Experiment 1

Density, Measurement, & Error

Introduction

Which is heavier, a one-liter container of water or a one-liter container of lead? This question can be answered quickly by just picking up each of the containers. While both containers contain the same volume of material, one liter, the one containing the lead is significantly heavier. The fact that there is a significant difference between the masses of the same volume of different materials demonstrates the concept of **density**.

Density

Density is a physical property that relates the mass of material to the amount of space it takes up, or volume. In mathematical terms, density is defined as the mass per unit volume.

$$Density\ =\ Mass \div Volume\ or\ D\ = \frac{m}{V}$$

Any units of mass and volume can be used to define density, but the most frequently used units in general chemistry are g/mL or g/cm^3. Thus, for the example above, you could determine the density of the water by measuring the mass of the water in the container and dividing by 1L, the volume of the container. The density of the lead would be found in the same manner. The results show clearly that the density of lead is approximately eleven times greater than that of water.

$$Density\ of\ water = \frac{1000g}{1L} = \frac{1g}{mL} = \frac{1g}{cm^3}$$

$$Density\ of\ lead = \frac{11000g}{1L} = \frac{11g}{mL} = \frac{11g}{cm^3}$$

$$Note: 1.000mL = 1.000cm^3$$

Density is an **intensive property** that is constant for a particular material regardless of how much material is present or how the material is manipulated to obtain the measurements. For example, if you measure out 25.0 mL of water in a graduated cylinder, the mass of water would be 25.0g. A cube of lead, with edge measurements of 1.00cm x 1.00cm x 1.00cm (1.00cm^3) would weigh 11.0g, since the tabulated density of lead is given as 11.000g /cm^3. Density is also a **physical property** that remains constant regardless of what form the material is in, so that a block of copper will have the same density as copper beads or pipe. Because of these two properties, density can be used as a means of identification.

Experimental Error

Experimentally determined values for measured or calculated quantities (the mass of an object, the heat produced by a reaction, the molar mass of a compound) almost always differ from their true values. The difference between the true value and a value that was determined through experimentation is called the **experimental error.**

Experimental error has many sources. Limitations in the measurements from the equipment used to obtain data are a source of experimental error that cannot be completely eliminated. However, if you know the limitations of your equipment you can make modifications that will allow you to minimize the error.

Example 1: Effect of Experimental Error on Density Measurements

To measure the density (the ratio of the mass to volume) of methanol, CH_3OH, you must measure both its mass and its volume. To measure its mass, you weigh an empty volumetric flask calibrated to hold 50.0 mL, fill it to the calibration mark with methanol, and weigh it again. After subtracting the mass of the flask to obtain the mass of the methanol only, the density is calculated by dividing the mass of the methanol by its volume.

Experimental error in the density calculation may arise from any or all of these three measurements: the mass of the flask, the mass of the flask and methanol, and the volume of the methanol. For example, if the sample and flask together actually weigh 60.8g, but the balance reads 61.2g, the density calculated from the masses will be higher than the actual density of methanol at 25°C which is 0.792g/ml.

Measurement	Correct values	Mass too high
Mass empty volumetric flask	21.2 g	21.2 g
Mass of flask + 50.0 ml CH_3OH	60.8 g	**61.2 g**
Mass of CH_3OH	39.6 g	**40.0 g**
Volume of CH_3OH	50.0 ml	50.0 ml
Density of CH_3OH	0.792g/ml	**0.800 g/ml**

Other sources of error arise from the way measurements are interpreted. For example, a contaminant in the methanol might alter its density, so that even an accurate measurement of the volume or mass of the liquid would not be an accurate measurement of the pure methanol. Error also arises from a failure to control experimental conditions. For example, since volume depends strongly on temperature, a density calculation may be less accurate if the temperature varies during the experiment. Knowing the source of an error can help you to minimize it.

Not all sources of error are significant. If you use a balance that can only measure the mass of the sample to within ±1 gram, many of your other sources of error would be insignificant relative to the error in the mass. The most important improvement to this experiment would be to simply use a more precise balance rather than trying to get perfect control of the room temperature. It is important to identify the sources of error in your experiment, and then identify those that are largest so that they can be minimized or eliminated. Smaller errors can then be ignored.

