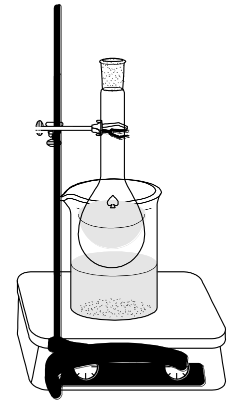
Chemistry M12
Laboratory Manual



Compiled by the Department of Chemistry at Moorpark College

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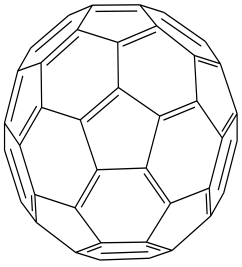

Robert Curl, Richard Smaley, and Harold Kroto were awarded the Nobel Prize in Chemistry in 1996 for the discovery of the soccer ball-shaped molecule C_{60} shown to the left, known as buckminsterfullerene. This fundamental molecule represents the first of a new series of carbon allotropes, a form of an element that differs in the way the atoms are linked.

Moorpark College Department of Chemistry
Chemistry 12 Lab Manual Table of Contents

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Experiment 1 – Separation of Copper(II) Sulfate from Sand

Discussion

Mixtures are a combination of substances in which the components keep their individual characteristics. Mixtures have variable proportions and can be separated by simple physical means. The mixture’s components have different physical properties like melting point, boiling point, or solubility that allow us to selectively remove individual components from the mixture. Once separated, the percentage of each component in the original mixture can be calculated.

In this experiment, you will separate a mixture of copper(II) sulfate and sand using the physical property of solubility. You will learn about certain methods of separation that include decantation, filtration, and evaporation. Finally, the Law of Conservation of Mass will be applied to check the validity of your final calculations.

Procedure

1. Weigh about 4 –5 grams of the $CuSO_4$/sand mixture in a 100 mL beaker on the laboratory balance by taring (your instructor will explain and demonstrate).

2. Add 10–15 mL of D.I. water to the beaker, and swirl. Next, weigh and record the weight of a piece of filter paper AND an evaporating dish separately. Then assemble the filter apparatus as demonstrated by the instructor, filter the mixture, and collect the filtrate (liquid) onto the evaporating dish. Use your wash bottle (filled with D.I. water) to transfer all the undissolved solid from the beaker to the filter paper. After all the liquid has drained through the filter, wash the filter with small portions of D.I. water from the wash bottle until the washings are colorless. Again, make sure you collect ALL the filtrate and washings in the evaporating dish. Try to use small amounts of water. You will be evaporating the liquid; the more water you add, the longer it will take!

3. Prepare a steam bath by placing a 250 mL beaker, 2/3 full of water with 4 – 5 boiling chips, on a wire screen on a ring stand. Place the evaporating dish carefully on the beaker and heat the water to boiling. Heat the steam bath until the filtrate has completely evaporated. Do NOT let the steam bath boil to dryness; you may have to refill the water in the beaker occasionally. If the beaker goes dry, it must be cooled before adding water to prevent the beaker from cracking!

4. Open the filter paper onto a large watch glass that has been labeled with your drawer number using a permanent marker. Dry the sand on the filter paper/watch glass in the drying oven, set at 150 °C, for 20 minutes. Afterwards, carefully remove the watch glass, allow it to cool, and weigh the filter paper with the sand by placing it into a TARED weighing boat on the analytical balance. Calculate the mass of your sand.

5. Once your filtrate has completely evaporated and your salt remains on the evaporating dish, weigh the evaporating dish with the salt sample. Compute the weight of the $CuSO_4$ sample by difference.

6. Calculate the weight percent $CuSO_4$ and sand in the sample.

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Data and Calculations for Experiment 1

Mass of $CuSO_4$/sand mixture ______________

Mass of empty evaporating dish ______________

Mass of evaporating dish and dry $CuSO_4$ ______________

Mass of $CuSO_4$ ______________

Mass of filter paper ______________

Mass of filter paper and sand ______________

Mass of sand ______________

Total mass of recovered sand and $CuSO_4$ ______________

Calculated total percent recovery ______________

Percent by mass of $CuSO_4$:

Show Calculation ______________

Percent by mass of sand:

Show Calculation ______________

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Questions

1. Many students do NOT recover 100% of the original mixture. Describe at least TWO possible problems that could cause LESS than 100% recovery of the mixture.

2. A student obtained the following data:

Mass of beaker	25.87 g
Mass of beaker with mixture sample	28.12 g
Mass of evaporating dish	146.36 g
Mass of evaporating dish with dried salt	147.10 g
Mass of beaker with dried sand	???

However, this student spills her sand sample out of the evaporating dish before weighing it. If the student believes in the Law of Conservation of Mass, what should have been the weight of the beaker with the dried sand in it? Show all your work.

3. A student receives a sample of a mixture with three components: (1) solid iodine that is first removed from the mixture by evaporation, (2) solid salt that is dissolved to separate it from the third component, and (3) solid sand. The salt and sand are dried and weighed, but the iodine escapes as a gas and is not recovered. The student starts with 4.25 g of the mixture and recovers 1.16 g of salt and 2.40 g of sand. What is the percent of each component in the original mixture? Show all your work.

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Experiment 2 – Measurements

Discussion

Experimental sciences, such as Chemistry, depend on making and using measurements properly. The SI system of units (sometimes called the metric system) is used almost exclusively. This system is very similar to our monetary system: $1 = 10 dimes = 100 cents = 1000 mils. In chemistry, the basic units of length, mass, and volume are the meter, gram, and liter, respectively. They all are divided the same way. For example, 1 meter = 10 decimeters = 100 centimeters = 1000 millimeters. The kilo is also commonly used; it equals 1000 of the basic unit. For example, 1 kilogram = 1000 grams.

Often you will be asked to compare you experimental or calculated value to an "accepted" or theoretical value. The closer you are to the accepted value, the greater the accuracy of your experiment. The accepted value could be located in a reference such as the Handbook of Chemistry and Physics. Percent error is a common method for reporting accuracy, where a smaller percent error represents a more accurate experimental value. Percent error is calculated as:

$$\text{Percent error} = \left|\frac{\text{Experimental} - \text{Accepted}}{\text{Accepted}}\right| \times 100\ \% \qquad (1)$$

In science, an experimenter is allowed to estimate one more digit past what can be measured exactly on an instrument. For example, if the smallest lines on a ruler are centimeters, and an object's length falls between 2 lines, more precision is gained by estimating between the lines. Therefore, the length of an object might be reported as 25.5 cm. The reported numbers are called "significant figures", and the more precise the instrument, the more significant figures it can produce.

A calculation cannot be any more precise than the least precise measurement. For example, density is calculated by dividing the mass of an object by its volume. Therefore, the density of an object might be 23.57 g/ 4.2 mL = 5.61190476 g/mL. But, the least precise measurement (the volume) only has a precision of 2 significant figures. Therefore, the density must be reported as 5.6 g/mL.

Procedure

Record your data on the report form as you complete the measurements.

A. Temperature

5 beakers with thermometers have been set up for you: (1) room temperature water, (2) boiling water, (3) a mixture of ice and water, (4) a stirred mixture of ice and water, and (5) a stirred mixture of ice, water and salt. Observe and record all temperatures to the nearest 0.1 °C.

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B. Mass

When using any measuring device, never round off your raw data. If the balance fluctuates on the last digit, estimate that value. Weigh a (1) 100-mL beaker, (2) a 250-mL Erlenmeyer flask, (3) a plastic empty weighing boat, and (4) and then add approximately 2 grams of sodium chloride to the weighing boat. Calculate the mass of the sodium chloride added.

C. Length

Using a metric ruler, measure the following in centimeters, remembering to estimate one extra digit: (1) the length of the double arrow on the report sheet, (2) the length of the external height of a 250-mL beaker, and (3) the length of a medium sized test tube.

D. Volume

The graduated cylinder is the most accurate equipment in your locker for measuring volume and can give a precision of 0.1 mL. Water is attracted to the glass sides of the cylinder, causing a curved effect called the meniscus. The cylinder should be read at eye level using the bottom of the meniscus. In theory, a 250-mL Erlenmeyer flask with a marking for 200 mL should have a volume of 200 mL at that mark! However the problem is that volumes marked on beakers and flasks are only approximate values. Therefore, fill a 250-mL Erlenmeyer flask to the 200 mL mark with water, transfer this volume of water to a 250-mL graduated cylinder, and determine the exact volume.

It is often convenient to estimate volumes of 5 and 10 mL simply by observing the height of a liquid in a test tube. Use your graduated cylinder to place 5 and 10 mL of water in a medium-sized test tube and measure the heights in cm.

E. Density

Density measures the "compactness" of material. For example, lead has a high density, and Styrofoam has a low density. Mathematically, this compactness is expressed as mass per unit volume. In chemistry, we use grams and milliliters: d = g/mL. Density is an intrinsic value; it does not depend on the amount of sample taken. We will take advantage of this by measuring the density of various sample sizes and averaging their densities:

1. Obtain 5 pieces of the same object and record its name on the data sheet.
2. Weigh each object and record the values on the data sheet.
3. Choose the appropriate size graduated cylinder (smallest size that will hold the object plus enough water). Add enough water to the graduated cylinder be able to cover your largest sample. Record the volume to the highest precision (0.1 mL or better).
4. Carefully add the sample to the graduated cylinder. There are two things to watch out for: breaking the cylinder and splashing water out. Tilting the cylinder and gently sliding the object in minimize both of these risks. Record the new volume.
5. Repeat with each sample piece. If the sample pieces are small, the pieces can remain in the graduated cylinder until all sample pieces have been added to the cylinder.
6. Determine the density of each piece and the average density.
7. Using Microsoft Excel®, graph the cumulative data. Use the largest values of mass and volume to determine your x and y scales.

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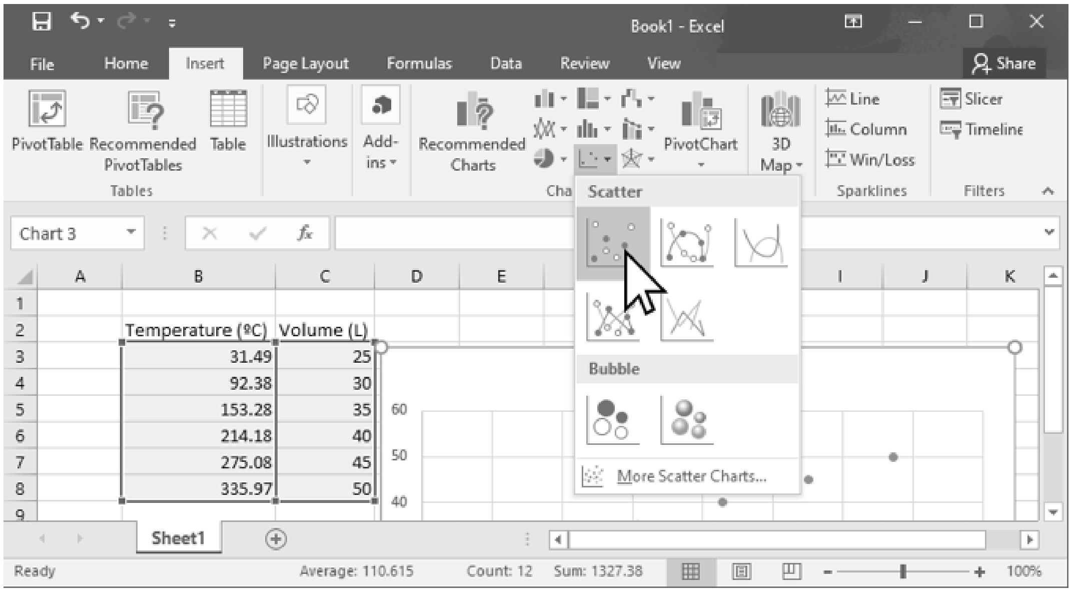
Excel® Procedure

Note that various versions of Excel® may function a bit differently from the directions outlined below (which work on department-owned laptop computers):

Put the title for your x-axis (include units) in one Excel® cell (box). In the cell to the right, put the title for your y-axis. Using these boxes as headings, input the numeric data (like a table) in the cells under these titles (each box should contain one number; each row represents one data point in x,y format). Click and drag your mouse to highlight just the numeric boxes. From the "Insert" tab, choose a "Scatter" plot. (See example, below.)



Temperature (ºC)	Volume (L)
31.49	25
92.38	30
153.28	35
214.18	40
275.08	45
335.97	50

Your graph must include a meaningful Chart Title and Axis Titles (with units). These Chart Elements can be added to your graph by clicking on the "+" icon in the upper right corner of your graph. Your instructor may request additional Chart Elements.

To add a Trendline, right click on any data point on your graph and choose "Display Trendline" from the menu that appears. The format trendline pane will appear on the right side of your screen. Linear should be selected by default. From this pane, you should check the box next to "Display Equation on chart." Your instructor may also ask you to check the box for "Display R-squared value on chart."

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Data and Calculations for Experiment 2

Measurements

A. Temperature
1. Water at room temperature ______________ °C

2. Boiling point ______________ °C

3. Ice water
Unstirred ______________ °C

Stirred ______________ °C

4. Ice water with salt added ______________ °C

B. Mass
1. 100 mL beaker ________________ g

2. 250 mL Erlenmeyer flask ________________ g

3. Weighing boat ________________ g

4. Mass of weighing boat + sodium chloride ________________ g

Mass of sodium chloride (show calculation setup) ________________ g

C. Length
1. Length of ⟷ ______________ cm

2. Height of 250 mL beaker ______________ cm

3. Length of test tube ______________ cm

D. Volume

1. 200 mL mark (from Erlenmeyer flask) water
transferred to graduated cylinder _____________ mL

2. Height of 5.0 mL of water in test tube _____________ cm

3. Height of 10.0 mL of water in test tube _____________ cm

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E. Data Sheet for Density of an Object

Name of Object: ________________________________

							Using Microsoft Excel®, graph the following:	
Sample #	Object Mass (g)	Initial mL H_2O	mL H_2O w/ Object	Volume object (mL)	Density (g / mL)	Cumulative Sample #s	Cumulative volume (mL) (x-axis)	Cumulative object mass (g) (y-axis)
1						1		
2						1 + 2		
3						1 + 2 + 3		
4						1 + 2 + 3 + 4		
5						1 + 2 + 3 + 4 + 5		

Average Density from Table = _________

Average Density from Graph (slope of line) = ___________

Be sure to show your properly formatted graph to your instructor to receive credit for this part of the experiment (or print your graph and attach it to this report).

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Questions

1. Which would work better in this experiment as an unknown solid whose density is to be determined, wood chips or small quartz rocks? Explain your choice.

2. Why is it best to use a smaller graduated cylinder as opposed to a larger graduated cylinder for this experiment?

3. How well does the average density from the table and density from the slope of the graph compare? Which value is closer to the accepted density of your metal? (Refer to the *Handbook of Chemistry and Physics*). Calculate the percent error between your better value and the handbook value.

4. What is the density of a 9.343 gram piece of metal that causes the level of water in a graduated cylinder to rise from 5.1 to 8.1 mL when the metal is submerged in the water? Consider significant figures when doing the calculation.

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Experiment 3 – Graphing Freezing Points

Discussion

The freezing and melting of a substance occur at the same temperature, which is measured when the solid and liquid phases are at equilibrium (therefore, the terms "melting point" and "freezing point" can be used interchangeably). Accurate measurement of this temperature is useful for identification purposes, since few chemicals share exactly the same melting point. The unofficial definition of freezing point is the temperature where crystals first start to form. Officially, freezing point is determined by graphing a "cooling curve", which is a plot of temperature versus time as a substance is being cooled. The temperature will decrease at first as the liquid is cooled; then the temperature will remain constant during the phase change from liquid to solid. This "plateau" is the freezing point. When all the liquid has frozen, the temperature will again decrease.

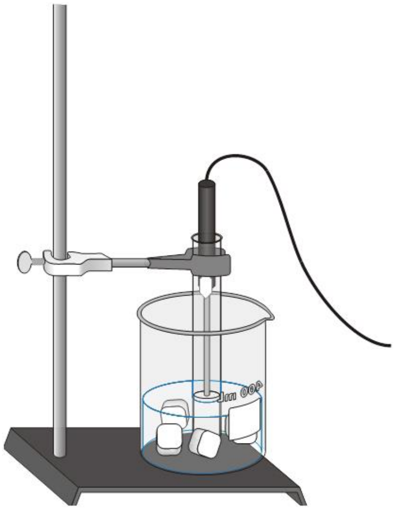
An impurity will disrupt the crystal structure of a substance; therefore, it must be cooled below its normal freezing point to achieve crystallization. This depression of the freezing point has some interesting applications: measuring the amount of impurity, identification by mixing a known with an unknown substance, lowering the freezing point of water using antifreeze, etc.

Occasionally, a substance will become colder than its normal freezing point and still remain a liquid. This unstable condition is called "supercooling". "Shocking" the system by stirring or jarring usually will cause crystallization and the substance will rapidly rise to its normal freezing point.

The goal of this experiment is to determine freezing points by graphing two cooling curves: (1) pure (or glacial) acetic acid with stirring – this usually eliminates supercooling; and (2) pure acetic acid without stirring – supercooling will probably occur.

Procedure

Trial 1: With Stirring



1. Fasten a test tube clamp to a stand. Take a small-sized test tube and a stopper to the fume hood. Add 5 mL of acetic acid to the test tube and close the test tube with the stopper before removing from the fume hood.
2. Place the test tube in the test tube clamp to minimize spillage. Adjust the temperature of the acetic acid to about 25 °C by warming or cooling the test tube in a beaker of water.
3. Prepare your cooling mixture by placing about 150 mL of ice in a 250 mL beaker; then add water to barely cover the ice.

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4. Obtain the following equipment: Chemistry Department laptop (with Logger Pro software), Vernier LabQuest Mini, USB cable, thermometer probe. (***Note:*** *an alternate procedure is available that uses LabQuest 2 devices instead of laptop computers.)*

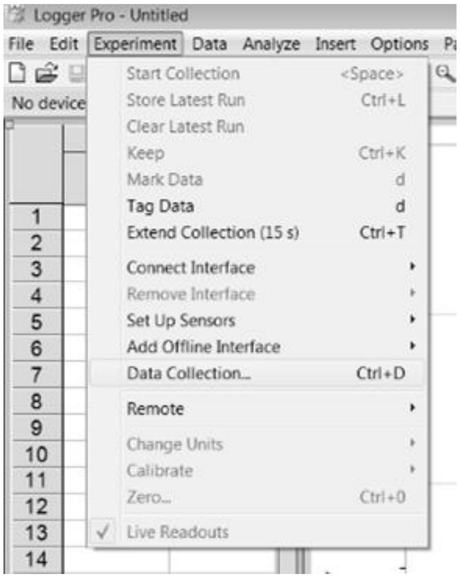
5. Carefully insert the thermometer probe into the stopper & be sure the metal probe is touching the acetic acid.

6. Connect the thermometer probe to the LabQuest Mini (any CH) and then use the USB cable to connect the LabQuest Mini to the laptop.

7. Open the Logger Pro software.

Logger Pro 3.11

8. Click on the Experiment menu and select Data Collection… Set the Duration: 10 minutes and the Sample Rate: 5 samples / minute and click Done.



Mode: Time Based

☐ Repeat

☑ Sample at Time Zero

Duration: 10 minutes

Triggering is disabled

☐ Continuous Data Collection

Sampling Rate: 5 samples/minute 0.2 minutes/sample

☐ Oversampling

Samples to be Collected: 51

9. When you are ready to start the Trial, lower the test tube clamp until the acetic acid in the test tube is below the surface of the ice water and click the green Collect button.

10. Use the thermometer as a gentle stirrer during this trial. Continue recording temperatures for 10 minutes where Logger Pro will stop automatically. If the temperature has stabilized for several minutes (flat part of the graph) and the acetic acid is completely frozen, you may click the red Stop button to end the run.

11. To prepare the acetic acid for your next run, raise the test tube clamp and remove from the ice bath. Replace with a beaker of warm water to melt back to about 25 °C.

12. On the displayed graph, analyze the flat part. Press the mouse button and hold it down as you drag across the flat part to select it. Click Statistics. The mean temperature value for the selected data is listed in the statistics box on the graph. Record this value as the freezing temperature (in the Post Lab Questions). To remove the statistics box, click the upper-left corner of the box.

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Trial 2: Without Stirring

1. Prepare Logger Pro for your next trial: From the Experiment menu, choose Store Latest Run (to save for later comparison). You may either leave Trial 1 on the screen (Trial 2 data will overlay in another color of line) or you may hide the curve by: click the Temperature label of the graph, click More, and uncheck the Run 1 Temperature box. Click OK.

2. Replenish the ice bath if necessary.

3. Repeat the same timing procedure as Trial 1 but do not stir.

4. If the temperature falls to about 4 ^{o}C without forming crystals, you are observing supercooling. If supercooling persists for more than 2 minutes, “shock” the solution by moving the thermometer until crystals form.

