

Chemistry 121 Laboratory Manual



Pierce College Puyallup
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Use of Volumetric Glassware: The Floating Egg Problem*

Background:

We are all familiar with direct measurement of things. If, for example, we weigh a 1.0 cm³ cube of iron, and find it has a mass of 7.86 g, we have determined directly from a piece of iron that the density of iron is 7.86 g/cm³. But in many cases we cannot measure a property directly. Therefore we resort to indirect measuring. We make a direct measurement and, using a known relationship, apply it in determining another quantity we cannot measure directly.

In this lab, you will see how we can indirectly determine the density of an object by directly measuring the density of a solution in which it just floats.

To carry out good measurements, chemists use special *volumetric* glassware in the chemistry laboratory rather than ordinary beakers and flasks. Understanding the proper use and limitations of such glassware is important in achieving reliable results. Experimental technique and the inherent accuracy of the glassware both affect experimental results.

In this experiment, you will practice with volumetric glassware to improve your technique and to solve an important historical problem in practical chemistry. Volumetric glassware is designed for two quite different purposes:

- To contain (TC) an accurate volume of liquid
- To deliver (TD) an accurate volume of liquid

Although a container meant to contain an accurately known volume could be used to deliver that volume, it is generally not used to deliver because the amount of liquid actually delivered would never exactly equal the amount originally in the container. Some liquid would inevitably remain clinging to the interior walls of the container.

When we must know accurately how much liquid is delivered, we use TD glassware. Some liquid also clings to the walls of TD glassware, but the calibration of the TD glassware takes this into account.

A Historical Question:

For many years, soap was made at home from a variety of recipes. Animal fat, usually from cattle (also called tallow), was cooked with a lye solution. Lye, though it was mostly sodium hydroxide, could not be made from purified chemicals, as it is now. Instead, the solution was obtained from ashes and water. The ashes were treated with hot water, and then the mixture was filtered to obtain a solution. But before this could be used in soap making, the concentration of the lye solution had to be checked.

One way to check the concentration was to determine the density of the solution. One simple test was to try to just float a raw egg in the solution. If the egg sank, the density of the lye solution was too low. If the egg floated too high, the density was too great, and water was added before adding the fat. To “just float” in this case means to make the top of the egg just touch the top of the solution, without any significant amount of the egg protruding above the surface of the

* This laboratory is from the Working with Chemistry Lab Manual, Donald Wink et al. Used by permission of the author.

solution. At that point, the density of the lye – and therefore its concentration – was correct. Thus, we can see how soapmakers had mastered the art of indirect measurement!

We will use this strategy in a different manner. We will use a solution that just floats an egg to determine the density of the egg.

Procedure:

Preparation of the Solution

Because lye is caustic and corrosive, the substitute sodium chloride (table salt) will be used in this lab. The solutions are still concentrated enough to pose a hazard to the eyes. Goggles must be worn. Salt, eggs, standard glassware, balances, and deionized water will be available. Other items may be available on request. The eggs are uncooked. But because they are used in a lab, and because they are at room temperature for long periods, they should *not* be reused as food!

Follow these guidelines to prepare a solution that will just float an egg:

- Make at least 1500 mL of the solution
- Use only one egg for testing. Note the code. At the end of the experiment, you will use this to find out the freshness of the egg from your instructor.

When you have prepared the solution, each individual should record the procedure in his or her notebook.

Measurement of Solution Density

You will use three different types of glassware to measure the density of the salt solution. For the equipment that dispenses solution (pipets and burets), you will determine the mass of the solution by collecting the solutions into a weighed Erlenmeyer flask. For the volumetric flask you will determine the mass of the solution in the flask itself. So, obtain the weight of a clean, dry Erlenmeyer flask and of the volumetric flask. Don't forget to take your notebook to the balance. Before weighing, be sure the balance reads zero when nothing is on the pan. Check with your instructor if this is not the case. Then place the flask on the balance. This is the mass of the empty container.

Carry out three repetitions of a determination of the density of the solution with each piece of glassware by measuring a volume close to 25 mL of the solution and then determining the weight of the flask plus the solution.

