. A NEW sensorium will be produced by this second reaction, one not made for the first test tube.

1. (a) Go to the hood with the second test tube. Add 2 droppers (not drops) of 6 M $HCl_{(aq)}$ to the it. Go back to your work space before doing step (b).

(b) Hold the test tube with a test tube holder (not a pair of tongs).

(c) Light a wooden splint with a lighter or match. Do this at least a foot away from the person holding the test tube. Student holding the splint: Do not insert the burning splint until step (d) has been done. When inserting the burning splint into the test tube, hold onto it. Do not drop it into the test tube. Do not let it get into the water at the bottom of the test tube.

(d) Place the piece of magnesium into the $HCl_{(aq)}$ in the test tube.

(e) Insert the burning wooden splint into the test tube.

(f) Record your NEW sensorium that did not happen in the first test tube.

Solubility Rules

NO_3^-	All nitrates are soluble.
Cl^-	All chlorides are soluble except those of silver, mercury (I), and lead (II)
SO_4^{2-}	Most sulfates are soluble except those of barium, lead (II) and strontium.
CO_3^{2-}	All carbonates are insoluble except those of ammonium and those of the Group 1 elements.
OH^-	All hydroxides are insoluble except those of the Group 1 elements, barium, and strontium.
S^{2-}	All sulfides are insoluble except those of the Group 1 and Group 2 elements and ammonium.

Pre-Laboratory Questions
Two Chemical Reactions
Single Replacement
Combustion

1. Define the following:

   (a) elemental form

   (b) single replacement reaction

   (c) combustion reaction

   (d) exothermic

2. Why do you hear a loud sound when something explodes?

3. Finish and balance these single replacement reactions. The compounds formed are ionic. Use the charges on the ions to determine the correct formulas for these ionic compounds. The elemental forms are single atoms, unless otherwise stated. Don't forget to balance the equations after writing the formulas of the products.

   (a) *The elemental form of hydrogen is a diatomic molecule.

   $Al + HCl \rightarrow$

   (b) $Ca + NaBr \rightarrow$

   (c) $Al + KNO_3 \rightarrow$

   (d) $Ca + Al(NO_3)_3 \rightarrow$

   (e) *The elemental form of iodine is a diatomic molecule.

   $Br_2 + KI \rightarrow$

4. After each of the following write soluble or insoluble to indicate if each would or would not dissolve in water. The solubility table is on the page after the instructions for Procedure B.

   a. $Ba(NO_3)_2$
   b. ZnS
   c. $(NH_4)_2CO_3$
   d. $PbCl_2$
   e. Ag_2SO_4
   f. $Ca(OH)_2$
   g. NH_4NO_3
   h. $NiSO_4$
   i. $AlCl_3$
   j. CuS

5. For magnesium, what would be present for the following elements on a GHS label?

    Pictogram:

    Signal Word:

    Hazard Statement:

    Precautionary Statement:

6. For hydrochloric acid, what would be present for the following elements on a GHS label?

    Pictogram:

    Signal Word:

    Hazard Statement:

    Precautionary Statement:

7. For magnesium chloride, what would be present for the following elements on a GHS label?

    Pictogram:

    Signal Word:

    Hazard Statement:

    Precautionary Statement:

8. For hydrogen gas, what would be present for the following elements on a GHS label?

    Pictogram:

    Signal Word:

    Hazard Statement:

    Precautionary Statement:

Report Sheet: Two Chemical Reactions
Single Replacement
Combustion

Name ___________________________________

Lab Partner's Name _________________________________

1. Write the balanced equation for the following reactants. Include (aq), (s) and (g).

   (a) hydrochloric acid and magnesium

   (b) hydrogen gas and O_2

2. What were your sensoria for the reactions in Question #1

   (a) hydrochloric acid and magnesium (3 sensoria)

   (b) hydrogen gas and O_2 (the NEW one)

3. (a) After the reaction between Mg and HCl, which product stayed in the water?

   (b) After the reaction between Mg and HCl, which product left the water?

4. For the reaction between hydrogen gas and O_2, where did the O_2 come from?

Hydrates

Some compounds when heated do not melt but undergo decomposition. In decomposing the compound can break down into two or more substances. If a reaction is reversible, the products can react to form the original reactants. Hydrates are examples of compounds which do not melt but which decompose upon heating. The decomposition products are an anhydrous salt and water. Anhydrous means without water. A salt is an ionic compound. The original hydrates can be regenerated by addition of water to the anhydrous salt.

The hydrate in today's experiment is copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$. As indicated by the formula, 5 waters of hydration are bound to the copper(II) ion in copper(II) sulfate. Notice how the formula is written, the waters of hydration are separated from the formula of the salt by a dot.

Heat can transform a hydrate into an anhydrous salt by making the water go to the gaseous state. The water can sometimes be seen escaping as steam. The blue crystals of copper(II) sulfate pentahydrate will lose the blue color as the water is driven off, leaving behind the white anhydrous salt, copper(II) sulfate.

(Eq. 1) $CuSO_4 \cdot 5H_2O_{(s)} + \text{heat} \rightarrow CuSO_{4(s)} + 5\ H_2O_{(g)}$

Whenever a reaction requires heat, the reaction is referred to as an endothermic reaction. The decomposition of a hydrate is an endothermic reaction.

This reaction is reversible; adding water to the white anhydrous copper(II) sulfate salt will rehydrate the salt and regenerate the blue pentahydrate.

(Eq. 2) $CuSO_{4(s)} + 5\ H_2O_{(l)} \rightarrow CuSO_4 \cdot 5H_2O_{(s)} + \text{heat}$

In the above equation for the regeneration of the hydrate, notice that (1) liquid water is used (not gaseous) and (2) heat is produced, not used. If we left the sample of $CuSO_4$ out, it would slowly turn blue as water vapor in the air came into contact with the anhydrous salt. We are going to speed up this regeneration of the hydrate by adding liquid water. A reaction that produces heat is an exothermic reaction.

One equation can be written to indicate the reversible nature of the reactions. In the equation below, the arrow pointing to the right is for the reaction that occurs when heating the hydrate. The arrow pointing to the left is for when the water is added to the anhydrous salt. Notice, there is no indication for the state of the water. This is because it changed in today's experiment. It was in the gaseous state when heating the hydrate and was in the liquid state when it was added to the anhydrous salt.

(Eq. 3) $CuSO_4 \cdot 5H_2O_{(s)} + \text{heat} \leftrightarrows CuSO_{4(s)} + 5\ H_2O$

One molar mass of a chemical is the mass in grams of one mole of a chemical. To calculate this mass, add the atomic weights of all atoms in a formula and add the unit grams. The molar mass of copper(II) sulfate pentahydrate is 249.6 grams. Below is the math done to get the sum of 249.6.

The atomic weights are obtained from a Periodic Table and rounded to the nearest tenth.

$CuSO_4 \cdot 5H_2O$

Cu	63.5 X 1 = 63.5	H	1.0 X 2 = 2.0
S	32.1 X 1 = 32.1	O	16.0 X 1 = 16.0
O	16.0 X 4 = 64.0	Sum for one H_2O	18.0
Sum for $CuSO_4$	159.6		

There are 5 H_2O in the formula; the total for water is 5 X 18.0 = 90.0

One mole of $CuSO_4 \cdot 5H_2O$ = 249.6 g

Since 90.0 grams of this is from the water, the mass of water is 36.0% of one molar mass. Below is the math done to get 36.0%.

$$\frac{90.0\text{ g}}{249.6\text{ g}} \times 100 = 36.0\%$$

36.0% of any mass of copper(II) sulfate pentahydrate is from the water. If the mass of the copper(II) sulfate pentahydrate is 85.0 g, then 30.6 g is from the water. Below is the math done to get 30.6 g.