5. Continue recording the temperature for the full 10 minutes or until completely frozen. When stopped, be sure to Store Latest Run.

6. On the displayed graph, analyze the flat part using the Statistics function again and record this value as the freezing temperature in the Post Lab.

Completing the Graphs for your Experiment:

1. If not already both displayed, click the Temperature axis label, click More, and check the Run 1 Temperature and Latest Temperature boxes. Click ok.

2. Label both curves by choosing Text Annotation from the Insert menu, and “Stirring” (or “Without Stirring”) in the edit box. Then drag each box to a position near its respective curve. Adjust the orientation of the arrowhead by clicking and dragging to the desired position.

3. You will need to show both Trial 1 & Trial 2 graphs to your instructor by bringing the CHEM Dept. laptop to them for credit on this experiment.

Hazardous Waste

Thaw the acetic acid with a beaker of warm water and then remove the temperature probe. Using your wash bottle filled with water, rinse the temperature probe off into a beaker from your drawer. Take the stoppered test tube, your wash bottle, and the beaker to the fume hood. Pour the acetic acid out of the test tube into the provided “Recycled Acetic Acid” bottle. Before you leave the fume hood, use your wash bottle to rinse the acetic acid out of the test tube into your beaker. ***Please do not contaminate the recycled acetic acid with water!*** You can then pour the water that is in your beaker down the sink drain and wash your test tube and beaker before you put them away.

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Post-Lab Questions

Use your graphs to answer the following questions. Note: The accepted freezing point of acetic acid, according to the *CRC Handbook of Chemistry and Physics*, is: 16.6 °C

1. a. What is the experimental freezing point of acetic acid in Trial 1? ________________

   b. What is the experimental freezing point of acetic acid in Trial 2? ________________

   c. Calculate the percent error in the Trial 1 measurement of the freezing point:

   SHOW CALCULATION:

   ____

2. What is supercooling? Did you observe supercooling in your experiment? Explain.

3. What is the difference between melting and freezing a substance?

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Experiment 4 – Hydrates and Anhydrous Salts

Discussion

When a salt incorporates a certain number of water molecules for each formula unit in order to create a stable complex substance, we have a hydrate. These compounds are not “wet”; the water molecules are incorporated into the solid structure in a fixed pattern. The following are some representative examples:

$MgCl_2 \cdot 2H_2O$, $Na_2SO_4 \cdot 10H_2O$, $CoCl_2 \cdot 6H_2O$, $BaCl_2 \cdot 2H_2O$, $CuSO_4 \cdot 5H_2O$

Notice that the formulas all begin with the salt and a dot separates the number of water molecules attached per formula unit. The water is chemically bound to the substance with weaker bonds than the ionic attraction within the salt or the covalent bonds within the water molecule.

In naming hydrates, you use the nomenclature rules you have learned to name the salt and add an appropriate prefix (mono = 1, di = 2, tri = 3, tetra = 4, penta = 5, hexa = 6…) to identify the number of water molecules ending the name with the word hydrate. For example, $CuSO_4 \cdot 5H_2O$ has the name copper(II) sulfate pentahydrate.

The water chemically bound in hydrates can generally be removed by heating without destroying the salt or the molecule of water. When the water has been completely removed, we call the remaining compound the anhydrous salt (*salt without water*). Anhydrous salts will readily absorb water from its surroundings and may be used as moisture detectors. A common example is anhydrous cobalt(II) chloride, $CoCl_2$ (blue in color). When in contact with moisture, this substance absorbs water to create cobalt(II) chloride hexahydrate, $CoCl_2 \cdot 6H_2O$ (pink in color).

In this experiment, hydrated salts will be heated. First, we will perform a qualitative test with the hydrate $CuSO_4 \cdot 5H_2O$ (blue in color), whereby heat is applied until the hydrate becomes anhydrous $CuSO_4$ (white in color). Next, the resulting liquid will be tested with anhydrous $CoCl_2$ to detect if that liquid is water. Finally, we will also perform a quantitative experiment with an unknown hydrate to determine the mass percent of water lost and solve for the formula and name of the unknown hydrate.

Procedure

A. Qualitative Determination of the Released Liquid after Heating $CuSO_4 \cdot 5H_2O$

1. Weigh out approximately 3 grams of copper(II) sulfate pentahydrate into a clean, dry large ignition tube (25 x 200 mm). Make sure the hydrate is not sticking to the sides; it should all be near the bottom of the tube.

2. Clamp the ignition tube at a 15–20 degree downward from the horizon so that the mouth of the tube is lower than the bottom of the tube where the hydrate is located. The clamp should be near the mouth of the test tube.

3. To collect a sample of the liquid that will be removed from the hydrate, place a medium test tube held upright in an Erlenmeyer flask just below the mouth of the ignition tube containing the hydrate. Your instructor will demonstrate this set-up.

4. Heat the hydrate with a Bunsen burner gently at first, and gradually increase the intensity of the heat. Note the changes that occur. Collect some liquid that condenses in the lower, cooler mouth of the tube. Continue heating until most of the blue color has disappeared. Your residue should now be pale blue or white (*not black, see note*). To remove the remaining liquid still inside the ignition tube, gently warm the entire tube to vaporize. Allow tube and contents time to cool. Observe and record the appearance and odor of the collected liquid. NOTE: at temperatures above 600 °C, the copper(II) sulfate decomposes; sulfur trioxide is driven off and black copper(II) oxide remains as residue.

5. Obtain a couple pieces of anhydrous cobalt(II) chloride test strips. Place distilled water on one and your experiment liquid on the other. Compare and record observations.

6. Empty the anhydrous salt from your cooled ignition tube onto a watch glass and divide it into two sections. To one section, add a few drops of distilled water; to the other, add your experiment liquid. Compare and record observations. Dispose of residue in the waste container provided.

B. Quantitative Determination of Mass Lost in a Hydrate

1. Obtain an unknown hydrate sample. Record the unknown number.

2. Weigh a clean dry crucible with its cover to the highest precision. Be sure that you use the same balance for all mass measurements within the same experiment to cancel out systematic errors.

3. Add between 2–3 grams of the unknown into the crucible with cover and weigh to the highest precision.

4. To a ring stand, attach a ring and place a clay triangle on the ring. Transfer the crucible onto the clay triangle and adjust the cover so it is slightly ajar to allow liquid vapor to escape. Very gently heat with a Bunsen burner for about 5 minutes. Readjust the flame to continue heating, more intensely now so that the bottom of the crucible glows red, for another 12 minutes.

5. After this two-step heating process (1st heating) is complete, turn off the burner, close the crucible lid, and allow the crucible to cool for about 10 minutes before weighing. It is important to have all mass measurements made at or near room temperature. Do not transfer the hot crucible to the tabletop; it is too hot and will permanently mark the table. Use the crucible cooling plates that are provided. Weigh your sample.

6. Heat the covered crucible and contents another 6 minutes at maximum temperature; cool and reweigh (2nd heating). If the results agree within ±0.050 g, you are done and will not need to have a 3rd heating. If the difference is greater than ±0.050 g, repeat the heating for another 6 minutes (3rd heating).

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Data and Calculations for Experiment 4

A. Qualitative Determination of the Released Liquid

1. Record observations regarding the solid before, during, and after heating the copper(II) sulfate pentahydrate.

2. Compare and record observations after adding liquid to the anhydrous cobalt(II) chloride test strips.

3. Compare and record observations after adding liquid to the residue on the watch glass.

4. What conclusions can you draw from the above observations?

5. Write the balanced chemical equation for the decomposition of copper(II) sulfate pentahydrate, include phases.

B. Quantitative Determination of Mass Lost in a Hydrate

Sample number: ____________________

1. Mass of crucible and cover ____________________
2. Mass of crucible, cover and sample ____________________
3. Mass of crucible, cover and sample after 1st heating ____________________
4. Mass of crucible, cover and sample after 2nd heating ____________________
5. Mass of crucible, cover and sample after 3rd heating ____________________
6. Mass of sample after final heating ____________________
7. Mass of original sample ____________________
8. Total mass lost by sample ____________________

9. Percentage of water in sample ____________________

Name: ______________________________ Section: ____________________

10. Ask your instructor for the name of the anhydrous salt of your residue and solve for the formula and name of your original unknown hydrate.

Formula: ____________________ Name: ______________________________

11. Is it possible that the decrease in mass from heating is something other than water?

Yes or *No* Explain and include an example.

Pre-Lab (to be completed before coming to lab)

A student heated a hydrated salt sample with an initial mass of 2.244 grams. After the second heating, the mass had decreased to 1.798 grams. Make the assumption that all the lost mass is water.

a) Solve for the mass lost.

b) Calculate the percentage of water in the original hydrated salt sample.

c) The instructor has informed you that the sample is a hydrate of $ZnSO_3$. Use the information above and molar masses from the periodic table to solve for the moles of $ZnSO_3$ in the residue and the moles of H_2O lost.

d) Solve for the X in the formula of the hydrate of $ZnSO_3 \cdot XH_2O$ and name it.

$$\text{Moles of water / moles of } ZnSO_3 = X$$

Formula: ____________________ Name: ______________________________

Name: ________________________________ Section: ____________________

Experiment 5 – Properties of Solutions

Discussion

In today's lab, you'll investigate the qualitative nature of solutions. The first step is learning some common terms.

Solute refers to a compound that **dissolves** in a solvent to form a solution. A **solution** can have one or more solutes, but only one solvent. The **solvent** is the compound that is predominant in the solution. A solute is said to be dissolved when it forms a clear, but not necessarily colorless, liquid. Thus, sugar dissolves in water, but fine sand and dust form **suspensions** which are not true solutions.

Solvents can be sorted by their **polarity**. Water is very polar, while benzene, decane, and gasoline are considered **non-polar**. The term organic solvent refers to most solvents other than water that are carbon-containing. Organic solvents can be either polar or non-polar, depending upon their structure. For example, methanol and ethanol are polar organic solvents, while ether and acetone are less polar, and decane and benzene are considered non-polar organic solvents.

Solubility is a measure of how much of a compound can eventually dissolve in a solvent. If a solid does not dissolve, the compound is said to be **insoluble**. It can also be described as slightly soluble, moderately soluble, or very soluble. If the compound is a liquid (not a solid) it can dissolve and is described as **miscible,** or instead forms two layers and is called **immiscible**. Ethanol and water are miscible, while oil and water are immiscible.

Concentration refers to the amount of solute relative to the total volume of solution. A dilute solution has little solute per 100 grams of solution, while a concentrated solution has more solute. A solution is considered saturated when no more solute can dissolve in that solution without it precipitating thereafter.

A supersaturated solution is a solution that holds more solute than it normally can hold at that temperature. Given time, some solute will precipitate out of solution. In other words, the solution is unstable over time.

Concentration can be measured using several terms. "Proof" is used to measure alcohol content in liquor and beer. Chemists tend to use mass percent and molarity which are defined below. Remember that mass percentages range from 0 to 100%, and molarities are generally less than 18 M. Very few compounds can form solutions with higher concentrations.

$$\text{Mass Percent of } \mathbf{X} = \frac{\text{mass of } \mathbf{X}}{\text{mass of } \mathbf{X} + \text{mass of solvent}} \times 100\ \%$$

Name: ________________________________ Section: ____________________

Procedure

A. Concentration of a Saturated Solution.

In this section, you'll figure out how many grams of potassium chloride per mL of solution were present in a pre-made saturated solution of KCl.

1. Weigh a clean, dry evaporating dish. In this dish, add 6.0 mL of solution and weigh again. Place the dish in a 250-mL beaker of boiling water. Evaporate the solution until a white solid is present in the dish. Don't let the boiling water bath go dry. This step will take approximately half an hour.

2. Remove the dish from the boiling water with tongs. Place the dish on a wire mesh and gently heat with a Bunsen burner. If you heat too strongly, the solid may "pop" and you will lose some.

3. Let the dish cool until it can be touched safely. Weigh the dish to find out how many grams of potassium chloride are present.

Cleanup: Wash the solid down the drain.

B. Relative Solubility of a Solute

In this section, you will determine whether iodine, a reddish solid, dissolves better in water or decane.

1. Take a test tube and add about 5 mL of water and 2 mL of decane. Stopper the test tube and give it a gentle shake. Note which layer was on top.

2. To this tube, add 5 mL of saturated iodine-water solution. Gently shake again and see which layer has more color.

Cleanup: Empty the test tube into the waste labeled "Decane Waste".

C. Miscibility of Liquids

In this section, you will find out what liquids are miscible with water.

Take three dry test tubes and add the following pair of liquids. Stopper the test tubes and gently shake them. Are there two layers or one?

1. 1 mL of kerosene and 1 mL of isopropyl alcohol
2. 1 mL of kerosene and 1 mL of water
3. 1 mL of isopropyl alcohol and 1 mL of water

Dispose of the first two kerosene mixtures in the "Kerosene Waste" container.

Name: ______________________________ Section: ____________________

D. Effect of Particle Size on Rate of Dissolution

1. Fill a test tube with about 0.5 cm of fine crystals of sodium chloride. Fill a second test tube with about 0.5 cm of coarse crystals of sodium chloride. Add 10 mL of water to each tube and shake both tubes an equal number of times. Shake both tubes equally. Time how long it takes to dissolve each.

These solutions can be disposed of down the sink.

E. Effect of Temperature on Dissolution

1. Weigh out two 0.5 g samples of fine sodium chloride crystals. Take two 250-mL beakers and add 50 mL of water to them. Heat one of the beakers to boiling, then let it cool for one minute.

2. Add the salt samples to each beaker and time how long it takes to dissolve each.

3. As soon as the salt dissolves, gently swirl the hot water and observe the denser salt layer in the bottom of the flask. Repeat the process with the cold water.

   These solutions can be disposed of down the sink.

F. Solubility versus Temperature; Saturated and Unsaturated Solutions

1. Weigh out 1.0 g of NaCl and 1.0 g of NH_4Cl and place them in separate, labeled test tubes and add 5 mL of water. Stopper the test tubes and shake the tubes until the salts dissolve.

2. Add another 1.4 g of NaCl to the NaCl solution, and another 1.4 g of NH_4Cl to the NH_4Cl solution. Stopper and shake the tubes for 3 minutes. Note whether or not the salts dissolved.

3. Remove the stoppers and place both tubes in a beaker of boiling water, gently shaking occasionally, and note the results after 5 minutes.

4. Remove the tubes and cool with running tap water for one minute and record your observations. Let the solutions stand for a few minutes and record your observations.

Pour the solutions down the drain.

Name: ________________________________ Section: ____________________

G. Ionic Reactions in Solution

1. Place a small lump of pea-sized quantities of a) barium chloride, b) sodium sulfate, c) sodium chloride, and d) barium sulfate into four separate labeled test tubes.

2. Add 5 mL of water, stopper the tubes, and shake them. Which sample(s) do(es) not dissolve?

3. Mix the barium chloride and sodium sulfate together and note the results.

4. Write an equation that describes the results of these test tubes being mixed.

Dispose of all solutions in the “Barium waste” container.

Name: ______________________________ Section: ____________________

Data and Calculations for Experiment 5

A. Concentration of a Saturated Solution (record all masses as x.xxx g)

1. a) Mass of evaporating dish ______________

   b) Mass of evap. dish and potassium chloride solution ______________

   c) Mass of evap. dish and residue ______________

2. Calculate: (show setups)
   a) Mass of potassium chloride solution ______________

   b) Mass of residue ______________

   c) Mass of water in potassium chloride solution

   ____

   d) Mass percent of potassium chloride in the solution

   ____

   e) Grams of potassium chloride per 100 g of water in the solution

   ____

B. Relative Solubility of a Solute in Two Solvents

1. a) Which liquid is denser, decane or water? ______________

   b) How did you decide which layer was water?

2. What is the color of iodine in water? ______________

   What is the color of iodine in decane? ______________

3. Which solvent dissolves more iodine? How did you decide this?

Name: ________________________________ Section: ____________________

C. Miscibility of Liquids

1. Which liquids were miscible with each other?

2. Which liquids were immiscible with each other?

D. Particle Size and Dissolution Rates

1. How long did it take the fine salt crystals to dissolve?

2. How long did it take the coarse salt crystals to dissolve?

3. Based on these observations, how does particle size affect the rate at which a substance is able to dissolve?

E. Temperature and Dissolution Rates

1. How long did it take the salt crystals to dissolve in hot water?

2. How long did it take the salt crystals to dissolve in cold water?

3. Based on these observations, how does temperature affect the rate at which a substance is able to dissolve?

F. Temperature and Solubility

1. Was the solution with 1.0 g of NaCl in 5.0 mL water saturated at room temperature?

2. Was the solution with 1.0 g of NH_4Cl in 5.0 mL water saturated at room temperature?

3. Was the solution with 2.4 g of NaCl in 5.0 mL water saturated at room temperature?

4. Was the solution with 2.4 g of NH_4Cl in 5.0 mL water saturated at room temperature?

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5. Which salt was least soluble at higher temperatures?

6. At the higher temperatures, was the NaCl solution saturated?

7. At the higher temperatures, was the NH_4Cl solution saturated?

8. What happened to the NaCl solution when it was cooled back to room temperature?

9. What happened to the NH_4Cl solution when it was cooled back to room temperature?

10. Solubility is defined as the amount of solute that can dissolve in a given quantity of solvent. Based on your observations in this part, how does temperature affect the solubility of solid solutes? Does it affect different substances in identical ways?

G. Ionic Reactions in Solution

1. Write the formulas for the following substances. Include states of matter (e.g. $_{(aq)}$ or $_{(s)}$) based on the results of your experiment:

   barium sulfate ________________________

   barium chloride ________________________

   sodium sulfate ________________________

   sodium chloride ________________________

2. Write the equation that shows the reaction of barium chloride and sodium sulfate. Use state indicators (e.g. $_{(aq)}$ or $_{(s)}$) for all compounds.

3. Which compound is the white precipitate? How do you know this?

Name: ________________________________ Section: ____________________

Experiment 6 – Decomposition of Potassium Chlorate

Discussion

The percentage composition of a compound is the percent by mass of each element in the compound. This can be theoretically calculated using atomic masses found on the periodic table, and experimentally calculated from lab data. The percent composition is calculated by taking the mass of the part divided by the mass of the whole times 100.

In this experiment, solid potassium chlorate will be strongly heated. This causes it to decompose into solid potassium chloride and oxygen gas. The relative amount of potassium chloride (residue left behind) and oxygen gas (mass lost) are experimentally measured and compared to the theoretical values.

Percent error is calculated by finding the absolute value of the difference between the theoretical and experimental values and dividing that number by the "true" theoretical value and last multiplying the answer by 100.

Following the quantitative test, we will perform a qualitative test on the residue and known compounds of $KClO_3$ and KCl using silver nitrate solution. $AgNO_3$ solution is used as a general test for chloride ions. A positive test forms white solid AgCl and indicates the presence of Cl^- ions in a solution. Tap water contains chloride ions, so it is very important to perform this test with clean equipment that has been rinsed out with deionized water and to use deionized water in the test.

To obtain accurate and precise data you should do the following:

1. Use the same balance for all mass measurements.
2. Record data directly and immediately into the lab report.
3. Record data completely (to the nearest 0.001 g); do not round off the data.
4. If an error is made in data collection, do not scratch out or erase but simply draw a single line through the error and write the correction beside it.
5. Duplicate samples should be analyzed so that results can be compared and averaged.
6. Handle the crucibles with tongs not fingers.
7. The crucible must be covered while heating to avoid any loss due to splatter.

Name: ________________________________ Section: ____________________

Procedure

A. Quantitative Determination of Percent Composition

1. Weigh a clean dry crucible and cover to the highest precision. Be sure that you use the same balance for all mass measurements within the same experiment to cancel out systematic errors.

2. Carefully read the label on the container and add between 1.0 to 1.5 grams of potassium chlorate, $KClO_3$, into the crucible and cover and weigh to the highest precision.

   NOTE: In part B of this lab, we will be using some potassium chloride. If you accidentally place KCl in your crucible, no mass will be lost on heating and you will have to start over and repeat the experiment correctly using $KClO_3$!

3. To a ring stand, attach a ring and place a clay triangle on the ring. Place the crucible with cover on the clay. Very gently heat with a Bunsen burner for about 8 minutes. Readjust the flame and continue heating more intensely now so that the bottom of the crucible glows red for another 10 minutes.

4. After this two step heating process (1st heating) is complete, turn off the burner, close the crucible lid, and allow to cool about 10 minutes before weighing. It is important to have all mass measurements made at or near room temperature. Do not transfer the hot crucible to the tabletop; it is too hot and will permanently mark the table. Use the crucible cooling plates that are provided. Weigh your sample. The cooling period is a great time to start the second sample.

5. Heat the covered crucible and contents another 6 minutes at maximum temperature, then cool and reweigh (2nd heating). If the results agree within ±0.050 g, you are done and will not need to have a 3rd heating. If the difference is greater than 0.050 g, repeat the heating for another 6 minutes (3rd heating).

