Experiment 1G

FV 8/17/2020

PHYSICAL AND CHEMICAL PROPERTIES

MATERIALS: $Cu(NO_3)_2$ (aq), 3 100 mL beakers, plastic dropper, small bottle with cap, 50 mL buret, buret clamp and stand, buret funnel, 10 mL graduated cylinder, 25 mL graduated cylinder, copper shot, zinc metal pieces.

PURPOSE: The purpose of this experiment is to become familiar with the common types of laboratory glassware and equipment, make observations, and analyze data.

LEARNING OBJECTIVES: By the end of this experiment, the student should be able to demonstrate the following proficiencies:

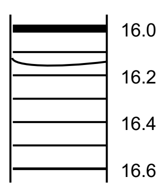
1. Understand the safety issues when working in the laboratory.
2. Know how to correctly use a buret, a pipet and a graduated cylinder.
3. Know how to use a top-loading balance.
4. Determine the density of a liquid and a solid.
5. Construct an Excel spreadsheet and graph with trendline using laboratory data.
6. Record laboratory data and observations.

BACKGROUND:

Laboratory glassware. There are two major categories of laboratory glassware: (1) those that <u>contain</u> a certain volume (volumetric flasks) and (2) those that <u>deliver</u> a certain volume (pipets, burets, and graduated cylinders). "To Contain" glassware (sometimes labeled TC) is typically used for preparing solutions of known volume. "To Deliver" glassware (sometimes labeled TD) is used to transfer known volumes between containers. Some glassware is very carefully designed and marked for high accuracy/precision work (burets, pipets, and volumetric flasks), while other glassware is not intended for such work (beakers, Erlenmeyer flasks, and graduated cylinders).

In high accuracy/precision work, the glassware must be <u>clean</u>. Not only does clean glassware avoid unwanted chemical contamination, but it also assures that delivered volumes of liquids will be correct. Even a small dirty spot inside a pipet or a buret, for example, can occupy volume and cause the liquid to stick to the contaminant resulting in the error of less volume being transferred than intended.

Measuring and recording data. The generally accepted rule for measuring volumes is to estimate <u>one</u> more digit beyond the digit associated with the closest spaced markings.



Portion of a buret

4 significant figures

Correct reading: 16.16 mL

3 Certain Digits

1 Uncertain Digit (different people may read different values for this digit)

For measurements obtained from devices which provide digital output, such as electronic mass balances or a Spectronic 20, all digits should be recorded. This should including any trailing zeroes.

Example readings from an analytical mass balance: 1.234 g (no trailing zeros) 0.0450 g 4.000 g (trailing zeros reported)

Appendix E (from the Plebe lab website) gives more detailed information about equipment precision, including the various types of glassware to be investigated in this experiment.

Properties. Any characteristic that can be used to describe or identify matter is called a *property*. Properties can be classified in many different ways. One common classification sorts properties as physical properties and chemical properties. *Physical* properties, like color or mass, are those which can be determined without changing the chemical makeup of the material. Weighing an object does not change it, so mass is a physical property. *Chemical* properties, on the other hand, rely on a chemical change. For example, iron reacts with oxygen to form rust. Rusting is a chemical property (because it is a chemical reaction that changes the chemical makeup of the metal).

Density is a quantity that can be used to describe a sample, or even identify one from a limited set of possibilities. This property can be used for both pure materials and mixtures. Density is defined as the ratio of the mass to volume (g/L). You may already be able to determine whether or not density is a physical or chemical property based upon what you already know. In this lab, we will deduce whether density is an 'extensive' or 'intensive' property.

Error Analysis. Appendix J is a basic introduction to the kinds of measurements most often encountered in the chemistry laboratory, along with methods for assessing the reliability of these measurements. *Accuracy* and *precision* are two different terms that are used to describe the validity of measurements. *Accuracy* refers to the agreement between a measured value and the true (or accepted) value. An accurate value is one that is very close to the true or accepted value. *Precision* refers to the degree of agreement among several measurements of the same quantity. Precision reflects the reproducibility of a given measurement. Precise values are very close to other values of the same measurement. For a set of data, the average may not be very accurate (far from its true value) but the data may be very precise (very similar values, reproducible). The goal of any experiment is to be both accurate and precise.