A. Buret

Obtain a buret and use a ring stand and a buret clamp to mount the buret vertically. Rinse the buret with a small amount of your salt solution. Fill the buret to about the zero level with your solution. Drain some out to ensure that no pure water or air gaps are in the tip of the buret. Record the level reading after it has stabilized.* The level should be recorded to the nearest 0.01 mL, but it does not have to be exactly 0.00 mL. Estimate the location of the bottom of the meniscus between the calibration marks which are 0.1 mL apart.

Deliver solution from the buret into the clean, dry, weighed Erlenmeyer flask. Deliver until the buret level is near the 24 mL mark. Record the reading after waiting for the level to stabilize.

* Readings that are obtained too quickly do not include the liquid that slowly trails down the side of the buret.

Now weigh the flask plus the solution and record the total mass. When you subtract the mass of the empty, dry flask, you will obtain the mass of the solution.

Repeat the entire procedure for a total of three trials. You want to develop your technique so that you can use a buret accurately and reproducibly (precisely)

B. Volumetric Pipet

Carry out the weighing procedure of Part A using a clean, pre-rinsed 25 mL volumetric pipet to deliver 25.00 mL of solution. If the pipet is marked TD, it is designed to give 25.00 mL by draining with gravity: Do not include the solution that remains in the tip after draining. The tips of the pipets are very fragile. Be careful not to exert pressure on them. Do a total of three trials.

C. Volumetric Flask

Fill your pre-weighed volumetric flask with the solution, stopping at the calibration mark. This will give you 25.00 ± 0.02 mL. Weigh the solution-filled container and record the mass. This is the mass of the flask plus the solution “contained”, and, by subtracting the mass of the empty flask, you can determine the mass of the solution contained.

Before you leave the lab, determine the density of your solution and place the results on the board.

Lab Report #1
The Floating Egg Problem

Name: _____

Partner's Name: _____

Date: _____

Data and Calculations:

Part A.

	Trial 1	Trial 2	Trial 3
Mass of Empty Erlenmeyer Flask			
Initial Buret Volume			
Buret Volume after Delivery of Solution			
Volume of Solution Added to Erlenmeyer Flask			
Mass of Erlenmeyer Flask and Solution			
Mass of Solution			
Density of Solution			

Part B.

	Trial 1	Trial 2	Trial 3
Mass of Empty Erlenmeyer Flask			
Volume of Solution Added to Erlenmeyer Flask			
Mass of Erlenmeyer Flask and Solution			
Mass of Solution			
Density of Solution			

Part C.

	Trial 1	Trial 2	Trial 3
Mass of Empty Volumetric Flask			
Mass of Volumetric Flask and Solution			
Mass of Solution			
Volume of Solution in Volumetric Flask			
Density of Solution			

Questions:

1. Compare the densities obtained for different eggs. What, specifically, was the effect of the freshness of the egg?
2. What were the advantages and disadvantages of each measuring device?
3. Calculate the average density obtained from each measuring device (Part A, B, and C). Which of the average densities do you believe was most accurate? Why?

The Separation of a Mixture into Pure Substances

The experiment is designed to familiarize you with some standard chemical techniques and to encourage careful work in separating and weighing chemicals.

In this experiment you will separate a mixture of three substances, sodium chloride (NaCl), benzoic acid (C₆H₅COOH), and silicon dioxide (SiO₂), into pure substances based on their solubility in water. The amount of a substance that will dissolve in water depends on the nature of the substance and on the temperature. Water is a polar substance so other polar substances will dissolve in water (“Like dissolves Like”). Sodium chloride is an ionic compound that is quite soluble in water. Silicon dioxide (sand) is a large macromolecule that is essentially insoluble in water. Benzoic acid is polar, but is much less polar than water so it dissolves in water only to a limited extent. The solubility of these three substances in water is given below. The solubility of each is expressed in terms of grams of the solid that dissolves in 100 g of water. It is evident that each of these has a different solubility/temperature relationship.