B. Qualitative Examination of the Residue

1. Place three clean test tubes in a rack. Put a pea sized quantity of KCl in test tube number 1 and a pea sized quantity of $KClO_3$ in the test tube number 2. Add 10 ml of deionized water to each and mix. Next, add 10 ml of deionized water to your residue in the crucible, mix, and transfer a portion to test tube number 3.

   NOTE: It is not necessary to dissolve and transfer the entire residue.

2. Add 3 drops of 6 M HNO_3 and 5 drops of 0.1 M $AgNO_3$ solution to each test tube. Record observations.

DISPOSE of solutions and precipitates containing silver in the waste container provided.

Name: ______________________________ Section: ____________________

Data and Calculations for Experiment 6

A. Quantitative Determination of Percent Composition

1. When solid $KClO_3$ is heated above 400 °C, it decomposes to solid potassium chloride and elemental oxygen gas. Write the balanced equation for the decomposition of $KClO_3$ solid.

2. What is the remaining residue in the crucible after heating?

3. What substance is lost during the heating?

	Sample 1	Sample 2
4. Mass of crucible and cover	____________	____________
5. Mass of crucible, cover and sample	____________	____________
6. Mass of crucible, cover and sample after 1st heating	____________	____________
7. Mass of crucible, cover and sample after 2nd heating	____________	____________
8. Mass of crucible, cover and sample after 3rd heating	____________	____________
9. Mass of original sample	____________	____________
10. Mass of the residue	____________	____________
11. Mass lost upon heating	____________	____________

12. Experimental percentage of KCl in the $KClO_3$ sample.

Sample 1:

Sample 2:

13. Experimental percentage of oxygen in the $KClO_3$ sample.

Sample 1:

Sample 2:

Name: ________________________________ Section: ____________________

14. Using the atomic masses from the periodic table, solve for the molar mass of $KClO_3$.

____

15. Theoretical percentage of KCl in the $KClO_3$ sample ________________

16. Theoretical percentage of oxygen in the $KClO_3$ sample ________________

17. Percent error in oxygen determination

    Sample 1:

    Sample 2:

B. Qualitative Examination of the Residue

1. Record what you observed when $AgNO_3$ solution was added to the following:
    - i. KCl
    - ii. $KClO_3$
    - iii. Residue

2. What does the evidence lead you to believe about the residue?

3. Does the evidence from the $AgNO_3$ test prove conclusively (without a doubt) that the residue is KCl? Explain.

Name: ________________________________ Section: ____________________

Pre-Lab Assignment (to be completed before coming to lab)

1. a) Write the balanced equation for the decomposition of $Mg(ClO_3)_2$ solid.

   b) A student heated 1.228 grams of $Mg(ClO_3)_2$ until a stable weight was determined. The remaining residue weighed 0.584 grams. Solve for the experimental percentage of oxygen.

   c) Calculate the theoretical percentage of oxygen in $Mg(ClO_3)_2$.

   d) Calculate the percent error in oxygen determination.

2. Given the mass percent of each element:

   18.8% Na 29.0% Cl 52.2% O

   Solve for the empirical formula and name it.

3. Predict the products and balance the equations for the following decomposition reactions:

   a) $NaClO_3(s) \rightarrow$

   b) $Ca(ClO)_2(s) \rightarrow$

   c) $Al(ClO_3)_3(s) \rightarrow$

   d) $Mg(ClO_2)_2(s) \rightarrow$

Name: ________________________________ Section: ____________________

Experiment 7 – Double Displacement Reactions

Discussion

In this experiment, double displacement reactions will be studied, where two water solutions, each containing positive and negative ions, will be combined. Consider the generalized reaction shown below:

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

where AB exists as A^+ and B^- ions in solution, and CD exists as C^+ and D^- ions in solution. Each of the positive ions can combine with the negative ion of the other compound as shown above. But the question then becomes: has there been a reaction? To answer this question, we look at the products. Is either one an insoluble compound giving a precipitate (information available from a solubility table), is either one a gas or producer of a gas, or would a temperature change be predicted? Has a weak electrolyte such as a weak acid been formed? If no to all of these, then no reaction occurs; this is simply a mixture. If yes to any one or more of these, then a reaction occurs.

Procedure

Each part of the experiment below consists of mixing equal volumes of two solutions from dropper bottles in a 24 well-plate. Place 5 drops of each indicated chemical in the well-plate. Write your observations on the report sheet. Note the formation of any precipitate or gas. If neither results, test the well-plate with a thermometer for any temperature change. If no change is noted, write NR (No Reaction) for the mixture.

1. Mix 5 drops of 0.1 M $NaCl(aq)$ with 5 drops of 0.1 M $KNO_3(aq)$.
2. Mix 5 drops of 0.1 M $NaCl(aq)$ with 5 drops of 0.1 M $AgNO_3(aq)$.
3. Mix 5 drops of 0.1 M $Na_2CO_3(aq)$ with 5 drops of dilute 6 M $HCl(aq)$.
4. Mix 5 drops of 10% $NaOH(aq)$ with 5 drops of dilute 6 M $HCl(aq)$.
5. Mix 5 drops of 0.1 M $BaCl_2(aq)$ with 5 drops of dilute 3 M $H_2SO_4(aq)$.
6. Mix 5 drops of dilute 6 M $NH_4OH(aq)$ with 5 drops of dilute 3 M $H_2SO_4(aq)$.
7. Mix 5 drops of 0.1 M $CuSO_4(aq)$ with 5 drops of 0.1 M $Zn(NO_3)_2(aq)$.
8. Mix 5 drops of 0.1 M $Na_2CO_3(aq)$ with 5 drops of 0.1 M $CaCl_2(aq)$.
9. Mix 5 drops of 0.1 M $CuSO_4(aq)$ with 5 drops of 0.1 M $NH_4Cl(aq)$.
10. Mix 5 drops of 10% $NaOH(aq)$ with 5 drops of dilute 6 M $HNO_3(aq)$.
11. Mix 5 drops of 0.1 M $FeCl_3(aq)$ with 5 drops of dilute 6 M $NH_4OH(aq)$.
12. IN THE HOOD: Add a small amount of $Na_2SO_3(s)$ to 5 drops of water and mix to dissolve. Add 5 drops of dilute 6 M $HCl(aq)$ to the first solution.

Dispose of all solutions in the appropriate WASTE CONTAINER in the hood.

Name: ___________________________________ Section: ______________________

Data for Experiment 7

Record your observations for each combination below. If a reaction occurs, write balanced MOLECULAR, IONIC, and NET-IONIC equations. If no reaction occurs, write NR. Make sure to include the physical states of all the products.

1. $NaCl(aq)$ and $KNO_3(aq)$

   Observations:

   Molecular:

   Ionic:

   Net-Ionic:

2. $NaCl(aq)$ and $AgNO_3(aq)$

   Observations:

   Molecular:

   Ionic:

   Net-Ionic:

3. $Na_2CO_3(aq)$ and $HCl(aq)$

   Observations:

   Molecular:

   Ionic:

   Net-Ionic:

4. $NaOH(aq)$ and $HCl(aq)$

   Observations:

   Molecular:

   Ionic:

   Net-Ionic:

Name: ___________________________________ Section: ______________________

5. $BaCl_2(aq)$ and $H_2SO_4(aq)$

Observations:

Molecular:

Ionic:

Net-Ionic:

6. $NH_4OH(aq)$ and $H_2SO_4(aq)$

Observations:

Molecular:

Ionic:

Net-Ionic:

7. $CuSO_4(aq)$ and $Zn(NO_3)_2(aq)$

Observations:

Molecular:

Ionic:

Net-Ionic:

8. $Na_2CO_3(aq)$ and $CaCl_2(aq)$

Observations:

Molecular:

Ionic:

Net-Ionic:

Name: ________________________________ Section: ____________________

9. $CuSO_4(aq)$ and $NH_4Cl(aq)$

Observations:

Molecular:

Ionic:

Net-Ionic:

10. $NaOH(aq)$ and $HNO_3(aq)$

Observations:

Molecular:

Ionic:

Net-Ionic:

11. $FeCl_3(aq)$ and $NH_4OH(aq)$

Observations:

Molecular:

Ionic:

Net-Ionic:

12. $Na_2SO_3(aq)$ and $HCl(aq)$

Observations:

Molecular:

Ionic:

Net-Ionic:

Name: ______________________________ Section: ____________________

Questions

1. For each of the reactions listed below, write balanced molecular, ionic, and net-ionic equations. If no reaction occurs, write NR. Assume all reactants are aqueous unless otherwise noted. Include all physical states.

   A. Lead(II) nitrate and magnesium sulfate solutions are combined.

   Molecular:

   Ionic:

   Net-Ionic:

   B. Barium chloride solution is poured into a solution of ammonium carbonate.

   Molecular:

   Ionic:

   Net-Ionic:

   C. Magnesium chloride solution is mixed with nickel(II) nitrate solution.

   Molecular:

   Ionic:

   Net-Ionic:

   D. Cobalt(II) sulfate and lithium sulfide solutions are combined.

   Molecular:

   Ionic:

   Net-Ionic:

   E. Hydrochloric acid solution is reacted with a solution of lithium carbonate.

   Molecular:

   Ionic:

   Net-Ionic:

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F. Hydroiodic acid and ammonium sulfite solutions are mixed.

Molecular:

Ionic:

Net-Ionic:

G. Sodium hydroxide solution is poured into a solution of cobalt(II) chloride.

Molecular:

Ionic:

Net-Ionic:

H. Ammonium chloride and potassium hydroxide solutions are reacted.

Molecular:

Ionic:

Net-Ionic:

I. Solid strontium bromide is mixed with a solution of potassium phosphate.

Molecular:

Ionic:

Net-Ionic:

J. Solutions of ammonium sulfate and sodium chloride are combined.

Molecular:

Ionic:

Net-Ionic:

Name: ________________________________ Section: ____________________

Experiment 9 – Ionization and the Nature of Acids, Bases, and Salts

Discussion

Compounds were defined by Sven Arrhenius to be acids if they release H^+ ions in solution when dissolved. This modern definition replaced older definitions based on taste (i.e acids tend to be sour tasting) or if they changed litmus paper's color. Bases (which tend to taste bitter) were defined as compounds that give up $OH^–$ (hydroxide) ions in water. This definition was limited to compounds in water and gives way to Brønsted-Lowry acid-base theory.

Brønsted-Lowry acid-base theory keeps the definition of an acid as something that donates an H^+ ion and defines bases as anything that accepts the H^+ ion. Acids become proton donors; bases become proton acceptors. In any acid-base equation, there will be one acid and one base *on each side* of the equation. Which compound is an acid depends on whether that compound is donating or accepting a proton.

$$\underset{\text{Base}}{NH_3} + \underset{\text{Acid}}{H_2O} \rightarrow \underset{\text{Conj. Acid}}{NH_4^+} + \underset{\text{Conj. Base}}{OH^-}$$

$$\underset{\text{Acid}}{HCl} + \underset{\text{Base}}{H_2O} \rightarrow \underset{\text{Conj. Base}}{Cl^-} + \underset{\text{Conj. Acid}}{H_3O^+}$$

Water can function as both an acid and a base, depending on the other reagents!

HCl(aq)	Hydrochloric acid	H_2SO_4	Sulfuric acid
HBr(aq)	Hydrobromic acid	$HC_2H_3O_2$	Acetic acid
HI(aq)	Hydroiodic acid	H_2CO_3	Carbonic acid
H_3PO_4	Phosphoric acid	HNO_3	Nitric acid

Many common strong bases contain hydroxides ($OH^–$) and a metal.

NaOH	Sodium hydroxide
KOH	Potassium hydroxide
$Ca(OH)_2$	Calcium hydroxide
$Mg(OH)_2$	Magnesium hydroxide
NH_4OH	Ammonium hydroxide (best written as $NH_3 \cdot H_2O$)

Solutions that contain bases are called *alkali* or *alkaline*, from an Arabic word for "ashes". Campfire ashes ("bitter ashes") contain hydroxides and carbonates of potassium and sodium, which form basic or alkaline solutions. Compounds from plants that dissolve in water to form alkaline solutions are called *alkaloids*. A common example of a bitter-tasting alkaloid is caffeine.

Name: ______________________________ Section: ____________________

The term pH is used to measure the concentration of an acid in water. Thus, it is important to remember that one acid can produce a range of pH values, depending upon the amount of acid relative to the volume of solution. pH is defined by the equation $pH = -\log [H^+]$. Therefore, a solution of 1.0 M HCl will produce 1.0 M H^+ ions, assuming the HCl breaks up entirely. Since $\log [1.0] = 0$, the pH of this solution is 0. The pH of pure water will be 7.0, while the pH of a very basic solution can be above 14.

$pH < 7$ acidic solutions	$pH = 7$ neutral solution	$pH > 7$ basic solution

When acids react with bases, the H^+ from the acid and the OH^- from the bases “cancel” each other and form water molecules (“HOH”). The anions of the acid and the cations from the base combine to form ionic compounds or salts. For example, consider the reaction of sulfuric acid with sodium hydroxide:

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Reactions of Oxides with Water

The oxides of elements often react with water to form new compounds. Depending upon which family the element is in, the new compound may be acidic or basic. For example, sulfur can be oxidized to form sulfur trioxide, which reacts with water to make sulfuric acid. Consider the following balanced equations:

$$S + O_2 \rightarrow SO_2$$

$$2SO_2 + O_2 \rightarrow 2SO_3$$

$$SO_3 + H_2O \rightarrow H_2SO_4$$

Carbon dioxide reacts with water to form carbonic acid as follows:

$$CO_2 + H_2O \rightarrow H_2CO_3$$

The metal oxides react with water to form basic compounds. Calcium oxide reacts with water to form calcium hydroxide, while magnesium oxide reacts with water to form magnesium hydroxide:

$$CaO + H_2O \rightarrow Ca(OH)_2$$

$$MgO + H_2O \rightarrow Mg(OH)_2$$

Name: ________________________________ Section: ____________________

Electrolytes

Surprisingly, pure distilled water does not conduct electricity. In order for a charge to pass through water, it needs to be carried by positive and negative charges. The more charges, the more current can pass. If the charges cannot move, as in solid salts with no water present, then electricity cannot be conducted.

Compounds can be divided into strong electrolytes, weak electrolytes, and non-electrolytes depending upon how well they conduct electricity when dissolved in solution. Remember that compounds that don't dissolve in the solvent shouldn't be called electrolytes at all. For example, iron bars, wood, or plastics are not electrolytes regardless of whether they conduct electricity or not.

In a strong electrolyte, the compound breaks up into cations or anions in a process called "dissociation". In a weak electrolyte, some of the compound dissociates into ions, even though the entire compound dissolves. In non-electrolytes, the compound dissolves but does not break up at all.

Procedure

A. Electrolytes

In this part of the experiment, your instructor will demonstrate the conductivity of various solutions and reactions.

B. Investigating Acids

1. Reactions of Acids with Metals

   a. Take four separate test tubes and place 5 mL of 6 M HCl in tube #1, 3 M H_2SO_4 in tube #2, 6 M HNO_3 in tube #3, and 6 M acetic acid in tube #4.

   b. Put roughly a 2 cm strip of magnesium metal into each tube. Record the results.

   c. As the metal is still bubbling, place a glowing piece of wood (splint) into the test tube.

2. Measurement of pH and Acidity

   a. Place 5 mL of water in a test tube and add 2 drops of a phenolphthalein indicator solution in it. Add a few drops of dilute hydrochloric acid and record what happens.

   b. There are three solutions of HCl prepared in front of the classroom. The most concentrated, 0.1 M HCl, is one hundred times more concentrated than the weakest solution, the 0.001 M HCl. Use the pH meter to record the pH's of the three solutions.

3. Reactions of Acids with Carbonates and Bicarbonates

    a. Take a 100 mL beaker and just cover the bottom with a thin layer of sodium bicarbonate (baking soda). Add about 4 to 5 mL of diluted (6 M) HCl to the beaker. Record the results. Lower a lit match into the beaker and record what happens.

    b. Try the above reaction again with a chip of calcium carbonate (limestone, marble). Let the reaction go for about 2 minutes before lowering a lit match into it. When completed, throw them in the labeled water container; DO NOT CLOG THE DRAIN!

4. Neutralizing Acids with Base: Using Indicators

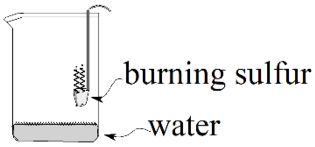
    In this experiment, you will make water acidic and then basic to see how the pH affects a common indicator solution.

    Add 25 mL of water and 3 drops of a phenolphthalein solution to a 100 mL beaker, and then add 5 drops of 6 M hydrochloric acid. To this solution, add 10 percent sodium hydroxide solution drop by drop until the indicator changes color. Once you've gotten this color change, reverse it by adding more dilute acid dropwise.

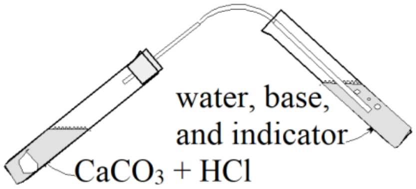
5. Reaction of a Non-Metal Oxide and Water

    In this section, you'll investigate what happens when an oxide of a non-metal, sulfur, reacts with water.

    a. This part of the experiment must be done in the fume hood! Place a small lump of sulfur in a deflagrating spoon (which looks like a ladle with a long handle) and set it on fire with a Bunsen burner. Once the sulfur is burning, lower the spoon into a bottle containing 15 mL of water; this will allow the fumes of combustion to fill the air space of the bottle. After 2 minutes, remove the sulfur and cover the bottle with a glass plate. Shake the bottle to mix the gas and water. Is the water acidic or basic?



    b. In a test tube, generate carbon dioxide gas by treating marble chips with hydrochloric acid (see section 3b). Bubble the gas into another beaker containing 10 mL of water, 2 drops of 10% sodium hydroxide, and a few drops of phenolphthalein indicator.



Name: ______________________________ Section: ____________________

C. Properties of Bases and Basic Solutions

1. Properties of ammonium and sodium hydroxides

   a. Place three drops of concentrated ammonium hydroxide (used in “Windex” cleaners) in 10 mL of water in a test tube. In another test tube, place three drops of concentrated sodium hydroxide (used in “Drano” pipe cleaners) in 10 mL of water. Rub a few drops of the diluted solution from each test tube onto your fingers. What is the difference in feeling between the two solutions? Wash your hands with water afterwards until your skin feels normal.

   b. Test the two solutions with red and blue litmus papers and record the changes you see.

   c. Add two drops of phenolphthalein indicator to each test tube and record the changes you see.

   d. Determine the pH of each solution using a pH meter. *Wash the electrode with dilute acetic acid and then distilled water to clean it between every reading and after you’re done.*

2. The Reaction of Metal Oxides and Water

   a. In three test tubes, place 10 mL of water, 2 drops of phenolphthalein, and a pinch of calcium hydroxide, magnesium hydroxide, or calcium oxide. Record the color changes.

   b. In this last section, you will explore the reaction that occurs when you heat limestone (“slaking lime”) to make a compound known as “quicklime”, which is used in the manufacture of concrete:

      Take a small piece of iron wire and wrap it around a small chip of calcium carbonate (marble chip). Heat the chip until it is white hot with a Bunsen burner, for about 2 minutes. Let the chip cool and drop it into a beaker with 15 mL of water and a few drops of phenolphthalein. Compare this result to an unheated chip.

Name: ______________________________ Section: ____________________

Data and Calculations for Experiment 9

A. Electrolytes and Instructor Demo

Place an "X" on the label that properly describes each compound below:

	Non-Electrolyte	Strong Electrolyte	Weak Electrolyte
1. Tap water			
2. Distilled water			
3. Sugar solution			
4. NaCl solution			
5a. Pure (glacial) acetic acid			
5b. Diluted acetic acid			
5c. Twice diluted acetic acid			
6a. 1 M acetic acid			
6b. 1 M HCl			
6c. 1 M NH_4OH			
6d. 1 M NaOH			
7a. $NaNO_3$			
7b. NaBr			
7c. $Ni(NO_3)_2$			
7d. $CuSO_4$			
7e. NH_4Cl			

1. What reaction occurs when barium hydroxide and sulfuric acid are mixed?
2. Explain why the light becomes dimmer as two strong electrolytes are mixed with each other.
3. Why does the light come back on after more of the electrolyte is added?
4. What happens to the glacial acetic acid as it is diluted? How does this explain the changes in light intensity?

Name: ________________________________ Section: _____________________

B. Properties of Acids

1. Reactions of Acids with Metals

a) Which acids reacted with the magnesium to produce H_2 gas?

b) Represent the reaction between the metal and ONE acid that occurred with an equation.