When reporting results, two useful quantities are the average (or mean) and standard deviation of a set of data. The standard deviation represents the spread in the data.

$$\text{Average} = \text{mean} = \bar{x} = \frac{\sum_{i=1}^{n} x_i}{n} \qquad \text{Standard deviation} = s = \sqrt{\frac{\sum_{i=1}^{n} (x_i - \bar{x})^2}{n-1}}$$

The variable n represents the number of measurements in a data set.
A common way to compare a result to a true value (x_t) is determine the percent error:

$$\% \text{ error} = (x_i - x_t) / x_t \times 100$$

where x_i is an experimental value, either an individual measurement or the average of a set of data. The % error can be a positive or negative value depending on whether x_i is larger or smaller than the true value (x_t). The smaller the magnitude of % error, the more accurate the set of data is.

PRE-LAB: Before lab, complete the pre-lab exercises. Also read the safety agreement and sign it. Bring the safety sheet, your pre-lab, your safety goggles, your laptop and this handout with you to lab.

PROCEDURE:

Safety: Safety for you and your classmates is the highest priority in the laboratory. Your instructor will discuss laboratory safety and give you a tour of the laboratory. Your signed safety sheet will also be collected.

You will be working with a lab partner but both students should record the data individually.

Part A: Making Observations in the Laboratory

1. Brass is an alloy, composed of approximately 30% zinc metal and 70% copper metal. You will observe reactions of the constituent metals with strong acids.
2. Your instructor will perform these demonstrations in the hood but you must also wear your goggles. Observe the reactions of copper with 7 M nitric acid (HNO_3) and copper with 7 M hydrochloric acid (HCl); then zinc with 7 M nitric acid (HNO_3), and zinc with 7 M hydrochloric acid (HCl).
3. Record your observations in the Data Section. Answer the in-lab questions.
4. Your instructor will perform these additional demonstrations in the hood. Observe the reaction of brass with 7 M nitric acid (HNO_3) and brass with 7 M hydrochloric acid (HCl). Record your observations.

Part B: Determination of the Density of a Copper Solution

Group A: Using a Graduated Cylinder

1. With a clean 100 mL beaker, obtain about 40 mL of the prepared aqueous copper solution from the labeled container in the room.
2. Using a <u>top-loading</u> balance, determine the mass of a small bottle <u>with</u> its cap (don't forget to tare the balance first). It is <u>not</u> necessary that the bottle is completely dry on the inside as long as it is capped. The bottle must be dry on the outside.
3. Using the 10 mL graduated cylinder, measure 5.00 mL of the copper solution and transfer it to the bottle.
4. Weigh the capped bottle on a top-loading balance and record the mass.
5. Repeat steps 3-4 for <u>four additional 5.00 mL</u> volumes. **DO NOT** empty the bottle between each addition. It's only important that you know the volume of solution added from the graduated cylinder and its corresponding mass.
6. When you are done, clean all the glassware.

Group B: Using a Pipet

1. With a clean 100 mL beaker, obtain about 40 mL of the prepared aqueous copper solution from the labeled container in the room.
2. Using a <u>top-loading</u> balance, determine the mass of a small bottle <u>with</u> its cap (don't forget to tare the balance first). It is <u>not</u> necessary that the bottle is completely dry on the inside as long as it is capped. The bottle must be dry on the outside.
3. Using the 5.00 mL pipet, transfer a 5.00 mL sample of the copper solution to the bottle.
4. Weigh the capped bottle on a top-loading balance and record the mass.
5. Repeat steps 3-4 for <u>four more 5.00 mL</u> volumes. **DO NOT** empty the bottle between each addition. It's only important that you know the volume of solution added from the pipet and its corresponding mass.
6. When you are done, clean all the glassware.