Remember: 36.0% means 36 out of 100. When calculating the percentage, you can use either 36.0/100 or 0.360 in the set up.

$$36.0\% = \frac{36.0}{100} = 0.360$$

Below are two set ups; one using the fraction and the other using the decimal number. Do not do both. They are the same. Choose the set up you like more.

$$85.0\text{ g} \times \frac{36.0}{100} = 30.6\text{ g} \qquad\qquad 85.0\text{ g} \times 0.360 = 30.6\text{ g}$$

Items Needed for Each Section of Lab (Stockroom)

- $CuSO_4 \cdot 5H_2O$: 90 grams

Items Needed for Each Pair of Students (Already on shelf or in drawer)

- Evaporating dish
- Wire screen
- Tripod
- Laboratory burner
- Striker, lighter, or match
- Stirring rod
- Tongs
- Wash bottle
- Thermometer

Procedure

(a) Mass of dish and hydrate	g
(b) Mass of empty dish	g
(c) Mass of hydrate [(a) – (b)]	g
(d) Mass of dish and anhydrous salt	g
(e) Mass of anhydrous salt [(d) – (b)]	g
(f) Mass of water in hydrate [(c) – (e)]	g

1. Zero the balance. Weigh the dish. Enter mass in table above. DO NOT ROUND.

2. Keep the dish on the balance. Add at least 5 grams of copper(II) sulfate pentahydrate to the dish. Record the mass of the dish and hydrate in the table. DO NOT ROUND.

3. Fill in the table for (c). DO NOT ROUND. If the number in (c) is not close to 5, then you did something wrong. Do step 1 and 2 and 3 again.

4. What is the color of the hydrate?

5. Place the dish on a wire screen on a tripod that is placed back toward the trough.

6. Light the laboratory burner and place it under the dish. The flame should be blue.

7. Heat with a medium-sized flame for 15 minutes; hold the dish with a pair of tongs and stir gently with a stirring rod. Do not place your face over the dish to look at the contents. Do not tilt the dish toward your face to look at the contents.

8. ***<u>Do not touch the tripod, wire screen, or dish with your fingers or any part of your body. All are hot.</u>***

Turn off the burner. With tongs, place the dish (don't drop it) on the bench top.

9. What is the color of the anhydrous salt?

10. When the dish is cool enough for you to handle without the use of the tongs, weigh and record the mass of the anhydrous salt and dish. Record this mass in the table. DO NOT ROUND.

11. Finish filling in the table. DO NOT ROUND.

<u>BEFORE CONTINUING, HAVE YOUR INSTRUCTOR CHECK ALL OF YOUR ENTRIES IN THE TABLE. IF THERE IS AN ERROR IT CAN POSSIBLY BE CORRECTED. HOWEVER, ONCE YOU DO STEP 15 CORRECTIONS ARE NOT POSSIBLE.</u>

12. Calculate the mass of water that should have been lost from heating the hydrate. This is the theoretical mass. Multiply 36.0% of the amount of hydrate that you heated.

13. Calculate your percent error. The || in the numerator means absolute. When subtracting the actual from the theoretical, if the difference is a negative number, switch it to positive. The %error will be a positive number with this set up.

$$\frac{|\text{Theoretical - Actual}|}{\text{Theoretical}} \text{ X } 100 = \ \% \text{ Error}$$

14. Get a thermometer. Do not stick it in anything, especially the dish. Do not hold the bulb.

Record the temperature of the thermometer __________ °C

15. Do not place your face over the dish when you do this next step. With a wash bottle, add just a small amount of water, **slowly**, to the anhydrous salt in the evaporating dish. Did you see steam?

16. Continue adding water ***and*** stir gently with the thermometer. Keep adding until you get enough water to cover the bulb of the thermometer. Monitor the change in the temperature until there is no more increase. What is the new temperature? _________° C

17. What is the color of the solid in the evaporating dish after adding the water?

18. Is the solid under the water the hydrate or the anhydrous salt?

19. Write the equation for what was occurring in step 7.

20. What are 2 pieces of evidence that this reaction occurred? One is an observation and one is a measurement.

    Observation:

    Measurement:

21. Write the equation for what was occurring in step 15.

22. What are 3 pieces of evidence that this reaction occurred? Two are observations and one is a measurement.

    2 Observations:

    Measurement:

Pre-Laboratory Questions
Hydrates

1. What is the formula of the hydrate used in today's experiment?

2. What is the hydrate's name?

3. What is the formula of the anhydrous salt formed in today's experiment?

4. What is the name of the anhydrous salt?

5. What does the word anhydrous mean?

6. What is the definition of a salt?

7. What are the definitions of endothermic and exothermic?

8. (a) What is the definition of a decomposition reaction?

   (b) Give an example of a decomposition reaction.

9. (a) What is the definition of a reversible reaction?

   (b) Is heating a hydrate a reversible reaction?

10. Write equations for the following reactions.

    (a) Decomposition of copper (II) sulfate pentahydrate by heating

    (b) Recombination of the products that formed in reaction (a)

11. Write one equation that can represent both reactions done in today's experiment.

13. (a) What is the molar mass of $BaCl_2 \bullet 2H_2O$?

    (b) What percentage of this molar mass is due to the water?

    (c) If you had 200.0 grams of $BaCl_2 \bullet 2H_2O$, how many grams would be due to the water?

14. For copper (II) sulfate pentahydrate, what would be present for the following elements on a GHS label?

Pictogram:

Signal Word:

Hazard Statement:

Precautionary Statement:

15. For copper (II) sulfate, what would be present for the following elements on a GHS label?

Pictogram:

Signal Word:

Hazard Statement:

Precautionary Statement:

Report Sheet: Hydrates

Name ___________________________________

Lab Partner's Name _________________________________

1. Fill in the table.

(a) Mass of dish and hydrate	g
(b) Mass of empty dish	g
(c) Mass of hydrate [(a) – (b)]	g
(d) Mass of dish and anhydrous salt	g
(e) Mass of anhydrous salt [(d) – (b)]	g
(f) Mass of water in hydrate [(c) – (e)]	g

2. Show calculations.

   (a) Molar mass of copper (II) sulfate pentahydrate

   (b) Percent of water in any mass of copper (II) sulfate pentahydrate

   (c) Theoretical amount of water in the mass of hydrate you heated

   (d) Percent error of experimentally measured amount of water lost

3. What are 2 pieces of evidence that a reaction occurred when heating the hydrate?

   Observation:

   Measurement:

4. What are 3 pieces of evidence that a reaction occurred when adding water to the anhydrous salt?

   2 Observations:

   Measurement:

5. Steam was produced when adding water to the anhydrous salt because this reaction is (endothermic, exothermic).

6. (a) When you first added water in step 15, the solid started turning blue.

Which of the following represents the blue solid that formed?

$CuSO_4 \cdot 5H_2O_{(s)}$ $CuSO_4 \cdot 5H_2O_{(l)}$ $CuSO_4 \cdot 5H_2O_{(aq)}$

(b) In step 16, you added more water and the water started turning blue because the blue solid was ___________________ in the water.

(c) Based on your answer in (b), which of the following represents the hydrate that you could no longer see as a solid?

$CuSO_4 \cdot 5H_2O_{(s)}$ $CuSO_4 \cdot 5H_2O_{(l)}$ $CuSO_4 \cdot 5H_2O_{(aq)}$

Stoichiometry of the Reaction of Sodium Bicarbonate with Hydrochloric Acid

Stoichiometry involves calculations to determine the quantities of reactants and products involved in chemical reactions.

The first step for these calculations is to write a balanced equation. In today's experiment, you will be reacting sodium bicarbonate with hydrochloric acid. Sodium bicarbonate is called baking soda at the grocery store. The products are sodium chloride, water, and carbon dioxide. Sodium chloride is called table salt at the grocery store.

$$NaHCO_{3(s)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

In the above equation, the coefficient for all reactants and products is one. This means that 1 mole of sodium bicarbonate will react with 1 mole of hydrochloric acid and form one mole of sodium chloride, one mole of water, and one mole of carbon dioxide.

Since 1 mole of anything is equal to its formula weight in grams, we can say that 84.01 g of sodium bicarbonate will react with 36.46 g of hydrochloric acid to form 58.44 g of sodium chloride, 18.02 g of water, and 44.01 g of carbon dioxide.

By writing all of this information under the equation, it is easy to keep track with the numbers.

$NaHCO_{3(s)}$	+ $HCl_{(aq)}$	$\longrightarrow$	$NaCl_{(aq)}$	+ $H_2O_{(l)}$	+ $CO_{2\,(g)}$
1 mole	1 mole		1 mole	1 mole	1 mole
84.01 g	36.46 g		58.44 g	18.02 g	44.01 g

Of course the starting quantity of $NaHCO_3$ may be more or less than 84.01 g, but a proportionate quantity of the hydrochloric acid will be used, and proportionate quantities of the products will be formed.

