**Honors Chemistry**

**Lab Manual**

**2012-2013**

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[Lab #1: Density and Measurement 4](#_Toc322673938)

[LAB #2: Significant Figures Lab 6](#_Toc322673939)

[Lab #3: Specific Heat of a Metal 7](#_Toc322673940)

[Lab #4: Calorimetry Lab - Food Energy in a Bugle 9](#_Toc322673941)

[Lab #5: Law of Definite Proportions 11](#_Toc322673942)

[Lab #6: Rutherford - Measuring What You Cannot See 14](#_Toc322673943)

[Lab #7 Isotopes and Mass Spectrometry 16](#_Toc322673944)

[Lab #8 : Separation of Mixtures Lab 20](#_Toc322673945)

[Lab #9: Periodic Table: “Families” 21](#_Toc322673946)

[Lab # 10: Properties of Metals and Nonmetals 22](#_Toc322673947)

[Lab # 11: The Baggie lab 24](#_Toc322673948)

[Lab #12 : Classifying Chemical Reactions (from Flinn ChemTopics) 25](#_Toc322673949)

[Lab #13: The Bean Lab (from World of Chemistry) 28](#_Toc322673950)

[Lab #14: Determining an Empirical Formula 30](#_Toc322673951)

[Lab #15 : Electrical Solutions from World of Chemistry 33](#_Toc322673952)

[Lab #16: Reactions in Solutions I - Precipitation 36](#_Toc322673953)

[Lab #17: Activity of Metals/Constructing an Activity Series 38](#_Toc322673954)

[Lab #18: Reactions in Solution – Acid and Base Lab 40](#_Toc322673955)

[Lab #19: Unknown solutions lab 43](#_Toc322673956)

[Lab # 20: MicroMole Rocket Lab 45](#_Toc322673957)

[Lab #21: Copper and Silver Lab 47](#_Toc322673958)

[Lab #22: Limiting Reactants – Sodium Bicarbonate and Vinegar 49](#_Toc322673959)

[Lab #23: Mass & Mole Relationships in a Chemical Reaction 51](#_Toc322673960)

[Lab #24: Heat of Solution and Heat of Reaction 53](#_Toc322673961)

[Lab #25: Heat of Reaction and Hess’s Law Lab 55](#_Toc322673962)

[Lab #26: Differences in Ionic and Covalent Bonding 57](#_Toc322673963)

[Lab #27: The Gas Labs – Investigating Properties Which Affect Pressure 59](#_Toc322673964)

[Lab #28: Determining the Universal Gas Constant 63](#_Toc322673965)

[Lab #29: The Effect of Temperature on Vapor Pressure 65](#_Toc322673966)

[Lab #30: Evaporation and Intermolecular Attractions 67](#_Toc322673967)

[Lab #31: Unsaturated, Saturated and Supersaturated Lab 70](#_Toc322673968)

[Lab #32: Solution Concentration 72](#_Toc322673969)

[Lab #33: Chloride Concentration in Tap Water Lab 73](#_Toc322673970)

[Lab #34: Determining Molar Mass from Freezing Pt. Depression 75](#_Toc322673971)

[Lab #35: Acid-Base Titration Lab 77](#_Toc322673972)

# Lab #1: Density and Measurement

In this lab we will be studying density, as well as the importance of measuring accurately. Your objective will be threefold: obtain the density of an unknown block of plastic, also calculate the density of an unknown rock (mineral) and finally, determine the thickness of a piece of aluminum foil. At the same time we will be looking at the importance of measuring accurately.

Before starting, read through the experiment to determine what measurements you will need to take and construct a data table in the appropriate space Your data table does not need to include calculations so please show only direct measurements like length or mass, do not include density in your data table. (Please use a ruler and include headings with units). Then answer the pre-lab questions in the space provided.

**Part One:**

In this section you will determine the density of an unknown block of plastic. Measure very carefully, you will be ask to identify this plastic based on its density.

1. Obtain a piece of plastic from your teacher. Make a note of the ID letter from your block of plastic and write it in your data table.
2. Using only a ruler and a balance, obtain the density of your block of plastic. (As a part of your pre-lab, you should have given some thought as to what measurements this would require. Record all measurements in your data table. When measuring with the ruler, remember to interpolate one decimal place.
3. There is some variety in the density of plastic based on the conditions under which the tree was grown. Because of this, we will be looking for an average density. Find two more pieces of plastic with the same ID letter and calculated the density for these blocks as well. Remember to record all measurements in your data table.

**Part Two:**

1. In this section you will determine the density of a cylinder of metal.
2. Obtain a sample of metal from your teacher. Make a note of the ID letter of your metal.
3. The cylinder might require a different method for calculating volume. Using only a graduated cylinder (and water) and a balance, measure as needed to obtain the density of your metal sample. When measuring with the graduated cylinder, remember to interpolate one decimal place.
4. While pure metal samples of the same metal should be consistent regarding density, please do at least one more sample of the same metal so that you can average the two densities. This will help with any human error.

**Part Three:**

1. In this section, you will determine the thickness, in mm, of a piece of aluminum foil. You may use a ruler and a balance. It may be helpful to know that the density of aluminum is 2.700 g/ml.

**Prelab Questions: Do these prior to coming to lab**

1. What is the formula for density?
2. Based on what is available in the lab, what measurements will you need to make to calculate the density of your block of wood.
3. What measurements will you need to make to calculate the density of your metal cylinder.
4. Suppose that in the lab you used a ruler marked in centimeters. If you were measuring the length of a block of wood, which of these measurements would be correct: 35.4 cm or 35 cm. Why?
5. What does it mean to interpolate?
6. When using a digital balance we do not interpolate. Is this the same as saying our measurement is exact? Explain.
7. In this lab we use two different methods for finding volume. If I were interested in finding the density for a rock I found in my backyard, which method would I use? Why?

## Data Table(s)

Read the lab to see what data is measured and then draw a table to include that data. Caution: In this lab, not all measurements are stated in the procedure. Some require you to think on your own.

# LAB #2: Significant Figures Lab

***Part One: Calculating Volume***

Obtain a four-sided meter stick from the supply table. Note that each side of the meter stick is marked with a different degree of precision.

1. You will be assigned an object to measure. You are to measure that object four times, once with each side of the meter stick.
	1. Remember when making your measurements to extrapolate one place past the marked unit, as discussed.
	2. Record your measurements in the table provided.
	3. Using your measurements, calculate the volume of the measured object and record.

|  |
| --- |
| Side One |
| Length |  |
| Width |  |
| Height |  |
| Side Two |
| Length |  |
| Width  |  |
| Height |  |
| Side Three |
| Length |  |
| Width |  |
| Height |  |
| Side Four |
| Length |  |
| Width |  |
| Height |  |
| Calculating Volume |
| Side One |  |
| Side Two |  |
| Side Three |  |
| Side Four |  |

# Lab #3: Specific Heat of a Metal

**Purpose**: Using three metal samples of the same mass, measure the heat transferred from each sample and compare.

**Background**: The movement of heat from one object to another always flows in one direction – from hot to cold. When two objects are the same temperature no transfer of heat will occur. The *amount* of heat transferred can be determined by measuring changes in temperature. In this experiment we will study the heat changes that take place in three metal samples.

While all objects will change temperature as heat is transferred, the overall change in temperature will not be the same for all objects. Some objects heat up (or cool down) more easily than others. An object’s ability to change temperature is called its heat capacity or specific heat. The specific heat tells us the energy (usually measured in joules) required to change 1 grams of a substance by 1 °C.

To measure the transfer of heat we will be using an object called a **calorimeter**. A calorimeter is an object with a known specific heat or heat capacity, in which the change in temperature can be monitored. We will be using a calorimeter to compare our metal samples. After heating the metal samples, we will place each one separately in a calorimeter (in this case a cup of water) and measure the change in the water’s temperature. From this change in temperature, we can then calculate the heat gained by the water.

You might wonder how the heat that the water gained is going to help us compare the metals. Well, calorimetry is based on the Law of Conservation of Energy. According to the Law of Conservation of Energy, energy cannot be created or destroyed. This in turn, tells us that the heat gained by the water must be equal to the heat lost from the metal. Because each metal we measure will have its own specific heat we will see a difference in how much each metal heats up the water.

**Procedure:**

1. Obtain three masses from your teacher and mark the type each metal in your data table.
2. Prepare a hot water bath according to the instructions given by your teacher.
3. Weigh each mass using the balance and record with your data.
4. Place all three masses carefully in the hot water bath. (You must lower the masses with your tongs to the bottom of the beaker. Do NOT drop them in or you risk breaking the beaker and spilling boiling water everywhere!). Allow the metals to stay in the boiling water for 5-10 minutes.
5. While you are waiting for the metal samples you will need you prepare your calorimeters. In this lab, the calorimeter is simply a coffee cup filled with water. Get three coffee cups from the supply table. Your teacher will give you a volume of water. Measure out this volume of cool tap water using a graduated cylinder (remember to interpolate!) and record the **exact** volume of water in your data table. Pour into your coffee cup being careful not to spill any water or leave any behind. Repeat until all three coffee cups are filled. After five minutes take the temperature of the water in each coffee cup. Record this as your initial temperature of water.
6. After the metals have been in the water for the 5-10 minutes, record the temperature of the boiling water. Because your metals have been in the boiling water, this is also the temperature of the metal. Record this value as your **initial** temperature of the metal.
7. Remove one metal from the boiling water, quickly touch it against a paper towel to remove excess water (do not touch it with your hands as it will be very hot!) and then immediately place it into your calorimeter (your cup of water).
8. With your thermometer monitor the temperature of the water until the temperature stops changing. When the temperature remains constant record the final temperature of the water and the metal.
9. Repeat steps 7 & 8 until all three metals have been tested.

## Prelab Questions: do these prior to coming to lab

1. You will be testing three metal samples. Each metal sample will be heated to 100°C. Following the heating, each metal will then be placed into a room temperature cup of water.
	1. What do you expect will happen to the temperature of the cup of water?
	2. How do you expect the final temperature of one cup to compare to the final temperature of another (cups with different types of metals). Explain your answer.
2. When will the temperature of the water stop rising? (For that matter, when will the temperature of the metal stop decreasing?). Explain your answer.
3. When the metal sample is placed into the cup of water heat is transferred from one object to the other. What can you tell me about the amount of heat lost by the metal and the amount of heat gained by the water? Explain how you know this.
4. In an earlier lab we learned about intensive and extensive properties. Is heat an intensive or extensive property? How about specific heat, is it intensive or extensive? Explain both answers.

**Data Table: Prepare a data table before coming to lab (remember to use a ruler!)**

# Lab #4: Calorimetry Lab - Food Energy in a Bugle

**Background:**

How do we know how much energy is stored in foods? Chemists can determine this by burning a known amount of food under controlled conditions and carefully measuring the quantity of thermal energy it releases. This procedure is called calorimetry and the measuring device is called a calorimeter.

All compounds, including foods, contain energy in the **bonds** of the chemicals they are composed of. When we digest our food changes in those chemical bonds take place and energy is released. Our bodies then capture that energy in a series of complex biochemical reactions to fuel our daily activities.

The energy found in chemical bonds cannot be measured directly. There is no instrument which can be inserted into a food for some sort of digital readout of calories. In order to determine the energy in a food we must consume the food and measure the changes that take place.

In this experiment you will burn the Bugle to release its energy. The oil in Bugles burns rapidly when ignited. As the Bugle burns it is releasing energy. As in the previous lab, this released energy will be captured by water. The temperature change in the water will be measured. From this the amount of heat transferred can be calculated.

One helpful note – In the previous lab we used joules as our unit of energy. In this lab we will want to use calories so that we can compare our results with the food label. Both joules and calories are units of heat however they are units of differences sizes. One calories is equal to 4.184 joules. To make our comparison slightly more difficult, a calorie in the lab is not the same as a dietary Calorie. The Calories you read about on food labels are actually ***kilo***calories, 1000 times greater in size than the calories we will be measuring in the lab. It takes 1000 laboratory calories to equal 1 food label Calorie.

Read the lab to determine what measurements you will be recording, answer the prelab questions and construct a data table. Because this lab calls for two trials make sure that your data table includes space for both trials.

**Procedure**

1. Make a simple stand for the Bugle using a paper clip and cork, as shown.
2. Measure (to the nearest milliliter) 100 ml of room temperature water in a graduated cylinder. Record the **exact** value of the water in your data table. Pour the water into the calorimeter’s flask being careful not to spill or leave any behind.
3. Set up the calorimeter as shown. Use a thermometer to measure the initial temperature of the water. Record this value in your data table. Remove the thermometer from the flask.
4. Place the Bugle on the support stand (paper clip and cork) and measure the mass of the entire assembly. Place the bugle assembly inside the calorimeter beneath the flask of water.
5. Use a kitchen match or lighter to light the Bugle directly. Do not keep the lighter under the flask for an extended period of time, as it will affect your results.
6. As soon as the Bugle stops burning, carefully stir the water with the thermometer. For a brief period the temperature may continue to rise. When the temperature has reached its highest value and begins to drop record the final (highest) temperature of the water in your data table.
7. Allow the Bugle residue to cool, and then measure the mass of the entire assembly. Record this value.
8. Repeat Steps 2-7 with a new Bugle as instructed. (Trial 2)

## Prelab Questions: Complete before coming to lab

1. If the objective of the lab is to determine the heat in the bugle, why do we need to burn the bugle?
2. What will happen to the temperature of the water as the Bugle burns? Why?
3. In our previous lab the energy transferred was measured in joules. In this lab we will use calories so that we can compare our value with the value given on the food label. What is the conversion from joules to calories?
4. Imagine that we complete our calculations and determine that one burned bugle releases 4,000 calories. This is A LOT of calories! Should I stop eating Bugles in fear of gaining massive amounts of weight? Explain your answer.

## Data Table: complete before coming to lab

# Lab #5: Law of Definite Proportions

**Purpose**: To increase understanding of and verify the Law of Definite Proportions

**Background:**

In this day and age, we are surrounded by a very sophisticated understanding of chemistry and we take for granted that **atoms** and **compounds** exist. You have known from a very young age that compounds such as water (H2O) form in definite, unchanging ratios. However, there was a time when the very fact that atoms existed was undetermined.