2. Measurement of pH and Acidity

a) Acids turned the red litmus paper ____________.

b) Acids turned the blue litmus paper ____________.

c) What is the color of phenolphthalein in acidic solution? ____________

d) What is the pH of the 0.1 M solution? ____________

What is the pH of the 0.01 M solution ____________

What is the pH of the 0.001 M solution? ____________

e) Which solution has the greatest concentration of H^+?

f) Calculate the H^+ concentration of a pH = 4.6 solution. Write the answer in scientific notation.

3. Reactions of Acids with Carbonates and Bicarbonates

a) What is the name and formula of the gas formed in this reaction?

b) What happened to the burning stick when it was placed in the beaker?

c) Write out the products of the reactions in a balanced equation:

$NaHCO_3 \quad + \quad HCl \quad \rightarrow$

$CaCO_3 \quad + \quad HCl \quad \rightarrow$

4. Neutralizing Acids with Base: Using Indicators

a) Write a balanced equation for the reaction of HCl and NaOH.

b) What happened when the acid was all neutralized?

5. Reaction of a Non-Metal Oxide and Water

a) Write a balanced equation for the reaction of sulfur and oxygen.

b) What happens when the product of the above reaction reacts with water? Write a balanced equation that represents this reaction.

c) Write a balanced equation for the reaction of carbon dioxide and water.

d) How do you know that the product in the reaction above is acidic?

Name: ___________________________________ Section: ______________________

C. Properties of Bases

1. Properties of ammonium and sodium hydroxides

   a) What did the sodium hydroxide feel like?

   b) What did the ammonium hydroxide feel like?

   c) Bases turned the red litmus paper _______________.

   d) Bases turned the blue litmus paper _______________.

   e) What is the pH of the ammonium hydroxide solution? ________

   f) What is the pH of the sodium hydroxide solution? ________

   g) Calculate the concentration of H^+ in the more basic solution ________

2. The Reaction of Metal Oxides and Water

   a) What is the color of phenolphthalein with CaO? _______________

      What is the color of phenolphthalein with MgO? _______________

      What is the color of phenolphthalein with $Ca(OH)_2$? _______________

   b) Write the balanced equations for the following reactions:

      $CaO + H_2O \rightarrow$

      $MgO + H_2O \rightarrow$

   c) Marble is calcium carbonate ($CaCO_3$). Write a balanced equation for the reaction that occurs when you heat the marble chip.

   d) Write a balanced equation for the reaction that occurs when you put the heated marble chip in water.

Name: ________________________________ Section: ____________________

Experiment 10 – Lewis Dot Structures and Molecular Geometry

Discussion

The Valence Shell Electron Pair Repulsion (VSEPR) Theory states that bonds and lone pairs are regions of high electron density in an atom that repel each other until they get as far apart as possible. This effect determines the atom's geometry and bond angles. Two regions will be 180° apart, three regions will be 120° apart, and four regions will be 109.5° apart.

Geometry Determination

A. Determine the Lewis dot structure of the molecule or ion.

B. For each central atom in the structure, determine the areas of electron density that lie directly on that atom. An area of electron density may be:
   - a lone pair that lies on the central atom. (Lone pairs on other atoms don't count.)
   - a single bond.
   - a double bond.
   - a triple bond.

C. Assign geometry according to the table on the next page.

Procedure

Do not build models for ionic compounds. For each of the polyatomic ions or molecules:

II. Draw the Lewis dot structure, including all resonance where appropriate.

III. Use the model kit to build the structure.
   - Use white or yellow balls for hydrogens.
   - Use black or blue or red balls for other atoms.
   - Use short sticks for nonbonded electron pairs (lone pairs).
   - Use long sticks for single bonds.
   - Use springs for double and triple bonds. Two springs form a double bond. Three springs form a triple bond.

IV. Sketch the shape of the structure in three dimensions. This is called the VSEPR structure.

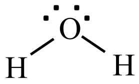
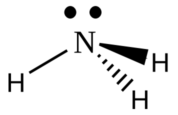
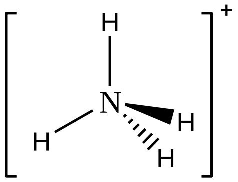
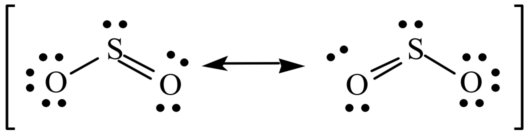
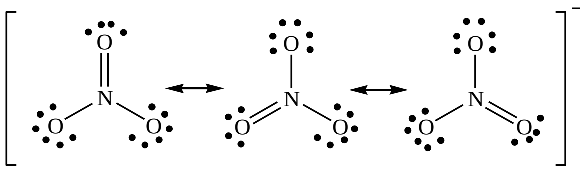
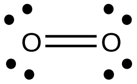
V. Draw dipole moments on the VSEPR structure to show all polar bonds.

VI. Give the name of the molecular geometry.

VII. State whether the molecule is polar, nonpolar, or ionic.

VIII. Determine the approximate bond angle on the central atom (if applicable).

Name: ________________________________ Section: ____________________

# of areas	**# of bonds**	**# of lone pairs**	**Geometry and bond angles**	**Example**
4	2	2	Angular or bent (109.5°)	
4	3	1	Pyramidal (109.5°)	
4	4	0	Tetrahedral (109.5°)	
3	2	1	Bent (120°)	 2 resonance forms
3	3	0	Trigonal (120°)	 3 resonance forms
2	2	0	Linear (180°)	$\mathrm{H{-}C{\equiv}N{:}}$
	1	any	Linear (Must be three or more atoms to form an angle.)	

Notice the convention for drawing bonds in 3-D space, where:

- the wedge (◀) represents a bond coming out of the paper, and
- the dash (||||||||) represents a bond going behind the paper.

Name: ___________________________________ Section: ______________________

Formula	Number of valence electrons	Lewis dot structure (including ALL resonance)	VSEPR structure (with dipole moments)	Molecular Geometry	Polar? Nonpolar? Ionic?	Bond angle on central atom(s)
I_2						
NO^{-1}						
CO						
BeH_2						

Name: ______________________ Section: ______________

Formula	Number of valence electrons	Lewis dot structure (including ALL resonance)	VSEPR structure (with dipole moments)	Molecular Geometry	Polar? Nonpolar? Ionic?	Bond angle on central atom(s)
H_2S						
PBr_3						
ClO_4^{-1}						
CS_2						

Name: ____________________________ Section: ____________________

Formula	Number of valence electrons	Lewis dot structure (including ALL resonance)	VSEPR structure (with dipole moments)	Molecular Geometry	Polar? Nonpolar? Ionic?	Bond angle on central atom(s)
$CHCl_3$						
PO_3^{-3}						
PO_4^{-3}						
CH_2O						

Name: ____________________ Section: ____________

Formula	Number of valence electrons	Lewis dot structure (including ALL resonance)	VSEPR structure (with dipole moments)	Molecular Geometry	Polar? Nonpolar? Ionic?	Bond angle on central atom(s)
SO_3						
SO_3^{-2}						
SO_4^{-2}						
SCN^{-1}						

Name: ___________________________________ Section: ______________________

Formula	Number of valence electrons	Lewis dot structure (including ALL resonance)	VSEPR structure (with dipole moments)	Molecular Geometry	Polar? Nonpolar? Ionic?	Bond angle on central atom(s)
NO_2^-						
KCl						
BrO_3^-						
IO_2^-						

Name: ____________________ Section: ____________________

Formula	Number of valence electrons	Lewis dot structure (including ALL resonance)	VSEPR structure (with dipole moments)	Molecular Geometry	Polar? Nonpolar? Ionic?	Bond angle on central atom(s)
CH_2Cl_2						
CO_2						
PH_3						
NaH						

Name: ________________________________ Section: ____________________

Experiment 11 – Boyle's Law

Discussion

Robert Boyle formalized his law around 1660. This is one of the earliest instances of the use of experiment and mathematics to describe physical systems as they are rather than using pre-conceived notions of what the system should be.

In this experiment, you will measure the pressure and volume of a sample of air trapped in a tube. Afterwards, you will examine the data and discover how close your data is to the ideal gas law.

Boyle's Law is simply stated as $P \propto 1/V$. That is, the pressure and volume are inversely proportional. Another way of writing this is $PV = k$. If the gas is "ideal", its pressure and volume will follow this equation at a constant temperature.

Procedure

1. Record the temperature (using a thermometer) and pressure (using a barometer) of the air in the room.

2. Obtain a gas pressure sensor, LabQuest mini, USB cable, chemistry department laptop computer, and a syringe.

3. Connect the gas pressure sensor to any channel on the LabQuest mini, then connect the USB cable from the LabQuest mini to the laptop computer

4. With the syringe disconnected from the pressure sensor, slide the plunger of the syringe so that the first black rubber gasket (the base of the rubber "cone") rests on the 10 mL line. Connect the syringe directly to the port on the pressure sensor with a gentle twisting motion – do not over-tighten (see Figure 1). If the syringe or the pressure sensor has been damaged from overtightening (by a previous student), you may use Parafilm to form an air-tight seal between the syringe and the pressure sensor. For the remainder of the experiment, do not unscrew the syringe from the gas pressure sensor or remove the piston from the syringe. The air that is trapped inside the syringe will serve as the "gas" for your experiment.

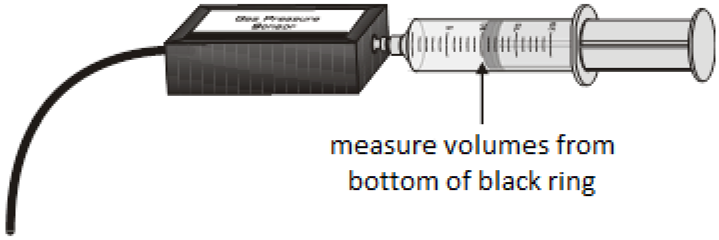


Figure 1. Syringe connected to pressure sensor.

5. Open the Logger Pro software on the laptop computer. The computer monitor will display the pressure of the air that is trapped inside of the syringe, measured by the gas pressure sensor. This pressure is shown in the lower left corner of the Logger Pro window (see Figure 2). As you adjust the volume, you will record the pressure of the air inside the syringe in your data table in the column labeled “Pressure”.

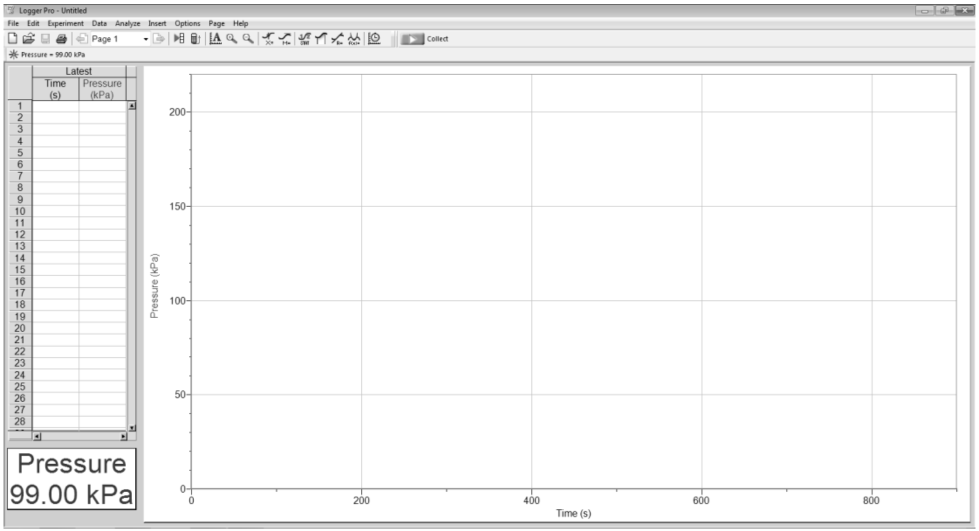


Figure 2. The Logger Pro window. In this example, the pressure is 99.00 kPa.

6. Gently depress the plunger until the first black gasket rests on the 5 mL line on the syringe. You will have to hold the plunger in. Do not push further than the 5 mL line – the maximum pressure the sensor can tolerate is 210 kPa.

7. Hold the plunger in this position for at least 15 seconds. If the pressure slowly drops during this time, then the seal between the syringe and the pressure sensor is not air-tight. Disconnect the syringe from the pressure sensor and go back to step 4! If the pressure is stable, then record the pressure in your data table.

8. Move the plunger to the 6 mL line on the syringe and hold it there. Record the pressure in the appropriate row of your data table. Repeat this procedure for each volume that is listed in your data table.

9. Open Microsoft Excel® on the laptop computer. Each of the “boxes” that you see are called “cells.” In one row of cells, type the headings (in the same order) that you see at the top of your data table.

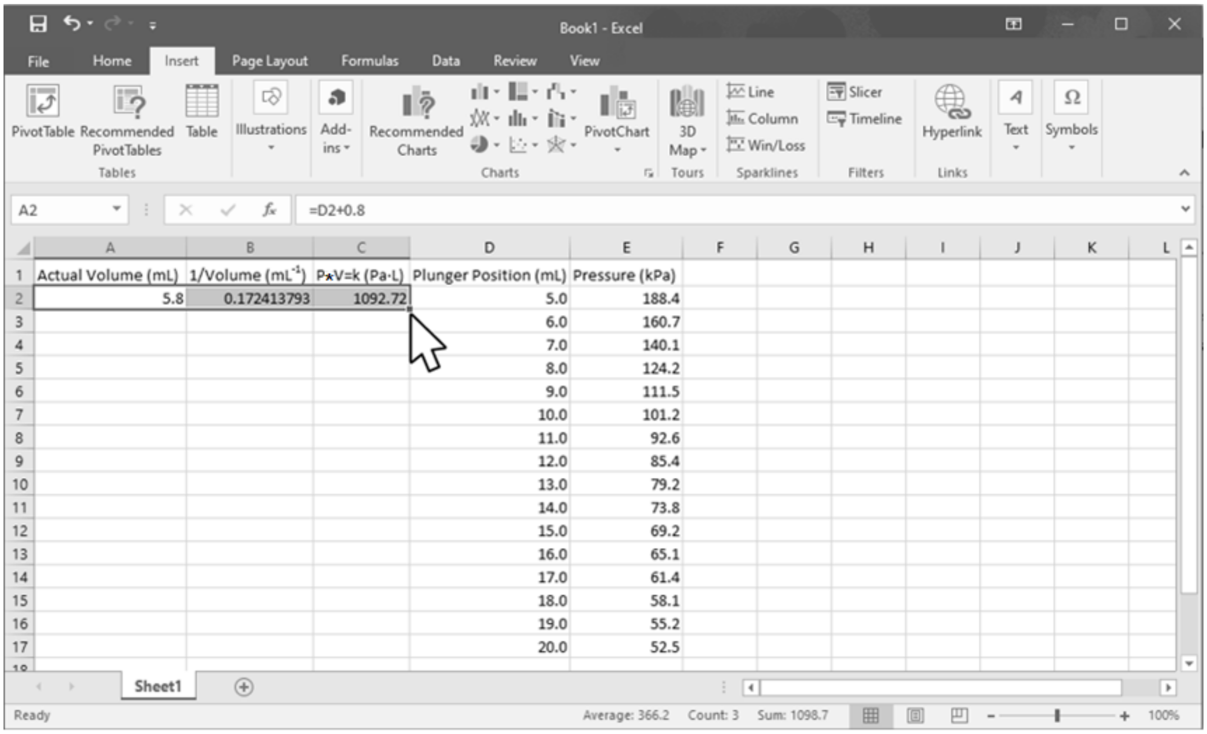
10. In the cells under the headings for “Syringe Volume (mL)” and “Pressure (kPa),” copy the numbers that you recorded in lab. Each cell should contain one number (without units).

11. Now, we are going to use Excel® to calculate the rest of our data for us. The volume printed on the side of the syringe does not account for the volume of air that is trapped in the tip of the syringe and inside of the gas pressure sensor itself. This volume has been measured to be 0.8 mL. To adjust for this volume, you will need to add 0.8 mL to each of the volumes listed in the data table. Click on the top empty cell under the “Actual Volume (mL)” heading. Type the equal sign (this tells Excel® to start a calculation), then click on the top

number in the “Plunger Position (mL)” column to use this number in your calculation. Type **+0.8** and hit enter. This cell should now display the number 5.8.

12. Repeat this process to calculate the top cell in each of the other columns. Be sure that you start each calculation by typing the equal sign and that use the actual volume (NOT the plunger position) for all of your calculations. “Click” on any number from the data table that will be a part of your calculation (do not enter any numbers by hand unless they are a part of a formula).

13. Once the first row of your data table is complete, click and drag to select the three cells that contain calculations. When these cells have been selected, a box will appear around them with a dark “dot” on the bottom-right corner (*Figure 3)*. Double click on this dot and Excel® will automatically calculate all the rest of the rows in your data table.



A2 =D2+0.8

	A	B	C	D	E
1	Actual Volume (mL)	1/Volume (mL^{-1})	P*V=k (Pa·L)	Plunger Position (mL)	Pressure (kPa)
2	5.8	0.172413793	1092.72	5.0	188.4
3				6.0	160.7
4				7.0	140.1
5				8.0	124.2
6				9.0	111.5
7				10.0	101.2
8				11.0	92.6
9				12.0	85.4
10				13.0	79.2
11				14.0	73.8
12				15.0	69.2
13				16.0	65.1
14				17.0	61.4
15				18.0	58.1
16				19.0	55.2
17				20.0	52.5

Average: 366.2 Count: 3 Sum: 1098.7

Figure 3. A partially complete Excel® data table.

14. To calculate the average value of “k”, find an empty cell away from your data table and type **=AVERAGE(** and then click and drag to select all of the numeric data in the P×V=*k* (Pa·L) column. Press enter on the keyboard and then copy the average onto your data page.

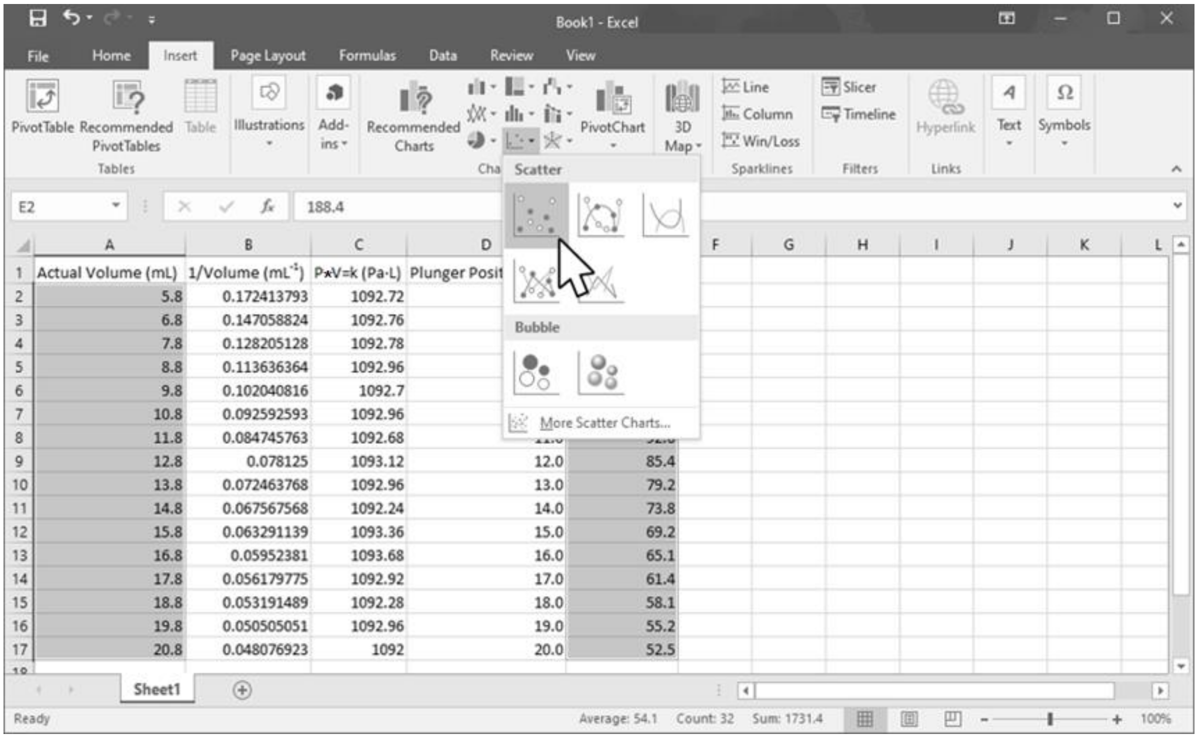
15. Create two graphs using Excel®. On the first graph, you will need to graph Pressure vs. Actual Volume (this is Pressure on the y-axis, Volume on the x-axis). The first graph is not linear, so it should not have a trendline. On the second graph, you will need to graph Pressure vs. 1 / Volume (again Pressure on the y-axis, 1 / Volume on the x-axis). Add a linear trendline to the second graph and include the equation of the line. Be sure to save a copy of the Excel® data table and graphs to include with your report (or show them to your instructor before the end of lab).