For example, let us start off with 100.00 g of $NaHCO_3$. By using an excess amount of hydrochloric acid, some of the acid will not react. But, we know that we have added more than enough. By doing this, the sodium bicarbonate will the reactant that we will use to calculate the amount of products that will form. It is referred to as the limiting reagent. Bubbles of carbon dioxide will form as long as a reaction is occurring. We will know we have added more than enough acid because the bubbling will stop.

Here is a typical stoichiometry question: 100.0 g of $NaHCO_3$ will produce how many grams of sodium chloride?

Use the information above to make a conversion factor (which is really a ratio or proportion).

$$100.0 \text{ g } NaHCO_3 \times \frac{58.44 \text{ g } NaCl}{84.01 \text{ g } NaHCO_3} = 69.56 \text{ g } NaCl$$

The 69.56 g of sodium chloride is called the theoretical yield. Theoretical yield is the maximum amount of product that can be produced from a given amount of reactant.

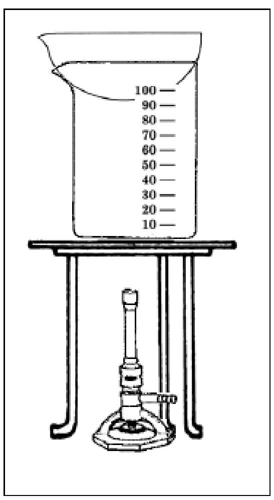
Getting the theoretical yield is not likely. Actual yield is the amount of product actually produced when the reaction is carried out. The actual yield is sometimes called the experimental yield because it is what is obtained by doing the experiment.

Percent yield is the percent of the theoretical yield that is obtained through laboratory procedures. Below is the set up for determining percent yield.

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

Using the previous calculation for a theoretical yield of 69.55 g, obtaining only 50.55 g would be a 72.67% yield.

In today's experiment, the only two products left in the reaction vessel will the water with the sodium chloride dissolved in it. The carbon dioxide bubbled out. By boiling off all of the water, the only thing left in the reaction vessel will be the sodium chloride. Any excess HCl will also go into the gaseous state as the water is boiled off.



Items Needed for Each Section of Lab (Stockroom)

- Sodium bicarbonate, 100 g
- Concentrated hydrochloric acid, 100 mL
- Beaker tongs

Items Needed for Each Pair of Students (Already on shelf or in drawer)

- Evaporating dish
- Scoopula
- 10-mL graduated cylinder
- Wash bottle
- Watch glass
- 250-mL beaker
- Wire screen
- Tripod
- Laboratory burner
- Lighter/striker
- Tongs

Procedure

1. Weigh an evaporating dish. Record its mass in blanks (b) and (e) in the data table. DO NOT ROUND THE MASS. Keep the dish on the balance.

2. With a scoopula, add about 5 g of $NaHCO_3$ to the dish. Record the mass in blank (a). DO NOT ROUND THE MASS. Go back to your work space.

3. Measure out 5 to 6 mL of water in a 10-mL graduated cylinder. Try to pour the water over the sodium bicarbonate evenly. Do not stir it; just pour the water on it. Cover the dish with a watch glass.

4. Go to the hood with the covered dish and the graduated cylinder. Measure out about 6 mL of concentrated hydrochloric acid. Stay at the hood. Move the watch glass aside slightly and add, in small portions, the acid. These small portions of acid of acid should be added so that the acid runs down the inside wall of the evaporating dish. Do you see the bubbles of carbon dioxide escaping?

5. Set up a water bath. Remove the watch glass and place the dish on top of the 250-mL beaker. Rinse off the watch glass since it probably has HCl on it. Light the laboratory burner and get the water to boil. Continue until all of the water has been boiled off.

6. After the water in the dish is gone, do the following. With a pair of tongs (NOT a test tube holder), remove the dish to the bench. With a pair of beaker tongs (NOT the tongs used on the dish) remove the beaker of HOT water. Place the HOT beaker on the bench toward the trough so as to prevent your accidently touching it. Now, with the tongs first used, place the dish on the wire gauze.

7. Heat the dish on the wire screen for 10 minutes. Turn off the burner. With the tongs, place the dish on the bench. Do not touch any of the hot items: tripod stand, wire screen, dish, beaker. AFTER the dish has cooled, weigh it. Record this mass in blank (d). DO NOT ROUND THE MASS.

8. Finish filling in the data table.

(a) Mass of dish and $NaHCO_3$	g
(b) Mass of empty dish	g
(c) Mass of $NaHCO_3$ [(a)–(b)]	g
(d) Mass of dish and residue	g
(e) Mass of empty dish (b)	g
(f) Mass of NaCl residue [(d)–(e)]	g

Calculations

1. (a) What are the two reactants used in today's work?

   (b) Which one do you know exactly how much was used?

   (c) Your answer in (b) will be used for this stoichiometry problem. Calculate the theoretical yield of sodium chloride. SHOW WORK.

2. Calculate the percent yield. SHOW WORK.

Extra Stoichiometry Problems. Work on these while the water is boiling. SHOW WORK. Answers for 1-3 are to have 4 significant figures.

1. If 35.00 g of NaCl is obtained, how many grams of CO_2 would be made?

2. If 60.00 g of sodium bicarbonate are used, how many grams of water will be made?

3. If 18.23 g of HCl are used, how many g of sodium bicarbonate are needed?

4. Calculate how many grams are in one mole of each of the following:

   (a) $Ba(OH)_2$ (b) H_2SO_4

5. A student calculated a theoretical yield of 2.80 g. He obtained 1.12 g. What was the percent yield? 3 significant figures

5. A student calculated a theoretical yield of 2.80 g. He obtained 1.12 g. What was the percent yield? 3 significant figures

Pre-Laboratory Questions

Stoichiometry of the Reaction of Sodium Bicarbonate with Hydrochloric Acid

Helpful information:

Based on the balanced equation for the experiment, the following are the two conversion factors that could be written involving sodium bicarbonate and water.

$$\frac{84.01 \text{ g } NaHCO_3}{18.02 \text{ g } H_2O} \quad \text{and} \quad \frac{18.02 \text{ g } H_2O}{84.01 \text{ g } NaHCO_3}$$

Example of question involving these two chemicals: If 100.0 g of sodium bicarbonate are used, how many grams of water will be made? Only the second conversion factor would be used in the set up. This would allow the canceling of grams of sodium bicarbonate and having only grams of water in the answer.

$$100.0 \text{ g } NaHCO_3 \times \frac{18.02 \text{ g } H_2O}{84.01 \text{ g } NaHCO_3} = 21.66 \text{ g } H_2O$$

1. How many grams are in one mole of each of the following? Show work.

   (a) $NaHCO_3$

   (b) HCl

   (c) NaCl

   (d) H_2O

   (e) CO_2

2. Write the two conversion factors that could be used for each of the following pairs of chemicals used in today’s experiment.

   (a) sodium bicarbonate and HCl

   (b) sodium bicarbonate and carbon dioxide

   (c) water and carbon dioxide

(d) water and sodium chloride

(e) sodium chloride and carbon dioxide

(f) HCl and sodium chloride

(g) HCl and water

(h) HCl and carbon dioxide

3. If 70.00 g of NaCl is obtained, how many grams of CO_2 would be made? (4 significant figures)

4. If 30.00 g of sodium bicarbonate are used, how many grams of water will be made? (4 significant figures)

5. A student calculated a theoretical yield of 2.80 g. He obtained 2.24 g. What was the percent yield? (3 significant figures)

6. For sodium bicarbonate, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

7. For hydrochloric acid, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

8. For sodium chloride, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

Report Sheet: Stoichiometry of the Reaction of Sodium Bicarbonate with Hydrochloric Acid

Name ___________________________________

Lab Partner's Name _________________________________

1. (a) What are the two reactants used in today's work?

   (b) Which one do you know exactly how much was used?

   (c) Your answer in (b) will be used for this stoichiometry problem. Calculate the theoretical yield of sodium chloride. SHOW WORK.

2. Calculate the percent yield. SHOW WORK.

Extra Stoichiometry Problems. Work on these while the water is boiling. SHOW WORK. Answers for 1-3 are to have 4 significant figures.