In the 19th century, The distinction between what was a compound and what was a mixture had not yet been determined. Scientists were struggling to define the concept of atoms and consequently compounds.

A French scientist, Joseph Proust, was the first to observe experimentally that when elements combine to form compounds they always do so in a consistent mass ratio. For example, when water is formed from hydrogen and oxygen, the mass of hydrogen to oxygen is always 1:8 or 1 grams of hydrogen for every 8 grams of oxygen.

An English scientist, John Dalton, emphasized the importance of Proust’s discovery when he developed the Modern Atomic Theory. In the modern atomic theory Dalton put forth the concept of an **atom** as the simplest form of nature and stated that for any given element the atoms for that element are identical in size and mass. Furthermore, when atoms come together they form compounds. This they always do so in a *consistent, definite ratio.*

The concept of Definite Proportions was not immediately accepted. A well established scientist, Claude Louis Berthollet proposed that compounds were not definite in nature but rather that they could vary in their composition based on the amounts of reactants (“ingredients”) present when the compound was formed. Because of his authority, Proust’s Law of Definite Proportions was only accepted after much experimentation.

In this lab we will be combining magnesium with oxygen by heating. After an initial ignition in which the magnesium burns brightly, the heat of the Bunsen burner will cause the magnesium to bond with oxygen from the air creating the compound magnesium oxide.

**Materials**:

|  |  |  |
| --- | --- | --- |
| Crucible with lid | Magnesium Ribbon | Bunsen burner  |
| Tongs | Clay triangle | Ring stand w/support |

**Procedure**:

Caution: Do not look directly at the burning magnesium. View indirectly. Sustained viewing can cause permanent eye damage! In this lab, the crucible you are working with will become quite hot and could cause a severe burn if handled improperly. Observe all precautions, especially the ones listed below.

1. Measure and record the mass of both your crucible and your lid separately. Record this data in your data table.

2. Obtain a piece of magnesium ribbon approximately 25 cm long. Roll into a coil. The coil should not be so tight that air cannot reach the inner coils; however, you should be able to fit the entire length into the crucible. Take and record the mass of your magnesium.

4. Over a high flame with a well defined blue cone, heat the **uncovered** crucible on the triangle until the magnesium ignites. Failure to have a well defined blue cone will result in a cooler than necessary flame. As a result your magnesium may not ignite. Caution: Do not inhale the smoke produced. Do not look directly at the white light, it can cause permanent eye damage!!!! When the magnesium begins to burn, immediately cover the crucible with your lid (using tongs) and turn off the burner.

1. After smoke production has ceased, restart the burner and continue heating the crucible. The lid should remain on at this time. Caution: Do not lean over the crucible.
2. After heating for a total of 10 minutes, carefully lift the lit with your tongs and check the magnesium. It should be wholly converted to a light gray powder, magnesium oxide.
3. Turn off and remove the burner. Allow the crucible to cool completely until it is cool to the touch. This will take several minutes. Be very careful not to burn yourself, as porcelain can get very hot! DO NOT attempt to weigh your crucible while it is still hot. This will result in inaccurate data.
4. When cool, measure and record the combined mass of the crucible, crucible lid and magnesium oxide.
5. Return your crucible to the ring stand for a second heating. This time you should heat for at least 5 minutes.
6. After this time is up, again turn off the burner and allow the crucible to cool (to the touch). Once it is cool, weigh the crucible/lid/magnesium oxide again.
7. If your two (after heat) weights are within 0.02 g of each other, then your reaction has run to completion and you will record your final mass. If NOT, then run a third heat/cool cycle and weigh again.
8. Continue to heat and weigh until the last two weigh cycles are within 0.02 g of each other. Remember to record your final mass.
9. Follow your teacher’s instructions for proper disposal of the materials.

## PreLab Questions (Must be completed prior to Lab)

1. State the Law of Definite Proportions (Constant Composition)

1. What do we take for granted regarding the formation of compounds?
2. Before the Law of Definite Proportions was experimentally proven and accepted, what was the accepted belief regarding the formation of compounds?
3. In this lab you will be combining magnesium and oxygen to determine their mass ratio. In order to confirm the Law of Definite Proportions, we will then need to compare your results with the rest of the class. Explain why.
4. What safety concerns are pointed out in the lab?
5. In this lab, your final mass of crucible and contents will be greater than your initial weight of crucible and magnesium. Explain why.

## Data Table: Construct a Data Table prior to coming to Lab.

# Lab #6: Rutherford - Measuring What You Cannot See

**Directions**: The keys to science are observation and measurement, which are often used together in experiments. Experiments are carried out to test hypotheses that attempt to explain the world around us. Also, experiments can lead to new hypotheses.

Chemistry experiments are often designed to gather information about what cannot be observed directly. The purpose of this activity is to demonstrate how an experiment can provide information about something that cannot be seen.

**Objectives**:

* Record date for repeated trials of an experiment
* Determine the unknown size of an object without direct measurement
* Box plot the size and analyze the graph
* Compare the difference between the calculated and the actual sizes

**Equipment**:

7 marbles

meter stick

masking tape

caliper

**Procedures:**

1. Make a masking tape line, 60 cm long, on the floor. Use the meter stick to mark the tape at 5, 15, 25, 35, 45, and 55 cm. Also, use a small piece of tape to mark a spot about 1 meter away from the center of the tape.
2. Place marbles along the long piece of tape, one at each of the marked spots. Place one marble on the small piece of tape.
3. With eyes closed, one team member will roll the single marble toward the line of marbles. The second team member will note whether this marble hits any of the other marbles or misses. The hit or miss will be tallied on the data sheet. (If the marble misses the entire line of tape completely, the trial is disregarded and another tallied.)
4. All of the marbles are returned to their original positions and the process is repeated for a total of 75 tallied trials.
5. After 75 trials, the team members trade places and 75 more trials are completed.

###### Data

|  |  |  |
| --- | --- | --- |
| **Trial** | **Team member 1** | **Team member 2** |
| **Hits** |  |  |
| **Misses** |  |  |

## Marble Activity Writeup:

1. Using team member 1’s data, calculate the diameter of one of the spheres to 3 significant figures using the following formula. Repeat using team member 2’s data.

 Diameter = Field width X Number of hits

 (2 X Number of target spheres X Number of trials)

1. Repeat step one using your partner’s data. Average the two values.
2. Your teacher will measure several marbles with an extremely accurate caliper and record the average of these values on the board. Calculate the percent error between your average value and that provided by direct measurement (the calipers).
3. Does the comparison between your number and the directly measured number support **indirect observation** (aka measuring what you cannot see). Why or why not?

# Lab #7 Isotopes and Mass Spectrometry

**Purpose** To study the laboratory method of mass spectrometry and apply it to our understanding of isotopes.

**Mass Spectra** A mass spectrometer is an instrument that measures the mass of a particle as it passes through. Particles (atoms or molecules) in the gas phase are injected into the instrument. They are then bombarded with a high-speed electron to give them a charge. The charged particles are pulled through the instrument by a strong electromagnetic field, until they reach a collection plate. At the collection plate, particles of like mass land in the same area and stack up. The instrument produces a mass spectrum -- a graph which shows the number of particles of each mass that hit the collection plate.

On the mass spectrum, each mass is represented by a peak. The height of each peak is proportional to the number of particles of that mass that were counted. The total height of all of the peaks is proportional to the total number of particles of all masses that were counted.

|  |  |  |
| --- | --- | --- |
| **Mass Number** | **Height of Peak w/ units** | **Percent Abundance** |
|  |  |  |
|  |  |  |
|  |  |  |

***Instruction Examples: Work to be done with the Teacher***

Use the mass spectrum for neon below to determine its fractional abundances. (Hint: Reread the second paragraph under Mass Spectra to help you.) Record your data in the table provided.

**More Mass Spectra** The mass spectrum for a molecule is a bit more complex than that of an atom. When a molecule of Hydrogen Chloride (HCl), for example, is injected, one of two outcomes can occur. The molecule can pass through the instrument intact, and the total mass of both H and I will be represented by a single peak. The other possible outcome is that the molecule will break apart, that is the Hydrogen and the Chlorine will separate from each other and will be represented by two different peaks, one at the mass of hydrogen and the other at the mass of chlorine.

Your instructor will use Hydrogen Chloride and its mass spectrum to show you how to use the spectrum of a molecule to determine its abundances.

Consider the mass spectrum for hydrogen chloride below. Make the required measurements (including units) and record them in the data table that follows the spectrum. Use the method that you devised for neon to calculate the percent abundance for each peak.

|  |  |  |
| --- | --- | --- |
| **Mass Number** | **Height of Peak w/ units** | **Percent Abundance** |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

Your Assignment: Complete the following including the questions that follow.

**Procedure** Consider the mass spectrum for mercury below. Make the required measurements (including units) and record them in the data table that follows the spectrum. Use the method that you devised for neon to calculate the percent abundance for each peak.

|  |  |  |
| --- | --- | --- |
| **Mass Number** | **Height of Peak w/ units** | **Percent Abundance** |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

1. How many isotopes of mercury are there? How do they differ from one another?
2. Using the data from your table, calculate the average atomic mass for mercury.

Now, consider the mass spectrum for hydrogen bromide below. Make the required measurements (including units) and record them in the data table that follows the spectrum. Use the method that you devised to calculate the percent abundance for each peak.

|  |  |  |
| --- | --- | --- |
| **Mass Number** | **Height of Peak w/ units** | **Percent Abundance** |
|  |  |  |
|  |  |  |
|  |  |  |
|  |  |  |

1. How many isotopes of hydrogen bromide exist? What accounts for the fact that there are more bars than there are isotopes of bromine?
2. Using your values from the table, calculate the average atomic mass for bromine.

# Lab #8 : Separation of Mixtures Lab

Background: In this chapter you will learn that there are different types of matter. Some of these types are mixtures, a collection of substances which are not chemically bonded to each other. Because they are not bonded, the components of these mixtures can be separated by physical means. In other words, we can separate them without changing their identity. In this lab, we will separate an iron, sand and salt mixture, using physical properties.

Procedure:

You will be giving a mixture containing iron, salt and sand. Your job is to figure out how we should separate these three components into their pure, uncombined forms. In order to successfully complete the lab, you must recover all three components. In other words, when you are done, I will need to see three piles: one iron, one salt, and one sand.

Your first step is to write a procedure. No lab work may begin until this procedure is approved by your teacher.

If you are wondering what supplies you may use, you may assume that any of your basic chemical supplies will be available. If you feel that you need some equipment not usually found in a lab, ask your teacher if such equipment would be made available.

# Lab #9: Periodic Table: “Families”

***Background***: We continue our look at the patterns in the periodic table. In this activity, we will be observing chemical reactions to see if there are any patterns in when they occur.

***Materials***:

|  |  |  |
| --- | --- | --- |
| Well Plate | Magnesium solution | Calcium solution |
| Barium solution | Sodium carbonate solution (Na2CO3) | Silver Nitrate (AgNO3) |
| Chlorine solution | Bromine solution | Iodine solution |

***Procedure***:

Part One:

1. Place several drops of each chemical (excluding silver nitrate and sodium carbonate) in your well plate. As you place each chemical in its well, make sure to mark the position of that chemical.

2. Add several drops of sodium carbonate to each of the chemicals. Records your observations in your periodic table “data table” by placing a dot in each block where a reaction occurred. In the space below, write down your general observations. In other words, if chlorine reacted with sodium carbonate, I would find the square on the periodic table that represents chlorine and I would place a dot in it.

3. Rinse out your well plate and fill each well with several drops of each chemical again (again, excluding silver nitrate and sodium carbonate. Again, remember to mark the well you place them in.

4. Place several drops of silver nitrate in each of the wells. Record your observations in your periodic table, by placing a triangle in each block where a reaction occurred.

5. Observe the demonstration performed by your teacher. Again, record in your periodic table.

# Lab # 10: Properties of Metals and Nonmetals

***Materials***:

|  |  |
| --- | --- |
| Conductivity tester | Wellplate |
| Forceps | Safety goggles |
| Hydrochloric acid | Sample elements |
| Blank periodic table | Copper chloride |
| Markers |  |

***Procedure***:

1. Dull or shiny?
	1. Observe each of the element samples provided. For each one, observe if it is shiny or dull. Remember that some elements will tarnish, so try to look past this when making your decision. As you did in the last lab, you will be keeping data on a periodic table. In the upper left corner, place a dot for each element that appears to be shiny and a star on each element that appears to be dull.
2. Malleable?
	1. Using tweezers or forceps, or your fingers, try bending each sample. Will it bend (is it malleable?) or does it snap or crumble into pieces, showing brittleness? Some of our samples are small and difficult to manipulate. Your teacher has larger samples of these elements that you can look at on the observation desk. Likewise, don’t be tricked by elements that are malleable (bendable) but are so strong you would need tools to help you. If a substance is able to bend w/o breaking it is called malleable. Draw a line in the lower left corner if the substance is malleable.
3. Conduct electricity?
	1. For this you will be using the conductivity testers. Be sure testing our elements we must make sure that the conductivity tester is working. Turn on the tester by moving the switch to on. Using your fingers, gently hold the two leads on the tester until they touch. At this time, both LED’s should light up. If they do not take the tester to your teacher for a replacement. Now, test the electrical conductivity of your sample by touching the two wire ends of your conductivity tester to the sample. Make sure that both leads are held against the sample. Note the following: Mark very bright lights with a triangle, dim lights (or one light) with a small square and no lights with a small X in the upper right hand corner of each block.
4. React with acid?
	1. Place a small sample of each element into a well in your well plate (each element should have it’s own well.) To each well add just enough dilute hydrochloric acid, to cover the solid; you do not need to fill the well. Look for any evidence of chemical reaction. If a reaction is observed, place a + sign in the lower right corner of each block.

***Clean Up*** – Shake your well plate into the garbage to remove the solid waste. Then clean with soapy water and set upside down to dry.

# Lab # 11: The Baggie lab

**Assignment:**

You will be given two baggies, one containing sodium bicarbonate, the other sodium chloride. Your job is to use your knowledge regarding the mole to determine which baggie contains more atoms of sodium. In the space below, write a plan for how you will accomplish this.