Solubility (g of substance per 100 g H₂O as a Function of Temperature

Substance	Temperature °C					
	0	20	40	60	80	100
NaCl	36.00	36.00	37.00	37.00	38.00	40.00
C ₆ H ₅ COOH	0.17	0.29	0.42	1.20	2.80	7.00
SiO ₂	0	0	0	0	0	0

In this experiment, you will use decantation and two types of filtration. Decantation is a process of separating a liquid (called supernatant) from a solid residue by gently pouring off the liquid from the solid. It is easier to decant when you pour the liquid down a stirring rod. Filtration separates a solid and a liquid using a porous material such as filter paper which allows the liquid to pass but retains the solid. Gravity filtration uses the force of gravity to achieve separation. Vacuum filtration uses the suction created by a vacuum to speed up the separation.

Pre-lab questions:

- Describe what types of mixtures the following techniques are best used to separate.
 - Decantation
 - Crystallization
 - Sublimation
 - Filtration
- Look at the table above. Compare the solubility of sodium chloride (NaCl) to the solubility of benzoic acid (C₆H₅COOH) at 0 °C.
- Based on the table above, which technique might be best used to separate the sand, SiO₂? Why would you choose this technique?
- Give an example of an “out of the vacuum of school” use for filtration.

Procedure:

1. Carefully weigh a clean, dry 250-ml beaker (beaker 1). Transfer about 5 to 6 grams of the mixture, which contains sodium chloride, benzoic acid, and silicon dioxide into beaker 1. Be sure to record which unknown you have. Weigh beaker 1 containing the mixture. From the two weights, obtain the weight of the mixture. Record your results in your lab notebook.
2. Add about 50 ml of distilled water to beaker 1. Place beaker 1 with its contents on a wire screen, which is resting on an iron ring connected to a ring stand.
3. Heat the mixture to the boiling point (stop heating when boiling is observed), and stir the mixture with a stirring rod to make sure that all soluble material is dissolved. At the boiling point temperature, all benzoic acid and sodium chloride should be in solution. Thus, they have been separated from sand. (This is an extraction process.) Take about 15 ml of distilled water in a beaker and bring it to a boil. This boiling water will be used in step 5.
4. Decant the liquid while it is hot into another 250 ml beaker (beaker 2). Use Hot Hands to hold the hot beaker. Do not let the sand get into beaker 2 or wait too long before decanting. Use a glass rod to aid the decantation process.
5. To dissolve as much benzoic acid and sodium chloride as possible, wash the sand in beaker 1 with about 10 ml or less of boiling water and decant the washing into beaker 2.
6. Place beaker 2 in an ice bath to let it cool. Observe carefully how the benzoic acid crystallizes out of the solution. Set beaker 2 aside to ensure maximum crystallization.
7. Place your beaker of sand in the drying oven so that it can dry completely before weighing. Make sure you mark the beaker with a grease pencil so you can identify it with your initials or some other mark. Leave the beaker in the oven until you have finished the rest of the lab to make certain that it is fully dried. When you take it out of the oven, let it cool, and weigh. You should move on to the next step and come back to weigh the beaker at the end of the experiment.
8. Assemble the Buchner funnel with a filter flask as directed by your instructor.
9. Connect a piece of vacuum rubber tubing from the vacuum to the filter flask. Vacuum rubber tubing is used so that it will not collapse when the pressure is reduced.
10. Place the proper size of filter paper in the Buchner funnel and wet it with distilled water using a wash bottle. This allows the filter paper to sit down on the grating. **Make sure to weigh your filter paper before doing so.**
11. Turn on the vacuum. Make sure to turn it on just enough to have proper suction occurring.

12. Pour the contents of beaker 2 (benzoic acid crystals) into the Buchner funnel.
13. Wash the solid (benzoic acid) with about 5ml or less of distilled ice water to ensure that the benzoic acid is free of sodium chloride.
14. Continue to suction until the liquid no longer drips from the funnel.
15. Turn off the vacuum and disconnect the rubber tubing from the filter flask
16. Using a spatula transfer the suction dried filter paper with benzoic acid on it onto a pre-weighed marked watch glass. Place the filter paper and watch glass into the drying oven for a few minutes to dry. When dry, take it out and let it cool before weighing.
17. To a large beaker add three or four boiling stones and weigh the beaker.
18. Transfer the filtrate (liquid) from the filter flask into the large beaker you just weighed in #17. Using a wash bottle rinse the filter flask with as little water as possible. Add the washings to the beaker. Place the beaker (which now contains sodium chloride and water) on a wire screen, which is resting on an iron ring. Start to heat the solution gently with a Bunsen burner so that the liquid does not boil over. Later, reduce the flame to avoid overheating, which may cause splattering of solid sodium chloride or shatter your beaker. When the water is evaporated entirely, let the sodium chloride dry, and the beaker cool before weighing. Weigh the beaker and its contents and calculate the weight of sodium chloride.
19. Calculate the percentage of each substance in the mixture by using the following formula:

$$\% \text{component} = \frac{\text{grams of component (salt or benzoic acid or sand)}}{\text{Total grams of mixture}} \times 100$$

Lab Report #2

Separation of a Mixture

Name: _____

Partner's Name: _____

Date: _____

Data and Calculations:

1. Weight of beaker 1 _____
2. Weight of beaker 1 and mixture _____
3. Weight of mixture _____
4. Weight of beaker 1 and dry sand _____
5. Weight of dry sand _____
6. Weight of filter paper _____
7. Weight of watch glass _____
8. Weight of watch glass, filter paper, and benzoic acid. _____
9. Weight of benzoic acid _____
10. Weight of large beaker and boiling chips _____
11. Weight of beaker, boiling chips and sodium chloride _____
12. Weight of sodium chloride _____
13. Percentage of sand in mixture _____
14. Percentage of benzoic acid in mixture _____
15. Percentage of sodium chloride _____
16. Total percentage of sample recovered _____

Questions: (answer on the back or a separate piece of paper)

1. Based on the data provided in the introduction, graph (on graph paper or a computer) the solubility of benzoic acid as it changes with temperature. Attach your graph to this sheet. Build the graph by hand or in excel (google sheets ok) – but make it a full page graph. Properly label you axes. Give it a title.
2. From your graph, determine the number of grams of benzoic acid that will dissolve in 100 grams of water at 25 °C.
3. If a solution is made up by mixing 6.0 g of benzoic acid in 1.00 L of water at 42 °C, would the solution be saturated? Would you expect to see solid on the bottom of the container?
4. Explain why the total percentage of sample recovered (#16) did not equal exactly 100%.

Physical Properties of Substances

In this experiment, you will develop your skills in scientific measurement by investigating of two basic physical properties, heat and melting point.

The Specific heat (C_p) of a substance is defined as the amount of heat energy required to raise the temperature of 1 gram of that substance 1 °C. The specific heat of a substance can be easily measured by heating the substance to approximately 100 °C and then adding it to cool water. The warm substance transfers some of its heat energy to the water, heating up the water. The temperature change of the water is related to the specific heat of the substance being studied.

The melting point of a compound is the temperature at which the solid becomes a liquid. To determine the melting point a small amount of material is heated slowly in a Melt-Temp apparatus. In this determination, two temperatures are recorded. The first temperature is recorded when the first drop of liquid forms among the crystals. The second temperature is the point at which all the crystals turn to a clear liquid. This means that the melting point is recorded as a range. The purity of a compound can be checked through its melting point. The purer the compound the closer the temperature will be to the actual melting point of the compound and the smaller the temperature range will be.

Pre lab questions:

1. How are temperature and heat different?
2. State the law of conservation of energy.
3. As you read the procedure and think about the law of conservation of energy, would you want to leave your test tube with the metal in it in the hot water bath for a shorter or longer period of time? Why?
4. In looking at the procedure, you add a lot of heat to the metal. When you put the metal in the calorimeter, where does the heat go?
5. Given your answer to number 2 above, what can you say about the heat lost and the heat gained by the substances in your experiment?

Procedure:

A. Specific Heat

Since efficient calorimeters are expensive, you will use an inexpensive calorimeter consisting of two Styrofoam cups, one inside the other. Weigh the cups as accurately as possible. Add approximately 50 ml of water and weigh again. Record the temperature of the water to the nearest 0.1 °C. Weigh a cylinder of your “specific heat unknown” and be sure to record the identifying letter. Place the unknown sample in a test tube and put the test tube in a beaker of water. The water in the beaker should be at a level equal to or higher than the sample in the test tube. Heat the water to boiling. Let the test tube stay in the boiling water for ten minutes. Measure the temperature of the boiling water to the nearest 0.1 °C.