1. If 35.00 g of NaCl is obtained, how many grams of CO_2 would be made?

2. If 60.00 g of sodium bicarbonate are used, how many grams of water will be made?

3. If 18.23 g of HCl are used, how many g of sodium bicarbonate are needed?

4. Calculate how many grams are in one mole of each of the following:

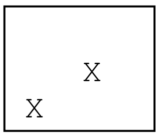
   (a) $Ba(OH)_2$ (b) H_2SO_4

Charles’ Law

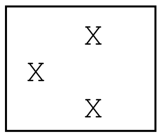
The pressure of a gas inside a container results from the force that is expressed when the gas molecules hit the sides of the container. The pressure can be changed by (1) changing the mass of gas, (2) changing the volume of the container (thus, changing the volume occupied by the gas), and (3) changing the temperature of the gas.

All containers below are sealed. No gas molecules can escape. No molecules from outside the container can enter. Each example is for altering one of the factors listed in the above paragraph. Each x represents a molecule of a gas.

(1) Changing the mass of gas.
The volumes of the cans are the same.
The temperatures are the same.

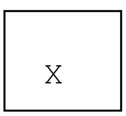


Less gas
Less hitting
Less pressure

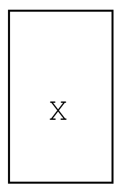


More gas
More hitting
More pressure

(2) Changing the volume of the container (also the gas)
The masses of gas are the same.
The temperatures are the same.

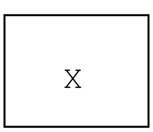


More hits/second
More pressure

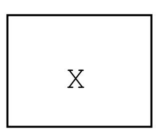


Less hits/second
Less pressure
(In the larger container, the gas particles have to move a longer distance, so fewer hits/second)

(3) Changing the temperature of the gas.
The masses of gas are the same
The volumes are the same.



Lower temperature
Molecules move slower
Fewer hits/second
Less pressure



Higher temperature
Molecules move faster
More hits/second
More pressure

Charles' law describes how the volumes of gases changes as the temperature changes. A statement of Charles' law is: The volume of a given mass of a gas is directly proportional to its temperature (in Kelvins) if the pressure and the mass of gas remain constant; that is, the volume of the gas increases or decreases by the same factor as its temperature. This law explains how a gas expands as the temperature increases; conversely, a decrease in temperature will lead to a decrease in volume.

The law can be written as:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad \text{or} \qquad V_1 T_2 = V_2 T_1$$

Explanation of Mathematical Equation

The Universal Gas Law is PV = nRT. This has been developed from years of experiments. The n stands for moles of a gas. It depends on how many grams of a gas is being used. R is a constant. It needs to be in the equation so that the product of the terms on the left equals the product of the terms on the right. R was determined from years of experiments.

If an experiment is done in which any or all of the factors are changed, the Universal Gas Law would be written $P_2V_2 = n_2RT_2$. The little twos represents a change in one or all factors. Notice R does not have a little 2; it is a constant. Constants do not change.

Rearrange the equations for the initial conditions and the final conditions.

$$\frac{P_1V_1}{n_1T_1} = R \qquad \frac{P_2V_2}{n_2T_2} = R$$

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Under the conditions for Charles' Law, P is the same on each side and n is the same. These terms cancel and the result is the same as the one at the top of this page.

$$\frac{\cancel{P_1}V_1}{\cancel{n_1}T_1} = \frac{\cancel{P_2}V_2}{\cancel{n_2}T_2}$$

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

Items Needed for Each Section of Lab (Stockroom)

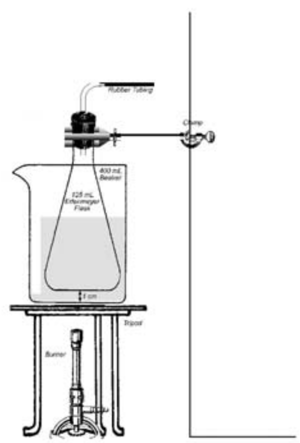
- Twelve one-hole rubber stoppers fitted with a short piece of glass tubing and a short piece of rubber tubing attached
- Twelve pinch clamps

Items Needed for Each Pair of Students (Already on shelf or in drawer)

- 125-mL Erlenmeyer flask
- 400-mL beaker
- Tripod
- Wire screen
- Laboratory burner
- Lighter/striker
- Ring stand
- Clamp
- 100-mL graduated cylinder
- Thermometer

Procedure

1. Use a 125-mL Erlenmeyer flask for this experiment. Fit the flask with a one-hole rubber stopper that has a short piece of glass tubing and rubber tubing attached. Assemble the apparatus as shown in the figure, using a 400-mL beaker. Be sure that the stopper fits tightly in the flask, the glass tubing fits tightly in the rubber stopper, and the rubber tubing fits tightly on the glass tubing. Leave a gap between the bottom of the flask and the beaker.



2. Pour water into the beaker until most of the flask is covered, but not to the very top of the beaker to avoid water spilling over the top during heating.

3. Using a laboratory burner, heat the water in the beaker until it boils and then continue to heat for 5 minutes. At this point the temperature of the water is 100 ° C, the boiling point temperature of water. It is also the temperature of the air inside the flask. This is T_1 for the air. The volume of air in the flask at this time is V_1. The system is open to the atmosphere, so the pressure of the air in the flask is atmospheric pressure.

4. (a) After the water in the beaker has been boiling for about 5 minutes, place a pinch clamp on the rubber tubing to close the rubber tubing. This is done to seal the system. No air in flask can and no air can enter the flask. This keeps n (mass constant.

   (b) Turn off the laboratory burner.

5. Loosen the ring stand clamp at the ring stand connector (NOT the part around the flask). Raise the clamp up to remove the flask from the beaker. Do not touch the flask; it is HOT. Carry the clamp still attached to the flask and immerse ALL (100%) in the water in a stoppered sink. It should be cooled in the sink for about 5 minutes, with the flask stopper and tubing completely immersed in water throughout this period.

6. At the end of this time, while everything is still ***completely submerged*** loosen the pinch clamp and let water be forced into the flask. BE VERY CAREFUL. WHILE LOOSENING THE PINCH CLAMP, HOLD THE RUBBER STOPPER IN THE FLASK TO KEEP IT FROM BEING ACCIDENTLY WIGGLED LOOSE.

   Explanation of what is happening. Back at step 3, you were heating an open flask. Before heating, there was a certain mass of air in the flask that could have been measured. Let us say there was 1 gram of air in the flask. You started heating the flask. Some of the gas was driven out of the flask. Let us say there was 0.5 g of air in the flask when you closed it at step 4. Even though there was less air, it was still at atmospheric pressure because it was moving faster. Then, you placed the closed system under water. This cooled the air in the sealed flask. The air starts moving slower and the pressure in the flask decreased. Charles' Law states the pressure at the end of the experiment has to be the same as at the beginning. To get the pressure inside the flask back to atmospheric pressure, you loosen the clamp to re-open the system. With less pressure inside the flask, the water (which is at atmospheric pressure) enters the flask. The water keeps entering the flask until the pressure inside the flask is equal to atmospheric pressure.

7. Now, ***holding the rubber tubing closed with your fingers***, remove the flask from the sink, and determine the volume of the water in the flask by pouring it into a 100-mL graduated cylinder. Record this volume as entry (a) in the data table.

8. Determine the temperature of the water in the sink and record it as entry (d) in the data table.

9. To determine the volume of gas (air) used in the beginning, it is necessary to determine the volume of the Erlenmeyer flask. ***Completely*** fill the flask with water and place the rubber stopper in its previous position. Remove the stopper and measure the volume of the water in the flask with a 100-mL graduated cylinder. Record this volume as entry (b) in the data table. STUDENTS: Do not go past the 100-ml line on the graduated cylinder when putting water into it. The Erlenmyer is a 125-mL flask. What will you need to do?

(a) Volume of water forced into the flask	mL	
(b) Initial volume of the air (measured volume of the flask/stopper)	mL	
(c) Initial temperature of the air (temperature of the hot water)	°C	K
(d) Final temperature of the air (temperature of the water in the sink)	°C	K
(e) Final volume of the air [(b)–(a)] (measured volume of the flask minus the volume of water forced into the flask)	mL	

10. (a) Using the recorded data and Charles' Law, calculate the theoretical final volume.

(b) What was the experimental final volume?