# Lab #12 : Classifying Chemical Reactions (from Flinn ChemTopics)

Concepts:

* Chemical reactions
* Single vs. double replacement
* Combination vs. decomposition
* Combustion reactions

Table 1. Symbols in Chemical Equations

|  |  |
| --- | --- |
| Symbol | Translation |
| 🡪 | Yields or produces (separates reactants from products |
| + | Reacts with or forms alongside (separates two or more reactants or products) |
| Δ | Reaction mixture is heated (written over the arrow) |
| NR | No reaction takes place when reactants are mixed |
| (s) | Pure substance (reactant or product) is a solid |
| (l) | Pure substance is a liquid |
| (g) | Pure substance is a gas |
| (aq) | Aqueous solution (reactant or product is dissolved in water) |

Pre-Lab Questions

1. Which reactants used in the experiment are flammable? Discuss the safety precautions that are necessary when working with flammable materials.
2. Summarize the following description of a chemical reaction in the form of a balanced chemical equation.

“When solid sodium bicarbonate is heated in a test tube, an invisible gas, carbon dioxide, is released into the surrounding air. Water condenses at the mouth of the test tube and a white solid residue, sodium carbonate, remains behind in the bottom of the test tube.”

1. Common observations of a chemical reaction are described in the Introduction section. For each observation, name a common or everyday occurrence that must involve a chemical reaction. Example: When a candle burns, it gives off light and heat. The production of light and heat is evidence for a chemical reaction.

Materials

|  |  |  |
| --- | --- | --- |
| Ammonium carbonate | Calcium carbonate | Copper(II) chloride solution |
| Ethyl alcohol | Hydrochloric acid | Magnesium ribbon |
| Phenolphthalein indicator | Sodium hydroxide solution | Sodium phosphate solution |
| Water, distilled | Zinc, mossy or zinc shot | Bunsen burner |
| Butane lighter | Evaporating dish | Forceps or tongs |
| Heat resistant pad | Litmus paper | Pipets |
| Spatula | Test tubes, small | Test tube clamp |
| Test tube rack | Wash bottle | Wood splints |

Procedure:

For each reaction, record the color and appearance of the reactant(s), the evidence for a chemical reaction, and the properties of the product(s) in the data table.

Reaction #1

1. Obtain a 3-4 cm strip of magnesium metal ribbon. Hold the piece of magnesium with forceps or crucible tongs and heat the metal in a laboratory burner flame. Caution: Do not look directly at the burning magnesium – ultraviolet light that is produced may damage your eyes.
2. When the magnesium ignites, remove it from the flame and hold it over an evaporating dish until the metal has burned completely. Let the product fall into the evaporation dish.
3. Turn off the laboratory burner and observe the properties of the product in the evaporating dish.
4. Record observations in the data table.

Reaction #2

1. Using a pipet, add about 40 drops of hydrochloric acid solution to a small test tube.
2. Obtain a 2-3 cm strip of magnesium metal ribbon and coil it loosely into a small ball. Add the magnesium metal to the acid in the test tube.
3. Carefully feel the sides of the test tube and observe the resulting chemical reaction for about 30 seconds.
4. While the reaction is still occurring, light a wood splint and quickly place the burning splint in the mouth of the test tube. Do not put the burning splint into the acid solution.
5. Record observations in the data table.

Reaction #3

1. Obtain a clean and dry test tube and place a small amount (about the size of a jelly bean) of ammonium carbonate into the test tube.
2. Use a test tube clamp to hold the test tube and gently heat the tube in a laboratory burner flame for about 30 seconds.
3. Remove the test tube from the flame and place a piece of moistened litmus paper in the mouth of the test tube. Identify any odor that is readily apparent by wafting the fumes toward your nose. Caution: do not sniff the test tube!
4. Test for the formation for a gas: Light a wood splint and insert the burning splint halfway down into the test tube.
5. Record observations in the data table.

Reaction #4

1. Place a small amount (jelly bean size) of calcium carbonate in a clean and dry test tube.
2. Using a pipet, add about 20 drops of hydrochloric acid to the test tube. Feel the sides of the test tube and observe the reaction for 30 seconds.
3. Quickly light a wood splint and insert the burning splint about halfway down into the test tube. Do not allow the burning splint to contact the reaction mixture.
4. Record observations in the data table.

Reactions #5

1. Add about 40 drops of copper II chloride solution into a small test tube.
2. Add several pieces (4-5) of zinc shot to the test tube and observe the resulting chemical reaction.
3. Record observations in the data table.

Reactions #6

1. Add about 40 drops of copper chloride solution into a small test tube.
2. Add about 25 drops of sodium phosphate solution to the test tube.
3. Record observations in the data table.

Reaction #7

1. Add 20 drops of sodium hydroxide into a small test tube.
2. Add one drop of phenolphthalein indicator (phth) to the test tube and mix the solution by gently swirling the tube. It might help to put a finger over the tip of the tube and shake gently. Hint: Phenolphthalein is called an “acid-base indicator”.
3. Add hydrochloric acid one drop at a time to the test tube. Count the number of drops of acid required for a permanent color change to be observed.
4. Record observations in the data table.

# Lab #13: The Bean Lab (from World of Chemistry)

**Background:**

The atomic masses found on the periodic table are relative masses determined by comparison to a standard. In the case of the periodic table, the standard used is the carbon-12 atom. The mass for every other atom was determined by comparing quantities of that element to quantities of carbon-12. The unit for relative atomic mass is the atomic mass unit or amu.

While the relative masses of each atom can be helpful in theoretical ways, an atomic mass unit is not a very useful day to day value. An atomic mass unit is a very small thing and, anyway, the ability of a scientist in the laboratory to isolate and weigh out just one atom is nonexistent, not to mention impractical.

What scientists needed was a way to relate a quantity of atoms to a mass measured in grams, a practical unit of mass. It was from this need that the concept of the “mole” was born.

A mole is defined as the **number of atoms found in the a sample (of an element) whose mass is equal to that element’s average atomic mass in grams.** A scientist named Amadeo Avogadro was the first to propose a valid number for the mole. Working with gases, Avogadro discovered that a sample of any element with a mass equal to its average atomic mass will contain 6.02 X 1023 atoms.

In this lab we will determine the relative mass for a series of beans so that you can see how relative mass might be calculated. We will then determine the number of beans in a “pot”, so that you can get a sense of what a **mole** represents.

**Procedure, Part 1: Determining the relative mass of each type of bean**

1. Obtain a plastic cup from your teacher. Place the plastic cup on the balance and zero the balance.
2. Obtain one type of bean from your teacher. Record the type of bean in your data table.
3. Determine the mass for 50 beans of this type. Do not use any broken or half beans in your measurement. Record the mass of 50 beans.
4. Repeat steps 2 and 3 for each type of bean until all five types have been measured. (Remember to zero the balance each time you measure a new type of bean in case the balance has been used by someone else in your absence.)
5. To calculate the relative mass we will use the lightest bean as our standard. Determine which bean was the lightest. Divide the mass (of 50 beans) for each type of bean by the mass for the standard (lightest) bean. This is the bean’s relative mass. (Note: If you are doing this properly, the relative mass for your standard will be 1). Record the relative mass of each bean in your data table.

**Procedure, Part 2: Determining the number of beans in a “pot” of beans.**

A pot of beans is a fictional **quantity** defined as the number of beans found in a sample whose weight is equal to the relative mass for that bean.

1. Obtain a plastic cup from your teacher and zero the balance as you did in Part 1.
2. Obtain one type of bean and record this type in a SECOND Data table.
3. Using one type of bean, place beans into the cup on at a time, counting the beans as you go, until the mass recorded on the balance is the same as the relative mass for that type of bean. (You may not be able to get exactly the same mass as your calculated relative mass. You may choose to be one bean over the mass or one bean under the mass but whichever you choose, stay consistent for each type of bean.)
4. Record the number of beans that you counted in your second data table.
5. Repeat steps 3 and 4 until you have measured all five types of beans. Record all values in your data table.

## Prelab Questions:

1. How are atomic mass units and grams similar? How are they different?
2. What makes a “relative” mass relative? In other words, what is it relative to?
3. For the relative atomic masses found on the periodic table, what is the standard they are calculated from?
4. In this lab you will be calculating the number of beans in a “pot”. What is a pot?
5. How will you determine the number of beans in a pot?

## Data Table(s):

# Lab #14: Determining an Empirical Formula

We have seen in an earlier lab (Law of Definite Proportions) that when elements combine to form compounds they always do so in a definite mass ratio. The reasoning behind this observation is that elements are made of atoms which, for any given element, are consistent in mass. In addition, atoms are indivisible, meaning that they cannot be split into pieces. Knowing that atoms combine in a definite fashion and that atoms are always found in whole pieces, not split in halves or thirds, we can represent the resulting compound with a formula showing a **whole** **number** **ratio** of the number of each type of atom. The simplest form of this ratio, with the lowest possible ratio, is called the empirical formula.

An experimentally determined mass ratio for a compound can be used to determine the compound’s empirical formula. Following the experiment, the mass for each element in the compound is converted into moles of that element. Because the number of atoms in a mole is always the same, the ratio of moles will reveal the ratio of atoms.

In this lab, you will experimentally determine an empirical formula by oxidizing tin (creating tin oxide) through a reaction with nitric acid. You will then calculate the mole ratio of the two elements (oxygen and tin) in one of the products and predict a formula for the resulting compound

**Objectives**: In this experiment, you will

* React a carefully determined amount of tin with excess nitric acid
* Form a tin-oxygen crystalline product
* Calculate the mole ratio of tin and oxygen in the crystalline product
* Predict an empirical formula for the tin oxygen product

Equipment:

|  |  |  |
| --- | --- | --- |
| Goggles and apron | Evaporating dish | Watch glass |
| Balance | Hot plate or burner with ring stand | Forceps |
| Beaker | Stirring rod |  |

Procedure

1. Safety goggles and lab apron must be worn at all times during this experiment. The smallest amount of nitric acid can cause permanent eye damage. You will be heating a mixture containing nitric acid. Popping and spitting may occur.
2. Wipe out an evaporating dish and watch glass with a dry paper towel.
3. Weight out about 2 g of 30 mesh granulated tin in a weight dish. Record this value in your data table.
4. Place the tin inside your evaporating dish, cover with the watch glass, and measure the mass of the dish, glass and tin together. Record this value in your data table.
5. Bring your evaporating dish to the hood along with a half sheet of paper you’ve written your name on. Place the dish in the hood on top of the piece of paper, making sure that you can still see your name. **CAUTION**: This next part of the lab must always be performed under a fume hood or in a well ventilated room. HNO3 causes burns; avoid skin and eye contact. Rinse spills with plenty of water.
6. The nitric acid has been placed in a labeled beaker to facilitate easy pouring. **Wearing goggles**, measure out 5 ml of 8M nitric acid, HNO3, into a small graduated cylinder. (This value does NOT need to be recorded in your data table). Remove the watch glass from your evaporating dish, add the acid and replace the watch glass, curved side facing down. Leave the evaporating dish in the hood until the reaction has stopped and no more bubbling or smoking can be seen. Remove your evaporating dish back to your lab station. Note: Your dish will feel warm to the touch.
7. Heat the dish on a hot plate or over a hot water bath as shown by your teacher. (An excessive amount of popping and spattering indicates you are heating too rapidly. If this occurs remove the flame until the spattering ceases. ) Continue to heat slowly until the contents are nearly dry.
8. When the popping and spattering no longer occur, remove the evaporating dish from the heat source. Remove the watch glass with your tongs, taking care not to lose any of the product (if any water droplets are seen on the inside of the evaporating dish allow them to roll into the dish, not on the counter). Set the watch glass aside but do not clean the watch glass until all the measurements in Step 10 have been made.
9. Break up the solid with a stirring rod. This is important because if there are any pockets of moisture still inside the solid they will pop when we begin our second heating and your reactants can go flying everywhere! If any of the mixture remains on your stirring rod afterward be sure to remove it by scraping along the inside edge of the evaporating dish. (The idea is to not lose any of your product.)
10. Place the dish (minus the lid) on the hot plate or position the dish on a clay triangle supported by a ring stand as shown. Heat carefully with a hot flame until the solid becomes pale yellow. Remove the dish from the heat source and allow to cool.
11. After the dish has cooled, replace the watch glass cover and measure the mass of the dish, contents and cover. Record this value in your data table. Reheat the dish for 2 to 3 minutes and allow it to cool again. Determine the mass of the dish, contents, and cover a second time. There should be very little difference between the two measurements. If this mass value does not agree within 0.02 g with the last mass reading, you must reheat and measure the mass until the last two measurements are within that range. When this happens, record your final mass in your data table.
12. Discard the solid material into the waste container designated by your teacher.

## Prelab Questions: answer prior to coming to lab

1. What is an empirical formula?
2. Explain why we will never see a formula such as As1.5O2.
3. When the experiment is finished you will be left with a light yellow powder in your evaporating dish. What is this powder?
4. In the lab I will be measuring my quantities of tin and oxygen in grams. In order to determine the empirical formula what unit must I convert my mass into?
5. In order to determine the empirical formula, I will need to know how much oxygen reacted. Yet, I never measure the amount of oxygen directly. At the beginning of the lab I measure my dish, cover and tin together. At the end of the lab I measure my dish, cover and product. How do I use these two values to determine the amount of oxygen that reacted?

## Data Table: Prepare this prior to coming to lab

# Lab #15 : Electrical Solutions from World of Chemistry

**Background**

 Many chemical reactions occur in solution. The behavior of chemicals in solution can be very different from that of the dry chemical. This is because of interactions (not necessarily reactions) between the water molecules and the chemical being dissolved. In this lab we will study how two types of compounds, ionic and covalent behave differently in water.