Systematic and Random Errors

Systematic Errors

Systematic errors are errors that occur each time an experiment is repeated. In a density measurement, the repeated use of an improperly calibrated volumetric flask or balance would lead to systematic errors. These errors can usually be eliminated or their effects factored into the calculations upon discovery. For example, if a balance consistently gave a reading that was exactly 1.000g too high, you could deduct 1.000g from every measurement made with that balance and your data would then be consistent with the actual measurement.

Once a systematic error has been identified, and hopefully eliminated, the remaining random error, or uncertainty in a measurement, can be determined. It is important to know the amount of error in your measurement, as that is an indication of how confident you should be in your results.

Random Errors

Unlike systematic errors, **random errors** affect a measurement in positive and negative directions with equal probability, such as the uncertainty inherent in a balance or buret. Another type of random error may arise from the conditions of the analysis, such as error that arises from the vibration of a table supporting an analytical balance, or air movement around the balance when a measurement is being taken. Unlike systematic errors, random errors can be reduced by averaging many measurements. Random error can also be minimized by using equipment that can provide more numerical digits in a measurement, such as a balance that measures mass to 1.0000g versus a balance that measures mass to 1.00g.

Example 2: Determining Random Error in Measurements

Uncertainty can be expressed as the **deviation** from a true measurement. The true value can be either a standard or an accepted tabulated value. For example, you weigh a standard 10.000-gram mass on an analytical balance and get three measurements of 9.96, 10.03, and 9.98 grams. The **uncertainty**, or error, of the measurement is calculated based on the average deviation of several measurements since a single measurement is not a good indicator of consistent behavior. The deviation of a measurement is the amount the measurement varies from the true value.

Deviations:	9.96 – 10.000 =	|– 0.04grams|
	10.03 – 10.000 =	|+0.03grams|
	9.98 – 10.000 =	|– 0.02grams|
Average deviation:		0.09g/3 = 0.03g

Since the average deviation in the measurements above is 0.03 grams, the uncertainty is reported as ±0.03g. The absolute values of each deviation are used to prevent positive and negative values from cancelling each other when averaging.

For most laboratory equipment, the manufacturer has already determined the uncertainty. Any additional error arises from the user. The uncertainties for some of the equipment used in this course are listed below:

Digital balance	± 0.001 g
50 mL beaker	± 2 mL
25 mL buret	± 0.02 mL
25 mL pipet	± 0.01 mL

Significant figures

The precision of a calculated result is based on the error in the experimental data. For example, if you weigh a mass to 1.00g and the volume of that mass is 1.30mL; you cannot report that your density is exactly 0.769230769g/mL, just because that is the value displayed by your calculator. You need to round off the number to something more reasonable. For a single measurement, using the balance listed above, the error in the balance is considered to be ± 0.01g. That means that if you have an error on the high side, you could actually record a measurement of 1.01g. The density calculation would now be 1.01g/1.30mL giving a density of 0.7<u>76</u>923076g/mL. If the measurement were too low, your density would be 0.99g /1.30mL with a resulting density of 0.7<u>61</u>538461g/mL. Thus, both values round up to approximately 0.77g/mL. While it is fairly easy to get a ballpark estimate of a number, such as guessing a value of 0.77g/mL in the problem above, systematically determining the correct number of significant figures provides a more scientific approach than "rounding".

<u>Determining the Number of Significant Digits in a Number</u>

Significant figures are the number of digits that should be reported when recording data or performing calculations. They are determined through a very rigid set of rules. The number of significant digits in your final answer is based on the least precise data value. To determine the number of significant figures in your data, first you must apply the following rules.