(c) Determine the percent error of the experiment.

$$\frac{|\textbf{Theoretical - Actual}|}{\textbf{Theoretical}} \textbf{ X 100 = \% Error}$$

Pre-laboratory Questions
Charles' Law

1. List 3 factors that affect the pressure of a gas.

2. Answer decrease or increase for each of the following effects on the pressure of a gas. Assume that the other factors are not different.

   (a) Temperature is decreased
   (b) Mass is increased
   (c) Volume is decreased
   (d) Volume is increased
   (e) Mass is decreased
   (f) Temperature is increased

3. (a) If you throw a sealed can into a fire, the inside of the can will get (colder, hotter).
   (b) The pressure inside the can will (decrease, increase).
   (c) The change of pressure inside the can will cause the can to (explode, implode).
   (d) What does implode mean?

4. Fill in the blanks.

   (a) 85° C = _____________ K.

   (B) 273 K = ___________ ° C

5. Work the following problems. Remember, the temperature needs to be in Kelvins. Also, remember to include the unit in the answer. The containers are sealed, but are able to have a decrease/increase in volume.

   (a) A container holds 50.0 mL of nitrogen at 25° C and a pressure of 736 mm Hg. What will be its volume if the temperature increases by 35° C and the pressure remains constant?

   (b) A sample of hydrogen has an initial temperature of 50° C. When the temperature is lowered to -5.0° C, the volume of hydrogen becomes 212 cm^3. What was the initial volume of the hydrogen? No change in pressure occurs.

Report Sheet: Charles' Law

Name ___________________________________

Lab Partner's Name _________________________________

1. For this experiment, there were several temperatures involved: room temperature, the boiling point of water, and the temperature of the water in the sink.

   (a) When did you measure the initial temperature?

   (b) At that point, the Erlenmeyer flask was not closed. You had an open system. What was the pressure of the air inside the flask?

   (c) Before you turned off the burner, you closed the flask. Why?

   (d) Then, you placed the sealed flask into the water to lower the temperature. According to Charles' Law, what happens to the volume of the gas?

   (e) But, the air in the sealed flask is still occupying the same volume as when it was hot. What is different is the pressure of the air in the flask. It is less than the pressure when it was hot. So, while the flask is cool and sealed, the conditions for Charles' Law have not been met. The initial pressure and the final pressure have to be the same. What do you do to change the pressure in the flask so that is back to atmospheric pressure?

2. The boiling point of water = __________°C.

3. (a) Using the recorded data and Charles' Law, calculate the theoretical final volume.

   (b) What was the experimental final volume?

   (c) Determine the percent error of the experiment.

$$\frac{|\textbf{Theoretical - Actual}|}{\textbf{Theoretical}} \textbf{ X 100 = \% Error}$$

ACIDS AND BASES
pH Indicators

In water, the following reaction always takes place to a very slight extent. Two H_2O molecules collide and are converted to the hydronium ion (H_3O^{1+}) and the hydroxide ion (OH^{1-}). This reaction is called a self-ionization reaction, because water is reacting with water and making ions.

$$H_2O + H_2O \longrightarrow H_3O^{1+} + OH^{1-}$$

In this reaction a hydrogen of one water molecule loses its electron to the oxygen to which it was bonding, forming a hydroxide ion (OH^{1-}). The hydrogen without its electron is the hydrogen ion, H^{1+}. It covalently bonds to the oxygen of another water molecule, forming a hydronium ion, H_3O^{1+}.

You will notice from the equation that every time water self-ionizes that the concentration of hydronium ions is the same as the concentration of hydroxide ions (a 1:1 ratio). In pure water at 25°C, self-ionization occurs to the extent that:

$$[H_3O^+] = 1 \times 10^{-7}\ M$$
$$[OH^-] = 1 \times 10^{-7}\ M$$

There are several definitions of acids. One of the definitions of acids and bases states that an acid is a substance which, when dissolved in pure water, causes an increase in $[H_3O^{1+}]$ and that a base is a substance which, when dissolved in pure water, causes an increase in $[OH^{1-}]$. The [] around a formula means ***the concentration of***.

The following are equations showing how some compounds are acidic because hydronium ions are made when reacting with water. Thus, the concentration of the hydronium ions increases.

1. $HCl + H_2O \longrightarrow H_3O^{1+} + Cl^{1-}$

2. $CuSO_4 + 5\ H_2O \longrightarrow Cu(H_2O)_3OH^{1+} + H_3O^{1+} + SO_4^{2-}$

The following are equations showing how some compounds are basic because hydroxide ions are made when reacting with water. Thus, the concentration of hydroxide ions increases.

3. $NH_3 + H_2O \longrightarrow NH_4^{1+} + OH^{1-}$

4. $NaC_2H_3O_2 + H_2O \longrightarrow HC_2H_3O_2 + OH^{1-} + Na^+$

Pure NaOH (sodium hydroxide) is a solid ionic compound. When dissolved in water, the ions are separated from each other. No hydroxide ions are formed; they already existed before dissolving the sodium hydroxide in the water. There is no chemical reaction involved. Adding hydroxide ions to the water increases the $[OH^{1-}]$; therefore, water-soluble ionic compounds containing the hydroxide ion are bases.

We can also define an acid as a substance that, when dissolved in pure water, causes the pH to drop below 7; a base causes the pH to rise above 7. By definition,

$$pH = -\log[H_3O^+].$$

The log of a number that is written in scientific notation and that has the following restrictions can be easily determined.

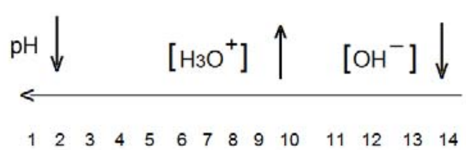
(1) 1 is the number being multiplied by a power of ten
(2) The exponent is a whole number

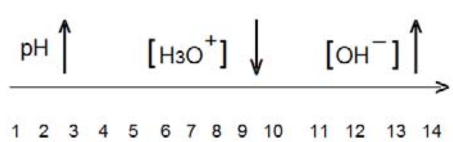
The log can be determined by looking at the exponent. Since pH is -log, then the sign of the exponent has to be reversed. For instance, if $[H_3O^+] = 1 \times 10^{-8}$ M, the pH = 8.

neutral solution	$[H_3O^+] = [OH^-] = 1 \times 10^{-7}$ M	pH = 7
acidic solution	$[H_3O^+] > [OH^-]$ $[H_3O^+] > 1 \times 10^{-7}$ M	pH < 7
basic solution	$[H_3O^+] < [OH^-]$ $[H_3O^+] < 1 \times 10^{-7}$ M	pH > 7

*Memory aid: A comes before B; 1 comes before 14. Acidic solutions have lower number pH.

If $[H_3O^+] = 1 \times 10^{-3}$, the pH is 3. If $[H_3O^+] = 1 \times 10^{-5}$, the pH is 5. Notice 1×10^{-3} is a larger number than 1×10^{-5}. A solution with a lower pH (3) has more hydronium ions than a solution with a higher pH (5). A solution with a pH of 3 is more acidic than a solution with a pH of 5.





If we know the pH of a solution, we can calculate $[H_3O^+]$:

$$[H_3O^+] = 1 \times 10^{-(pH)}$$

For instance, if the pH = 2, then the $[H_3O^+] = 1 \times 10^{-2}$ M.

In pure water:

$$[H_3O^+][OH^-] = (1 \times 10^{-7}\ M)(1 \times 10^{-7}\ M) = 1 \times 10^{-14}\ M^2$$

In aqueous solutions at 25^0C, the product of the hydronium ion concentration and the hydroxide ion concentration is 1×10^{-14}, always. If we know the pH of a solution, we can calculate the $[OH^-]$. For instance, if the pH is 3 the $[OH^-] = 1 \times 10^{-11}$.

An indicator is a substance that changes color at a certain pH. Paper that has been treated with a mixture of indicators can be used to test for a wide range of pH values. When using indicators, there is no pH change. The indicator's pH does not change. The pH of the solution being tested does not change.

An acid will react with a base. Before the reaction, an aqueous solution of the acid would have a pH below 7 and an aqueous solution of the base would have a pH above 7. The pH of the solution after adding the acidic solution to the basic solution would be 7 or would be closer to 7. Since a neutral solution has a pH of 7, the reaction between an acid and a base is called a neutralization reaction. A neutralization reaction always makes a salt. A salt is an ionic compound.