 Let me start by clearing up a common misunderstanding. Have you noticed that your bathroom appliances such as your electric hair dryer or electric razor caution you to never use them in the bathtub? Or that your toaster actually comes with a warning that it should not be submerged in water (wouldn’t that create soggy toast anyway?)? After all, everyone knows that water conducts electricity, right? Wrong. Water does not conduct electricity. Now before you decide to cut down on the time of your morning routine by drying your hair or fixing your breakfast while still in the shower, let me be very clear: **PURE** water does not conduct a current. Pure water consists of only H2O molecules and because H2O molecules are neutral, there is nothing there to carry a current. It is the **dissolved particles** in the water that conduct a current and only some of the dissolved particles at that. (Please realize that even if you were to bath in distilled (pure) water, the minute you step into the water it would become contaminated with all sorts of dissolved particles and carry a current quite well. Therefore, sadly, we must abandon the idea of fixing our hair or toast while bathing altogether.)

 To understand the next part you must remember that, covalent substances are compounds containing only nonmetals. For example sugar (C12H22O11), carbon dioxide (CO2) and rubbing alcohol (C2H5OH) are all covalent compounds. On the other hand there are ionic compounds. These contain metals and nonmetals. They might also contain a polyatomic ion. Examples are salt (NaCl) and sodium bicarbonate (NaHCO3).

 Now let’s talk about how substances dissolve. Imagine I put a sugar cube in a glass of water. You know from experience that the sugar will dissolve and mix with the water. But why does it do that? Why doesn’t it just sink to the bottom of the water and sit there?

 As we have discussed, the atoms and molecules which compose all matter in nature are in **constant motion**. The glass of water we are discussing may appear to be absolutely still; however the molecules inside the glass are moving and colliding with each other at a fairly rapid pace. **Collisions** between the water molecules and the surface of the sugar cube cause individual molecules of sugar to break free of the solid, then the **constant motion** of the water molecules carries the particles away. Even if you never stir the solution, with enough collisions, eventually the entire solid cube will dissolve.

 In a solid **covalent compound**, such as sugar, the individual particles are neutral molecules. When water collides with sugar, a complete neutral sugar molecule breaks off from the solid. This is different from **ionic compounds**. In an ionic compound, the individual particles are ions. Remember an ion is a charged particle. When water collides with solid salt, charged ions break off and float into solutions.

 This leads to a key difference in the two types of solutions. A solution containing a covalent compound will contain only neutral particles. Because there are only neutral particles, covalent compounds will not conduct a current. A solution containing an ionic compound will contain charged particles. These dissolved particles in ionic solutions **will** conduct a current. Solutions which conduct a current are called electrolytes.

 There is a small group of covalent compounds which “break the rules” and conduct a current. Acids are covalent compounds in which the hydrogen atoms will break apart causing ions. Some acids break apart completely just as an ionic substance would. This type of substance is called a **strong electrolyte** because it releases a great deal of ions into solution. Other acids only break apart slightly. These weak acids are called **weak electrolytes** because the small amount of ions breaking off results in a much weaker current.

In this lab you will observe the conductivity (or lack thereof) for both covalent and ionic solutions.

**Procedure**:

1. Half fill six beakers or cups with **distilled** water.
2. Add the following to each cup: (When filling make sure to avoid contamination by rinsing and drying spoons and graduated cylinders completely).
	1. Nothing
	2. Spoonful of sugar
	3. Spoonful of sodium chloride
	4. Spoonful of potassium nitrate
	5. 10 ml of 0.10 M HCl
	6. 10 ml of vinegar
3. Use the conductivity tester provided to test each cup. Between each solution be sure to rinse the conductivity tester completely with distilled water and then dry with paper towels. As you make your observations record them in your data table. Be as descriptive as possible in your observations indicating which solutions light both bulbs completely or not at all, or perhaps light up one bulb but not the other. Any variation in how bright or dim each bulb was should also be noted.
4. After testing all six solutions, choose what solution which DID cause the light bulbs to burn brightly. Pour a small amount of this solution (about 1 or 2 ml) into another cup and then fill that cup almost full with distilled water. Test this solution for conductivity and record your observations.

## Prelab: Complete prior to coming to lab.

1. What causes a solid to dissolve in water?
2. Explain why dissolved covalent substances will not conduct a current while dissolved ionic substances will.
3. What type of substance is the exception to the difference you describe in question 1? Why?
4. For each compound listed below, indicate whether you would expect it to be an electrolyte or a nonelectrolyte:
	1. KF
	2. C12H24
	3. (NH4)2S
	4. H2O
5. How does a strong electrolyte differ from a weak electrolyte? What type of substances are weak electrolytes?

## Data Table: Prepare prior to coming to lab.

# Lab #16: Reactions in Solutions I - Precipitation

## Prelab Questions: answer prior to coming to lab

1. What is a precipitate?
2. I could say that whether or not a precipitate forms is all about attractions. What would I mean by this?
3. If I were to mix Ag+ ions with Cl- ions, I would observe the formation of a white precipitate. What does this tell me about the silver and chloride ions?
4. The reaction between the sodium iodide and the lead nitrate could be illustrated with the equation: Pb(NO3)2 + NaI 🡪 PbI and NaNO3. In addition to being a precipitation reaction, what other type of classification does this reaction fall under (ie: synthesis, combustion etc.).
5. Write the formulas for the following ionic compounds
	1. Zinc sulfide
	2. Chromium (III) hydroxide
	3. Lead (II) phosphate

6. Write the ions that are present when each of the following compounds are placed into water.

* 1. Sodium chloride
	2. Copper (II) chloride
	3. Iron (III) sulfate

7. Write a molecular equation for that reaction that occurs when a solution of cobalt (II) nitrate reacts with aqueous potassium phosphate.

**Materials**

Well plate

Pipets

Wash bottle

Soap

Safety goggles

Lab apron

**Reagents**

Sodium iodide

Sodium carbonate

Sodium phosphate

Sodium sulfate

Sodium hydroxide

Cobalt (II) chloride

Copper (II) chloride

Aluminum chloride

Barium chloride

Nickel (II) chloride

**Procedure**

1. Use 4-5 drops of each reagent in each test. Note and record any sign of a reaction. Test all possible pair combinations. Record your observations in a data table similar to the one below.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | Cobalt (II) chloride | Copper (II) chloride | Aluminum chloride | Barium chloride | Nickel (II) chloride |
| Sodium Iodide |  |  |  |  |  |
| Sodium Carbonate |  |  |  |  |  |
| Sodium Phosphate |  |  |  |  |  |
| Sodium Sulfate |  |  |  |  |  |
| Sodium hydroxide |  |  |  |  |  |

# Lab #17: Activity of Metals/Constructing an Activity Series

**Materials**

**Metals Solutions Containing Metals**

Copper Copper chloride

Magnesium Magnesium nitrate

Tin Tin (II) chloride

Zinc Zinc nitrate

**Procedure**

**Hint: Some reactions occur more slowly than others. Give each possible reaction a little time before making your final observations. Be as complete in your observations as possible.**

1. Using label tape, label four droppers with the names (or formulas) for each solution. Fill each dropper with the appropriate solution.
2. Get three different pieces of each type of metal. Be sure to keep track of which metal is which.
3. Place a single piece of magnesium into each of three different well plates.
4. We are going to add zinc nitrate to one of these wells. Use just enough zinc nitrate to cover the piece of magnesium. Observe any changes. After a few minutes it might be helpful to remove the metal piece with your tweezers and rub it with a paper towel. Record your complete observations.
5. We will now add tin chloride to the next well with magnesium. Again, just enough to cover. Observe and record. (Again, remove the metal from solution if necessary.)
6. We will now add copper chloride to the final well containing magnesium. Just enough to cover. Observe and record.
7. You will repeat this procedure with each of the remaining metals, testing them with each of the solutions. The exception will be the solution containing the same metal. For example, magnesium was not tested with magnesium nitrate. Tin does not need to be tested with tin chloride and so forth.
8. When you are finished testing all four metals, use your tweezers to remove all metal pieces to the trash can. Wash and dry your well plate and clean the rest of your lab area.

## Prelab Questions: Answer prior to coming to lab

1. What is the purpose of the lab?
2. In this chapter we have been studying the permanent changes or driving forces that take place when a chemical reaction has occurred. In an earlier lab we saw that formation of a precipitate was one type of change. In a **redox** reaction, what type of change is occurring?
3. Consider the net ionic equation written below. Has the iron metal lost or gained electrons in this reaction? Would we say that it was oxidized or reduced?

Fe + 3 Cu1+ 🡪 3 Cu + Fe3+

1. A chunk of calcium is placed into a beaker containing sodium nitrate. A reaction is observed. Is calcium more reactive or less reactive than sodium?
2. Consider the following partial activity series. If I were to place a strip of chromium into a beaker containing a zinc solution, would a reaction occur? Explain your answer.

Zinc

Chromium

Iron

## Data Table: Construct prior to coming to lab

# Lab #18: Reactions in Solution – Acid and Base Lab

**Materials**

Well plates

pH indicator paper

stirring rod

conductivity tester

goggles

apron

* 1. M Acetic acid (HC2H3O2)

0.1M NH3 (a weak base)

phenolphthalein

universal indicator

litmus paper, red

litmus paper, blue

pH paper

0.1 M Hydrochloric Acid, HCl

0.1 M Sulfuric Acid, H2SO4

0.1 M sodium hydroxide, NaOH

Distilled water

Magnesium, zinc, and copper

**Important Safety Notes:**

1. Acids and bases are toxic. They are also corrosive to skin and clothing. Wipe up all spills with large volumes of water. If either an acid or a base gets on your skin or clothing, rinse the affected area thoroughly for 5 minutes and notify the teacher.
2. Wear safety goggles and a lab apron at all times in the laboratory.

**Procedure**:

***Property #1: Acids react with metals.***

1. Place a small amount of each metal in the bottom of a well. (Each metal should be in its own well.)
2. Add several drops of hydrochloric acid to each metal. Observe and record your reactions.
3. Repeat the procedure, placing each metal in a new well. This time test each metal with the sulfuric acid. Observe and record your reactions.
4. Repeat the procedure, placing each metal in a new well. This time test each metal with the acetic acid (vinegar).
5. Finally, repeat one more time, placing each metal in a new well. This time test each metal with the sodium hydroxide.
6. Remove any remaining metal with your forceps/tweezers and place in the trash. Clean your well plate and rinse with distilled water. Dry completely.

***Property #2: Acids and bases are electrolytes.***

1. Place 8-10 drops of each of the following in your well plate. (Each solution should be placed in its own well.
	1. Hydrochloric acid
	2. Sulfuric acid
	3. Acetic acid
	4. Sodium hydroxide
	5. Ammonia
2. Test each well using the conductivity tester. IMPORTANT: Be sure to rinse the conductivity tester with distilled water and dry between each well. Observe and record your observations.

***Property #3: Acids and Bases will affect indicator molecules.***

1. Place 8-10 drops of hydrochloric acid into two wells. In one well place 1 drop of phenolphthalein. In the second well place 1 drop of universal indicator. Observe and record your results.
2. Place 8-10 drops of acetic acid into two wells. In one well place 1 drop of phenolphthalein. In the second well place 1 drop of universal indicator. Observe and record your results.
3. Place 8-10 drops of sodium hydroxide into two wells. In one well place 1 drop of phenolphthalein. In the second well place 1 drop of universal indicator. Observe and record your results.
4. Finally, place 8-10 drops of ammonia into two wells. In one well place 1 drop of phenolphthalein. In the second well place 1 drop of universal indicator. Observe and record your results.
5. Now we will test indicator papers. Place 8-10 drops of each of the following solutions into your well plate (each solution should have its own well).
	1. Hydrochloric acid
	2. Acetic acid
	3. Sodium hydroxide
	4. Ammonia
6. Tear off a small piece of blue litmus paper. Dip one end of the blue litmus paper into the hydrochloric acid and immediately remove it. Record any changes that occur. Then dip new pieces of blue litmus paper into each of the remaining solutions. Record all observations.
7. Repeat step 14 using the red litmus paper. Observe and record.
8. Repeat step 15 using the pH (pHydrion) paper. Observe and record.

***Property #4 Acids and Bases will react to neutralize each other.***

1. Carefully place 10 drops of hydrochloric acid into a well of your well plate. Because the hydrochloric acid and the sodium hydroxide are the same concentration, an equal amount of sodium hydroxide should neutralize our acid. Add 10 drops of sodium hydroxide to your acid. Can you see the neutralization reaction that occurs? Record your results.
2. Now place 10 more drops of hydrochloric acid into a new well in your well plate. To this acid add 1 drop of phenolphthalein. Now add 10 drops of sodium hydroxide to your acid. What change do you observe? (If at first you don’t observe a change, add a few more drops of sodium hydroxide). Record your observations.

Clean up by disposing of solid materials in trash container and washing liquid material in your well plate down the drain. Clean well plates with AMPLE amounts of soap and water. Dry completely.

**Prelab: Complete this before coming to Lab**

1. List the four properties that we will be testing in lab.
2. When acids react with metals, bubbling is always observed. What gas is being formed in this reaction?
3. Explain why acids and bases are electrolytes.
4. What difference would you expect when testing the conductivity of hydrochloric acid versus vinegar? Why?
5. What is an indicator? List two indicators that we will be using in lab?
6. What product (s) is/are formed in neutralization reaction?
7. Explain why a neutralization reaction is called a neutralization reaction.

**Data Table**

# Lab #19: Unknown solutions lab

In this lab you will receive the following nine solutions in numbered bottles. Your job will be to determine which bottle is filled with which solution.

|  |  |
| --- | --- |
| **Compound** | **Compound** |
| Table Sugar | Acetic Acid |
| Sodium Hydroxide | Sulfuric Acid |
| Potassium Nitrate | Copper Nitrate |
| Lead Nitrate | Sodium Chloride |
| Aluminum Nitrate |

You will have the following materials available for the task:

The nine solutions

Phenolphthalein

Conductivity tester

Well plate

Plastic pipettes

Procedure: Obtain nine pipettes and a strip of numbers from your teacher. Using label tape label a pipette with each number. Then bring the pipettes to your teacher for filling. Be sure to record your nine numbers on this page for future reference.

## Prelab:

The following questions are intended to provide hints on how to identify your unknown solutions.