Take the test tube out of the boiling water and quickly pour the unknown solid from the test tube into the calorimeter. Agitate the water with a glass rod. Watch the temperature as it increases and record the maximum temperature reached to the nearest 0.1 °C.

It is assumed that the heat given up by the metal is transferred to the water. The specific heat of water is 4.18 J/g°C, which means that 4.18 Joules of heat energy is required to raise the

temperature of 1 gram of water 1 °C. Use the table below to determine which metal your unknown is.

Metal	Specific Heat (J/g°C)
Aluminum	0.902
Brass	0.380
Copper	0.385
Gold	0.126
Iron	0.449
Nickel	0.444
Silver	0.235

B. Melting Point

For this part of the experiment, 5 or 6 substances will be on the front counter. Select one of these substances and determine its melting point.

The melting point of a substance can easily be determined by using the Mel-Temp apparatus. Before finding the melting point it is necessary to place a small amount of the substance being studied in a capillary tube. This is most easily accomplished by placing some of the substance on a watch glass and then pushing the open end of the tube into the substance. Then turn the tube over and tap the closed end on the bench top. This is normally all that is required to move the substance from the open end of the tube to the closed end. If this does not work in your case, please see your instructor for further suggestions.

Once the capillary tube has been filled with 1 to 3 mm of the substance, place it in the Mel-Temp apparatus. (Be sure that the apparatus has cooled to a temperature below the expected melting point of your sample.) Turn on the heater of the Mel-Temp. Determine the Mel-Temp range of the compound by recording the temperature when the solid first begins to melt and the temperature when it is completely melted.

Use the Merck Index to determine the melting point of the compound you tested.

C. Periodicity

Use the CRC to look up the melting point data for the elements in period 3 on the periodic table. Arrange this data into a data table of your own design. Then look up the melting points for the elements in group 1, (omitting Hydrogen) and collect this information in a data table. This information will help you answer the questions on the lab report.

Lab Report #3

Physical Properties of Substances

Name: _____

Partner's Name: _____

Date: _____

Specific Heat

Unknown Letter	_____
Mass of metal	_____
Mass of cups and water	_____
Mass of cups	_____
Mass of water used	_____
Initial temperature of water in calorimeter	_____
Temperature of boiling water	_____
Maximum temperature of metal + water	_____
Heat gained by water	_____
Heat lost by metal unknown	_____
Specific Heat of metal unknown	_____
Identity of Unknown Metal	_____

Melting Point

Substance used	_____
Experimental melting point range	_____
Melting point from the Merck Index	_____

Periodicity

Place your table here.

Questions:

1. Do you see a trend that continues through all the melting points of the period 3 elements?
If so describe that trend.

2. Do you see a trend that continues through all the melting point of the group 1 elements?
If so describe that trend.

3. Explain why the melting point you determined was different from that given in the Merck Index.

4. Calculate the amount of heat required to raise the temperature of 1000.00 g of mercury from 25.0 °C to 37.0 °C. The specific heat of mercury is 0.139 J/g °C.

5. Consider 10 g of H₂O and 10 g of mercury each in a separate container at 0 °C. If 100 J of heat energy was added to each, which of these statements would be true?
 - a. the water would be at a higher temperature than the mercury
 - b. the mercury would be at a higher temperature than the water
 - c. the temperature of water and mercury would be the same

Empirical Formula of a Compound

In research laboratories all over the world, it is the job of the chemist to figure out the chemical formulas of unknown compounds. The chemist's first task is to find out what elements compose the compound. The next step is to find the ratio of the atoms of each element in a molecule of the compound. This enables the chemist to determine the empirical formula of the compound. The empirical formula is the simplest whole-number ratio of atoms in a molecule of the compound. Once the chemist knows the empirical formula of a compound, he or she can try to discover its molecular formula. The molecular formula of a compound is the number of atoms of each element in a molecule of the compound.