Vinegar is an aqueous solution of acetic acid. Baking soda is sodium bicarbonate. Sodium bicarbonate is a base. When they react, the products are water, carbon dioxide, and sodium acetate. Sodium acetate is the salt; it is composed of sodium ions and acetate ions.

CH_3COOH	+	$NaHCO_3$	$\longrightarrow$	H_2O	+	CO_2	+	$NaCH_3COO$
Acetic Acid		Sodium Bicarbonate						Sodium Acetate

Below is a list of the solutions used in today's experiment. You will first determine the pH of each by using pH paper bought from a chemical supply company. Then, you will use the recorded pH to assign pH numbers to colors produced by an indicator bought from a grocery store. Purple cabbage (sometimes called red cabbage) contains a water-soluble pigment called anthocyanin that changes color when it is mixed with an acid or a base.

Aqueous Solutions Used:

- HCl, hydrochloric acid
- NaOH, sodium hydroxide
- CH_3COOH, acetic acid (Vinegar is a 5% solution of acetic acid.)
- $NaHCO_3$, sodium bicarbonate (baking soda)
- NH_3, ammonia (Household ammonia in grocery stores is a 5-10% solution.)
- Sprite (or 7-UP)
- Tap water (tap water is not pure water)

Items Needed for Each Section of Lab (Stockroom)

- 100 strips (each 1 in. long) of Hydrion pH paper, 0.0-12.0
- $NaHCO_3$, enough to fill an evaporating dish about half full
- All of the solutions below should be in bottles with droppers
  - HCl, 0.1 M, 100 mL
  - NaOH, 0.1 M, 100 mL
  - Vinegar, 100 mL
  - $NaHCO_3$, 0.1 M, 100 mL
  - Household ammonia, 100 mL
  - Sprite (or 7-UP), 100 mL
  - Tap water, 100 mL
  - Purple (Red) cabbage broth, 100 mL (made by adding a few leaves to distilled water and boiling for a few minutes)

Items Needed for Each Pair of Students (Already on shelf or in drawer)

- 50-mL beaker
- 7 test tubes
- Stirring rod
- Wash bottle

Procedure:

1. Get 8 strips of pH paper.

2. Put a drop of an aqueous solution on one of the pieces of the pH paper. There is a pH/color chart for this indicator paper. Record the pH below. The pH values should be whole numbers, no decimals. Do the same for each of the other solutions.

	pH	Color of pH paper
HCl		
NaOH		
CH_3COOH		
$NaHCO_3$		
NH_3		
Sprite		
Tap Water		

BEFORE CONTINUING, HAVE YOUR INSTRUCTOR CHECK YOUR pH VALUES.

3. Fill in the following chart for the solutions. Show the chart to your instructor.

	pH	$[H_3O^{+1}]$	$[OH^{-1}]$	Acidic, Basic, Neutral
HCl				
NaOH				
CH_3COOH				
$NaHCO_3$				
NH_3				
Sprite				
Tap Water				

4. (a) Add 10 drops of cabbage juice to 7 test tubes. If the color of the cabbage juice changes when you add it to a test tube, empty the test tube and rinse it several times. Then add 10 drops again. If the cabbage juice changes, this means that something was in the test tube that should not be there.

   (b) Add 5 drops of one of the 7 tested solutions to one of the test tubes. Now do the same for the other 6 tested solutions; mark the test tubes to remember which test solution is in which test tube.

   (c) With a stirring rod, stir the contents of each test tube. Rinse off the stirring rod between each test tube.

   (d) Fill in the table below based on the cabbage juice. Show your table to your instructor.

	Cabbage Juice Color
HCl	
NaOH	
CH_3COOH	
$NaHCO_3$	
NH_3	
Sprite	
Tap Water	

5. Fill in the table below.

	pH	pH paper color	Cabbage Juice Color	pH
HCl				
HaOH				
CH_3COOH				
$NaHCO_3$				
NH_3				
Sprite				
Tap Water				

6. (a) Does the color of the indicator depend on the pH, or does the pH depend on the color of the indicator?

   (b) Does your answer in (a) agree with your table in #5?

7. Neutralization procedure:

   (a) Get a 50-mL beaker. Add a dropper full of vinegar to the beaker. Add a small strip of pH paper.

   Record the pH of the vinegar: _____________

   (b) Add two good pinches of $NaHCO_3$ to the beaker.

   (c) What are two observations indicating a chemical reaction is occurring?

8. (a) Wait until most of the bubbling has subsided. What is the pH of the solution in the beaker?

   (b) Was the increase of pH evidence of a neutralization reaction?

Pre-Laboratory Questions
Acids and Bases
pH Indicators

1. When a water molecule reacts with a water molecule, what are the two ions that form?

Name	Formula
________________	___________
________________	___________

2. What is the concentration of each of the following in pure water at 25 °C?

   (a) H_3O^+

   (b) OH^-

3. Identify the substance that is reacting with water in the following equations by drawing a circle around it. Classify it (the substance) as an acid or a base.

   (a) $HI + H_2O \rightarrow H_3O^+ + I^-$
   (b) $Mg + 2\ H_2O \rightarrow Mg^{2+} + H_2 + 2\ OH^-$
   (c) $Na_2CO_3 + 2\ H_2O \rightarrow H_2O + CO_2 + 2\ Na^+ + 2\ OH^-$
   (d) $H_2SO_4 + H_2O \rightarrow HSO_4^- + H_3O^+$

4. Write less than, greater than, or equal to on the blanks.

   (a) In an acidic solution, the $[H_3O^+]$ is ____________ the $[OH^-]$

   (b) In a basic solution, the $[H_3O^+]$ is ____________ the $[OH^-]$

   (c) In a neutral solution, the $[H_3O^+]$ is ____________ the $[OH^-]$

5. State whether each of the following solutions is acidic, basic, or neutral:

   (a) blood, pH = 7.4

   (b) vinegar, pH = 2.8

   (c) drain cleaner, pH = 11.2

   (d) pure water, pH = 7.0

6. Arrange the pH values of the following groups in order from the most acidic to the most basic.

   (a) 14, 5, 10, 7

   (b) 3.5, 9.7, 8.8, 11.4

7. Calculate the pH of each of the following solutions with the indicated $[H_3O^+]$.

(a) 1×10^{-4} M

(b) 1×10^{-8} M

(c) 1×10^{-7} M

8. Calculate the $[H_3O^+]$ if the pH is as follows. (Do not write just 10^{-x}, be sure to include the 1 x in front of the 10^{-x}.)

(a) 2.0

(b) 10.0

(c) 7.0

9. Calculate the $[OH^-]$ if the $[H_3O^+]$ is as follows.

(a) 1×10^{-4} M

(b) 1×10^{-8} M

(c) 1×10^{-7} M

10. Complete the following table.

	$[H_3O^+]$	$[OH^-]$	pH	Acidic, Basic, Neutral
(a)	________	1×10^{-6}	____	_______________
(b)	________	________	2.0	_______________
(c)	1×10^{-5}	________	____	_______________
(d)	________	________	10.0	_______________
(e)	________	________	____	Neutral
(f)	1×10^{-12}	________	____	_______________
(g)	________	1×10^{-8}	____	_______________

11. A pH of 1 is how many times more acidic than a pH of

(a) 2? (b) 3?

12. The answers for #12 are such because the pH system is based on powers of the number _______.

13. Answer decreases or increases for each of the following.

(a) As the $[H_3O^+]$ increases, the pH ________________.
(b) As the $[H_3O^+]$ decreases, the pH ________________.
(c) As the $[OH^-]$ increases, the pH _________________.
(d) As the $[OH^-]$ decreases, the pH _________________.
(e) As the pH increases, the $[H_3O^+]$ _________________.
(f) As the pH decreases, the $[H_3O^+]$ _________________.
(g) As the pH increases, the $[OH^-]$ _________________.
(h) As the pH decreases, the $[OH^-]$ _________________.

14. An acid added to water increases which ion's concentration, the hydronium or the hydroxide?

15. What does [] around a formula mean?

16. Aqueous means that something is dissolved in ________.

17. (a) What is the grocery store name for $CH_3COOH_{(aq)}$?

(b) What is the grocery store name for $NaHCO_3$?