Group C: Using a Buret

1. With a clean 100 mL beaker, obtain about 50 mL of the prepared aqueous copper solution from the labeled container in the room.
2. Using a <u>top-loading</u> balance, determine the mass of a small bottle <u>with</u> its cap (don't forget to tare the balance first). It is <u>not</u> necessary that the bottle is completely dry on the inside as long as it is capped. The bottle must be dry on the outside.
3. Pour the copper solution into the buret and make sure there are no air bubbles in the buret tip.

4. Record the initial buret reading with units and the correct number of significant figures. Buret volumes should be read to the hundredths place in volume (for example, 1.26 mL).
5. From the buret, add about 5 mL of the copper solution to the bottle. **DO NOT** empty the bottle between each addition. It's only important that you know the volume of solution added from the buret and its corresponding mass.
6. Record the final buret reading.
7. Weigh the capped bottle on a top-loading balance and record the mass.
8. Repeat steps 5-7 for four more different volumes. The bottle does not need to be emptied between each addition. It's only important that you know the volume of solution added from the buret and its corresponding mass. The buret does not need to be refilled (unless necessary).
9. When you are done, clean all the glassware.

Part C: Determination of the Density of a Metal

Density of Zinc and Copper

1. Using the equipment available in your hood and the dry metal pieces, design an experiment for measuring the density of your assigned metal.
2. Have your instructor initial your proposed procedure and assign your group a metal. Carry out your described measurements three times. Record your data with appropriate units and to the appropriate number of significant figures in the Data Section.

Clean-up:

Rinse all glassware and return them to your lab bench area (hood). Organize your lab bench area. Clean up any spills. At the end of any lab experiment, wash your hands to remove traces of chemicals.

Name ____ Alfa ______________________

Partner ______________________________________

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DATA SHEET: Report all values with units and the proper number of significant figures.

Part A: Making Observations in the Laboratory

	Observations – describe what happened	Did a chemical reaction occur?
1. Cu and HNO_3		
2. Cu and HCl		
3. Zn and HNO_3		
4. Zn and HCl		

In-lab Question 1. Are zinc and copper elements, compounds or mixtures?

In-lab Question 2. Make a simple sketch of the particles present in zinc and copper.
Please write your alfa here ____________

Zinc: Copper:

In-lab Question 3. Brass is an *alloy* of zinc and copper. Is brass an element, a compound or a mixture?

In-lab Question 4. Make a simple sketch of the particles present in brass.

Brass:

In-lab Question 5. Predict what will happen when brass is treated with 7 M HNO_3 and when it is treated with 7 M HCl.

HNO_3 prediction	HCl prediction

	Observations – describe what happened	Did a chemical reaction occur?
1. Brass and HNO_3		
2. Brass and HCl		

Part B: Determination of the Density of a Copper Solution

Group A: Using a Graduated Cylinder

Copper solution sample	**Volume of Solution added to bottle**	**Mass of bottle and solution**
0	**-------------------------**	
1		
2		
3		
4		
5		

Group B: Using a Pipet

Copper solution sample	**Volume of Solution added to bottle**	**Mass of bottle and solution**
0	**---------------------**	
1		
2		
3		
4		
5		

Group C: Using a Buret

Copper solution sample	Initial Buret Reading	Final Buret Reading	Mass of bottle and solution
0	---------------	---------------------	
1			
2			
3			
4			
5			

Part C: Determining the Density of a Metal

1. Design an experiment for measuring the density of a metal. Write enough detail so that someone else could follow your procedure.

Instructor's Initials__________________________________ Assigned Metal ________________________________

DATA for Density of a Metal

DATA ANALYSIS AND QUESTIONS

Part B: Determination of the Density of a Copper Solution

1. Using experimental data, complete the following tables and determine the density of the copper solution. Obtain data from other groups to complete the tables.
   a. Show one sample calculation here.

Graduated Cylinder

Copper solution sample	Volume of solution added to vial	Total Volume of solution in vial	Mass of vial and solution	Mass of solution	Density
0	---------------------	------------------------			-----------------------
1					
2					
3					
4					
5					

Pipet

Copper solution sample	Volume of solution added to vial	Total Volume of solution in vial	Mass of vial and solution	Mass of solution	Density
0	----------------------	-----------------------			------------------------
1					
2					
3					
4					
5					

Buret

Copper solution sample	Initial Buret Reading	Final Buret Reading	Volume of solution added to vial	Total Volume of solution in vial	Mass of vial and solution	Mass of solution	Density
0	--------------	----------	--------------	------------			-----------------------
1							
2							
3							
4							
5							

2. Determine the average density of the copper solution for each of the three methods, reported with the correct number of significant figures. Show your work here.