Some things you will learn by doing this experiment:

- You will learn what an empirical formula of a compound is, and what a molecular formula of a compound is.
- You will learn how to obtain data experimentally, and how to use these data to determine the empirical formula of magnesium oxide.

Suppose that you are a chemist, and someone gives you 10.00 grams of an unknown compound, asking you to find out what it is. What do you do?

You perform an elemental analysis of the compound and discover that it is composed of Carbon, Hydrogen, and Oxygen. Upon further analysis, you find that from your 10.00 grams of compound you can obtain 4.00 grams of Carbon, 0.67 g of Hydrogen, and 5.33 g of Oxygen. Now look up the atomic weight of each element in the periodic table. You will find that Carbon's atomic weight is 12, Hydrogen is 1.0 and Oxygen's is 16. Using all this information, you find the empirical formula of the compound as follows.

$$\text{Moles of C atoms} = (4.00 \text{ g}) \left(\frac{1.0 \text{ mole}}{12 \text{ g}} \right) = 0.33 \text{ mole}$$

$$\text{Mole of H atoms} = (0.67 \text{ g}) \left(\frac{1.0 \text{ mole}}{1.0 \text{ g}} \right) = 0.67 \text{ mole}$$

$$\text{Mole of O atoms} = (5.33 \text{ g}) \left(\frac{1.0 \text{ mole}}{16 \text{ g}} \right) = 0.33 \text{ mole}$$

Thus, $\text{C}_{0.33}\text{H}_{0.67}\text{O}_{0.33}$ or in terms of simplest whole numbers
 CH_2O (obtained by dividing by the smallest number of moles)

The empirical formula of the compound is CH_2O .

If you wish to find the molecular formula of the compound, you must obtain its molecular weight. Let's say that the molecular weight of this compound has been determined and found to be 180. How can we use this information, along with the information about the empirical formula to obtain the molecular formula of the unknown compound? First, we must determine the formula weight of the CH_2O .

note: The atomic weight of C is 12, H is 1, O is 16.

Formula weight of $\text{CH}_2\text{O} = (1)(12) + (2)(1) + (1)(16) = 30$

The molecular formula is a whole-number multiple of the empirical formula.
Determine what it is for this compound.

$$30x = 180; \text{ therefore } x = 6$$

What did we do here? We used the following logic:

1. The empirical formula is CH_2O
2. The molecular formula must have this same carbon-hydrogen-oxygen ratio; in other words, 1 to 2 to 1, but it must have a molecular weight of 180. This means $(\text{CH}_2\text{O})_x = 180$, $30x = 180$, $x = 6$.
3. Since the formula weight of CH_2O is 30, the molecular formula must have six times as many carbon, hydrogen, and oxygen atoms per molecule as the CH_2O unit. Therefore the molecular formula of the compound is $(\text{CH}_2\text{O})_6$ or $\text{C}_6\text{H}_{12}\text{O}_6$.

In this experiment, you are going to react a piece of magnesium ribbon with oxygen to form a compound. You will do this by heating the magnesium ribbon in the air. Then you will determine the empirical formula of the resulting compound, magnesium oxide, using a method analogous to the one we have just described.

Procedure:

1. Obtain a crucible and cover. Be sure that it is clean and dry. (If it is not, your instructor will explain a procedure to you called "Heating to constant weight.")
2. Weigh the crucible plus cover and record the weight.
3. Get a piece of magnesium ribbon from you instructor. Clean it with a dry paper towel to remove moisture and oil that may have gotten on it from handling. (If the ribbon is not a shiny silver color, you may clean the magnesium ribbon with sandpaper to remove any oxide coating on the surface of the ribbon.)
4. Roll the magnesium ribbon into a loose coil. Do this neatly and carefully. The size of the coil should be such that it will fit the bottom of the crucible.
5. Place the magnesium coil in the crucible and weigh the crucible plus cover plus magnesium. Record this weight in your lab notebook.
6. Compute the weight of the magnesium by subtraction.
7. Obtain a ring stand, an iron ring, a clay triangle, and a gas burner.