18. At the grocery store, vinegar is usually a 5% solution. Hydrogen peroxide is usually a 3 % solution.

(a) Vinegar is _________% water and ________% acetic acid.

(b) Grocery store hydrogen peroxide is ________% water and _________ % hydrogen peroxide.

19. Will the pH of the tested solutions change when you use indicators to determine pH?

20. Which of the following statements is correct?

(a) The pH of a solution depends on the color of an indicator.

(b) The color of an indicator depends on the pH of a solution.

21. (a) What type of a substance reacts with an acid to bring the pH closer to 7 (the pH of a neutral solution)?

(b) An acid reacting with a base is called a _________________ reaction.

(c) A neutralization reaction always makes a _______________.

(d) A salt is an ________________ compound.

(e) Why is a reaction between an acid and a base called a neutralization reaction?

22. Write the chemical equation for the reaction between sodium bicarbonate and acetic acid. Circle the formula of the product that is a salt.

23. In this laboratory, several reagents are found in grocery stores. These are vinegar, baking soda, sprite, and household ammonia. You do not buy hydrochloric acid or sodium hydroxide at the grocery store. Your stomach makes HCl for the process of digesting proteins. Sometimes sodium hydroxide is called lye. Also, sometimes a solution of sodium hydroxide is referred to as a caustic solution.

(a) What are the two compounds in household ammonia?

(b) What are the two compounds in vinegar?

(c) Why does your stomach make HCl?

(d) Name an acid in sprite. (Google it)

(e) What is an alternative name sometimes used for NaOH?

(f) What does caustic mean? (Look it up)

24. What is the type of compound found in red cabbage leaves that can be used as an indicator?

25. For hydrochloric acid, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

26. For sodium hydroxide, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

27. For acetic acid, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

28. For sodium bicarbonate, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

29. For ammonia, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

Report Sheet: Acids and Bases
pH Indicators

Name ___________________________________

Lab Partner's Name _________________________________

1. (a) Which of the basic solutions was the most basic?
   (b) Why do you say this?

2. (a) Which of the acidic solutions was the most acidic?
   (b) Why do you say this?

3. (a) What were your two observations indicating a chemical reaction between the sodium bicarbonate and the acetic acid.

   (b) What gas was in the bubbles that you saw?

   (c) Is the gas that forms in this reaction water-soluble?

   (d) Explain your answer for (c).

4. You used two different types of indicators today. One was the pH paper bought from a chemical supply business. The other was the red cabbage that can be bought at a grocery store.

   (a) Will the pH of a tested solution change when you use indicators to determine pH?

   (b) Circle the following statement which is correct.

   The pH of a solution depends on the color of an indicator.

   The color of an indicator depends on the pH of a solution.

   (c) Do your answers for (a) and (b) agree with the two tables (#3 and #5 in the Procedure)? In other words, is the pH for each test solution the same in the two tables, no matter which indicator is used?

5. Fill in the following chart for the solutions.

pH	$[H_3O^{+1}]$	$[OH^{-1}]$	Acidic, Basic, Neutral
HCl			
NaOH			
CH_3COOH			
$NaHCO_3$			
NH_3			
Sprite			
Tap Water			

Acid-Base Titration

The following are definitions of acids and bases.

- An acid is a substance which causes an increase in the $[H_3O^{1+}]$. A base is a substance which causes an increase in the $[OH^{1-}]$. Below are examples of reactions to illustrate this definition.

  $$HCl + H_2O \rightarrow H_3O^{1+} + Cl^{1-}$$

  $$NH_3 + H_2O \rightarrow NH_4^{1+} + OH^{1-}$$

- An acid causes the pH to decrease; a base causes the pH to increase.

- An acid donates (loses) a hydrogen ion, H^+. A base gets a hydrogen ion. The equations above illustrate this definition also. HCl is an acid. Cl keeps the electron of H, so the H becomes H^+ and the Cl becomes Cl^-. HCl is donating a hydrogen ion. Water gets the H^+, so it is reacting as a base. In the other equation, NH_3 gets a hydrogen ion from water and forms NH_4^+, so NH_3 is a base. In this reaction, water gives up a hydrogen ion and it is reacting as an acid. The oxygen in water keeps the hydrogen's electron, thus a hydroxide ion forms.

A neutralization reaction is a reaction between an acid and a base. The pH of the resulting solution is 7 OR closer to 7 than the pH of the acid or of the base. Since a neutral solution has a pH of 7, the reaction is called a neutralization reaction. A neutralization reaction always makes a salt. A salt is an ionic compound. Below is an example of a neutralization reaction.

$HCl_{(aq)}$	+	$NaOH_{(aq)}$	$\rightarrow$	$NaCl_{(aq)}$	+	H_2O
acid		base		salt		water
pH =1		pH=13		pH of $NaCl_{(aq)}$=7		pH of water=7

As explained above, HCl is an acid because it donates a hydrogen ion. NaOH is an ionic compound. The hydroxide ion, OH^-, accepts the H^+. It is the basic part of NaOH. The sodium ion, Na^+, is not changed by the reaction. The hydroxide ion and hydrogen ion combine to form water. Water IS NOT an ionic compound. THERE IS NO IONIC CHARGE ON WATER. If the water was boiled off, NaCl would be left in the beaker. NaCl is composed of Na^+ and Cl^- ions, so it is a salt.

Here is another neutralization equation.

$NH_{3(aq)}$ +	$HCl_{(aq)}$	$\rightarrow$	$NH_4Cl_{(aq)}$
base	acid		salt
pH=10	pH=1		pH=6

As explained above, NH_3 is a base and HCl is an acid. The salt produces a solution with a pH of 6. A pH of 6 is closer to 7 than a pH of 10 or 1. NH_4Cl is composed of NH_4^+ and Cl^- ions, so it is a salt.

Titration is the measurement of the volume of one reagent required to react with a volume of another reagent. For the titration of an acid-base reaction, the solution of a base is added from a burette to the solution of an acid in a flask until the acid is neutralized. This is called reaching the end-point. Then, one more drop of base is added, so that the pH of the solution in the flask becomes basic. This process requires the use of an indicator so that you will know when the end-point is reached. The indicator phenolphthalein is colorless in an acidic solution and pink in a basic solution. Phenolphthalein is added to the acidic solution in the flask before the titration is begun. This colorless solution turns pink when, as a result of base addition from the burette, the solution changes from acidic to barely basic. When the color change appears, the burette is read to obtain the volume of basic solution added.

Below is the equation for the reaction being done today. The acid is acetic acid. The base is sodium hydroxide. The hydrogen ion (H+) is donated by the acid to the hydroxide ion (OH-) to form water. The sodium ion (Na^+) from sodium hydroxide and the acetate ion ($CH_3CO_2^-$) formed from the acid make the salt sodium acetate. The balanced equation tells you that one mole of acid is neutralized by one mole of base. Since the ratio of reacting moles is 1:1, then if you know the moles of one reactant being used, you also know the number of moles of the other reactant.

$$CH_3CO_2H_{(aq)} + NaOH_{(aq)} \rightarrow H_2O + NaCH_3CO_{2(aq)}$$

Molarity (M) expresses solute concentration. Molarity is defined as moles of solute per liter of solution. For simplicity: solute is the substance dissolved in water; solution is water and what is dissolved in it.

$$M = \frac{\text{moles of solute}}{\text{1 liter of solution}}$$

If a concentration is given as 0.225 M, this means that in 1 liter of that solution, there is 0.225 moles of solute. If you know the formula weight of the solute, then the 0.225 moles can be converted to grams.

This equation can be rearranged.

$$M = \frac{\text{moles}}{\text{volume}} \qquad \text{moles} = M \times \text{volume}$$

From the balanced equation (as stated earlier): $\text{moles}_b = \text{moles}_a$

Therefore: $M_bV_b = M_aV_a$

In this experiment, you are going to be given the M_b. You will measure the volumes of the acid and the base. Then you will solve for M_a.