1. All non-zero digits are always significant 51.759 5 significant
2. For numbers containing zeros, use the decimal point to determine significance of the zero values. A zero is not significant if it is only a place holder and is not a measured number.
   a. A zero between 2 nonzero numbers is significant <u>50.002</u> 5 significant
   b. A zero before a decimal point is not significant 0.<u>502</u> 3 significant
   c. A zero before the first digit is not significant 0.00<u>52</u> 2 significant
   d. A zero at the end after the decimal point is significant <u>5.0200</u> 5 significant
3. If no decimal point is present, ex. 500, zeroes at the end of a number are not significant. If the zeroes are known to be significant, a decimal point should be included (500.) or scientific notation should be used (5.00×10^2) to indicate that the zeroes are significant.
4. Other numbers that may be used in calculations are called "exact numbers". These numbers are inherently whole numbers and are not going to limit the number of significant figures. You discount these values when determining the correct number of significant figures in your final results. Examples of exact numbers are:

4 sides to a square: inherently an integer	56 people in a room: counted
½ of a pie: inherently a fraction	12 eggs in a dozen: defined

Example 3:Determining the Number of Significant Digits in a Calculated Result

Use the following rules to determine the number of significant figures that should be reported for your results. However, use all the numbers in your calculator to actually do the calculation. Round off using rules governing significant figures for your final value after you have finished all calculations. Remember to discount any defined numbers when determining significant figures.

Addition & Subtraction

1. Convert all common numbers to the same unit (cm and mm, convert mm to cm). You cannot add/subtract values with different units!
2. Determine the number of significant figures from the number of digits after the decimal point.

$$\begin{array}{r} 3.572914 \\ \underline{-3.232} \\ 0.340\cancel{914} \end{array}$$

The values of the 4th, 5th, and 6th digits after the decimal point are questionable because you do not know the corresponding digits in the second number. Therefore, the value becomes 0.341 with only 3 significant figures. **If this number is to be used in future calculations, use the entire number (0.340914) but remember that there are actually only 3 significant figures/3 decimal places when you round your final answer.**

Multiplication & Division

1. To determine the number of significant figures in a multiplication or division problem, count the number of significant figures in each of the values used in the calculation. If the calculation uses a number calculated from a previous problem, remember to refer to the previous problem when deciding the number of significant figures in the final answer.

0.340914 x 2.3156 = 0.789420458 needs 3 sig. fig. due to previous calculation

The intermediate value 0.340914 when rounded to 0.341 has the fewest number of significant figures (2.3156 has 5 sig figs), so you would report the result as 0.789.

Remember to keep all significant figures throughout the calculation and only assign significant figures to the final value.

Percent Error

Percent error is calculated using the following equation, where the experimental value is the data obtained from the experiment and the actual value is a known quantity, such as the tabulated density of a metal. It is a measurement of the accuracy of the experimental value.

$$\%Error = \frac{|Experimental\ Value - Actual\ Value\ |}{Actual\ Value} x100$$

The bars (|), on either side of the calculation are used to represent the absolute value of the number calculated and to show that the sign is not used. The final answer is usually written as positive, but may actually be either positive or negative

Ideally, for you to be confident that a measured result is acceptable, the percent error should be less than 5%.

Example 4: Determining % Error and Accuracy of Calculations

Density calculation from earlier example = 0.800g/mL
Tabulated density of methanol = 0.792g/mL

$$\%Error = \frac{|0.800g/mL - 0.792g/mL|}{0.792g/mL} = \frac{|\ 0.008g/mL|}{0.792g/mL} = 0.0101x100 = 1\%$$ (1 sig. fig. due to subtraction step)

Since the error is under 5%, the estimate of the density calculation is reasonable and any error in the data did not significantly impact the density calculation. On a practical basis, this allows you to confidently say that the liquid measured is likely to be methanol and that your measurements are accurate, as the calculated experimental density varies less than 5% when compared to the accepted density of methanol. For the measurements taken in this laboratory an error of 5% or less is considered to be a reasonable cutoff for concluding your results are accurate.

Accuracy and Precision

The **accuracy** of a measurement is a comparison of your experimentally measured data to the actual or tabulated value of the measurement. Ideally you have the actual value, but when this is not possible, the average of the measurements can be used instead. This now makes the percent error a function of the **precision** of the measuring. **Precision** is an estimate of how closely scientific measurements agree with each other. In general, the more measurements you make, the more precise your average measurements will become based on the random errors cancelling out and as a result, your average value is also likely to be more accurate as well.