Name: ________________________________ Section: ____________________

Excel® Graphing Procedure

Note that various versions of Excel® may function a bit differently from the directions outlined below (which work on department-owned laptop computers):

Click and drag your mouse to highlight the boxes containing the numeric data for the x-axis. Then, ***while holding CTRL on the keyboard***, click and drag your mouse to highlight the boxes containing the numeric data for the y-axis. From the "Insert" tab, choose a "Scatter" plot (*Figure 4*).



1	Actual Volume (mL)	1/Volume (mL^{-1})	P*V=k (Pa·L)	Plunger Posit	
2	5.8	0.172413793	1092.72		
3	6.8	0.147058824	1092.76		
4	7.8	0.128205128	1092.78		
5	8.8	0.113636364	1092.96		
6	9.8	0.102040816	1092.7		
7	10.8	0.092592593	1092.96		
8	11.8	0.084745763	1092.68		
9	12.8	0.078125	1093.12	12.0	85.4
10	13.8	0.072463768	1092.96	13.0	79.2
11	14.8	0.067567568	1092.24	14.0	73.8
12	15.8	0.063291139	1093.36	15.0	69.2
13	16.8	0.05952381	1093.68	16.0	65.1
14	17.8	0.056179775	1092.92	17.0	61.4
15	18.8	0.053191489	1092.28	18.0	58.1
16	19.8	0.050505051	1092.96	19.0	55.2
17	20.8	0.048076923	1092	20.0	52.5

Figure 4. Creating a Scatter Plot.

Your graph must include a meaningful Chart Title and Axis Titles (with units). These Chart Elements can be added to your graph by clicking on the "+" icon in the upper right corner of your graph. Your instructor may request additional Chart Elements.

To add a Trendline, right click on any data point on your graph and choose "Display Trendline" from the menu that appears. The format trendline pane will appear on the right side of your screen. Linear should be selected by default. From this pane, you should check the box next to "Display Equation on chart." Your instructor may also ask you to check the box for "Display R-squared value on chart."

Name: ______________________________ Section: ____________________

Data and Calculations for Experiment 11

Pressure of the air in the room: ____________________________________

Temperature of the air in the room: ____________________________________

Actual Volume (mL)	**1 / Volume (mL^{-1})**	**Pressure * Vol. = *k* (Pa·L)**	**Plunger Position**	**Pressure (kPa)**
These will be calculated in Excel®. Submit your spreadsheet and your graphs.			5	
			6	
			7	
			8	
			9	
			10	
			11	
			12	
			13	
			14	
			15	
			16	
			17	
			18	
			19	
			20	

Average *k* = __________

Name: ________________________________ Section: ____________________

Questions

1) On your linear graph, do any points deviate from the straight line?

2) Write down the equation of the trendline ($y = mx + b$) from your linear graph. How does the slope (m) compare to the average P*V=*k* value from the table of data?

3) Using the equation of your trendline, solve for the pressure at a volume of 2.0 mL.
Hint: $x = 1/V$ in your equation!

4) Why must the temperature be constant during this experiment? Use observations from your experiment and the graphs to support your answer!

5) If you repeated this experiment at a higher temperature, how would the P vs. V curve obtained differ from the curve on your 1st graph?

6) You have a 1.00 L sample of Argon gas at 700.0 mmHg. You decrease the pressure to 500.0 mmHg. What is the new volume?

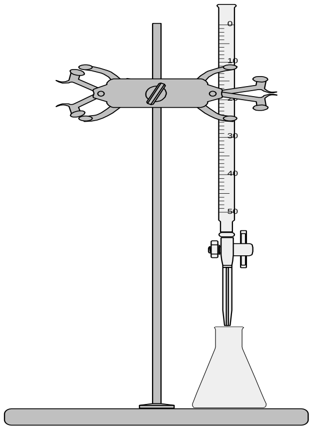
7) Describe (quantitatively) what you would do to the volume of a container of gas if you wanted to double the pressure inside.

Name: ________________________________ Section: ____________________

Experiment 12 – Acid/Base Titrations

Discussion

This experiment demonstrates an analytical technique known as titration, where a solution is delivered from a buret until it completely consumes another solution in a flask. Consider the following:



Acid-base titrations are an example of **volumetric analysis**, a technique in which one solution is used to analyze another. The solution used to carry out the analysis is called the **titrant** and is delivered from a device called a **buret**, which measures the volume accurately. The point in the titration at which enough titrant has been added to react exactly with the substance being determined is called the **equivalence point** (or **stoichiometric point**). This point is often marked by the change in color of a chemical called an **indicator**. The titration set-up is illustrated in the schematic shown.

First, the concentration of a base solution (standard) will be determined; this data will be used to determine the concentration of unknown acid solutions. To standardize the NaOH solution, it will be reacted with potassium hydrogen phthalate, $KHC_8H_4O_4$ (abbreviated KHP). The molar mass of KHP is 204.2 g/mol. The reaction of KHP with NaOH is known to be:

$$KHC_8H_4O_4(aq) + NaOH(aq) \rightarrow KNaC_8H_4O_4(aq) + H_2O(l)$$

Starting with a known mass of KHP and recording the volume of NaOH required to reach the endpoint, the molarity of the base can be determined. The indicator to be used, phenolphthalein, is colorless in acidic solution and rosy pink when slightly basic. Even though the endpoint is slightly overshot in order to make the color change, the goal is to use as little excess base as possible. Therefore, the titrated solution should be very pale pink, not bright rosy red, at the endpoint.

Once the concentration of the NaOH solution is known, one could determine the concentration of an acid solution by doing an additional titration.

Name: ______________________________ Section: ____________________

Procedure

1. Measure out 1.000 to 1.200 g of KHP and add to a 125 mL Erlenmeyer flask.

2. Add approximately 30 mL of D.I. water to the flask. If some KHP is sticking to the walls of the flask, rinse it down with D.I. water from a wash bottle.

3. Take a clean, dry 100 mL beaker to the hood and obtain approximately 60 mL of NaOH base. Be sure to keep the beaker covered with a watch glass when not in use.

4. Obtain a clean 50 mL buret. Carefully fill the buret with base, making sure that no air bubbles are present. Run some of the base solution through the buret tip to remove the air pocket in the tip.

5. Record the initial buret reading (x.xx mL) in the data section. When you read a buret, the line of sight must be level with the BOTTOM of the meniscus to avoid error. The top of the buret reads 0.00 mL; the bottom reads 50.00 mL. Notice the numbers increase going down. Take note of this when reading the numbers. Your instructor will demonstrate.

6. Add 2 to 3 drops of phenolphthalein indicator solution to the 125 mL Erlenmeyer flask containing KHP and water. Swirl your acid solution until the KHP is completely dissolved before moving on to the next step of the procedure.

7. Place the flask under the tip of the buret. A piece of white paper under the flask makes it easier to see the pale pink color at the endpoint. Open the valve and allow base to drip from the buret into the flask. Swirl continually to mix the solutions. As you get closer to the endpoint, the solution will begin to show pink color that goes away when you mix. Slow the rate of base addition to one drop at a time, mixing the solutions well after every drop. If you splash the solution up onto the sidewalls of the flask, spray a stream of water from your wash bottle over the inside of the flask. The extra water that mixes into your acid sample will not affect your results. Once the addition of ONE drop of base changes the solution from colorless to pale pink, close the buret valve and make certain that the pale pink color lasts for at least 30 seconds while swirling. If so, record the final buret reading. If not, carefully add one more drop of base from the buret valve until the pale pink color persists for 30 seconds of swirling or longer. If at the end of your trial the color is bright rosy red, you have overshot the endpoint. Make a note in your data if you overshoot.

8. Discard the titrated solution into the sink, rinse the flask with D.I. water, and then titrate another new sample of KHP following the same procedure above. Do at least TWO successful titrations that achieve a pale pink color of the indicator.

9. Calculate the average molarity of the base from your two successful trials and check with your instructor for verification.

Name: ______________________________ Section: ____________________

Data and Calculations for Experiment 12

	Sample 1	**Sample 2**
Mass of flask and KHP		
Mass of empty flask		
Mass of KHP		
Initial buret reading		
Final buret reading		
Volume of base used		

1. Moles of acid (KHP, Molar mass = 204.2)

   Sample 1:

   Sample 2:

2. Moles of base used to neutralize acid

   Sample 1:

   Sample 2:

3. Molarity of base (NaOH)

   Sample 1:

   Sample 2:

4. Average Molarity of Base:

Name: ________________________________ Section: _____________________

Questions

1. A titration required 13.42 mL of 0.1638 M NaOH solution. How many moles of NaOH were in this volume?

2. A student weighed a sample of KHP and found it weighed 1.396 g. Titration of this KHP required 21.36 mL of base (NaOH). Calculate the molarity of the base.

3. Write and balance the equation for the neutralization of a sulfuric acid solution of unknown concentration by sodium hydroxide. Calculate the molarity of an unknown sulfuric acid solution if a 25.0 mL sample of the acid solution consumes 27.2 mL of 0.138 M NaOH solution in a titration.

4. What might happen to your calculated NaOH molarity if you used tap water instead if D.I. water to dissolve the KHP crystals or to rinse down the walls of the flask during the titration? *Hint: Tap water contains some calcium carbonate*.

Name: ______________________________ Section: ___________________

Experiment 13 – Weighing by Difference

The purpose of this experiment is to learn how to correctly and accurately use the analytical mass balance.

Procedure

NOTE: Always use the same balance during an experiment and leave it clean.

A. Check to see if the balance is level by looking at the bubble in the level gauge. If the bubble is not centered, adjust the legs of the balance until the bubble is centered. Use the same balance throughout the experiment.

B. Place a weighing boat on the balance pan. Zero the balance by pushing down on the tare button or zero button and waiting a few seconds before placing anything on it. Place approximately 3 grams of salt on the boat. Record the exact mass of the salt. Remove the weighing boat and salt from the balance and save them for step D.

C. Zero the balance. Place a clean dry evaporating dish on the balance pan. Record its mass.

D. Pour the salt from the weighing boat into the evaporating dish.

E. Record the mass of the evaporating dish with the salt sample.

F. Return the salt to its original container. Wipe clean and return the evaporating dish.

Data and Calculations for Experiment 13

1. Mass of salt sample (from step B)	
2. Mass of evaporating dish	
3. Mass of evaporating dish and salt	
4. Mass of salt in evaporating dish (calculate)	
5. Difference between 1 and 4	

Show how you determined 4 and 5.

Name: ______________________________ Section: ____________________

Experiment 14 – Atomic Spectra

The purpose of this experiment is to show that different elements give off unique colors of light when atoms of the elements are excited by heating. By identifying the unique colors the element can be identified.

Part I

There will be three gas discharge tubes set up in the lab. Observe the color of light given off by each discharge tube. Record the colors in the data table below. After recording the color of the light, observe the light through the diffraction grating (look off to one side). Draw a picture of the spectral lines. Identify the color of each line.

1. Element __________________________ Color of light ______________________

Spectral diagram:

Violet Red

2. Element __________________________ Color of light ______________________

Spectral diagram:

Violet Red

3. Element __________________________ Color of light ______________________

Spectral diagram:

Violet Red

Name: ________________________________ Section: ____________________

Part II-A:

There will be seven containers with wooden splints soaking in salt solutions. Each solution will be labeled with its chemical name.

Procedure

Light the bunsen burner and adjust its flame until you see a blue inner cone. Use crucible tongs to remove a splint from the soaking solution and place it in the flame. Observe the color of the flame and record the color in the table below. If two solutions give colors which seem similar, repeat the experiment until you can notice the differences in color well enough that you can describe the differences. Always use the same burner with a given chemical. If you mix up burners, you can contaminate the colors.

Chemical	Color of Flame
lithium chloride	
calcium chloride	
potassium chloride	
copper(II) nitrate	
strontium chloride	
sodium chloride	
barium chloride	

Part II-B.

There will be seven containers having unknown chemicals. Repeat the procedure that you used for known chemicals and identify the unknown chemicals. Record results below.

Unknown #	Color of Flame	Chemical
1		
2		
3		
4		
5		
6		
7		

Name: ______________________________ Section: ____________________

Experiment 19 – Specific Heat Capacity

The heat absorbed by any sample of matter when its temperature rises by 1 ^oC is called its heat capacity. Clearly, the heat capacity of a sample depends on its mass; the greater the mass of a sample, the greater the amount of heat it must absorb to increase its temperature by one temperature degree. The heat capacity of one gram of a substance is called its specific heat capacity, which is measured in J/g ^oC or J/g K.

The molar heat capacity is the quantity of heat that must be absorbed by one mole of a substance to raise its temperature by 1 ^oC or 1 K. Its units are J/mol ^oC or J/mol K. The SI unit of energy is the Joule (J), where 4.184 Joules of heat are equal to 1 calorie. A calorie is defined as the energy required to raise 1 gram of water by 1 ^oC.

The relationship between the temperature change and the heat associated with the change is given by the equation: $Q = ms\Delta T$, where Q = heat, m = mass, s = specific heat, and ΔT represents the change in temperature ($T_{final} – T_{initial}$). In any experiment involving calorimetry, the amount of heat lost by the solid is equal to the amount of heat gained by the water. However, since heat loss is given a negative sign, we state that:

$$Q(gain) = –Q(loss)$$

This can also be represented by: $Q(gain) + Q(loss) = 0$.

Procedure

1. Weigh and record the mass of a DRY metal sample. Please note that the mass of the attached string WILL NOT affect your results.

2. Place the loop of the string over an appropriate clamp to suspend the metal into a boiling D.I. water bath and allow the metal enough time to equilibrate to the temperature of the boiling water (about 5 minutes). Make certain that the metal is completely immersed at ALL times and does not touch the bottom of the beaker. After a constant temperature is observed, measure the temperature of the boiling water with a thermometer obtained from the lab cart, noting how to correctly read its scale. This is the initial temperature of the metal sample and should be recorded to within 0.01 ^oC. Afterwards, remove the thermometer from the water bath, and allow it to cool gradually.

3. Weigh an empty calorimeter (two styrofoam coffee-cups nested together with a plastic cover as the top). Use this same mass of the empty calorimeter for ALL TRIALS unless you use a different calorimeter.

4. Add between 50 – 100 mL of water to the calorimeter (less for a small sample, more for a large sample). Weigh and record the mass of the calorimeter and water using the SAME analytical balance.

5. Place the styrofoam calorimeter into a glass beaker for added stability. Insert the thermometer through a hole in the calorimeter lid. Use a thermometer clamp to hold it in place while water equilibrates to a constant temperature (should remain constant for approximately 2 minutes). Occasionally swirl the water to obtain its average temperature. Record this equilibrium temperature to within 0.01 ^{o}C. Be careful not to place the calorimeter near the Bunsen burner.

6. Quickly and carefully place the hot metal into the water in the calorimeter. Carefully return the lid and thermometer to the calorimeter. Gently swirl the calorimeter while observing the temperature. Remove the thermometer from the clamp for this part and hold by hand so that the cup can be constantly swirled. Note the temperature immediately and continuously observe until a constant, HIGHEST stable temperature is reached. Record this final, highest temperature to within 0.01 ^{o}C.

7. Repeat the experiment. Use the same calorimeter and piece of metal. Pour out the original water from the calorimeter and add a fresh 50 – 100 mL portion of water. Weigh and record this new mass. <u>Note</u>: If your temperature change was less than 1.2 ^{o}C, use less water in your second experiment.

8. When completely finished, dry the calorimeter, thermometer, and metal sample. Return them to the lab cart. Be sure to place the metal in its correct container.

Name: ___________________________________ Section: ______________________

Data and Calculations

	Trial 1		Trial 2		Trial 3
Mass dry calorimeter	___________	=	___________	=	___________
Mass calorimeter + volume H_2O	___________		___________		___________
Initial temperature of water in calorimeter	___________		___________		___________
Mass of metal	___________	=	___________	=	___________
Initial temperature of hot metal (before adding it to calorimeter)	___________		___________		___________
Final temperature of water + metal in calorimeter	___________		___________		___________
Δt_{water}	___________		___________		___________
Δt_{metal}	___________		___________		___________

1. Calculate the specific heat of the metal from each trial and find the average value. If the two values do not agree to within 0.06 J/g ^{o}C, a third trial must be run. SHOW CALCULATIONS:

   Trial 1 __________ Trial 2 __________ Trial 3 __________

   Average Specific Heat__________ J/g ^{o}C

2. Find the actual value for the specific heat of your metal in a reference book. Give this value in J/g ^{o}C or J/g K. Calculate the % error of your average value.

Name: ________________________________ Section: ____________________

Post-lab Questions

1. Do objects that have the same temperature have the same amount of heat? Briefly explain.

2. What is the difference between something which is hot and something which has a lot of heat?

3. How much heat would it take to raise the temperature of 645 g of water by 25^{o}C? SHOW CALCULATIONS.

4. When a 15.411 gram sample of metal gains 128.0 J of heat, its temperature changes from 18.55 ^{o}C to 83.00 ^{o}C. What is the specific heat of the metal? SHOW CALCULATIONS.

5. A metal sample weighing 71.9 g and at a temperature of 100.0 ^{o}C was placed in 41.0 g of water in a calorimeter at 24.5 ^{o}C. At equilibrium, the temperature of the water and metal was found to be 35.0 ^{o}C.

    A. What was Δt_{water}?

    B. What was Δt_{metal}?

    C. How much heat flowed into the water?

    D. Calculate the specific heat of the metal.

Name: ________________________________ Section: ____________________

Experiment 20 – Precipitation of Strontium Sulfate

In this experiment, you will study a precipitation reaction between sodium sulfate and strontium chloride. You will collect, dry, and weigh the precipitate and compare this experimental yield to the theoretical yield.

Procedure

Weigh a clean, dry, 100-mL beaker. Add about 0.25 g (0.350 g max!) of solid sodium sulfate to the beaker and weigh it again. Dissolve the sodium sulfate in about 20 mL of D.I. water. Add 5.0 mL of 0.50 M strontium chloride solution and heat for fifteen minutes. Try to keep the mixture from boiling.

After the heating period has passed for the mixture, set it aside so as to return to room temperature, and then cool it further by putting the beaker in a cold water bath. Your precipitate should settle to the bottom, leaving a relatively clear solution above it. Obtain a piece of filter paper and weigh it on the analytical balance. Set up a vacuum filtration apparatus with a Büchner funnel and your weighed filter paper (your instructor will show you how). Using a stirring rod to guide the stream of liquid, pour the contents of the beaker into the Büchner funnel. Use your wash bottle (filled with D.I. water) to rinse any solid out of the beaker and into the filter. Make sure no precipitate remains in the beaker or on the stirring rod. Fill the beaker with 15 mL of D.I. water, swirl it around, and then pour it into the filter. Repeat the washing process, and then draw air through the funnel for a few minutes to help dry the crystals.

Turn off the vacuum, carefully remove the filter paper containing your precipitate with a spatula, and place it over a watch glass. Fill a 100-mL beaker half-way with water, place the watch glass with filter paper over the beaker, and heat to boil for twenty minutes to dry the precipitate (alternatively, you can place the watch glass with filter paper in a drying oven at 130 °C for twenty minutes). Allow to cool, then determine the mass of your precipitate. Heat for another five minutes, cool, and reweigh. The two weights should agree within ±0.05 g or a third heating should be done.

Data and Calculations for Experiment 20

1. Weight of empty beaker ____________________

2. Weight of beaker and sodium sulfate ____________________

3. Weight of sodium sulfate

   Show Calculation ____________________

Name: ______________________________ Section: ____________________

4. Moles of sodium sulfate:

   Show Calculation ____________________

5. Moles of strontium chloride

   $$\text{moles } SrCl_2 = 5.0 \text{ mL } SrCl_2 \left(\frac{10^{-3} \text{ L } SrCl_2 \text{ solution}}{1 \text{ mL } SrCl_2 \text{ solution}}\right)\left(\frac{0.50 \text{ mol } SrCl_2}{1 \text{ L } SrCl_2 \text{ solution}}\right) =$$

   Solve the Equation Shown ____________________

6. Write a balanced MOLECULAR equation for the reaction:

7. Write a balanced NET-IONIC equation for the reaction:

8. Weight of empty filter paper ____________________

9. Weight of filter paper and dried precipitate (first time) ____________________

   Weight of filter paper and dried precipitate (second time) ____________________

   Weight of filter paper and dried precipitate (third time) ____________________

10. Weight of precipitate:

    Show Calculation ____________________

11. Determine the limiting reactant and excess reactant for your reaction. Also, calculate the theoretical yield (in grams) of strontium sulfate.