1. Write the formulas for each compound listed above. (Exception: the formula for sugar is C12H22O11)
2. What types of substances can a conductivity tester distinguish between?
3. What does phenolphthalein help to identify? Explain how this works. Hint: you can identify more than one thing with phenolphthalein.
4. Write equations for each of the following reactions and indicate (circle) any precipitates that occur:
	1. Lead nitrate and sodium chloride
	2. Aluminum nitrate and sodium chloride
	3. Aluminum nitrate and potassium nitrate
	4. Sulfuric acid and lead nitrate
	5. Sulfuric acid and aluminum nitrate
	6. Sodium hydroxide and aluminum nitrate
	7. Sodium hydroxide and lead nitrate
	8. Sodium hydroxide and copper nitrate
	9. Sodium hydroxide and potassium nitrate

## Data Table: In this lab, the data table does not need to be prepared in advance. Instead use your data table to keep track of your work as you go. Whatever tests you run need to be recorded on this sheet of paper. Your records should be neat enough for me to tell what you are testing and what your results are.

# Lab # 20: MicroMole Rocket Lab

**Purpose**: to investigate the optimum ratio of hydrogen to water when water is created.

**Background**: In this lab we will generate hydrogen gas and oxygen gas through separate reactions. We will then mix the generated hydrogen and oxygen gas to produce water. The formation of water is an exothermic process, releasing energy in several forms including sound. Our objective is to find the combination of hydrogen and oxygen that makes the most sound. We will then test our optimum ratio to see how far our “rockets” will fly.

**Prelab:**

1. Write an equation for the reaction that occurs when zinc is mixed with hydrochloric acid.
2. Write an equation for the reaction that occurs when hydrogen peroxide decomposes.

**Procedure**:

**Filling the Bulb**

1. Your basket contains a pipet bulb which has been calibrated. Take this bulb to the water bin in your sink and fill it completely with water. It is important that no bubbles remain in the bulb.

**Generating Gases**

1. Your teacher will show you the appropriate set up for the gas collector tubes.
2. You will be generating two gases in two separate test tubes, as follows:
	1. In one test tube place several pieces of zinc. Then fill this test tube with HCl until the liquid level is about 2 cm below the mouth of the test tube. Allow the test tube to sit in the empty beaker. ***Write the gas generated in this reaction in the cloud to the left.***
	2. In the second test tube, you will be using yeast to decompose hydrogen peroxide. Fill your test tube about ¼ full with yeast suspension. Add hydrogen peroxide until the liquid level is about 2 cm below the mouth of the test tube. ***Write the gas generated in this reaction in the cloud to the left.***
	3. Place a gas collecting stopper on each of your test tubes.

**Making Mixtures**

1. **Mixture One**: We will now be mixing oxygen and hydrogen in the same bulb. For our first mixture let’s use 1:5 of oxygen and hydrogen. Go to the sink and fill your bulb with water.
	1. Place the bulb on the oxygen gas generator and allow the bulb to fill with gas until the bulb 1 segment is full with oxygen and the rest of the bulb is still water. Remove the bulb being careful to hold the bulb with the opening straight down so that the gas cannot escape.
	2. Immediately place the bulb on the hydrogen generator and collect hydrogen gas until the bulb is completely empty of water. Remove the bulb, cover the opening with your finger as before and invert the bulb, keeping the opening covered.
	3. Light a match and test flammability as you did earlier for hydrogen alone.
	4. ***Record your observations in the margin to the left.***
2. **Mixture 2**: This time we will put go for a 2:4 mixture of oxygen to hydrogen. Fill a bulb completely with water. Place it over the oxygen generator and allow oxygen gas to fill 2 segments of your bulb. Remove the bulb, being careful not to let the gas escape. Then place the bulb on the hydrogen gas generator. Fill the bulb completely with hydrogen gas. Test flammability as you did before.
3. You will continue to work with mixtures of oxygen. Remember to always put oxygen in your bulb first. By now your realize that mixtures of oxygen and hydrogen produce a decent sized “pop” Your objective is to find the ratio of hydrogen to oxygen that gives you the best pop. Keep track of your mixtures in the space below.

# Lab #21: Copper and Silver Lab

**Background**

Earlier in this year you learned about the activity of metals and that more reactive metals will transfer their electrons to less reactive metals. You observed this from a qualitative perspective in the Activity Series Lab. In this lab we will take a quantitative perspective, measuring the amount of silver produced and then using **stoichiometry** to determine how much copper was consumed.

The metals we will be using in this lab are copper and silver. We will introduce a copper strip into a solution containing silver nitrate. Because copper is more reactive, it will transfer its electrons to silver. In the process copper will dissolve into solution and silver will precipitate out producing solid silver for us to weigh.

**Procedure:**

1. Obtain a clean and dry 50 ml beaker. Label the beaker with your name.
2. Tare the balance with the beaker on the balance pan and then carefully add 1.4-1.6 grams of silver nitrate to the beaker. Caution: Use a spatula to transfer the solid – do not touch the silver nitrate. Carefully clean up any silver nitrate spills in the balance area or on the bench top.
3. Measure and record the exact mass of silver nitrate in your data table.
4. Fill the beaker with approximately 30 ml of distilled water and stir the solution with a stirring rod until the entire solid has dissolved. Rinse the stirring rod with distilled water. (In this lab your quantities of distilled water do NOT need to be recorded in your data table).
5. Cut a 25 cm piece of copper wire and loosely coil it as shown by your teacher. Cut and coiled pieces may already be provided.
6. Find the initial mass of the copper wire and record it in your data table.
7. Use a wood splint to suspend the copper wire in the silver nitrate solution as demonstrated. The copper wire should not be touching the bottom or sides of the beaker.
8. Carefully add 3 drops of 3M nitric acid to the beaker. Do NOT stir the solution.
9. Allow the beaker to sit undisturbed on the lab bench for 15 minutes. Try not to jostle or shake the suspended copper wire during this time.
10. Observe any signs of a chemical reaction occurring in the beaker and record all observations in the data table.
11. **While the reaction is taking place, label a 100 ml beaker with your name and class/period. Measure and record the mass of this labeled beaker in the data table.**
12. After 15 minutes, gently lift the wooden splint to remove the copper wire from the solution. Be careful not to lose the silver that has accumulated on the copper wire. (If any silver does fall into the 50 ml beaker use your stirring rod to move this silver into the 100 ml beaker.
13. Holding the wire with the wooden splint, place the copper wire above the clean, 100 ml beaker. Rinse the wire with a steady stream of distilled water from the wash bottle. The silver crystals should easily fall of the wire into the beaker. Gently shake the wire and rinse with water until no more silver adheres to the wire. Note: Use a total of about 40 ml of distilled water.
14. When all of the silver has been removed, lift the copper wire out of the beaker and rinse with acetone. The acetone will clean the wire surface and allow it to dry more quickly.
15. Place your copper wire on a piece of paper labeled with your name and then turn your copper wire into your teacher as instructed.
16. Examine the beaker containing the silver product in your 100 ml beaker. Most of the silver should settle into a dense mass at the bottom of the beaker. Carefully decant (pour) the liquid into a waste flask (125 ml Erlenmeyer) to remove most of the water. Note: Try not to lose any of the solid in the process.
17. Rinse the solid with 5-10 ml of distilled water from a wash bottle. Decant the rinse water into the waste flask as well.
18. Repeat the rinsing/decanting cycle with a second portion of distilled water.
19. Dispose of the waste solution as directed by your instructor.
20. When all of the liquid has been decanted, take the 100 ml beaker containing your silver , labelled with your name, to your teacher for drying.
21. Allow the solid to dry overnight.

Day Two

1. Retrieve your dried silver from your teacher. (DO NOT PICK UP YOUR COPPER WIRE!) Weigh your beaker of silver and record your total mass in your data table.
2. Go to the lab writeup for further instructions.

## Prelab: Complete prior to coming to lab

1. Write a balanced equation for the reaction between silver nitrate and copper producing silver and copper (II) nitrate.
2. In this experiment what two substances are we trying to relate stoichiometrically?
3. What is the mole to mole ratio of copper to silver in this lab?
4. Imagine that we have complete the experiment and that 5.4 grams of silver were produced? Use stoichiometry to determine how many grams of copper were consumed in this reaction.

## Data Table: Prepare prior to coming to lab

# Lab #22: Limiting Reactants – Sodium Bicarbonate and Vinegar

**Materials Needed**

8 tablets Alka Seltzer

Acetic acid

2 250 ml beakers

10 ml graduated cylinder

50 ml graduated cylinder

Glass stirring rod

**Procedure:**

1. Measure 30 ml of water into a 250 ml beaker. Measure out 5 ml of acetic acid and add to the beaker. Weigh the beaker with water and acid and record the mass.
2. Weigh and record the mass of an Alka Seltzer tablet.
3. Add the mass of the Alka Seltzer to the mass of the beaker (and liquid). Record this in your data table as the total mass of reactants.
4. Place the Alka Seltzer tablet into the water in the 250 ml beaker.
5. When the bubbling has ceased, stir the solution to ensure complete dissolution of the tablet and to remove any visible bubbles of carbon dioxide. Gently tap the stirring rod against the inside of the beaker to remove any drops of liquid before weighing the beaker and its content.
6. Weigh the beaker and its contents. Record this mass as the “mass of products minus the carbon dioxide”
7. Dispose of the solution down the drain. Wash and dry the beaker.
8. You will now repeat the experiment 4 times, each time increase the amount of acetic acid by 5 ml and decrease the amount of water by 5 ml as shown in the table below. Remember to record all information in your data table and to clean and dry your beaker between trials.

|  |  |  |
| --- | --- | --- |
| **Trial #** | **Amount of Water** | **Amount of Acetic Acid** |
| 1 | 30 | 5 |
| 2 | 25 | 10 |
| 3 | 20 | 15 |
| 4 | 15 | 20 |
| 5 | 10 | 25 |

## Prelab:

1. What is a limiting reactant?
2. In this reaction what reactant does the Alka Seltzer contain?
3. What gas is produced in this reaction?
4. One of the measurements in this lab will be the “total mass of reactants” before the reaction. Take a look at the total amounts of liquid used in each step of the procedure. Considering the fact that you are using 1 tablet of Alka Seltzer each time, what should you expect to see when looking at the “total mass of reactants” for each trial?
5. Read over the lab procedure, look at the measurements you will be taking and describe how you will be determining the amount of carbon dioxide produced in each trial.
6. If the amount of carbon dioxide does not change, even though the amount of vinegar is increasing, which of your reactants is the limiting reactant? Explain your answer.

## Data Table: This data table has been prepared for you.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Trial** | **Mass of Alka Seltzer tablet** | **Mass of liquids & beaker before reaction** | **Total Mass before reaction** | **Total Mass after reaction (minus the carbon dioxide)** | **Mass of Carbon Dioxide which was produced.** |
| **1** |  |  |  |  |  |
| **2** |  |  |  |  |  |
| **3** |  |  |  |  |  |
| **4** |  |  |  |  |  |
| **5** |  |  |  |  |  |

# Lab #23: Mass & Mole Relationships in a Chemical Reaction

**Background:** Due to the widespread use of sodium bicarbonate (commonly called baking soda) in many food products, the thermal decomposition reaction has been studied extensively by food chemists. Baking soda is used to prepare cakes in order to insure that cakes “rise” as they bake.

As the temperature of the cake batter reaches approximately 50 °C, the baking soda decomposes and carbon dioxide is released. The use of baking soda is especially popular in pancakes and waffles since the high cooking temperatures of 350-450°C (175-230°C) cause the carbon dioxide to be liberated before the dough has set. Thus, the batter rises before it sets, and we get a light and tasty finished product.

**Possible Decomposition Reactions:**

Equation 1: Sodium bicarbonate 🡪 sodium hydroxide (s) + carbon dioxide (g)

Equation 2: Sodium bicarbonate 🡪 sodium oxide (s) + carbon dioxide (g) + water (g)

Equation 3: Sodium bicarbonate 🡪 sodium carbonate (s) + carbon dioxide (g) + water (g)

**Objective**: The goal of this lab if for you to experimentally determine which of these three reactions is correct. You will need to eliminate reactions that do not fit your experimental results when you conduct the thermal decomposition of baking soda.

**Prelab:**

1. In the space provided above, write a balanced equation for each of the possible decomposition reactions.

**Materials**: Ring Stand, Iron Ring, Evaporating Dish, Clay triangle and Bunsen burner.

**Chemicals**: Sodium bicarbonate,

**Procedure**:

As with many of our experiments, the accuracy of mass measurements is of the utmost importance in order to collect useful data.

1. Determine the mass of a clean, dry evaporating dish and watch glass. Record the mass in the data table.
2. Add approximately 15 g of sodium bicarbonate to the evaporating dish. Record the exact mass of the sodium bicarbonate in the data table.
3. Set up your ring stand as follows: Attach the ring to the ring-stand. Place the clay triangle on the ring and Bunsen burner beneath it. Put the evaporating dish on the clay triangle.
4. GENTLY heat the evaporating dish and its contents with a small flame for 10 minutes. Use a spatula to carefully break up any clumps that form as the heating proceeds. The clumps only have to be broken up once during the heating process.
5. Turn off the burner and allow the apparatus to cool for about 5 minutes. When it is cool, determine the mass of the entire assembly. Record this mass in the data table.
6. Reheat the substance for another 5 minutes and again allow to cool.
7. Measure the mass of the assembly again. Record the final temperature in your data table. If the two masses are within 0.02 grams you are finished. If not repeat the heating process for a third time.
8. Rinse the residue down the sink. Wash your hands and clean up your area before leaving the lab.

|  |
| --- |
| **Data Table** |
| **Item** | **Mass (grams)** |
| Empty Evaporating Dish  |  |
| Evaporating Dish with Sodium bicarbonate |  |
| Evaporating Dish and Product (after heating and cooling first time) |  |
| Evaporating Dish and Product (after heating and cooling second time) |  |
| Evaporating Dish and Product (after heating and cooling third time, if necessary) |  |

# Lab #24: Heat of Solution and Heat of Reaction

A note: the reactants in this lab exercise are extremely corrosive and can be harmful to tissue and clothing. Be careful to keep your goggles on at all times and consider wearing an apron to protect your clothing.