The deviation of 0.03g calculated in example 2 was an example of percent error based on accuracy. The experimental values were compared to the accepted value of the mass of 10.000g. However, if you do not have the true measurement available for your comparison when determining the deviation, you can use the average of all your values as a substitute for the true value. We now report the error as a **standard deviation** rather than a percent error and it is a measure of the precision of the measurements rather than the accuracy.

Example 5: Determining Standard Deviation and Precision of Measurements

Average mass: (9.96g+ 10.03g+ 9.98g)/3 = 9.99g

Deviation:
$|9.96 - 9.99| = 0.03$grams
$|10.03 - 9.99| = 0.04$grams
$|9.98 - 9.99| = 0.01$grams

Average deviation: 0.08g/3 = 0.026667g = 0.03g (1 sig. fig.)

Standard deviation: The error that is reported for experimental measurements in scientific journals is often the standard deviation. Standard deviation can be calculated from the following formula (add up all of the deviations, divide by the number of measurements, then take the square root). Essentially, it is the square root of the average deviation.

$$\sigma = \sqrt{\frac{\Sigma|x_1 - \mu|}{N}}$$

Where σ = standard deviation
Σ = sum
x = individual measurements
μ = average value
N = number of measurements

For the example above,

$$\sigma = \sqrt{\frac{(|9.96 - 9.99| + |10.03 - 9.99| + |9.98 - 9.99|)}{3}}$$

$$\sigma = \sqrt{\frac{(0.03 + 0.04 + 0.01)}{3}} = 0.163 = 0.2$$

$$\sigma = \sqrt{0.026667} = 0.163 = 0.2$$

Error can be a measurement of either accuracy or precision of a measurement. If the percent error is based on a comparison to the true value of the measurement it is an estimate of the accuracy. If it is based on comparing individual measurements to the average of those measurements, it is a measure of precision.

In Your Experiment

You will be using copper to show that density can be used to tentatively identify a material. You will measure the density in two different ways. First, you will mathematically determine the density by determining the volume of a copper cylinder by using the mathematical relationship:

$V = \pi r^2 h$,	V = volume of cylinder	(cm^3)
	r = radius of cylinder	(cm)
	h = height of cylinder	(cm)

Once you have determined the volume of the cylinder in cm^3, then you will determine the mass of the cylinder on a balance and calculate the density by dividing the mass by the volume.

In the second part of the experiment, you will confirm the density of the copper by water displacement. To do this, you will place the copper cylinder you used in part 1 into a graduated cylinder containing a known amount of water and measure the increase in the volume. The increase in volume of the displaced water is equal to the volume of the copper. Ideally, when you divide the mass of the copper by the volume of the water it displaced, you will come out with the same value for the density that you calculated from the linear measurements of the copper cylinder.

You will identify the copper and check the accuracy of your measurements by comparing your calculated density to the known density of copper, 8.96g/mL. The quality of your results will be evaluated by calculating the percent error in your measurements.

Finally, as a test of your laboratory skills, you will be given a test solution of unknown concentration. You will be graded on the accuracy of your mass and volume measurements by comparing the percentage solute determined from the density of the solution to actual mass of the solute in the prepared solution.

Safety Precautions and Chemical Disposal

Chemical Hazards

Test Solutions
See label on test solutions

Chemical Disposal

Used Test Solutions
Dispose in the waste container provided

Laboratory Equipment Procedures

Using the Balance

A **balance** is a piece of laboratory equipment used to measure mass. Balances are **calibrated** at the factory to accurately measure a mass by placing an object of known mass on the pan and setting the balance to display the known mass. They do not usually need to be checked frequently unless a problem is suspected. However, anytime you use a balance, you must keep in mind the following factors that can impact the balance and the accuracy of your measurements:

1. Air currents will affect your measurements and make it difficult for the balance to give a steady reading. Minimize movement near the balance. Be sure to close all doors (both sides and the top) when recording a mass.

2. Only weigh chemicals and equipment that are at room temperature. Air currents from hot or cold objects will affect the measurements.

3. Chemicals must be contained in a weigh boat, on weighing paper, or in a beaker. **Never** place chemicals directly on the balance. Avoid putting liquids on a balance whenever possible.