    Limiting Reactant: ____________________ Excess Reactant: ____________________

    Show Calculation (theoretical product yield) ____________________

Name: ______________________________ Section: ____________________

12. Determine the percentage yield of your reaction.

Show Calculation ____________________

13. Calculate the theoretical yield (in grams) of strontium sulfate if you had used half as much $SrCl_2(aq)$?

Show Calculation ____________________

14. Calculate the theoretical yield (in grams) of strontium sulfate if you had used twice as much $SrCl_2(aq)$?

Show Calculation ____________________

15. Briefly describe how you could have improved your percentage yield in this experiment.

Name: ________________________________ Section: ____________________

Workshop 1 – Math Review

Algebra is an essential skill in solving scientific problems. The following problems review the type of math you will need to use in this course.

1. Given the following equation: **$y = 3x^2 + 7$**

   a) Show your work using algebra (symbols only) to solve for x:

   b) If y = 100, solve for x by entering into your re-arranged equation:

   Write your numerical answer from the calculator: ________________

2. a) Given the equation: $M_1V_1 = M_2V_2$, solve for M_2 (hint: rearrange the symbols)

   b) If $M_1 = 0.100$, $V_1 = 5$, and $V_2 = 250$, then what is the numerical value of M_2?

   M_2 = ________________

Name: ____________________________________ Section: _______________________

3. For the equation $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, use algebra to solve for T_2:

Ask yourself: Did you *actually* solve for T_2 or is your answer in terms of T_2^{-1}?

Check your math: If $V_1 = 950{,}000$, $V_2 = 115{,}600$, $P_1 = 0.980$, $T_1 = 298$, $P_2 = 0.905$, what does $T_2 = ?$

$T_2 =$ ____________

4. Use Power Rules to solve for the following (without a calculator!):

a) $\frac{10^8}{10^2} =$ (a) _____________

b) $(10^4)(10^3) =$ (b) _____________

c) $(10^1)(10^{-2}) =$ (c) _____________

5. a) Write 10^{-7} as a fraction.

b) Write 10^{-2} as a decimal.

6. Average the following numbers:

3.75 4.23 4.95 3.80 4.41 4.72 ________________

Name: ________________________________ Section: _____________________

Workshop 2 – Scientific Notation and Scientific Calculators

1. Write each of the following numbers in proper scientific notation:

(a) 587 (a) ___________________

(b) 0.0077 (b) ___________________

(c) 9,200 (c) ___________________

(d) 406.0 (d) ___________________

(e) 13,800,000 (e) ___________________

(f) 0.0004 (f) ___________________

2. For each of these problems, complete the answer with a 10 raised to the proper power. Note that each answer is expressed to the correct number of significant figures.

(a) $(1.73 \times 10^3)(2.0 \times 10^3) =$ (a) 3.5 × ______________

(b) $\dfrac{6.477 \times 10^5}{3.62 \times 10^3} =$ (b) 1.79 × _____________

(c) $(5.7 \times 10^3)(2.6 \times 10^5) =$ (c) 1.5 × ______________

(d) $\dfrac{2.75 \times 10^{-6}}{2.3 \times 10^3} =$ (d) 1.2 × ______________

(e) $\dfrac{5.80 \times 10^4}{9.53 \times 10^7} =$ (e) 6.09 × _____________

3. Solve each of the following problems, expressing each answer to the proper number of significant figures. Use scientific notation.

(a) $(7.55 \times 10^2)(2.83 \times 10^8) =$ (a) ___________________

(b) $\dfrac{(6.51 \times 10^{-2})(7.07 \times 10^{-5})}{2.92 \times 10^3} =$ (b) ___________________

Name: ________________________________ Section: _____________________

Scientific Calculator

4. Write down the Brand and series number of your Scientific Calculator.
   Example: Texas Instruments TI-30XIIS

   Find at least one other student with the same brand and version of calculator (you may consider working with them for the remainder of this assignment).

5. Enter Avogadro's Number (6.022×10^{23}) into your calculator. Write down the sequence of buttons you used to enter a number in scientific notation on your calculator:

6. The diameter of a penny is 0.01905 meters. Convert this number into scientific notation and then enter into your calculator. Write below what button(s) would allow you to convert the number back to "standard" notation on your calculator.

7. Use this number for all the questions in this problem: $\mathbf{10^{-3}}$

   Write it as a fraction ________________

   Write it as a decimal ________________

   Enter into your scientific calculator. Which button(s) did you use to input?

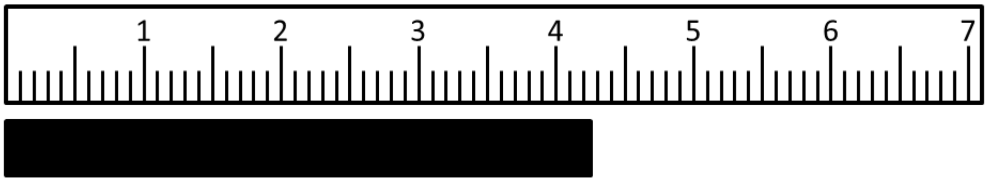
Name: ________________________________ Section: ____________________

Workshop 3 – Significant Figures

Show calculation setups and answers for all problems below.

1. Using the ruler shown on the page, what is the length of the dark rectangle to the correct number of significant figures?

____________ cm



2. How many significant figures are in each of the following numbers?

(a) 8.57 _________ (b) 9.3 ___________ (c) 6.20 ___________ (d) 74,000 ________

(e) 0.058 ________ (f) 0.0085 ________ (g) 0.0790 _________ (h) 0.4020 ________

3. How many significant figures should be in the answer to each of the following calculations? (You may need to solve the math to answer the question, but your final answer is the number of significant figures, NOT the numerical answer to the problem.)

(a) 16.20
+ 0.87

(a) ___________________

(b) 46.837
–9.5

(b) ___________________

(c) 23.3 × 1.73

(c) ___________________

(d) 3.1 × 6.4215

(d) ___________________

(e) $\dfrac{0.4273}{0.1853}$

(e) ___________________

(f) $\dfrac{4.73 \times 8.31 \times 337}{834 \times 357}$

(f) ___________________

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Workshop 4 – Dimensional Analysis

Show your calculation setup for the following problems. Make certain to express the appropriate units and round-off your answers to the proper number of significant figures.

1. Convert 25 ^{o}F to degrees Celsius. _______________

2. Convert –75 ^{o}C to degrees Fahrenheit. _______________

3. A ruler is 48.0 in. long. How long is this in centimeters? _______________

4. A bowling ball weights 15.3 lbs. Calculate its mass in grams. _______________

5. 125 mL of water are contained in a beaker. Convert this to quarts. _______________

6. A baseball bat is 95.9 cm long. How long is this in:

   (a) Millimeters? _______________

   (b) Feet? _______________

7. An object has a mass of 35.8 g and a volume of 40.5 cm^3. Calculate the density of the object in g/mL.

____

Name: ___________________________________ Section: ______________________

8. A rubber stopper weighing 65.4 g is immersed into a graduated cylinder filled with 30.0 mL of liquid. The liquid level then rises to 48.8 mL. Calculate the density of the stopper.

____

9. If the density of the liquid in Problem 8 is 0.785 g/mL, calculate the mass of the liquid in the graduated cylinder.

____

10. A flask contains 365 mL of water. The density of water is 1.00 g/mL. Calculate:

    (a) The mass of the water in grams. ________________

    (b) The volume of the water in liters. ________________

11. The density of CCl_4 is 1.57 g/mL. Calculate the volume of 135 g of CCl_4.

____

12. What is the density (g/mL) of a rectangular block of wood if it measures 4.0 cm thick, 120 mm long, and 0.57 in wide and has a mass of 0.0620 kg? Will the block sink or float in water?

____

<u>Circle one</u>: sink or float

Name: ______________________________ Section: ____________________

Workshop 5 – Nomenclature

Hint: The names and formulas of a variety of polyatomic ions (including ones that your instructor many not have previously given) can be found in the appendix of this lab manual.

A. Provide a chemical name for the following formulas:

1. NaBr ____________________
2. $MgBr_2$ ____________________
3. H_2O ____________________
4. $Ca(NO_3)_2$ ____________________
5. $Fe(NO_3)_2$ ____________________
6. Na_2SO_4 ____________________
7. SO_3 ____________________
8. $(NH_4)_3PO_4$ ____________________
9. $Fe_3(PO_4)_2$ ____________________
10. Cu_2CO_3 ____________________
11. Na_2O ____________________
12. KOH ____________________
13. $Mg(OH)_2$ ____________________

Name: ______________________________ Section: ____________________

B. Provide a formula for the following names:

1. Sodium fluoride ______________________________

2. Calcium iodide ______________________________

3. Sodium phosphate ______________________________

4. Barium phosphate ______________________________

5. Chromium(III) nitrate ______________________________

6. Gold(I) carbonate ______________________________

7. Potassium hydrogen carbonate ______________________________

8. Nickel(I) bicarbonate ______________________________

9. Cobalt(II) acetate ______________________________

10. Ammonium hydrogen sulfate ______________________________

11. Calcium oxide ______________________________

12. Barium hydroxide ______________________________

13. Copper(II) chloride ______________________________

Name: ________________________________ Section: ____________________

C. Harder Set! Provide a chemical name for the following formulas:

1. Na_2S ________________________________

2. $Ca(C_2H_3O_2)_2$ ________________________________

3. $Fe(NO_2)_2$ ________________________________

4. $MgSO_3$ ________________________________

5. $NaHSO_3$ ________________________________

6. Na_2CrO_4 ________________________________

7. $Na_2Cr_2O_7$ ________________________________

8. CCl_4 ________________________________

9. $KClO_3$ ________________________________

10. $Ca(ClO)_2$ ________________________________

11. HNO_3 ________________________________

12. $HBr_{(aq)}$ ________________________________

13. HBr ________________________________

Name: ________________________________ Section: ____________________

D. Harder! Provide a formula for the following names:

1. Sodium permanganate ______________________________________

2. Beryllium chromate ______________________________________

3. Sodium sulfite ______________________________________

4. Calcium hydrogen phosphate ______________________________________

5. Chromium(III) chlorate ______________________________________

6. Sodium perchlorate ______________________________________

7. Sulfur pentachloride ______________________________________

8. Chlorine trioxide ______________________________________

9. Cobalt(III) cyanide ______________________________________

10. Potassium permanganate ______________________________________

11. Potassium carbonate ______________________________________

12. Hydrochloric acid ______________________________________

13. Phosphoric acid ______________________________________

Name: ________________________________ Section: ____________________

E. Still harder set! Provide a chemical name for the following formulas:

1. $HClO_{4(aq)}$ ________________________________

2. Na_2O_2 ________________________________

3. $HI_{(aq)}$ ________________________________

4. $HC_2H_3O_{2(aq)}$ ________________________________

5. NaH ________________________________

6. $TiCl_4$ ________________________________

7. $Cu(MnO_4)_2$ ________________________________

8. NH_4HSO_3 ________________________________

9. $MgSO_4{\cdot}5H_2O$ ________________________________

10. $Ca(ClO_3)_2$ ________________________________

11. $H_2Cr_2O_{7(aq)}$ ________________________________

12. $H_2CO_{3(aq)}$ ________________________________

13. CO_2 ________________________________

Name: ________________________________ Section: ____________________

F. Still harder! Provide a formula for the following names:

1. Hydrogen peroxide ______________________________________
2. Arsenic trichloride ______________________________________
3. Potassium chromate ______________________________________
4. Chromic acid ______________________________________
5. Potassium hypochlorite dihydrate ______________________________________
6. Carbon disulfide ______________________________________
7. Ammonia ______________________________________
8. Iron(III) dichromate ______________________________________
9. Chloric acid ______________________________________
10. Copper(II) permanganate ______________________________________
11. Sodium hydrogen phosphate ______________________________________
12. Magnesium sulfide ______________________________________
13. Methane ______________________________________

Name: ________________________________ Section: ____________________

Workshop 6 – Writing and Balancing Equations

Balance the following reactions. If given words, write the formulas and balance reactions in the space below the words. Remember which elements are diatomic. Include phases.

1. $Al(s) + O_2(g) \rightarrow Al_2O_3(s)$
2. $Fe(ClO_3)_3(s) \rightarrow FeCl_3(s) + O_2(g)$
3. $Ag(s) + HI(aq) \rightarrow AgI(s) + H_2(g)$
4. $H_2O(l) + N_2O_5(g) \rightarrow HNO_3(aq)$
5. $NH_3(g) + O_2(g) \rightarrow NO_2(g) + H_2O(l)$
6. $C_3H_8(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
7. Aqueous sodium hydroxide and sulfuric acid react to form aqueous sodium sulfate and liquid water
8. Methane gas (CH_4) and oxygen gas react to form carbon dioxide gas and water.
9. Solid calcium oxide and water create aqueous calcium hydroxide.
10. Solid sodium bicarbonate decomposes when heated to form solid sodium carbonate, carbon dioxide gas and liquid water.
11. Aqueous potassium sulfide and lead(II) nitrate react to produce solid lead(II) sulfide and aqueous potassium nitrate.
12. Aqueous acetic acid and potassium sulfite react to form aqueous potassium acetate, water and sulfur dioxide gas.

Name: ________________________________ Section: ____________________

Predict products and Balance the following reactions. If no reaction takes place, write NR for no reaction. Include phases.

13. Combustion reactions: *nonmetals + O_2 → nonmetal oxides* *(CO_2, H_2O)*

a) $C_7H_{16}(s)$ + $O_2(g)$ →

b) $C_6H_6(l)$ + $O_2(g)$ →

c) $C_4H_{10}O(l)$ + $O_2(g)$ →

d) $C_7H_6O_2(s)$ + $O_2(g)$ →

14. Double displacement reactions: *AB + CD → AD + CB*

a) $AlCl_3(aq)$ + $Pb(NO_3)_2(aq)$ →

b) $HC_2H_3O_2(aq)$ + $Ba(OH)_2(aq)$ →

c) $K_2CrO_4(aq)$ + $SnF_4(aq)$ →

d) $Ca(HCO_3)_2(aq)$ + $HBr(aq)$ →

15. Mixed reactions: Classify, Predict products, and Balance. Write the formulas and balance reactions in the space below the words. Identify all types of reactions for each in the margin.

a) $HCl(aq)$ + $Sr(OH)_2(aq)$ →

b) $AlCl_3(aq)$ + $NaNO_3(aq)$ →

c) $C_2H_4(g)$ + $O_2(g)$ →

d) $HNO_3(aq)$ + $Li_2SO_3(aq)$ →

16. Word reactions: Write formulas and balance the reactions.

a) Crude gunpowders often contain a mixture of potassium nitrate (KNO_3) and charcoal (solid carbon). When heated until a reaction occurs, a solid residue of potassium carbonate (K_2CO_3) is produced. The explosive force of the gunpowder comes from the fact that two gases are also produced, carbon monoxide and nitrogen, which increase in volume with great force and speed.

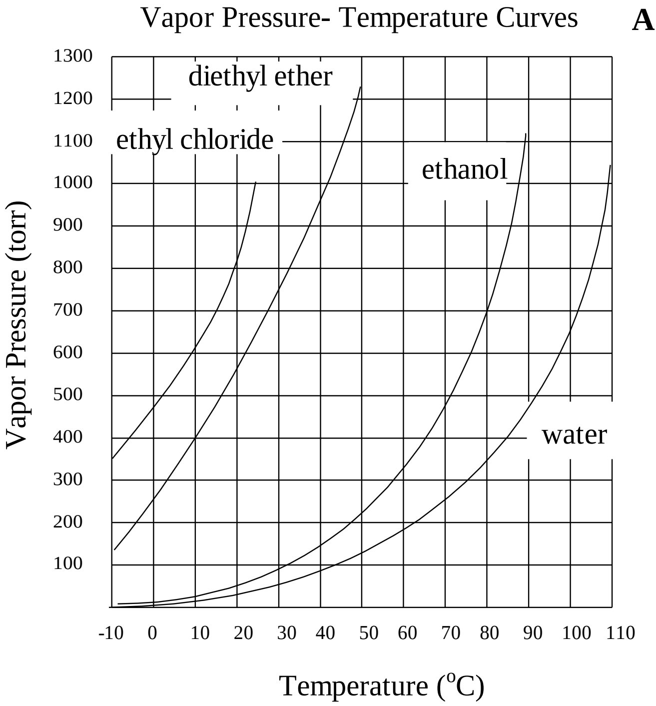
b) A method of preparing pure iron involves heating iron(III) oxide and carbon monoxide together; they react to produce solid iron and carbon dioxide gas.

c) The following reaction takes place in termites as they digest wood. Solid glucose, $C_6H_{12}O_6$, and liquid water react to produce aqueous acetic acid ($HC_2H_3O_2$), carbon dioxide, and hydrogen gas. Write a balanced chemical equation for the reaction including phases. (There are several correct answers possible, try to come up with more than one.)

Name: ________________________________ Section: ____________________

Workshop 7 – Graphical Representation of Data

Answer the following questions by plotting and interpreting the data respectively.



A. Reading a Graph

From the figure at the left, read values for the following:

1. The vapor pressure of water at 70 °C.

____

2. The temperature at which diethyl ether has a vapor pressure of 600 torr.

____

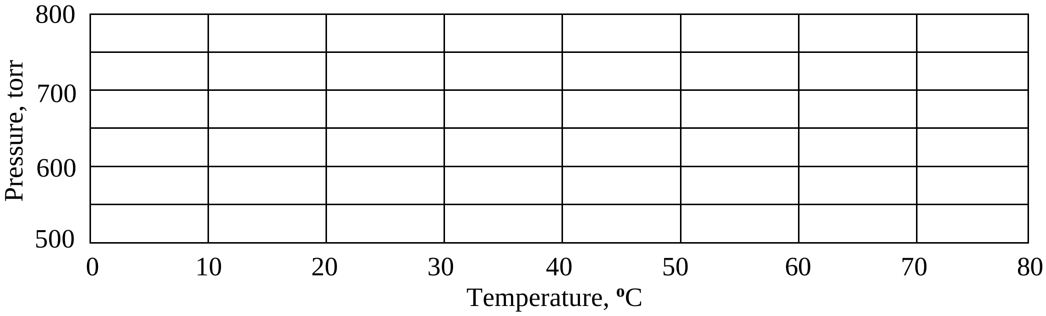
3. The temperature at which ethyl chloride has the same pressure ethanol has at 80 °C.

____

B. Plotting Graphs

1. Plot the following pressure-temperature data for a gas on the graph. Draw the best possible straight line through the data.

Temperature, °C:	0	20	40	60	80
Pressure, torr:	550	605	665	720	775



2. Solve for the slope of the graph above. Slope is defined as rise/run ($\Delta y/\Delta x$).

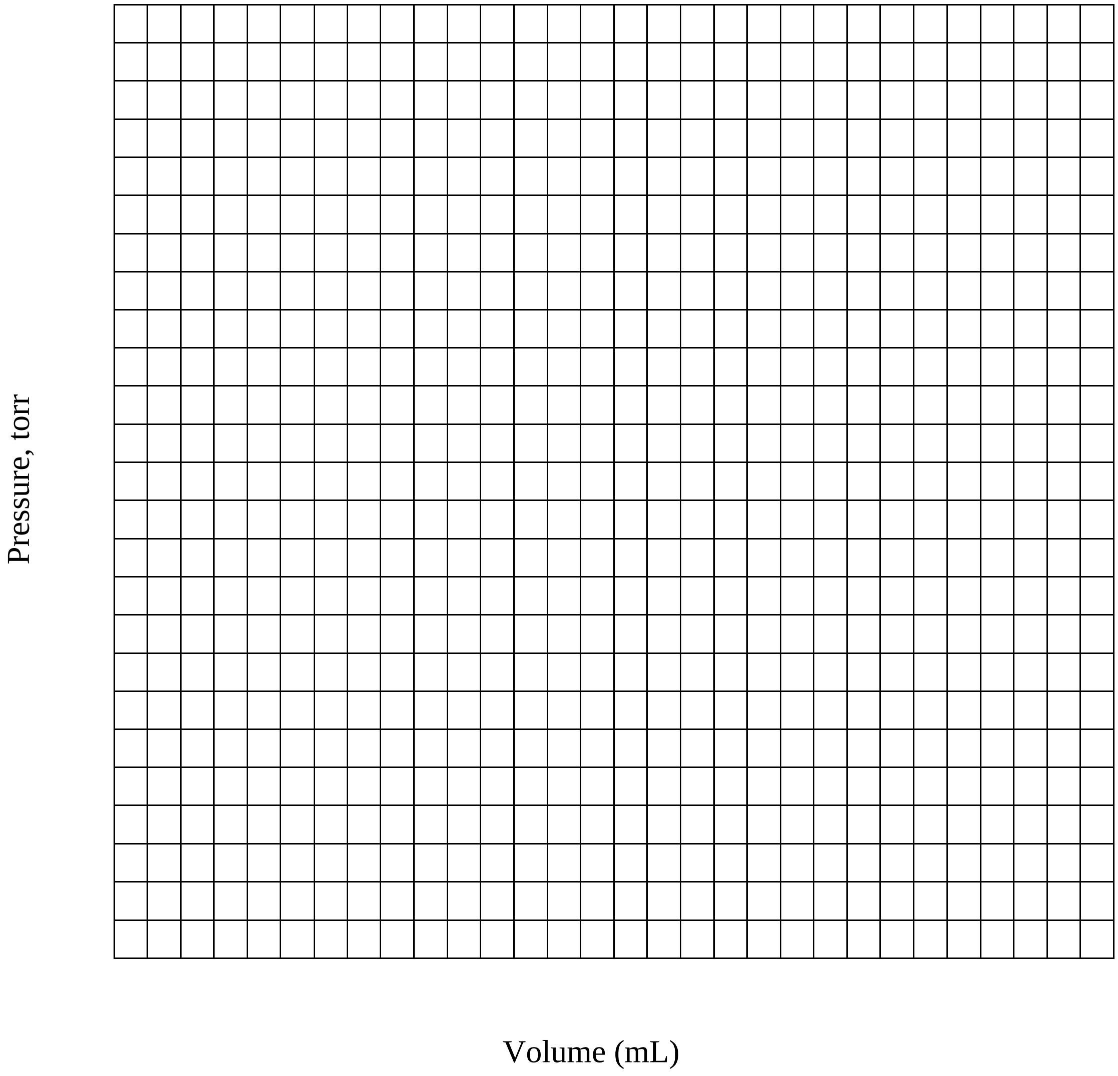
Slope = _____________ (include units)

Name: ______________________________ Section: ____________________

3. (a) Study the data given below; (b) determine suitable scales for pressure and for volume and mark these scales on the graph; (c) plot the eight points on the graph; and (d) draw the best possible CURVE through these points.