Graduated cylinder _____________

Pipet: ________________________

Buret: ________________________

3. **Using Excel to determine the density.**
   a. Enter *your* experimental data into Excel and add columns corresponding to the tables above. Use the program to calculate values. For help, see the Excel tutorial on the General Chemistry website:
   https://www.usna.edu/ChemDept/_files/documents/excel-tutor/Excel%202016%20Tutorial.pdf
   Calculate the average density for each method and calculate the standard deviation for your densities. Show values with the correct number of significant figures.
   b. Create a graph of total mass vs. total volume added. Unless directed otherwise, graphs for this course are Scatter charts and the independent variable is shown on the x-axis. For this experiment, total volume added is the independent variable. Label the axes.
   c. Add a trendline to your graph. Select the box to display the equation.
   d. The equation is in the form of the equation of a line: $y = mx + b$, where m = slope and b = y-intercept.
      i. Slope is rise over run. Since the numbers on your graph have units, slope will also have units. What are the units for your slope? ____________________
      ii. What physical property does the slope represent? ___________________________
   e. For your trendline, select the box to display the R^2 value. This value helps to evaluate how close your data points are to the trendline, with a perfect line having an R^2 value of 1.00000.
   f. Print your graph and spreadsheet onto one page.

Part C: Determining the Density of a Metal

1. Calculate the density of your assigned metal from your experimental values. Show your work here.

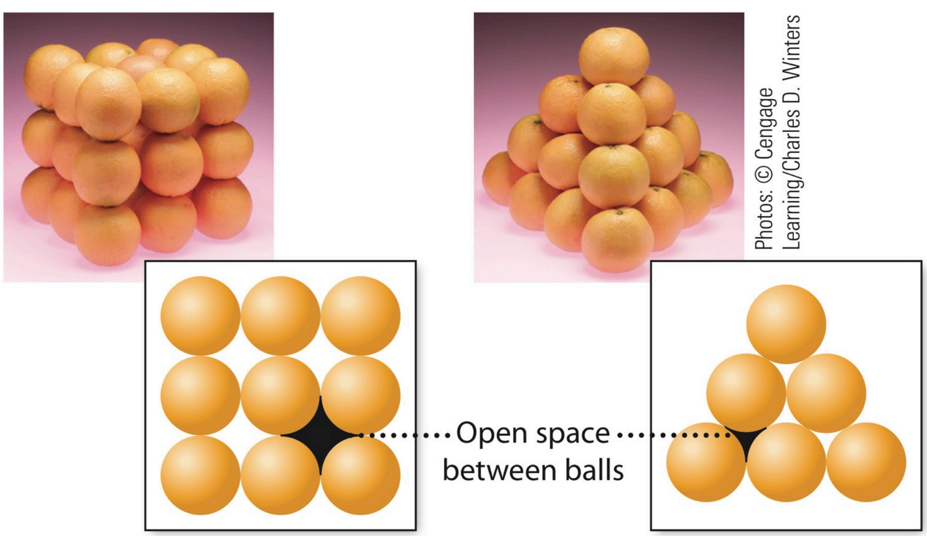
2. The accepted value for the density of copper is 8.960 g/cm^3 at 293K and the accepted value for the density of zinc is 7.133 g/cm^3 at 293K. Determine the % error in comparison to your experimentally determined value.

Approximate volume and radius of zinc and copper atoms.

1. From the density of your metal and the atomic mass of your metal, determine the number of atoms of your metal in a 1.00 cm^3 sample. Use the accepted values given above for the density of your metal.

2. Assuming that atoms are spherical, we can visualize them as packing in different possible arrangements. Two examples of these possible packings are shown below. Examining these two possibilities, which one is more efficient, that is, which one has less empty space between the spheres?