8. Set up the equipment so that the crucible is directly over the flame of the Bunsen burner. Check your setup with your instructor before you begin heating.
9. Heat the crucible and its contents, using the hottest part of the burner flame. Be certain that the cover is on the crucible. Heat for 15 minutes.
10. After you have heated the crucible for 15 minutes, obtain crucible tongs and adjust the crucible cover so that a slight opening exists to let more air enter the crucible. Continue heating for an additional 15 minutes.
11. Shut off the gas burner and allow the crucible to cool for 10 minutes. During this initial heating, magnesium metal combined with oxygen in the air to form magnesium oxide. However, also during this initial heating, magnesium metal combined with nitrogen in the air to form magnesium nitride. In other words, the crucible at this point contains a mixture of magnesium oxide and magnesium nitride. This could cause a serious error in the determination, unless the magnesium nitride is converted to magnesium oxide. The procedure which follows solves this problem.
12. Using an eyedropper, add about 40 drops (2 ml) of water to the contents of the crucible.
13. Heat the crucible again with the cover fully on for an additional five minutes. Heat gently at first.
14. Turn off the gas burner. Allow the crucible to cool for 10 minutes and weigh it. Record the weight of the crucible plus cover plus magnesium oxide in your lab notebook. (Note: The crucible should be cool enough to hold in your hand before you weigh it.)
15. During steps 12 and 13, the magnesium nitride converted to magnesium oxide in the following manner:
 - A. Magnesium nitride + water -----> magnesium hydroxide + ammonia gas
 - B. Magnesium hydroxide + heat -----> magnesium oxide + water vapor
16. From the data obtained, calculate the empirical formula of magnesium oxide.
17. Obtain a second set of data from another lab group and record it in the data table. Compare your results with the other lab group's results.

Lab Report #4**Empirical Formula of a Compound**

Name: _____

Partner's Name: _____

Date: _____

	<u>Your Group's Data</u>	<u>Another Group's Data</u>
1. Weight of dry crucible and cover	_____	_____
2. Weight of dry crucible, cover, and metal	_____	_____
3. Weight of metal	_____	_____
4. Weight of dry crucible, cover, and metal oxide	_____	_____
5. Weight of metal oxide	_____	_____
6. Weight of oxygen in metal oxide	_____	_____
7. Metal to oxygen mass ratio	_____	_____
8. Moles of metal	_____	_____
9. Moles of oxygen	_____	_____
10. Metal to oxygen mole ratio	_____	_____
11. Empirical formula	_____	_____

Questions:

1. How do your results support the Law of Constant Composition?
2. What experiment could someone perform to test the Law of Constant Composition further?

3. If you had used a different weight of magnesium metal, would you expect the empirical formula of the compound to be the same? Explain.

4. From the following raw data, calculate the empirical formula of the compound. A chemist heats 50.00 g of sulfur under controlled conditions to produce a sulfur-oxygen compound. The weight of the sulfur-oxygen compound is 100.00 g. What is the empirical formula of the sulfur-oxygen compound?

Unlabeled Bottles

The objective of this lab is to determine, using precipitation and gas formation, the identity of several unlabeled bottles that the lab technician has unfortunately forgotten the contents of. Therefore, you must discover the mysterious contents of the 10 bottles through deductive reasoning and clues. There will be two sets of bottles (A and B) consisting of five bottles each. Each bottle contains a different solution. The original labels have fallen off, but we do know what was in each set. Each set of bottles are numbered A-1, A-2 and so on. To determine what is in each of these bottles you will solve by testing for ionic reactions between various pairs in a given set. You will only use a drop or two of from each bottle placed in the wells of a spot plate. Please do not contaminate, with the dropper or in any other way, the solutions. This could lead to a very poor grade for you, if not bodily injury from the lab instructor.

Hints:

- Wait to see if there is a reaction
- Look closely for precipitates
- Only work with one set at a time
- Be sure to make observations about the unknown solutions.
- Write out all the possible chemical equations for the combinations within a set. Look at a solubility table for clues for which combinations will form solid precipitates.

SET A Contains:

KCl	AgNO ₃	Ba(NO ₃) ₂	CuCl ₂	CuSO ₄
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SET B Contains:

NaCl	ZnCl ₂	NaOH	Na ₂ CO ₃	HCl
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Lab #5**Name:** _____**Pre-Lab Assignment****Date:** _____

Write complete reactions of all the possible combinations within each set. Make sure to indicate which combinations produce precipitates, gases, or any other indications of a reaction. If you need more room, you can finish writing the reactions on the back of this paper.