Pressure-Volume data for a gas

Volume, mL	107	76.4	55.7	45.6	35.2	29.7	24.3	20.1
Pressure, torr	25	35	48	60	76	90	110	133



Read from your graph:

(a) The pressure at 100 mL ________________________

(b) The volume at 70 torr _________________________

Name: ______________________________ Section: ____________________

Workshop 8 – Quantum Mechanics

Show calculation setups and answers for all problems below.

1. An FM radio station has a frequency of 88.9 MHz (1 MHz = 10^6 Hz). Determine the wavelength (in nm).

____

2. Violet light has a wavelength of about 410 nm. What is its frequency (in Hz)?

____

3. An advertising sign gives off red light and green light.

   A. Which light has the higher energy? Briefly explain below.

   ____

   B. One of the colors has a wavelength of 680 nm, and the other has a wavelength of 500 nm. Identify which color has which wavelength. Explain your identifications below.

   Red = __________________

   Green = __________________

   C. Which light has the higher frequency? Briefly explain below.

   ____

4. Write the symbols for three cations and three anions *isoelectronic* with neon:

________ ________ ________ ________ ________ ________

Name: ______________________________ Section: ____________________

5. Write complete and abbreviated electron configurations for each of the following atoms/ions:

   A. S
      Complete: ______________________________

      Abbreviated: ______________________________

   B. Nb
      Complete: ______________________________

      Abbreviated: ______________________________

   C. Sb^+
      Complete: ______________________________

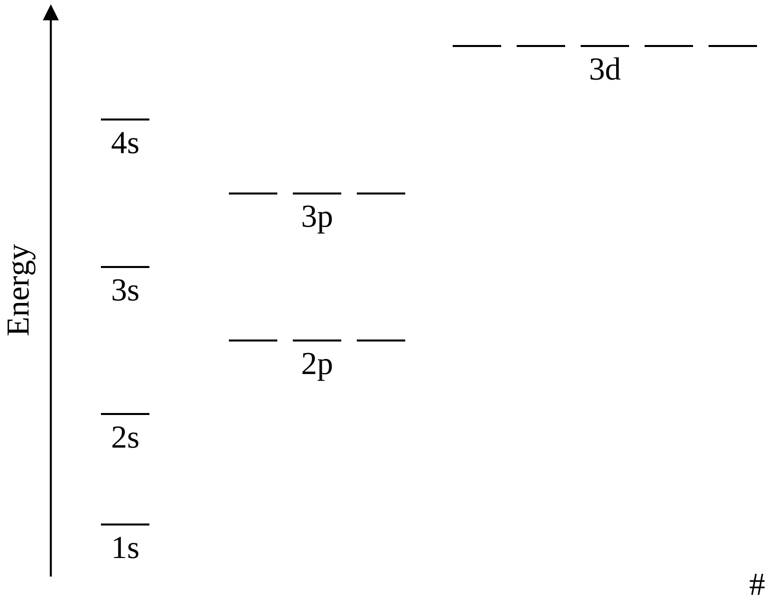
      Abbreviated: ______________________________

6. Arrange the following forms of electromagnetic radiation in order of increasing energy:

   A. gamma rays from a supernova
   B. infrared rays from a hot plate
   C. ultraviolet light from the sun
   D. radiowaves from an MP3 player
   E. green light from chlorophyll

____

7. Complete the orbital energy diagram below for Co. How many unpaired electrons does the Co atom have?



unpaired electrons ________

Name: ________________________________ Section: ____________________

Workshop 9 – Mole Conversions

Show calculation setups and answers for all problems below. Use scientific notation for very large or very small numbers.

1. Find the molar mass of (a) carbonic acid, H_2CO_3; (b) aluminum sulfate, $Al_2(SO_4)_3$; and (c) ammonium dichromate, $(NH_4)_2Cr_2O_7$.

(a) ______________________________

(b) ______________________________

(c) ______________________________

2. A sample of nickel(II) phosphate, $Ni_3(PO_4)_2$, weighs 114 g. How many moles are in this sample?

____

3. What is the mass (in kg) of 35.6 moles of methane gas, CH_4?

____

4. Calculate the molecules of copper(II) nitrite, $Cu(NO_2)_2$, in 0.92 mol $Cu(NO_2)_2$.

____

5. How many molecules of water, H_2O, are present in 28.4 g of H_2O?

____

Name: ____________________________________ Section: _______________________

6. Find the weight (in mg) of one atom of gold, Au.

____

7. Determine the weight (in g) of nitrogen atoms in 6.14×10^{30} molecules of dinitrogen tetroxide, N_2O_4.

____

8. Calculate the percent composition by mass of aluminum hydroxide, $Al(OH)_3$.

Al __________________________________

O __________________________________

H __________________________________

9. Caffeine, a compound found in coffee, tea, and cola drinks is found to contain 49.47% C, 5.19% H, 28.86% N, and 16.48% O by mass. Its experimentally determined molar mass is 194 g/mol. What is the empirical formula of caffeine? What is its molecular formula?

Empirical __________________________________

Molecular __________________________________

10. How many mL of liquid mercury (Hg) with a density of 13.6 g/mL must you dispense to have 1.56×10^{-3} mol?

____

Name: ______________________________ Section: ____________________

Workshop 10 – Stoichiometry I

Show calculation setups and answers for all problems below.

1. Ammonia gas will react with oxygen gas to yield nitrogen monoxide gas and water vapor.

   (a) Write the balanced chemical equation for this reaction.

   (b) How many moles of ammonia will react with 6.73 g of oxygen?

   ____

   (c) If 6.42 g of water is produced, how many grams of oxygen gas reacted?

   ____

   (d) If the reaction uses up 9.43×10^5 g of ammonia, how many kilograms of nitrogen monoxide will be formed?

   ____

   (e) When 2.51 g of ammonia react with 3.76 g of oxygen, 2.27 g of water vapor are produced. What is the percentage yield of water?

   ____

Name: ________________________________ Section: ____________________

2. Use the balanced equation below to solve the following problems:

$$2\ KMnO_4 + 16\ HCl \rightarrow 5\ Cl_2 + 2\ KCl + 2\ MnCl_2 + 8\ H_2O$$

(a) How many moles of HCl are required to react with 28 g of $KMnO_4$?

____

(b) How many Cl_2 molecules will be produced using 1.5 mol $KMnO_4$?

____

(c) To produce 29.0 g of $MnCl_2$, what mass (in g) of HCl will need to react?

____

(d) How many moles of water will be produced when 5.0 mol of $KMnO_4$ are consumed?

____

(e) What is the maximum mass of Cl_2 that can be produced by reacting 65.9 g of $KMnO_4$ with 18.0 g of HCl?

____

Name: ________________________________ Section: ____________________

Workshop 11 – Gas Laws

Show calculation setups and answers for all problems below.

1. You have a sample of 2.0 L of oxygen gas at 3.0 atm pressure. If you reduce the pressure to 0.50 atm, what is the volume of the gas?

____

2. A sample of argon gas occupies 2.50 L at 25.0 ^{o}C. If the gas is heated at constant pressure, what will the volume be at 99.9 ^{o}C?

____

3. A 252 mL sample of nitrogen gas is at 715 torr and 25.0 ^{o}C. What volume would the sample occupy at 760. torr and 0 ^{o}C?

____

4. How many moles of methane (CH_4) are present in a 10.0 L sample at STP?

____

5. How many liters would 14.0 grams of chlorine gas occupy at 300.0 K and 1.51 atm?

____

6. How many grams of CH_4 at STP would fill a 1.00 L flask?

____

7. A gas has a pressure of 1.07 atm, a volume of 13.7 L, and a mass of 28.0 g at a temperature of 294 K. What is the molar mass of this gas?

____

8. A sample of O_2 gas is stored at 30.0 ^{o}C and 755 torr. If the volume was 125 mL, how much did the oxygen weigh?

____

9. Small quantities of hydrogen gas can be prepared in the laboratory by the addition of aqueous hydrochloric acid to metallic zinc according to the following balanced equation:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Suppose 240. mL of hydrogen gas is collected at 40.0 ^{o}C and has a pressure of 1.030 atm by this process. How many grams of zinc must have reacted to produce this quantity of hydrogen?

____

Name: ______________________________ Section: ____________________

Workshop 12 – Stoichiometry II

Show calculation setups and answers for all problems below.

1. Consider the balanced chemical equation to solve the following problems:

$$6\ KI + 8\ HNO_3 \rightarrow 6\ KNO_3 + 2\ NO + 3\ I_2 + 4\ H_2O$$

(a) If 26.0 g of KI are reacted, how many grams of I_2 will be formed?

____

(b) What volume of NO gas, measured at STP, will be produced if 39.0 g of HNO_3 are reacted?

____

(c) How many milliliters of 6.00 M HNO_3 will react with 26.0 g of KI?

____

(d) When the reaction produces 0.500 g of NO, how many molecules of I_2 will be produced?

____

(e) How many grams of iodine can be obtained by reacting 25.0 mL of 0.350 M KI solution?

____

Name: ________________________________ Section: ____________________

2. Consider the Haber Process for the synthesis of ammonia shown below. Use the given equation to solve the following problems:

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

(a) If 4.0 g of H_2 react, how many grams of NH_3 will be formed?

____

(b) When 3.25 mol of N_2 react, what volume of NH_3, measured at STP, will be formed?

____

(c) What volume of NH_3 will be formed when 16.0 L of H_2 are reacted at STP?

____

(d) How many molecules of NH_3 will be formed when 20.0 L of N_2 react at STP?

____

(e) What volume of NH_3, measured at 35 ºC and 720. torr, will be produced from 12.0 g of H_2?

____

(f) If a mixture of 14.0 L of N_2 and 24.0 L of H_2 are reacted, what volume of NH_3 can be produced at STP?

____

Name: ________________________________ Section: ____________________

Workshop 13 – Solution Concentrations

Show calculation setups and answers for all problems below.

1. What is the percent composition by mass of a solution made by dissolving 25.0 g of sodium phosphate, Na_3PO_4, in 50.0 g of water?

Na_3PO_4 ________________________

H_2O ________________________

2. How many moles of magnesium hydroxide, $Mg(OH)_2$ are required to prepare 2.50 L of a 0.350 M solution?

____

3. Determine the molarity of a solution if 2.75 g of potassium hydroxide, KOH, are dissolved in water to make 250. mL of solution.

____

4. How many milliliters of a 0.250 M solution can be prepared by dissolving 4.00 g of NaCl in water?

____

5. How many grams of lithium bromide, LiBr, could be recovered by evaporating 550. mL of 20.0 percent LiBr solution to dryness (d = 1.34 g/mL)?

____

Name: ________________________________ Section: ____________________

6. How many milliliters of 6.0 M HCl is needed to prepare 500. mL of a 0.150 M HCl solution?

____

7. A sample of potassium hydrogen phthalate, $HKC_8H_4O_4$, weighing 0.512 g was dissolved in water and titrated with 24.82 mL of an NaOH solution. Calculate the molarity of the NaOH solution.

____

8. How many grams of hydrogen nitrate are in 75. mL of concentrated (18 M) HNO_3 solution?

____

9. A sulfuric acid solution has a density of 1.49 g/mL and contains 32 percent H_2SO_4 by mass. What is the molarity of this solution?

____

10. Oxalic acid reacts with sodium hydroxide according to the following equation:

$$H_2C_2O_4 + 2\ NaOH \rightarrow Na_2C_2O_4 + 2\ H_2O$$

A 25.00 mL sample of the $H_2C_2O_4$ solution required 19.62 mL of 0.341 M NaOH for neutralization. Calculate the molarity of the acid.

____

Name: ________________________________ Section: ____________________

Workshop 14 – Trends on the Periodic Table

Exercise I

This chart represents the main group (representative elements) portion of the periodic table.

A. Several trends are listed to the sides and below the chart. Use a periodic table with proper values to determine the direction of these trends. Convert the underlines into arrows by adding heads (i. e. → or ←) to each underline to indicate the direction of each trend.

B. In each box, write the electronic configuration of the valence electrons of that element. See the box containing element 84 (polonium) as an example.

IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA or 0
3	4	5	6	7	8	9	10
11	12	13	14	15	16	17	18
19	20	31	32	33	34	35	36
37	38	49	50	51	52	53	54
55	56	81	82	83	84 $6s^2 6p^4$	85	86
87	88						

Atomic Radii Increase

Metallic Properties Increase

Ionization Energy Increase

Electronegativity Increase

Nonmetallic Properties Increase

Atomic Radii Increase

Ionization Energy Increase

Electronegativity Increase

Exercise II

Fill in the blank spaces.

Group Number	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
Number of valence electrons				4				
Electronic configuration of valence electrons. Omit principle quantum number.				s^2p^2				
Common oxidation states				±4				

Name: ____________________________ Section: ____________________

Exercise III

Fill in all the boxes on this periodic table with the atomic numbers of the elements and the electronic configurations of the last ground state electron to be added to the element.

1 $1s^1$																	
3	4 $2s^2$											5 $2p^1$					
															$3p^4$		
			$3d^2$														
$5s^1$													$5p^2$				
	56	57	72					$5d^7$									
	88	89	104														

58													
90									$5f^{10}$				

Name: ________________________________ Section: ____________________

Workshop 15 – Simple Nomenclature of Inorganic Compounds

I. Ionic Compounds (Compounds composed of a metal and a nonmetal or a metal and a polyatomic ion.)

A. Monatomic cations (positive ions formed from one atom) from Groups IA, IIA, IIIA and hydrogen take the name of the element from which they were derived.

H^+	hydrogen	K^+	potassium
Mg^{+2}	magnesium	Al^{+3}	aluminum

B. When a metal forms more than one ion (if it has variable ionic charge), it is necessary to distinguish between the possible ions. We will use the Stock method which gives the charge of the ion as a Roman number in parentheses immediately after the name of the metal. This will occur with most of the transition metals and the metals of groups IVA and VA.

Fe^{+2}	iron(II)	Ni^+	nickel(I)	Pb^{+2}	lead(II)
Fe^{+3}	iron(III)	Ni^{+2}	nickel(II)	Pb^{+4}	lead(IV)

C. Monatomic anions (negative ions formed from one atom) are named by adding the suffix **–ide** to the stem of the name of the nonmetal from which they are derived. These names should be memorized.

F^-	fluoride	O^-	oxide	N^{-3}	nitride
Cl^-	chloride	S^{-2}	sulfide	H^-	hydride
Br^-	bromide	Se^{-2}	selenide	I^-	iodide

D. The names and formulas of these polyatomic ions must be memorized. The charge is an essential part of the formula.

$C_2H_3O_2^-$	acetate	ClO^-	hypochlorite
CO_3^{-2}	carbonate	ClO_2^-	chlorite
HCO_3^-	bicarbonate (hydrogen carbonate)	ClO_3^-	chlorate
OH^-	hydroxide	ClO_4^-	perchlorate
NO_3^-	nitrate	MnO_4^-	permanganate
NO_2^-	nitrite	SO_4^{-2}	sulfate
CrO_4^{-2}	chromate	SO_3^{-2}	sulfite
$Cr_2O_7^{-2}$	dichromate	HSO_3^-	bisulfite (hydrogen sulfite)
PO_4^{-3}	phosphate	HSO_4^-	bisulfate (hydrogen sulfate)
HPO_4^{-2}	hydrogen phosphate	O_2^{-2}	peroxide
NH_4^+	ammonium	CN^-	cyanide

Name: ______________________________ Section: ____________________

E. An ionic compound is a combination of one or more cations, and one or more anions. To name the compound, name the cation, then name the anion.

K_2S	potassium sulfide	NH_4Cl	ammonium chloride
$AlCl_3$	aluminum chloride	NH_4NO_3	ammonium nitrate

F. If the cation is a metal with variable ionic charge, you must determine the charge on the metal so that you know what number to put in the parentheses. To do this, look at the anion(s). The charge on the anion(s) multiplied by the number of anions gives the total negative charge. Since the number of positive charges and negative charges in a compound must equal, the total positive charge must be the absolute value of the total negative charge. The total positive charge should be divided by the number of metal ions in the formula to give the charge on an individual ion. This is the number which goes in the parentheses.

CuBr	copper(I) bromide	$Pb(C_2H_3O_2)_4$	lead(IV) acetate
$CuBr_2$	copper(II) bromide	$PbSO_4$	lead(II) sulfate
CuS	copper(II) sulfide	$NiCl_2$	nickel(II) chloride
$Fe(OH)_3$	iron(III) hydroxide	Cu_2SO_3	copper(I) sulfite

G. The formula of an ionic compound must contain equal numbers of positive charges and negative charges. When you are given a name and you need to write a formula, you may need to use several cations and/or anions for the number of charges to be equal. Often, you can use the charge on the **cation** as the number of **anions** and the **absolute value** of the charge on the **anion** as the number of **cations**. However, sometimes when you try to do this you will get a formula where the number of cations and the number of anions have a common factor. In this case you must divide both numbers by that common factor to give you the correct empirical formula.

aluminum iodide	AlI_3	sodium sulfide	Na_2S
tin(IV) chloride	$SnCl_4$	iron(III) oxide	Fe_2O_3
chromium(VI) oxide	CrO_3	magnesium oxide	MgO

II. Binary Compounds of Two Nonmetals (Covalent Compounds)

A. A compound composed of two nonmetals is a covalent compound. The compound's name is written by taking the less electronegative element first, writing the name of that element, then taking the more electronegative element and adding the **–ide** suffix to the stem of the name of the element. If more than one atom of an element is in the formula, prefixes are used to indicate the numbers. These prefixes are:

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

The prefix for one is mono–, but it no longer has to be used.

P_4O_{10}	tetraphosphorus decoxide	NF_3	nitrogen trifluoride
CCl_4	carbon tetrachloride	SiO_2	silicon dioxide
CS_2	carbon disulfide	NO	nitrogen oxide
Cl_2O	dichlorine oxide	N_2O_3	dinitrogen trioxide

B. There are several binary compounds that have common names.

H_2O	water	NH_3	ammonia	CH_4	methane*

*For organic hydrocarbons (carbon and hydrogen), a different method is employed for naming.

III. Acids

A. Acids are compounds that contain hydrogen and which, under certain conditions, ionize to form hydrogen ions and a negative nonmetal ion or a negative polyatomic ion. You can always recognize the formula of an acid because in acids, but not in other compounds, the H is written first. (Water is a very weak acid.)

B. Binary acids: Binary acids contain hydrogen and one other element. They are formed from hydrogen ions and a monatomic nonmetal ion. They are given two different names depending on whether or not they are dissolved in water. If they are dissolved in water their names are formed by taking the name of the anion, dropping the –**ide** suffix, and then adding both the **hydro–** prefix and the **–ic acid** suffix. If binary acids are not dissolved in water, they are named as ionic compounds.

HCl(aq)	hydrochloric acid	HCl(g)	hydrogen chloride
HF(aq)	hydrofluoric acid	HF(g)	hydrogen fluoride
H_2S(aq)	hydrosulfuric acid	H_2S(g)	hydrogen sulfide

C. Oxyacids: Oxyacids are acids that contain oxygen in addition to hydrogen and another element. They can be formed by combining hydrogen ions and polyatomic anions. For Chem 12, you only need to worry about the acids of polyatomic ions that end in **–ate**. For these acids, the acid is named by taking the name of the polyatomic ion and changing the **–ate** ending to **–ic acid**. Note that the “hydro–” prefix is only used for binary acids, not for oxyacids.

The names of oxyacids are the same whether or not they are dissolved in water.