Materials:

Foam coffee cups

Sodium hydroxide pellets

2 M hydrochloric acid

2M NaOH

thermometer

Procedure, Part I:

1. Obtain 4 foam cups and label them cup 1, 2, and 3
2. Measure 100 ml of cool tap water with your graduated cylinder and pour into coffee cup #1.
3. Wait several minutes for the tap water to reach room temperature and then take the temperature of the water. Record in your data table.
4. Obtain approximately 4 grams of sodium hydroxide pellets. Determine the exact mass of your sample and record in your data table. NOTE: DO NOT handle the sodium hydroxide pellets with your fingers.
5. Pour the solid NaOH into coffee cup #1 with the water.
6. Place your thermometer into the solution and swirl gently to mix the solution. You will notice that the temperature of the solution is increasing.
7. When the temperature of the solution reaches its highest point it will begin to fall again. When this occurs, record the highest temperature of the solution.
8. The solution that you have created is very caustic and cannot be poured down the drain. Please dispose of the solution into the waste container provided and then rinse your coffee cup well and return it to the supply station.

Part II:

1. Measure 50 ml of 2.0 M HCl solution with the graduated cylinder and place it into cup #2.
2. Take the temperature of this solution.
3. Now measure out 50 ml of 2.0 M NaOH and place it into cup #3.
4. Pour the cup containing the NaOH (cup #3) into the cup containing the HCl (cup #2)
5. As you did in part 1, use your thermometer to swirl your solution/reaction. Again wait until the temperature begins to fall and then record the highest temperature reached.
6. This solution can disposed of down the drain with lots of water.

**Prelab Questions**

1. Some of the changes in a chemical reaction absorb (require) energy, while others release energy. What type of change would require an input of energy?
2. What type of change would involve a release of energy?
3. When the overall energy change for a reaction shows an excess of energy, is this reaction exothermic or endothermic. Explain.
4. We are monitoring the temperature change that occurs in a solution when a reaction occurs. If the temperature of the solution increases (in other words, the water warms up), is this reaction endothermic or exothermic? Explain.
5. Based on what you have learned in class, in a thermochemical equation, would the energy value show a positive or negative sign for an endothermic reaction?

## Data Table

# Lab #25: Heat of Reaction and Hess’s Law Lab

Materials

|  |  |
| --- | --- |
| Hydrochloric acid (1M), 60 ml | Magnesium ribbon, Mg, 7-cm strip |
| Magnesium oxide, MgO, 0.4 g | Balance |
| Foam Calorimeter | Digital thermometer |
| Forceps | Graduated cylinder |
| Metric ruler | Scissors |
| Spatula | Stirring rod |
| Wash bottle with distilled water | Weighing dish |

Safety Precautions

Hydrochloric acid is toxic by ingestion and inhalation and is corrosive to skin and eyes. Magnesium metal is flammable solid. Keep away from flames. Do not handle magnesium metal with bare hands. Wear chemical splash goggles and chemical-resistant gloves and apron. Wash hands thoroughly with soap and water before leaving the lab.

**Procedure**

*Record all data for parts A and B in a data table*.

Part A: Reaction of magnesium with hydrochloric acid

1. Obtain a 3 to 4 cm strip of magnesium ribbon. Obtain the mass for your strip and record it in your data table.
2. Mass a clean, dry foam calorimeter to the nearest 0.01 g. Record.
3. Using a graduated cylinder, add 15 mL of 1 M hydrochloric acid to the calorimeter and measure the combined mass of the calorimeter and acid.
4. Using a digital thermometer or a temperature sensor, measure the initial temperature of the hydrochloric acid solution. Record.
5. Add the piece of magnesium ribbon to the acid and stir the solution until the magnesium has dissolved and the temperature of the solution remains constant.
6. Record the final temperature of the solution to the nearest 0.1 ºC.
7. Rinse the contents of the calorimeter down the drain with excess water.
8. Dry the calorimeter and mass it again to the nearest 0.01 g.
9. Repeat steps 5-9 using the second (larger) piece of magnesium ribbon.

Part B. Reaction of Magnesium Oxide with Hydrochloric Acid

1. Mass a clean, dry calorimeter to the nearest 0.01 g.
2. Using a graduated cylinder, add 15 ml of 1 M HCl to the calorimeter and measure the combined mass of the calorimeter and hydrochloric acid.
3. Using a small weighing dish, add about 0.20 g of magnesium oxide. Record the exact mass of magnesium oxide to the nearest 0.01 g.
4. Using a digital thermometer, measure the initial temperature of the hydrochloric acid solution to the nearest 0.1 ºC.
5. Using a spatula, add the magnesium oxide to the acid. Stir the reaction mixture with the spatula until all the magnesium oxide dissolves. Watch the temperature until it remains constant for about 30 seconds or starts to decrease. Record the highest temperature of the solution.
6. Pour the reaction mixture down the drain with excess water. Rinse and dry the calorimeter.

## Prelab Questions:

1. What reaction are we interested in studying in this lab?
2. Explain why we can’t just measure the heat from this reaction directly?
3. What is a state function?
4. If we can’t measure the heat from a reaction directly, what does Hess’ Law say we can do instead?
5. Give the equations for the two reactions we will be performing in this lab.

## Data Table (s)

# Lab #26: Differences in Ionic and Covalent Bonding

**Objectives:** In this lab, you will make observations about melting point, solubility and conductivity to observe which substances have similar properties. This data will then be used to predict the type of bonding in each substance.

**Safety:**  Safety goggles and aprons must be worn at all times during the lab. Do not touch any chemicals to your skin. When using an open flame, be sure to tie back your hair as well as any loose clothing. In the event of an accident, call the teacher immediately. **Cyclohexane is flammable. Do not use this substance while anyone in the room is using an open flame!**

**Chemicals we will be testing**:

 Sodium chloride p-Dichlorobenzene

 Calcium chloride Potassium iodide

 Sucrose Benzoic Acid

**Materials needed:**

 Cyclohexane Conductivity tester

 24-well microplates Candle

 Aluminum pie pan Plastic pipets

 Microscope Hand Lens

 Iron ring Ring stand

**Procedure:**

1. Before beginning, wash a microplate and pie pan with soapy water. Dry it thoroughly.
2. Use a microscope or hand lens to observe 1 or 2 crystals of each solid. Write a brief description of each on your data table. (e.g. large or small crystalline, fine grained or coarse grained powder) Make sure you include enough detail to tell them apart later, if necessary. As you remove each chemical from the jar, take note of whether or not the chemical has an odor. Record this on the data table.
3. Place a pie pan on an iron ring attached to a ring stand. Position the ring so that it is just above the tip of the candle flame. Light the candle for a moment to check that you have the correct height.
4. Place a few crystals of each substance in separate locations on the pan. Do not allow the samples of the crystals to touch each other.
5. For this experiment, it is not necessary to have the exact melting temperature for each substance. We are only interested in comparing each substance to the others, so the order of melting will be recorded. Arrange the pan so that one substance is directly over the candle flame. Count the number of seconds until you notice the first sign of melting. Record the number of seconds. Repeat this process with the other five substances.
6. If after 3 minutes a substance has not begun to melt, write an N in the data table, indicating that it did not melt. Extinguish the candle flame and allow the pan to cool while you complete the remainder of the experiment.
7. Place a few (one or two) crystals of each of the white solids in the top row of your well plate. Repeat with the second row. Add 10 drops of distilled water to each well in the top row. Wait 3-5 minutes and look to see which substances dissolved (this is called solubility). Next, test the conductivity of each solution in the top row by dipping both electrodes of a conductivity meter into each well of the well plate. Be sure to rinse the electrodes with distilled water and dry them with a paper towel after each test. There is a “key” on the back of the conductivity meter, which explains how to record the results. Record the results in the data table.
8. For this next part we will go to the hood. Remember, cyclohexane is flammable and we don’t want to expose it to any possible flame. In the hood, add 10 drops of cyclohexane to each well in the second row of the well plate. As we did with the water, wait several minutes and then check to see which substances dissolved. Record the solubility of each substance in the data table. Also - test the conductivity of each solution in the bottom row by dipping both electrodes of a conductivity meter into each well of the microtitration plate. Again, be sure to rinse the electrodes with cyclohexane and dry them with a paper towel after each test. Record the results in the data table.
9. Clean the microplate with soapy water. Rinse and dry. If any wells are difficult to clean, use a small test tube brush. When you are finished, put away all equipment, clean your area and wash your hands before leaving the lab area.

## Prelab Questions:

1. Explain, in your own words, how ionic compounds form.
2. Now explain, in your own words, how covalent compounds form.
3. List four ways in which ionic substances can be distinguished from covalent substances in the lab. Be specific in your explanation by explaining how each types of compound will behave. (In other words, don’t just list four ways).

## Data Table:

# Lab #27: The Gas Labs – Investigating Properties Which Affect Pressure

In the next two days we will carry out three labs to investigate the affects of temperature, volume and amount on the pressure exerted by a gas. This will be divided over two days. You will do one prelab and one lab writeup for all three days. You will need to prepare three (3) data tables, one for each part of the experiment.

**Part 1: Temperature and Pressure**

**Procedure:**

1. Obtain and wear goggles.
2. You are going to be working with four ranges of temperature:
	1. Boiling water bath
	2. Hot water bath
	3. Room temperature bath
	4. Ice water bath
3. First we must set up our flask/pressure sensor equipment and our CBL. You will be using the gas pressure sensors as illustrated by your teacher. For this lab you will also use a temperature sensor. Prepare the Gas Pressure Sensor and sample of gas (air) for data collection according to the following instructions:
	1. Obtain a rubber-stopper assembly with a piece of plastic tubing connected to one of its two valves. Attach the connector at the free end to the Gas Pressure Sensor. At this time, be sure that the two way valve open. (Blue valve stem should be parallel to the valve.)
	2. Insert the rubber stopper assembly into a 125 ml Erlenmeyer flask. **IMPORTANT: Twist the stopper into the neck of the flask to ensure a tight fit.**
	3. Now close the two-way valve above the rubber stopper- do this by turning the blue valve handle so it is perpendicular with the valve stem. The air sample to be studied is now confined in the flask.
4. Now, prepare the calculator and CBL for data collection.
	1. Place the calculator into the CBL. Connect with appropriate wire.
	2. Plug the Gas Pressure Sensor into Channel 1
	3. Plug the Temperature Probe into Channel 2
	4. Plug in the CBL unit and turn on the calculator.
	5. Press the APPS button. Choose EASY DATA or DATA MATE
	6. The calculator should register in Channel 1: Gas Pressure (kPa) and Channel 2: Temperature (C)
5. We will first test our flask in ice water.
	1. Prepare an ice water bath by filling about 300 ml of tap water into a 600 ml beaker. Then add a handful of ice.
	2. Place the flask into the ice-water bath. Make sure the entire flask is covered.
	3. Place the temperature probe directly into the ice water bath.
	4. Give the flask several minutes so that the gas inside can reach the same temperature as the ice water bath.
	5. Then, when the readings on the CBL are stabilized, record the pressure and temperature in your data table.
6. We will now test our flask in a room temperature bath. Your teacher has prepared a flask of water which has been sitting out so that it may reach room temperature. Fill your 600 ml beaker with about 400 ml of this water. As you did in Step 5, immerse the flask, giving the air several minutes to adjust and then measure the temperature and pressure. Record your data in a data table.
7. We will now test our flask in a hot water bath. Using your sink, fill your 600 ml beaker with approximately 400 ml of hot tap water. If your sink is not cooperating because we have run out of hot water you may use a mixture of the boiling water with ice to get hot water. Once you have hot water, immerse the flask as you did in Step 5, giving the air several minutes to adjust and then measure the temperature and pressure. Record your data in a data table.
8. Finally we will test our flask in a boiling water bath. Your teacher has prepared a boiling water bath on a hot plate. Special caution must be taken to prevent burns to yourself and your classmates. In addition – the tubes and wires connecting your sensors will melt if they come into contact with the hot plate. Take extra care to hold the wires away from the heating unit. Now, as you did in Step 5, immerse the flask, giving the air several minutes to adjust and then measure the temperature and pressure. (Beware: if the air in the flask expands enough, the stopper on your flask may pop!) Record your data.

 **Part 2: Volume and Pressure**

**Procedure**

1. You will be working with an air filled syringe. The syringe will hold the air sample we will be testing. (You will not be using the flask in this part).
2. Pull the plunger of the syringe out until it reads 20 ml; then attach the tip of the syringe to the pressure gauge as illustrated by your teacher and make sure it is on tight. (The tubing for the sensor is threaded, as is the syringe tip. Be sure to screw the two pieces together.)
3. Plug your CBL into the wall
4. Plug the gas sensor into Channel 1
5. Turn on the calculator and press the APPS button.
6. Choose DATAMATE or EASY DATA
7. The calculator should read that you have the gas pressure senor in channel 1 and should be reading a pressure in kPa.
8. WITHOUT moving the plunger, record the volume and the pressure on the data sheet.

*NOTE: The volume that is in the syringe is not the total volume. You need to make sure to add 0.8mL to your syringe readings. For example: If you are at the 5.0 ml mark on the syringe, the actual volume is 5.8ml, which is what should be recorded on your data table.*

1. We are now going to decrease the volume by a series of movements. Move the syringe in by approximately 2 ml. Record your new volume as accurately as possible (remembering to estimate one place). Also record the new pressure at this volume.
2. Continue changing the volume in increments of approximately 2 ml. You should be able to end up with a total of 7 readings.

**Part 3: Moles and pressure**

In this part of the lab we are going to carry out a reaction between sodium bicarbonate and an excess of vinegar which produces a gas. In each step of the reaction we will increase the amount of sodium bicarbonate and as a result, increase the amount of gas produced. This way we can study the effect that greater amounts of gas will have on pressure.