Items Needed for Each Section of Lab (Stockroom)

- 12 burettes
- Acetic acid solution, 1 L (concentration not known to students)
- Sodium hydroxide solution, 1 L (concentration written on container)

Items Needed for Each Pair of Students (Already on shelf or in drawer)

- Ring stand
- Burette clamp
- Beaker
- Wash bottle
- 100-mL graduated cylinder
- Funnel
- Two 125-mL Erlenmeyer flasks

Procedure

1. Get a ring stand and a burette clamp. Get a burette and place it in the clamp.

2. Place a small beaker (this is going to be the waste beaker) under the burette. Using a wash bottle, rinse out the burette by squeezing out water and directing the water around the interior of the burette. Drain out as much of the water as best you can. You will probably not be able to get all of it out. At this time, you can practice learning how to use the burette. See if you can open and close the stopcock well enough to allow only one drop at a time.

3. Get a 100-mL graduated cylinder. Rinse it out with some water. Go to the hood and get about 60 mL of acetic acid solution.

4. Place a funnel in the top of the burette. Hold the funnel up a little when pouring liquids into the burette to allow the air to escape. The waste beaker is still under the burette.

5. Add acetic acid solution to the burette until it is slightly above the 0 mL line. Remove the funnel. AFTER REMOVING THE FUNNEL, CHECK TO MAKE SURE THE BOTTOM OF THE MENISCUS IS STILL ABOVE THE 0 mL LINE. Open the stopcock SLOWLY and let the acetic acid solution come down so that the bottom of the meniscus is exactly at the 0 mL line. You may have to lower the burette to get the meniscus at your eye level.

6. Get two Erlenmeyer flasks. Rinse them out with water. Pour out as much of the water as possible.

7. Place one Erlenmeyer flask under the burette. Lower the burette so that the tip is in the top of the flask.

8. Open the stopcock slowly and add exactly 15.0 mL of acid solution to the Erlenmeyer flask.

9. Remove the first Erlenmeyer flask to the side and place the second one under the burette. Add exactly 15.0 mL to the second Erlenmeyer flask. Remove this second flask to the side.

10. Add 2-3 drops of phenolphthalein indicator to each flask. Swirl to mix. There should be no color produced since this indicator is colorless at an acidic pH. IF YOU SEE ANY PINKISH COLOR IN ANY OF THE ERLENMEYERS, then empty the Erlenmeyer and repeat steps 6-8 for that flask. If there is no color in a flask, it is good to go and the solution does not need to be replaced.

11. Empty the burette as much as possible into the sink/trough. Discard any acetic acid solution remaining in the graduated cylinder. Thoroughly rinse the graduated cylinder.

12. Using a wash bottle, rinse out the burette by adding a little water along the sides. Drain out as much of the water as best you can. You will probably not be able to get all of it out.

13. Go to the hood with your thoroughly rinsed out graduated cylinder and get about 60 mL of sodium hydroxide solution.

14. Place a funnel in the top of the burette. However, remember to hold the funnel up a little when pouring liquids into the burette. Place the waste beaker under the burette.

15. Add sodium hydroxide solution to the burette until it is slightly above the 0 mL line. Remove the funnel. AFTER REMOVING THE FUNNEL, CHECK TO MAKE SURE THE BOTTOM OF THE MENISCUS IS STILL ABOVE THE 0 mL LINE. Open the stopcock SLOWLY and let the NaOH solution come down so that the bottom of the meniscus is exactly at the 0 mL line. You may have to lower the burette to get the meniscus at your eye level.

16. Place one of the Erlenmeyer flasks under the burette. Place a piece of white paper under the flask. You can use a paper towel. Lower the burette until the tip is in the flask.

17. Open the stopcock SLOWLY and add 2 or 3 drops of NaOH solution. You will see some pinkish color where the NaOH is hitting the acid solution. Swirl the flask until the pink color is gone. Add 2 or more 3 drops again. Swirl the flask until the pink color is gone. Do not add any more drops of NaOH until swirling removes all of the color. Keep doing this process (drops and swirl) until the swirling does not remove the color. In the table (step #21) record to the nearest tenth the amount of base added. You may have to lower the burette to get the meniscus at your eye level.

    HAVE YOUR INSTRUCTOR CONFIRM THIS MEASUREMENT. THIS IS THE STEP WHERE MOST STUDENTS MAKE A MISTAKE!

18. Place the second Erlenmeyer flask under the burette. Place a piece of white paper under the flask. You can use a paper towel. Lower the burette until the tip is in the flask.

19. This second titration will be done based on your first titration. The procedure will be slightly different. Open the stopcock SLOWLY and add 2 or 3 drops of NaOH solution. Swirl until the pink color disappears. Continue doing this process until you get to within 2 mL of your first addition. Then, start adding 1 drop of NaOH solution at a time and swirl after each drop. Watch carefully for any evidence of a color change. Get the lightest pink possible.

20. In the table record to the nearest tenth the amount of base added. You may have to lower the burette to get the meniscus at your eye level.

AGAIN HAVE OUR INSTRUCTOR CONFIRM THIS MEASUREMENT.

21. Complete the data table below. To convert mL to L, divide by 1000. This is because 1 L = 1000 mL.

(a) NaOH used for first titration ________ mL; ________ L

(b) NaOH used for second titration ________ mL; _______ L

(c) Acetic acid used for each titration ________ mL; _______ L

22. Use the MV=MV equation to calculate the molarity of the acid for each titration. The molarity of the base (M base) is written on the bottle of NaOH. Then calculate the average of the molarity of the acid solution. DO NOT ROUND TOO MUCH. Have at least 3 decimal places.

	Titrations		
	1	**2**	
M base			
L $_{base}$			
L $_{acid}$			
M $_{acid}$			Average M $_{acid}$

Questions

1. (a) How many moles of acetic acid are in 1 liter of the acetic acid solution? Use the average molarity of the acid solution for your answer.

(b) How many grams are in one mole of CH_3CO_2H? SHOW WORK.

(c) Calculate the grams of acetic acid in one liter of solution. SHOW WORK.

2. Write the name of the indicator with correct spelling.

Be sure to clean your burette before you return it to the burette box. Fill it with water and drain.

Pre-Laboratory Questions
Titration

1. An acid is a substance which causes an increase in the __[   ]_______.

2. A base is a substance which causes an increase in the ____[   ]______.

3. An acid causes the pH to _________; a base causes the pH to __________.

4. An acid donates a ______________ion, which has the formula __________.

5. A base gets a ________________ion.

6. What is the definition of a neutralization reaction?

7. True/False: A neutralization reaction always produces a neutral pH.

8. A neutralization reaction always makes a __________.

9. What is the definition of a salt?

10. What is the definition of titration?

11. (a) Write the equation for the reaction being done in your titration lab.

    (b) Write the names of the reactants and the products.

12. What is the definition of molarity?

13. Do you use M or m for molarity?

14. What are the simple definitions for solute and solution?

15. What is the mathematical formula for molarity?

16. If M = 1.023, how many moles of a solute are dissolved in one liter of solution?

17. Rearrange the mathematical formula for molarity to calculate moles of solute.

18. Rearrange the mathematical formula for molarity to calculate the volume of solution.

19. All of the questions below (a-c) are based on the following neutralization reaction. The coefficients let you know that the moles of NaOH (base) and the moles of HCl (acid) are equal.

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

(a) How many moles of HCl are needed to neutralize 0.10 L of 2.0 M NaOH? Express your answer with 2 significant figures.

(b) If it takes 0.025 L of 0.050 M HCl to neutralize 0.034 L of NaOH solution, what is the molarity of the NaOH solution? Express your answer with 2 significant figures.

(c) If it takes 0.0540 L of 0.100 M NaOH to neutralize 0.125 L of an HCl solution, what is the molarity of the HCl? Express your answer with 3 significant figures.

20. (a) What is the name of the indicator that is going to be used for the titration? Correctly spelled.

(b) What color is this indicator in an acidic solution?

(c) Is clear a color?

20. For acetic acid, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

21. For sodium hydroxide, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

22. For phenolphthalein, what would be present for the following elements on a GHS label?

Pictogram(s):

Signal Word:

Hazard Statement:

Precautionary Statement:

Report Sheet: Titration

Name ____________________________________

Lab Partner's Name __________________________________

	Titrations		
	1	**2**	
M base			
L base			
L acid			
M acid			Average M acid

Questions

1. (a) How many moles of acetic acid are in 1 liter of the acetic acid solution? Use the average molarity of the acid solution for your answer.

   (b) How many grams are in one mole of CH_3CO_2H? SHOW WORK.

   (c) Calculate the grams of acetic acid in one liter of solution. SHOW WORK.