1. In the set of solutions labeled A, how many will form precipitates when mixed with KCl?
2. In the set of solutions labeled B, which of the solutions will form precipitates when mixed with only two of the other solutions?
3. In Set B, one of the solutions will cause a bubbling reaction when it is mixed with Na_2CO_3 . Which solution is that?

Lab Report #5
Unlabeled Bottles

Name: _____

Partner's Name: _____

Date: _____

Data

In each square indicate your observations such as precipitate, etc.

	A-1	A-2	A-3	A-4
A-5				
A-4				
A-3				
A-2				

				B-2
				B-3
				B-4
				B-5
B-4	B-3	B-2	B-1	

Give the formula of the compound that is in each of the bottles.

1 2 3 4 5

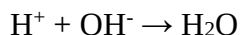
SET A _____

SET B _____

What will happen to the results of this experiment if you, or one of your classmates, contaminate the solutions?

Acid – Base Neutralization

In this lab, we will be mixing acids and bases in an attempt to create a neutral (neither acidic nor basic) solution. The point at which the solution is neutralized is called the equivalence point. At the equivalence point, the number of moles of hydrogen ions from the acid equals the number of moles of hydroxide ions from the base. The equation for the ion reaction is:



If your acid and base only give off one ion (H^+ or OH^-) per molecule, then you can use some simple equations to determine concentration, volume and number of moles. Taking the volume in liters and the concentration in molarity (M) the number of moles of your acid can be determined by the equation:

$$\text{moles} = \frac{(\text{molarity})}{M} \times \frac{(\text{volume})}{V}$$

At the equivalence point the moles of the acid equals the moles of base so the equation:

$$M_A \times V_A = M_B \times V_B$$

holds true. You must use these three equations and a little algebra to determine the concentration of your acid in this experiment.

To determine when you reach the equivalence point we use a chemical called an indicator. The indicator changes color when the pH (a measure of H^+ concentration) changes. Our specific indicator is called phenolphthalein and since all indicators are weak acids we will only use two drops of it so that we do not affect the primary acid-base reaction. This indicator turns from clear to pink when the pH goes from acidic to basic.

Procedure:

1. Obtain a buret and rinse it with distilled water three times. Be sure to handle the buret with great care as they are a precise and expensive instrument.
2. After cleaning, rinse out the buret with the base solution provided one time. Then fill the buret with the base solution.
3. After the buret is full, get rid of any air bubbles trapped at the tip of the buret by opening the stopcock fully. The level of the buret must be below zero. Record the initial (starting) level of your burets to the nearest .01 ml.
4. Place 25.00 ml of acid into a clean (rinsed with distilled water) erlenmeyer flask using a 25.00 mL volumetric pipet. Place two drops of your indicator (phenolphthalein) into the erlenmeyer flask.

5. Titrate (add) base into your erlenmeyer flask until the solution in the flask turns a very, very light pink. Record the level of the base in the buret to the nearest .01 ml. The total volume of base added is the difference between the initial reading and the final reading.
6. Pour the neutralized solution down the drain and rinse your flask with distilled water.
7. Repeat steps 3 through 6 for two more trials. Do not forget to add phenolphthalein.
8. Clean up.
9. Calculate the concentration of your acid using the volume totals for each titration. Average the three concentrations and get your best calculated figure. Compare this answer with the true value which is on the board.

Lab Report #6**Acid – Base Neutralization**

Name: _____

Partner's Name: _____

Date: _____

Data and Calculations:*Show all of your calculations!*

Concentration of Base _____

	<u>Trial One</u>	<u>Trial Two</u>	<u>Trial Three</u>
Acid volume	_____	_____	_____
Initial base reading	_____	_____	_____
Final base reading	_____	_____	_____
Total base added	_____	_____	_____
Acid concentration	_____	_____	_____

Average Acid Concentration _____

True Acid Concentration _____

Percent Error _____

Provide an explanation for why your calculation is different from the true acid concentration.