**Procedure:**

1. Set up your CBL system. This time we will be using both the flask and the syringe.
	1. Plug the CBL into the wall.
	2. Plug the pressure sensor into Channel 1
	3. Plug the plastic tubing that is attached to one end of the rubber stopper into the pressure sensor.
	4. The syringe will eventually be placed on the stopcock. (The stopcock is the piece with the blue valve.) For now, the valve should be open.
	5. Turn on your calculator. Press the APPS button.
	6. Choose DATAMATE or EASYDATA. It should be reading the pressure sensor in channel 1.
2. On a piece of weighing paper, measure out approximately 0.05 g of sodium bicarbonate. Try to be as close to 0.05 as possible. Record the mass and then empty it into a 125ml Erlenmeyer flask.
3. Fill the syringe to the 5 ml mark with acetic acid. Screw the syringe tip snugly into the stopcock on the rubber stopper. (Yes this is an unfortunate choice of wording!)
4. Make sure that valve is open and carefully add the vinegar to the flask.
5. You must be careful to not release the syringe plunger or gas will leak into the syringe and your measurements will be meaningless. When all the vinegar is emptied from the syringe, keep holding the plunger down (this will take some pressure as the gas is pushing back) until the valve on the stopper can be closed. Close the valve and you may then let go of the plunger.
6. The vinegar will react with the sodium bicarbonate to create gas. When all bubbling stops in the flask, the reaction is complete. You will also see that the gas pressure levels off and remains constant. Record the pressure in your data table alongside the amount of sodium bicarbonate used.
7. Remove the rubber stopper from the flask. Rinse the flask with tap water.
8. We are going to run the experiment three more times, increasing the mass of the sodium bicarbonate each time. Each mass should be between 0.05 – 0.10 g greater than the preceding mass. DO NOT EXCEED 0.4 g of sodium bicarbonate. For each reaction record the final pressure and the amount of sodium bicarbonate used in your data table.
9. Clean your flask and return all equipment to its proper place.

## Prelab: This prelab covers all three experiments

1. Name two ways in which gases differ from liquids and solids.
2. The Kinetic Molecular Theory tells us that individual gas particles do NOT interact with one another. How does this affect the behavior of gases?
3. The low density of gases tells us that the particles in a gas have a great deal of space between them. What special feature of gases does this lead to?
4. Explain how the pressure in a gas is created.
5. Name three factors which can affect the pressure exerted by a gas.
6. If you have a sample of gas and you increase the temperature of that gas, what effect will that have on the pressure of the gas sample? Explain why?
7. If you have a sample of gas in a flexible container and you wanted to increase the pressure, would you increase or decrease the volume? Why?
8. Explain why a greater amount of gas particles will lead to greater pressure exerted by the gas.

## Data Tables for Gas Labs

# Lab #28: Determining the Universal Gas Constant

**PROCEDURE**

1. Fill a 400 ml beaker 2/3 full with room temperature water. Obtain a piece of magnesium ribbon about 1.5 cm in length and determine the exact mass. Do not take more magnesium than this or the reaction will produce too much gas and you will need to start over!
2. Roll the magnesium ribbon into a loose coil and tie it with one end of a piece of thread, approximately 25 cm in length. Tie it so that all of the loops of the coil are fastened together.
3. Carefully pour approximately 15 ml of 6M hydrochloric acid into a 50 ml beaker. Then pour the HCl into the eudiometer.
4. While holding the eudiometer at an approximately 45º angle, very slowly pour water from the 400 ml beaker into the eudiometer, being careful to layer the water over the acid so that they do not mix. Add enough water to fill the eudiometer completely, so that when a stopper is inserted, some water will be displaced.
5. Lower the magnesium coil into the water in the eudiometer to a depth of about 5 cm, holding on to your thread. Insert the rubber stopper into the open end of the eudiometer to hold the thread in position. There should be no air in the tube at this point.
6. Cover the hold of the stopper with your finger, and invert the eudiometer in the 400 ml beaker of water. Clamp the eudiometer into position on the ring stand. The acid will flow down the tube and react with the magnesium. Allow this to continue until all of the magnesium is gone.
7. Your eudiometer will now contain both a small amount of liquid and a large volume of gas. The gas should not extend past the calibrated portion of the eudiometer. If it does, you will need to start over and use a smaller piece of magnesium.
8. At this point we could read the eudiometer and record our volume, however, do we know the pressure inside the eudiometer at this point? The answer, of course, is no. To determine the pressure inside the tube (and the corresponding volume) we will be using a large graduated cylinder. Cover the stopper hole of your eudiometer with a finger and carefully transfer the eudiometer to a 4L graduated cylinder that has been filled with water. Once the opening in the stopper is underwater you can remove your finger. Move the eudiometer up and down within the graduated cylinder and note that the volume of the gas in your tube changes as the amount of pressure from the water changes. Read the next step to see how to determine the volume at atmospheric pressure:
9. Move the eudiometer within the graduated cylinder until the level of the water inside the eudiometer is the same as the water level in the graduated cylinder. Line up the water level inside the eudiometer so that it is at the same level as the water in the 4L graduated cylinder. When the two water levels the same this tells us that the pressure of the gas pushing down inside the tube is equal to the atmospheric pressure pushing down on the graduated cylinder. In other words, the pressure inside the eudiometer is equal to atmospheric pressure outside the tube. Read, as accurately as possible, the **volume** of gas at this pressure. Also, your teacher has checked the atmospheric **pressure** and written it on the board. Remember that this is the same as the pressure within your eudiometer right now. Record this as your pressure.
10. At the same time, note the temperature in the room using the thermostat. Convert from Fahrenheit to Kelvin and record this temperature.
11. Remember that this gas was collected over water. This means that the measured pressure inside the eudiometer includes some water molecules which “hitched a ride”. Using the table in your book/notes, subtract out water vapor pressure and record the hydrogen gas pressure in your data table.

## PRELAB QUESTIONS

1. List the four properties which can be used to describe any gas.
2. How are these four properties related to the Universal Gas Constant? (This can be answered with an equation.)
3. What is the objective in this lab?
4. Write the equation for the reaction between magnesium and hydrochloric acid that we will be performing in this lab.
5. From your equation, what gas will we be collecting?
6. In this experiment we are collecting our gas “over water”. What does this mean?
7. When a gas is collected over water, water molecules get caught up in the gas creating a mixture. How does this affect the pressure of the gas? How do we correct this problem?
8. Read through the procedure and explain how we determine the pressure of the gas inside the eudiometer.

## Data Table

# Lab #29: The Effect of Temperature on Vapor Pressure

**Procedure**

1. This lab will be done in groups of 4 to facilitate the number of large beakers needed.
2. Obtain and wear goggles. CAUTION: The methanol used in this experiment is flammable. Avoid inhaling the vapors. Avoid contact with your skin or clothing. Be sure there are no open flames or sparks in the lab during this experiment.
3. Prepare the temperature probe and gas pressure sensors in the following fashion:
	1. Plug the gas pressure sensor into CH 1 and the temperature probe into CH 2.
	2. Insert the rubber stopper assembly into a 125 ml Erlenmeyer flask. Make sure the rubber stopper is on tight.
	3. Press the APPS button on your calculator.
	4. Choose Easy Data or DATA MATE.
	5. The temperature and pressure readings should be displayed.
4. In this experiment we will be measuring the vapor pressure in the flask at four different temperatures to see how temperature affects our vapor pressure. For each of the four temperature ranges follow these steps.
	1. Draw 3ml of methanol up into the syringe. With the two way valve CLOSED, screw the syringe onto the two way valve.
	2. Open the 2 way valve and push the plunger so that all of the methanol goes into the flask.
	3. Quickly pull the plunger back to the 3 ml mark on the syringe and then close the 2 way valve.
	4. Remove the syringe.
	5. Using a mixture of ice and tap water, in a 600 ml beaker, prepare a water bath whose temperature is in the range of 0 to 5°C.
	6. Place the temperature probe in the water bath.
	7. Place the flask into the water bath and hold it so that the majority of the flask is submerged. (Do not allow the water level above the stopper.)
	8. Wait at least five minutes so that the temperature inside the flask can reach the same temperature as the water bath.
	9. Make sure that the temperature and pressure readings are stabilized and record the data in your data table. (Hint: While you are waiting for it to stabilize your partner can be preparing the next water bath.)
5. You now have three more water baths to test in. For each test we will use the same sample of methanol so DO NOT remove the stopper on your flask or open the stopper valve.
6. Repeat steps **e – i,** each time with a water bath of a different temperature. The temperature of each bath should be in the following ranges:
	1. 10 to 15° (this should need a little less ice than the last bath)
	2. 20 to 25° (this should be close to the temperature straight from the tap)
	3. 30 to 35° (you will probably need to use your warm tap or a little boiling water to achieve this temp. range)
7. When you are done, gently loosen and remove the Gas pressure sensor so the Erlenmeyer flask is open to the atmosphere. Remove the stopper from the flask and dispose of the methanol in the container provided by your teacher.
8. Clean and dry your Erlenmeyer flask with soap and water. Draw air in and out of the syringe enough times that you are certain that all methanol has evaporated from it. Then clean with soap and water.

## Prelab:

1. If I have a glass of water, what two phases (states of matter) of H2O are present?
2. Explain how evaporation occurs.
3. What is equilibrium?
4. What is vapor pressure?
5. Based on what you have read, what do you expect to observe in your vapor pressure reading, as the temperature is increased. Explain your answer.

## Data Table

# Lab #30: Evaporation and Intermolecular Attractions

In this experiment, temperature probes are placed in various liquids . The liquids are then allowed to evaporate and the resulting temperature change is measured. You may remember that evaporation is an endothermic process. In order to break the intermolecular attractions that hold the molecules close to each other, energy must be absorbed. As energy is absorbed from the surroundings, the temperature of the surroundings will drop. As you may infer, the magnitude (size) of a temperature drop, and thus the energy taken up, is proportional to the number of molecules which evaporate. By monitoring the change in temperature we can tell whether a lot of molecules evaporate or only a few.

The rate of evaporation is dependent on the strength of **intermolecular forces** of attraction. In a compound which has weak intermolecular forces, you would expect a great deal of evaporation and therefore a greater drop in temperature. On the other hand, for molecules with strong intermolecular forces evaporation occurs more slowly and the change in temperature will not be as great.

In this experiment, you will observe temperature changes caused by the evaporation of several liquids whose intermolecular forces vary in strength. You will then relate the differences in temperature change to the comparative strength of the intermolecular forces.

You will encounter two types of organic compounds in this experiment—alkanes and alcohols. Remember that **alkanes** are nonpolar molecules containing only carbon and hydrogen. In addition to carbon and hydrogen atoms, the **alcohols** also contain the polar -OH functional group.

**MATERIALS**

CBL/graphing calculator setup

Temperature Probe

2 Vernier adapter cables

6 pieces of filter paper (2.5 cm X 2.5 cm)

paper clip

methanol (methyl alcohol)

ethanol (ethyl alcohol)

1-propanol

1-butanol

n-heptane

n-hexane

n-pentane

**PROCEDURE**

Obtain and wear goggles! **CAUTION:** *The compounds used in this experiment are flammable and poisonous. Avoid inhaling their vapors. Avoid contacting them with your skin or clothing. Be sure there are no open flames or sparks in the lab during this experiment. Notify your teacher immediately if an accident occurs.*

1. Obtain a test-tube holder/test tubes containing compounds.
2. Roll a piece of filter paper around the temperature probe and secure with a paperclip.
3. Place the filter paper in the test tube containing methanol and allow it to sit for 45 seconds. During this time the filter paper should become saturated with liquid.
4. After 45 seconds, read the temperature displayed on your CBL/calculator. Be sure that the temperature is stable and not changing. If not, continue to monitor the temperature for about 15 seconds more to establish the initial temperature of the liquid.
5. You will be **timing** the evaporation rate. Begin timing when the paper/probe assembly has been removed from the liquid. Remove the probe from the liquid and hold it off the edge of the counter so that probe tip extends approximately 5 cm over the edge of the countertop.
6. As the filter paper dries, you should observe a drop in temperature. The temperature should stop dropping when the liquid has completely evaporated. **Watch carefully, as the temperature will quickly begin to rise once the paper is dry!** You should stop timing as soon as the temperature begins to rise. Record your lowest temperature and the time interval in the table provided.
7. Subtract the minimum temperature from the maximum temperature to determine Δt, the temperature change during evaporation. Then determine Δt /second. **NOTE: Some substances will continue to drop temperature from a long time. Rather than wait, record the temperature observed after two minutes has passed, then record 120 seconds as your time interval. Calculate Δt /second as instructed above.**
8. Remove the filter paper and dispose of in the proper waste container as directed by your teacher.
9. Repeat steps 3-9 with the remaining liquids provided. Record your values in the space provided.

## Prelab Questions:

* 1. What is an endothermic process?
	2. Explain why evaporation causes the temperature of the surroundings to drop.
	3. How are intermolecular forces related to evaporation rate? In other words, what does evaporation have to do with intermolecular forces?
	4. The evaporation rate is measured for two different substances and the temperature change for Substance A is significantly lower than the change for Substance B? How must the intermolecular forces for Substance A compare to those for Substance B? In other words, are they weaker or stronger? How do you know?
	5. Of the two types of molecules we are testing, alkanes and alcohols, which have the stronger IMF? Explain your answer.

## Intermolecular Forces Data Sheet (This time the data table has been prepared for you!)

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Liquid tested** | **T** | **Time elapsed (in seconds)** | **T/seconds** | **Mass of the molecule** | **Type of IMF present** |
| **Methanol****CH3OH** |  |  |  |  |  |
| **Ethanol****C2H5OH** |  |  |  |  |  |
| **Propanol****C3H7OH** |  |  |  |  |  |
| **Butanol****C4H9OH** |  |  |  |  |  |
| **Pentane****C5H12** |  |  |  |  |  |
| **Hexane****C6H14** |  |  |  |  |  |
| **Heptanes****C7H16** |  |  |  |  |  |

# Lab #31: Unsaturated, Saturated and Supersaturated Lab

**Procedure, observations, and conclusions:**

**Situation #1**

1. Weigh out 5.0 grams of sodium thiosulfate pentahydrate. (Na2S2O3\* 5H2O)
2. Place the sodium thiosulfate pentahydrate into a medium sized test tube.
3. Measure out 1.0 ml of distilled water in your 10 ml graduated cylinder.
4. Add the water to the test tube of sodium thiosulfate pentahydrate.
5. What do you note about the solution? Did all of the substance dissolve?
6. Make a note regarding the appearance of your test tube in your data table.