3. Both zinc and copper pack in ways closer to the picture on the right. It is known that 74% of the available space is taken up by atoms and 26% is empty space between the atoms. For 1.00 cm^3, calculate the volume of the atoms.



4. Determine the approximate volume of an atom of your metal, using your values from 2 and 3.

5. From your approximate volume, determine an approximate radius for your metal.

6. The accepted values for atomic radius are 127.8 pm for copper and 133.2 pm for zinc. Calculate the percent deviation for your metal.

Name ______________________________ Section ____________________

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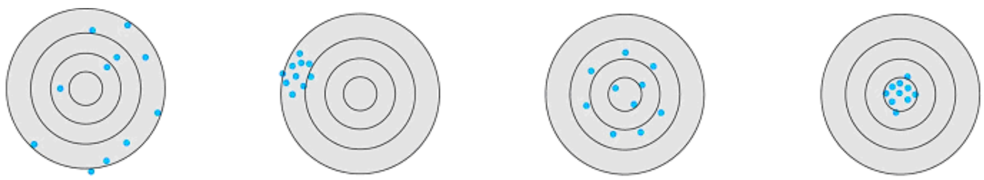
PRE-LAB EXERCISES

Complete this page prior to attending lab.

1. Based on the targets, which would be considered *inaccurate* but *precise*? Circle your choice.



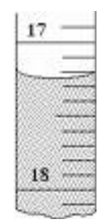
2. Given the following 3 values: 2.5 3.3 4.2

   a. Determine the average (or mean). Average = ____________

   b. Determine the standard deviation. Standard Deviation = ___________

3. Read the volume in this buret and report it with the correct number of significant figures. →

   Buret volume = __________________ mL



4. What is Archimedes' Principle? Search the web.

5. To the right is a picture of a volumetric flask. →

   Is it a TC (to contain) or TD (to deliver) piece of glassware? TC TD

6. Review the concept of significant figures in your textbook.

   a. How many significant figures do the following numbers contain?

      0.0129 has ________ significant figures. 2.0100 has ________ significant figures.

   b. Report these numbers to 3 significant figures.

      2.996 → ______________ 8000 → ______________________

   c. Perform this operation and report the answer with the proper number of significant figures.

      1.05 ÷ 12.00 = __________________

7. Circle the chemical property. Mass Flammability Color

CHEMISTRY LABORATORY SAFETY AGREEMENT

Before working in the chemistry laboratory, read carefully the safety precautions and techniques for handling chemicals described in https://www.usna.edu/ChemDept/_files/documents/SAFETY%20f17.pdf. Give this agreement, signed and dated, to your laboratory instructor on the first day of lab.

When you are in the laboratory, THINK about what you are doing at all times.

1. If any ventilation hood sash in your laboratory is up, you must wear your approved chemical splash goggles.
2. Do not attempt any unauthorized experiments.
3. Know the location and operation of safety equipment.
4. Bring only necessary materials into lab. Book bags, jackets, etc., are to be left in the hall.
5. Never work alone in the laboratory.
6. Never eat or drink in the laboratory
7. Use the fume hood when necessary.
8. Keep your work area uncluttered.
9. Use only equipment that is in good condition.
10. Dispose of waste and excess materials according to your laboratory instructor's directions.
11. No horseplay in the laboratory
12. Don't sit or lean on laboratory work surfaces.
13. Handle chemicals with caution.
    (a) Read labels carefully.
    (b) Use only the amount required.
    (c) Leave reagent containers in their proper places.
    (d) Clean up all spills immediately.
    (e) Label all chemical containers.
14. Thoroughly wash your hands any time you leave the laboratory.
15. Immediately report all accidents and physical/chemical injuries, no matter how minor, to your laboratory instructor. Be ready to take immediate action as needed to assist any injured classmate.
16. Don't leave the laboratory without your instructor's approval.

I have carefully read all the safety precautions on the pages at the website above and recognize that it is my responsibility to observe them throughout my chemistry course.

Name ________________________________ Signature ___________________________

Course ______________________________ Section _____________________________

Instructor ____________________________ Date _______________________________