HNO_3	nitric acid	$HC_2H_3O_2$	acetic acid
$HClO_3$	chloric acid	H_2CO_3	carbonic acid

There are several acids whose names are not quite what you would expect. Memorize these names.

H_2SO_4	sulfuric acid (not sulfic acid)
H_3PO_4	phosphoric acid (not phosphic acid)

D. To write the formula of an acid when given the acid’s name, determine the name of the anion that corresponds to the acid by using the rules in section B and C in reverse. For each negative charge on the formula of the anion, add one H^+ in order to give a neutral formula.

hydroselenic acid	H_2Se(aq)
perchloric acid	$HClO_4$

Name: ______________________ Section: ______________

Write the formulas and the names of the compounds formed by combining these ions.

	SO_4^{-2}	NO_3^{-}	ClO_3^{-}	CO_3^{2-}	PO_4^{3-}	HSO_4^{-}	S*	OH^{-}	I*
NH_4^{+}									
Na*									
Mg*									
Al*									
Cu^{+2}									
Fe^{+3}									

* Predict the oxidation number and write it on the symbol.

Name: ________________________________ Section: ____________________

Mini-Report 1 – How Do Scientists Report Data?

In order for science to progress forward, it is important for scientists to communicate their results to each other. This allows scientists to learn from each other's results, to notice and analyze trends across multiple sets of experimental data, to propose new theories to account for phenomena seen across multiple scientific areas, and to question and verify other scientists' experimental results.

As such, scientists strive to report their results clearly, concisely and broadly to other scientists. To assist in broad distribution of these results, a large number of scientific journals are published every year. Some journals are published monthly or quarterly, while others are published weekly. Some journals have very narrow target audiences, while others target entire disciplines or even cross-disciplinary audiences. While individual journals may have specific formatting requirements, most scientific articles are written in the same general format. The purpose of this assignment is to introduce you to the way that scientific articles are written.

Instructions:

1) Download the sample scientific report from the Chemistry Department website.

2) Write a short written summary (in paragraph form) of the way in which the scientific report is written. Your summary should be typed, 1–2 pages in length, 12 point font, and double spaced. Your summary should address the following questions (not necessarily in this order):

    a) There are six sections of the sample report, each of which is clearly identified with a bold title. What is each section called? What is the apparent purpose of each section? How do they differ in purpose? Why might other scientists want the report to be divided into clearly identifiable sections?

    b) Is the paper written in first–person ("Next, I added 3.21 grams of salt"), second–person ("Next, add 3.21 grams of salt"), or third–person ("Next, 3.21 grams of salt was added")? Is the paper mostly factual in nature, or does it contain many opinions? In which (if any) labeled sections are the author's opinions included?

3) Be sure to include your name on your written report.

Name: ________________________________ Section: ____________________

Mini-Report 2 – Abstract

The “abstract” section of an experiment report is essentially a miniature version of a complete experiment report. It contains a brief statement about the background of the experiment, includes a brief summary of the procedure, and thoroughly summarizes the results and conclusions of the experiment.

While it may sound easy to write a miniature report, it is actually very challenging, because the abstract is the shortest section of the report – usually around half of a page, and certainly no longer than an entire page. When writing a full experiment report, it is usually easiest to write the abstract last. That way, you already know what you will be writing about in your report and you can simply summarize what you’ve already written.

The key to writing a good abstract is to be very concise, omitting unnecessary information and focusing on the essential points of each section of your report.

As with any section of your report, if you reference any external sources, you must include those references in your abstract.

Instructions:

Write an abstract section for the lab that you just completed. It must be double spaced, 12 point font, and should have your name on it. The references that you use to write your abstract section should be cited in a separate “references” section.

Name: ________________________________ Section: ____________________

Mini-Report 3 – Introduction

The “introduction” section of an experiment report is essentially the first part of the formal written report. This section must convey two key pieces of information: 1) the background information that is necessary to place the experiment in context and 2) the purpose of this experiment.

Scientists rarely invent new areas of science. Instead, each scientific experiment uses information from previous experiments to place each new experiment in context. The key to determining what information must be included in the background portion of the introduction is to ask yourself: “What information did I need to know to understand this experiment?” You must then research all of this information and write a cohesive introduction describing the theories, laws, concepts, and equations that allowed your experiment to be conducted. Since all of this information was discovered in previous experiments, it is required that you reference the sources of information that you used to prepare your introduction. In the “real world,” a scientist will only use primary sources when doing this research – that is, they will only reference the original experiments conducted and the original authors that conducted each experiment. For this class, however, you are allowed to reference secondary (or tertiary) sources, such as your textbook. These non-primary sources were written by authors who examined primary source material and then summarized the key findings in an easy-to-use format. The most important rules to remember, however, when preparing your background are: 1) if a law, theory, equation, or concept was needed to complete the experiment, then you need to describe that law, theory, equation, or concept in the background section of your introduction and 2) if you get information from ANYWHERE (lecture, lab manual, textbook, etc.), you need to reference the source.

After you have written your background section, the introduction must conclude with a statement of purpose. The statement of purpose explains what the purpose of YOUR experiment is. In other words, how will your experiment use the material that you wrote about in your background in a new and interesting way? The statement of purpose leads directly into the method / materials section of your report, so it serves as a transition from background material (other people’s experiments) to your experiment.

The introduction section of a report is a bit challenging to write, because it requires the author to research the background material of the experiment. Because of this fact, authors typically write the background section of a report after they have completed the other parts of the report. That way, they will know which background material was necessary to understand their experiment and can focus their research on this material.

Instructions:
Write an introduction section for the lab that you just completed. It must be double spaced, 12 point font, and should have your name on it. The references that you use to write your introduction section should be cited in a separate “references” section.

Name: ________________________________ Section: ____________________

Mini-Report 4 – Method / Materials

The “method/materials” section of an experiment report is the section in which a scientist communicates the specific steps that were done in order to perform the experiment described in the report. This section must be clearly written so that another scientist would be able to duplicate the experiment exactly. This section may allude to data collected during the experiment, but the data is not typically given explicitly and no calculations are shown or performed in this section of the report. The data is not analyzed in any way within this section and conclusions are not drawn from the results of the experiment within this section.

The method/materials section serves as a record of the experiment that was performed. It is NOT an instruction manual or recipe for future students. Therefore, if your lab manual says “Add about 0.3 g of salt”, you would write “0.321 g of salt was added” in your report. The methods/materials section is typically written in the third-person and NOT in the first person (***incorrect:*** “I added 0.321 g of salt”) or in the second person (***incorrect:*** “Add 0.321 g of salt”). Also included in this section are the observations made by the experimenter as the experiment was performed. If the color of something changed or if something started to boil when you heated it, then this should certainly be mentioned in the procedure. ***Example:*** “The solution was brought to a boil and the blue precipitate changed to a dark black color over a period of six minutes.”

In order to ensure that future scientists who are following your procedure are not injured, it is imperative that you include any safety information that scientists should be aware of. If something is dangerous or toxic, this should be mentioned explicitly in your procedure, and your procedure should explain how to avoid any potential problems or accidents associated with this danger.

Your procedure should be written using language that could be understood by a future student who has taken as much chemistry as you have, but has never done or seen this experiment. That means that it is not necessary to explain HOW to weigh something or to use a thermometer, but it may be necessary to explain specific details about THIS experiment, particularly if you had never done an experiment before that was comparable to this one.

Instructions:
Write a method / materials section for the lab that you just completed. It must be double spaced, 12 point font, and should have your name on it. If you use any references to write your method / materials section, they should be cited in a separate “references” section.

Name: ________________________________ Section: ____________________

Mini-Report 5 – Results / Calculations

The "results / calculations" section of an experiment report is the section in which a scientist communicates the specific results (including all numeric data) of the experiments performed and also shows any values that can be calculated from the numeric data obtained. The "results / calculations" section should be free of opinion. This allows another scientist to view the data obtained from the experiment without being biased by the author's own opinion of the experiment performed. Any commentary about whether the experiment was successful or what should be done differently is NOT included in this section of the report. Any conclusions that are not immediately obvious from the data given should also be excluded from this section. (For example, it's alright to state that the percent yield was "98.6%," but it is not alright to state that "this percent yield proves that mass is conserved for this specific reaction." Not only is this an opinion, but there is no obvious equation for "proof".)

Experimental data obtained may be presented in a variety of ways. If an experiment or calculation is done only one time, the data is typically mentioned in the text. If an experiment or calculation is repeated, however, the data is typically given in a table. Sometimes graphs are also made from the data obtained during the experiment. Any graph or table must be numbered and a caption must be provided that explains what the graph or table represents. Graphs and tables are numbered independently. (Example: a report may include a "Table 1," a "Figure 1," and an "Equation 1." The first table is called "Table 1" and the first graph is called "Figure 1," even if these are unrelated).

Any equations that were solved are also explicitly shown (and numbered) within this section. You should also show a "sample calculation" in which your data has been plugged into the equation and the answer is shown so that a reader can see how your data is used to reach the final answer. If an equation is used more than one time, then only one sample calculation should be shown. If a calculation is very obvious (for example, taking an average), then it may be omitted completely. Consult your instructor as to what calculations are "very obvious." When in doubt, include the equation!

Any tables, figures, equations, etc. that you have numbered MUST be explicitly mentioned in the text. Imagine that these tables, figures, equations, etc. are NOT a part of your report. Instead, the text of the report should discuss the results and calculations and then refer to the tables, figures, and equations in which those results can be found. The reader is expected to "find" the table, figure, or equation as it is mentioned in the text. (Example: "The mass of each metal was converted to moles using *Equation 1*, and the results were graphed versus heat capacity, as shown in *Figure 1*.")

Instructions:

Write a results / calculations section for the lab that you just completed. It must be double spaced, 12 point font, and should have your name on it. If you use any references to write your results / calculations section, they should be cited in a separate "references" section.

Name: ________________________________ Section: ____________________

Mini-Report 6 – Discussion / Conclusion

The “discussion / conclusion” section of an experiment report is the only section in which a scientist communicates his or her opinions of the experiment performed. For example: Did the experiment work well? What should have been done differently? How does the scientist know whether or not the experiment was successful? These are all considerations that the scientist must make when writing the concluding section of the report.

There are basically three parts to this section: (1) Interpretation of the results – the author must explain how the theory (discussed in the “background” section of the report) allows for the numeric and observational results to be interpreted and must use the theory to draw conclusions about the experiment (was it successful, what were the identities of unknown substances or numeric quantities, etc.). Be sure to address any questions asked in the lab manual! (2) Discussion of error – no experiment is perfect. In fact, most experiments have some minor (or perhaps major) errors involved. The author must interpret these errors to determine their source and significance. Common (minor) errors include transfer loss (each time you transfer a substance from one container to another a small amount is left behind), equipment calibration (for example, a graduated cylinder is only accurate to ~0.5% of the total volume), estimation of “significant figures” (since you estimated the last digit, this digit could be inaccurate by a small amount), etc. Less common (major) errors result from things like incomplete reaction (you didn’t allow the reaction enough time to reach its conclusion), “wet products” (you were weighing a wet solid, so the extra mass is from your solvent), spillage (you spilled part of your product, so you got less than you expected), etc. A scientist must interpret the accuracy of the results and account for any errors with logical reasoning. For example, if you have a 50% yield, this is probably NOT due to transfer loss (you “accidentally” left half of your product behind?!?!) but MIGHT be due to incomplete reaction. (3) Room for improvement – a scientist’s job is never complete. If you were continuing to study this reaction, what would be done next? You should propose ways to improve the experiment and may also wish to suggest future experiments that could be conducted.

Instructions:

Write a discussion / conclusion section for the lab that you just completed. It must be double spaced, 12 point font, and should have your name on it. If you use any references to write your discussion / conclusion section, they should be cited in a separate “references” section.

Name: ________________________________ Section: ____________________

Solubility Rules

A compound is *soluble* in a particular liquid if it dissolves in that liquid. A compound is *insoluble* if it does NOT dissolve in the liquid. There is no easy way to tell whether a particular compound will be soluble or insoluble in water. For ionic compounds, however, there are empirical rules that have been deduced from observations of many compounds. Consider the following:

Compounds Containing the Following Ions Are Mostly Soluble[*]	**Exceptions**
Li^+, Na^+, K^+, NH_4^+	None
NO_3^-, $C_2H_3O_2^-$	None
Cl^-, Br^-, I^-	When any of these ions pairs With Ag^+, Hg_2^{+2}, Pb^{+2}, or Cu^+, it is *insoluble*
SO_4^{-2}	When SO_4^{-2} pairs with Sr^{+2}, Ba^{+2}, Pb^{+2}, or Ca^{+2}, it is *insoluble*
Compounds Containing the Following Ions Are Mostly Insoluble[*]	**Exceptions**
OH^-, S^{-2}	When either of these ions pairs with Li^+, Na^+, K^+, or NH_4^+, it is *soluble*
S^{-2}	When S^{-2} pairs with Sr^{+2}, Ba^{+2}, or Ca^{+2}, the compound is *soluble*
OH^-	When OH^- pairs with Sr^{+2}, Ba^{+2}, or Ca^{+2}, it is *slightly soluble*[**]
CO_3^{-2}, PO_4^{-3}	When either of these ions pairs with Li^+, Na^+, K^+, or NH_4^+, it is *soluble*

[*] adapted from Tro, Nivaldo J. Introductory Chemistry, 2nd ed. Upper Saddle River: Prentice Hall, 2006.

[**] For our purposes, these can be considered *insoluble*

Name: ________________________________ Section: ____________________

Vapor Pressure of Water

Temperature (ºC)	Vapor Pressure (mm Hg)
0	4.6
5	6.5
10	9.2
15	12.8
16	13.6
17	14.5
18	15.5
19	16.5
20	17.5
21	18.6
22	19.8
23	21.2
24	22.4
25	23.8
26	25.2
27	26.7
28	28.3
29	30.0
30	31.8
31	33.7
32	35.7
33	37.7
34	39.9
35	41.2
40	55.3
45	71.9
50	92.5
60	149.4
70	233.7
80	355.1
90	525.8
100	760.0
105	906.1
110	1074.6

1 atm = 760 torr = 760 mmHg

Name: ________________________________ Section: ____________________

Names, Formulas, and Charges of Common Polyatomic Ions

Positive Ion (Cation)		
1+	Ammonium	NH_4^+
Negative Ions (Anions)		
1–	Acetate	$C_2H_3O_2^-$
	Bromate	BrO_3^-
	Chlorate	ClO_3^-
	Chlorite	ClO_2^-
	Cyanide	CN^-
	Hydride	H^-
	Hydrogen Carbonate (bicarbonate)	HCO_3^-
	Hydrogen Sulfate (bisulfate)	HSO_4^-
	Hydroxide	OH^-
	Hypochlorite	ClO^-
	Iodate	IO_3^-
	Nitrate	NO_3^-
	Nitrite	NO_2^-
	Perchlorate	ClO_4^-
	Permanganate	MnO_4^-
	Thiocyanate	SCN^-
2–	Carbonate	CO_3^{2-}
	Chromate	CrO_4^{2-}
	Dichromate	$Cr_2O_7^{2-}$
	Oxalate	$C_2O_4^{2-}$
	Peroxide	O_2^{2-}
	Sulfate	SO_4^{2-}
	Sulfite	SO_3^{2-}
3–	Arsenate	AsO_4^{3-}
	Phosphate	PO_4^{3-}
	Phosphite	PO_3^{3-}

Moorpark College Chemistry Department Laboratory Report Rubric

Name: ______________________________ Experiment: ____________________________ Total: _______

CATEGORY	4 – Accomplished	3 – Good	2 – Developing	1 – Beginning	0 – Substandard	Score
Abstract	Clear, concise (~½ page), and thorough summary of results with appropriate literature references.	Refers to most of the major results; some minor details are missing or not clearly stated.	Misses one or more major aspects of the results.	Missing several major aspects of the results and merely repeats information from the introduction.	None, unrelated, or plagiarized.	_____×2
Introduction	A cohesive, well-written summary (including relevant reaction chemistry) of the background material pertinent to the experiment with appropriate literature references (at least one scientific reference if required by your instructor) and a statement of purpose.	Introduction is nearly complete but does not provide context for minor points. Contains relevant information but fails to provide background for one aspect of the experiment, or certain information is not cohesive.	Certain major introductory points are missing (e.g., background, theory, reaction chemistry), or explanations are unclear and confusing. References are not scholarly.	Very little background information is provided, and information is incorrect. No references are provided.	None, unrelated, or plagiarized.	_____×2
Methods & Materials	Contains a complete listing of safety information, a narrative of experimental procedures followed, and materials used. Omits information that can be assumed by peers. Includes observations when appropriate and only important experimental details.	Narrative includes most important experimental details. Missing one or more relevant pieces of safety information or experimental procedure.	Narrative is missing several experimental details and safety information or includes insignificant procedural details.	Several important experimental details and safety information are missing. Procedural steps are incorrect, illogical, or occasionally copied directly from the laboratory manual.	None, unrelated, or plagiarized (including completely copied from the laboratory manual).	_____
Results & Calculations	All figures, graphs, and tables are numbered with appropriate titles and captions. Sample calculations are shown and correctly solved. All data is explicitly mentioned in the text.	All figures, graphs, and tables are correctly drawn, but some have minor problems or could still be improved. All data and sample calculations are mentioned in the text.	Most figures, graphs, and tables are included, but some important or required features are missing. Certain data and sample calculations are not explained in the text and/or solved incorrectly.	Figures, graphs, and tables are poorly constructed, have missing titles, captions or numbers. Certain data and sample calculations are not referenced in the text and solved incorrectly.	None, unrelated, or plagiarized.	_____×2

CATEGORY	4 – Accomplished	3 – Good	2 – Developing	1 – Beginning	0 – Substandard	Score
Discussion & Conclusion	Demonstrates a logical, coherent working knowledge and understanding of important experimental concepts, forms appropriate conclusions based on interpretations of results and/or spectrum (spectra) analysis, addresses any post-lab questions in paragraph format, includes applications of and improvements in the experiment, refers to the literature when appropriate, and demonstrates accountability by providing justification for any errors.	Demonstrates an understanding of the majority of important experimental concepts, forms conclusions based on results and/or spectrum (spectra) analysis but either lacks proper interpretation, does not answer post-lab questions in paragraph format, suggests inappropriate improvements in the experiment, refers to the literature insufficiently, or lacks overall justification of error.	While some of the results have been correctly interpreted and discussed, partial but incomplete understanding of results is still evident. Student fails to make one or two connections to underlying theory.	Does not demonstrate an understanding of the important experimental concepts, forms inaccurate conclusions, does not answer post-lab questions in paragraph format, suggests inappropriate improvements in the experiment, refers to the literature insufficiently, and lacks overall justification of error.	None, unrelated, insignificant error analysis and incorrect explanation, or plagiarized.	______×2
References* (see sample below)	All sources (information and graphics) are accurately documented in ACS format. At least one reference is taken from primary scientific literature relevant to the report if required by instructor.	All sources are accurately documented, but a few are not in ACS format. Some sources are not accurately documented.	All sources are accurately documented, but many are not in ACS format. Most sources are not directly cited in the text.	All sources are accurately documented but not directly cited in the text.	Sources are not documented nor directly cited in the text.	______
Miscellaneous (check all that apply) ☐ Mechanics, grammar, and appearance ☐ Appendix ☐ Lab Notebook ☐ VSEPR and Valence Bond drawings	Grammar and spelling are correct. All required components are included, complete, and/or illustrated correctly. Paper is not written in first person. Includes ChemSketch image(s) if required by instructor. For Chem 1A XY lab, see lab manual for more details.	Less than three grammatical and spelling errors are present. Missing one required component or features an improperly labeled molecular representation.	More than three grammatical and spelling errors are present or paper is written in first person. Features multiple errors with labeled molecular representation.	Frequent grammatical and spelling errors, and writing style lacks cohesion and fluidity. Paper is written in first person. Labeled molecule contains multiple errors.	None, unrelated, or plagiarized.	______×2

*Journal citations must include author or editor, *title (in italics)* followed by a period, **year (boldface)**, *volume (in italics)*, and page numbers. For example: Schrauzer, G.N.; Windgassen, R.J. *J. Am. Chem. Soc.* **1966**, *99*, 3738–3743. For additional examples, see the *ACS Style Guide* (summary can be found online).

Periodic Table of the Elements

I A 1	II A 2	III B 3	IV B 4	V B 5	VI B 6	VII B 7	---- 8	VIII B 9	---- 10	I B 11	II B 12	III A 13	IV A 14	V A 15	VI A 16	VII A 17	VIII A 18
1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.066	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.88	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.847	27 Co 58.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.22	41 Nb 92.906	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.03	89 Ac 227.03	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Ds (281)	111 Rg (272)	112 Cn (285)	113 Nh (284)	114 Fl (289)	115 Mc (288)	116 Lv (293)	117 Ts (294)	118 Og (294)

58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.92	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)