**Situation #2**

1. Attach a test tube clamp to the test tube of sodium thiosulfate solution from situation #1.
2. We are going to heat the solution by **passing** it through a flame. YOU ARE NOT TO HOLD THE TEST TUBE DIRECTLY IN THE FLAME. KEEP IT MOVING AND MAKE SURE THAT IT IS NOT POINTED AT ANYONE. Hold the test tube at an angle. Remember, keep it moving while you pass it through the Bunsen burner flame.
3. Continue heating until all the sodium thiosulfate has dissolved.
4. Place the hot solution of sodium thiosulfate into your test tube rack.
5. We will now test the solution with a seed crystal to determine whether it is saturated, supersaturated or unsaturated. Obtain a single small crystal of sodium thiuosulfate pentahydrate from the stock bottle.
6. Add the crystal to the hot solution of sodium thiosulfate.
7. Record what you observe.

**Situation #3**

1. Add cold tap water to your 250 or 400 ml beaker until its almost full.
2. Take the solution from situation #2 and gently reheat it in your flame until any solid that formed is redissolved.
3. Now place the hot test tube in the cold water and let it sit for at least 5 minutes in the cold water.
4. If, at the end of 5 minutes, the solution of sodium thiosulfate is still clear you may continue to step 19. If not, repeat steps 16 and 17.
5. Remove the cooled solution of sodium thiosulfate from the cold water. Place the test tube in your test tube rack.
6. Once again, we will test the solution with a seed crystal. As before, place one seed crystal in the solution and observe what happens.
7. Record your observations in your data table.
8. If the seed crystal dissolved then you need to continue to step 22. If you observed the formation of crystals instead, skip ahead to step 23.
9. Add one or two more seed crystals to your test tube and then reheat it until everything dissolves. Repeat steps 16 through 21 until you have observed the extra crystals which form when a supersaturated solution is disturbed. Then proceed to step 23.
10. Gently reheat your test tube of sodium thiosulfate so that it is once again clear.
11. Pour the hot solution into the waste beaker in the fume hood.
12. Clean up your equipment and work station .

## Prelab Questions:

1. **What is solubility?**
2. **What precautions must you take when heating a test tube?**
3. **What steps are involved in creating a supersaturated solution? Explain why the second step is necessary.**
4. **What is observed when a seed crystal is placed into a saturated solution? An unsaturated solution? A supersaturated solution?**

## Data Table

# Lab #32: Solution Concentration

Purpose: In Part 1 of the lab, we will prepare a solution of a given concentration. We will then test this concentration, in Part 2, by using it in a single replacement lab.

**Part One: In this part of the lab we will be asked to prepare a solution of a given concentration.**

You will prepare a solution 50 ml in volume, with a concentration of 0.1 M Cu(NO3)2

**Part Two: In this part, you will determine the expected grams of product in a reaction using your prepared solution.**

**Procedure:**

1. Part One of this lab will be up to do. You will need to determine the procedure according to what you have learned about solutions.
2. Part Two:
	1. Place your 50 ml of solution in a 100 ml beaker.
	2. Suspend a zinc strip into the copper nitrate solution and allow to sit for 10 minutes.
	3. Remove the zinc strip and wipe clean. See your teacher to weigh your zinc strip.
	4. Clean up your lab station and discard of your left over solution down the sink.

# Lab #33: Chloride Concentration in Tap Water Lab

Concentration is a measure of how much solute is dissolved in a given volume of solution (or solvent). In order to determine concentration in a lab setting we must find out how much solute is dissolved. One way to do this is through titration. In a titration a solution with an unknown concentration is reacted with a solution of a known concentration. By using our measurement we can calculate the concentration of the unknown solution.

In this experiment we will use silver nitrate as our solution with the **known concentration**. As we add silver nitrate to the water, the silver reacts with the chloride ions in the water to form a precipitate. In a reaction that looks like this:

AgNO3 + Cl- 🡪 AgCl + NO31-

The formation of silver chloride removes the chloride ions from solution.

A titration used an indicator to change color when the reaction is complete. In this experiment we will use fluorescein. When there are chloride ions in the water the fluorescein makes the solution appear a light green. When all of the chloride ions have been removed from the water, the solution turns pink. This tells us that our reaction is complete. (Note: because silver chloride is a precipitate the appearance of the solution will be cloudy.)

In this lab you will perform two titrations. In the first one you will test a solution whose concentration is known. This way you can compare your perform your own calculations and then compare your results with the actual concentration to make sure you are being accurate. In the second titration you will test a solution whose concentration is unknown. You will then calculate the concentration of the chloride ion in this unknown solution.

**Materials:**

Graduated cylinder

Ring stand 250 ml beaker

Buret clamp

Buret

Distilled Water

Water sample (UNKNOWN CONCENTRATION)

Fluorescein

Dextrin

Potassium chloride solution (0.005 M solution)

Silver Nitrate (0.005 M solution)

***Procedure (Part 1)***

1. Fill your buret with the 0.005 M silver nitrate. Record the initial reading.
2. Use the graduated cylinder to measure about 10.0 ml of the 0.005 M potassium chloride solution. Record your actual volume as accurately as possible in your data table.
3. Add 2-3 drops of the fluorescein and a small scoop of dextrin. (The purpose of the dextrin is simply to help the precipitate form. You do not need to record these amounts.)
4. Open the valve on the buret to slowly add silver nitrate to the solution. Continue to add silver nitrate until the solution changes from green to pink. (Add the silver nitrate slowly, closing the valve several times if necessary. You do not want to add more silver nitrate than is necessary to get to the pink endpoint.
5. Record the final volume of the silver nitrate in your buret.
6. Using your final and initial volumes of silver nitrate, calculate the volume of silver nitrate used in the experiment. Record this amount in your data table.

***Procedure (Part 2)***

1. If necessary, refill your buret with silver nitrate.
2. Using your cleaned graduated cylinder, obtain approximately 10 ml of the Unknown tap water provided. Record you actual volume in your data table.
3. Repeat steps 3 through 6 from Part 1, being sure to record all your data.
4. Repeat Part 1 (1 through 3) one more time with a new sample of the Unknown tap water.

## Prelab Questions:

1. What is a titration? What are we trying to determine when we do a titration?
2. What ion are we testing for in this experiment?
3. Why are we adding silver nitrate to our solution?
4. What is the purpose of the fluorescein in our experiment?
5. Using the balanced equation shown in the prelab, if I have one mole of chloride ions, how many moles of silver nitrate will it take to react completely?

## Data Table

# Lab #34: Determining Molar Mass from Freezing Pt. Depression

**Introduction**:

Pure substances have many distinguishing properties by which they can be identified. Among these would be the boiling or freezing/melting point of the liquid which is consistent for any pure substance. For example, pure water will always freeze or melt at 0°C and always boil at 100°. Note that I have made a point of saying pure substance. When a pure substance is then mixed to create a solution, such as water and salt used to make salt water, the properties of the solution are different from those of the pure substance. Salt water has a higher boiling point and a lower freezing/melting point than pure water alone. So what is the boiling point for salt water? Well, it varies. The boiling or freezing point of salt water is dependent on the amount of salt that I have added. In other words, the boiling point of salt water is dependent on its **concentration**. A property, such as boiling or melting point, which is affected only by the concentration of the solute is called a **colligative** **property**.

For ionic substances, when comparing the boiling point of a solution to the boiling point for the pure substance, the boiling point of the solution will always be higher. This is called boiling point elevation. In comparison, the freezing point for a solution will always be lower than the freezing point of the pure substance. This is called freezing point depression. Now remember, this change in temperature is only controlled by **how much** solute has been added. By observing **how high** the boiling point has been raised or **how low** the freezing point has been depressed we can calculate how much of the solute is present. In other words, we can calculate the concentration of the solution.

**Purpose**

You will measure the freezing point depressions that occur with solutions containing a solute whose empirical formula is known. Your data and those of the entire class will be used to calculate the molar mass of the solute (antifreeze).

**Materials**:

600 mL beaker

about 20 g NaCl or MgCl2

stirring rod

thermometer

aluminum foil

3 small test tubes

distilled water

300 mL ice

2 250mL beakers

approx. 20 mL antifreeze

large graduated cylinder

small graduated cylinder

**Procedure:**

1. **Preparing a cold water bath. This is where you will freeze your solutions.** Half fill a 600-mL beaker with ice and cover it with about 20 g of NaCl. The NaCl will, of course, lower the freezing point of water, creating an ice bath colder than 0°. Stir this ice-salt mixture with the stirring rod until it reaches a constant temperature at or below -10°C. For extra insulation, cover the outside of the beaker with aluminum foil with the reflective side in.

2. **Determine the freezing point of water *according to your thermometer****.* Half fill a test tube with distilled water and place it in the ice bath. Place your thermometer inside the test tube and use it to gently stir. Clamp the thermometer to the test tube. Note the temperature every 30 seconds while you are stirring (it might be helpful to keep track in the margin of your paper). When the freezing point is reached, crystals will start to form and the temperature will remain constant (the freezing point). Record your freezing point for pure water in your data table.

3. **Prepare your two antifreeze/water solutions.** Tare the mass of a *clean* 250-mL beaker so that the scale reads zero. Measure out *exactly* 5.00 g of antifreeze. When you are getting close to 5.00 g, use a plastic pipette to ensure accuracy. Add 50 mL of distilled water to the beaker and stir thoroughly with a stirring rod. In a similar fashion, prepare a second solution using 10.00 g of antifreeze. Label the beakers *Solution 1* and *Solution 2*.

4. **Determine the Freezing points of the two antifreeze/water solutions.** Half fill a test tube with *Solution 1* and place it in the ice/salt bath. Place your thermometer into the test tube and note the temperature every 30 seconds. Stir gently with your thermometer. When the solution reaches 0ۜC, add a small chip of ice to the solution. (This will eliminate the possibility of supercooling - when the liquid is temporarily cooled to below its freezing point). Keep stirring until the solution freezes and you can’t stir anymore. Continue to monitor the temperature until it remains constant for a period of time. This is your freezing point. Record in your data table.

5. Repeat step 4 with *Solution 2.*

6. Dispose of the antifreeze solutions in the waste container your teacher has set out and BE SURE TO WASH YOUR HANDS!! (Antifreeze is toxic!!)

## Prelab Questions:

1. What can you tell me about the boiling point of a pure substance?
2. How does the boiling point for a **solution** compare to the boiling point of the **pure** **solvent** it was created from?
3. How does the freezing point for a **solution** compare to the **pure solvent** it was created from?
4. What is freezing point depression?
5. How is freezing point depression related to the concentration of the solution?
6. What is a colligative property?
7. Using your new knowledge of colligative properties, explain why the city’s road crew puts salt down on the road during the winter months. (Be as detailed in your explanation as possible.)

## Data Table (s):

# Lab #35: Acid-Base Titration Lab

In an earlier lab we learned that a titration is a laboratory procedure where you react one solution that has an unknown solution with a second solution whose concentration is known. By comparing the quantities involved the experimenter can determine the concentration of the unknown solution.

The titration we will be doing in this lab involves and acid and a base. The reaction carried out in this lab involves vinegar (an acetic acid solution) and sodium hydroxide. The reaction we will observe is described in the equation:

HC2H3O2 + NaOH 🡪 H2O + NaC2H3O2

You may remember from our earlier study of acids and bases, that this reaction is called a neutralization reaction because the acid and base, when mixed in the proper proportions, will neutralize each other creating water and a salt.

As with the earlier titration, the point at which neutralization occurs (what we call the endpoint) will be made visible through the use of an indicator. The indicator phenolphthalein, which we used earlier this year, will be used. You may remember that phenolphthalein is clear in acid but turns pink as the solution becomes basic.

Vinegar is a solution of acetic acid, HC2H3O2, dissolved in water. The concentration of acetic acid in the vinegar is unknown. In this lab you will titrate a sample of vinegar against a 1.0 M NaOH solution. You will calculate the concentration (molarity) of the acetic acid in the vinegar.

**Procedure**

*Note: When using the buret you must make sure that you do not fill it higher than the 50 ml mark or let it drain lower than the 0 ml mark. If you find that you are running out, stop at zero and refill the buret with more solution.*

1. Rinse one buret with distilled water, and then rinse twice with small amounts of NaOH. Fill the buret with the 1.0 M NaOH. Using label tape, label the buret NaOH.
2. Do the same with the second buret, this time rinsing with a small amount of vinegar. Record the initial volume. Label this buret HC2H3O2.
3. Using the vinegar buret, transfer approximately 25 ml of vinegar into a clean, dry Erlenmeyer flask. Add 2-3 drops phenolphthalein to the flask as well as about 50 ml of distilled water. Swirl gently until the contents are well mixed. (NOTE: IT IS VERY IMPORTANT THAT THE PHENOLPHTHALEIN ONLY BE PLACED INTO THE FLASK. AT NO TIME SHOULD THE PHENOLPHTHALEIN BE PLACED INTO THE BURET ITSELF.)
4. Clamp the buret so that its tip is within the opening of the flask, but above the surface of the solution.
5. Record the initial buret reading in the NaOH buret.
6. Begin the titration by adding successive small portions of the NaOH solution. After each addition, swirl the flask to mix.
7. Continue adding NaOH until the solution turns the slightest shade of pink. Record the final buret reading.
8. If you go past the slight pink endpoint, change over to your vinegar buret and add, drop by drop, vinegar, until you go back to the final endpoint. Record the final buret reading.
9. It may be necessary to go back and forth between each buret several times to get the best endpoint. Just be that when you reach the desired endpoint, you note and record the final reading on each buret.
10. Dispose of the contents of your flask down the drain and rinse your flask well. Repeat this procedure, steps 3 through 9 at least two more times.
11. When you have finished, empty the NaOH and vinegar from the burets and **rinse each well** with distilled water, filling them and allowing them to drain (through the tip) at least two times. Store both burets, upside down, with the valve open.
12. Clean up your area and wash your hands.

## Prelab

1. What must you be careful of when filling and using your buret?
2. What acid is found in vinegar?
3. Why is this reaction called a neutralization reaction?
4. How will we known when neutralization (the endpoint) has occurred?
5. What are we trying to determine in this lab?

## Data Table