4. Never return excess chemical that you have taken out of a container back to the same container. Excess material needs to be treated as waste.

5. Don't forget to weigh the empty container as well so you can find the weight of the chemicals alone. (This step can be eliminated by "taring" the balance first)

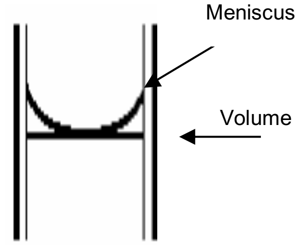
6. Clean up any spills carefully and promptly. Chemicals left on the balance will cause corrosion.

The balances have a button labeled "**O/T**". This button resets the balance to read 0.000g which is called "taring" the balance. You can **tare** the balance with an empty pan (this is called "zeroing" the balance) or you can tare it with an object on the pan (such as an empty beaker that you intend to fill with a chemical). Taring a balance with the container already on the pan eliminates the need to subtract the mass of the container from the measurement of the contents.

Volumetric Glassware (burets, pipets and graduated cylinders)

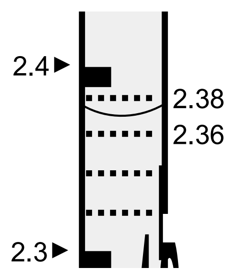
Using a Meniscus to Estimate a Volume

A **meniscus** is the curved top of a liquid often seen in glassware. It results from the edge of the liquid surface adhering to the glass and makes it more difficult to accurately read the volume. The general rule for reading the volume from a piece of glassware where the meniscus is present is to consider the **bottom** of the meniscus, the point where the liquid touches the measurement line on the glassware. The illustration below shows how to use the meniscus when measuring volume for the glassware you will be using in this lab.

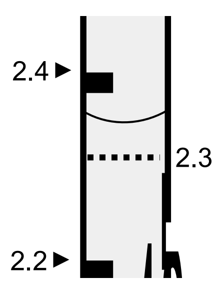


Estimating Significant Figures from a Scale Etched on Graduated Glassware
You will notice a numerical scale on the side of the graduated cylinder. When measuring an exact number of milliliters, you fill the glassware until the bottom of the meniscus touches the desired etched line. However, if you are measuring the amount of liquid collected or used during an experiment, the volume may not fall exactly on one of the lines of the scale. This means you need to estimate between the lines to get the highest number of significant figures from your measurement. The more significant figures you can record in your measurement the higher the precision of the measurements will be for the equipment you are using. You can often determine the number of significant figures in a measurement directly from the laboratory equipment. For a balance, it is usually +/- 1 in the last digit to the right. For calibrated glassware with lines, such as a beaker, buret or pipet, you measure as far as you can on the marked lines and then mentally divide the remaining space between that line and the next line into either 5 or 10 parts. Your last significant figure is based on your estimate of the space in between the lines.

Example 6: Estimating Volume Measurements
Example A: In the graduated cylinder shown below, the bottom of the meniscus lies between 2.3 and 2.4 mL. Therefore, you have at least 2.3 mL, but less than 2.4 mL. To estimate your last significant figure, mentally divide the space between 2.3 mL and 2.4 mL into an equal number of parts, usually either 5 or 10. For this example, divide the space into 5 parts of 0.02mL each. Count the number of parts until you reach the meniscus. For this measurement, your meniscus lies between your mental marks of 2.36 and 2.38. Your estimate is then 2.37. Therefore, your recorded measurement would be 2.37 mL with an error of +/-0.01 mL. Since you really don't know just how precise your mental division was, you really shouldn't estimate the volume from this particular example any further than +/-0.01mL.



Example B: In the graduated cylinder shown below, the meniscus lies between 2.2 and 2.4 mL. Therefore, you have at least 2.2 mL, but less than 2.4 mL. Unfortunately, you must estimate the 2.3 mark, so in this case, you can estimate that the value is at least 2.3, but can only estimate the next decimal to +/-0.05ml. This means your actual value should be reported as 2.35 mL as the meniscus is approximately halfway between your estimate of 2.3 mL and your etched value of 2.4mL. You cannot be more precise than this due to the estimate of the value of 2.3 mL. As a general rule when using glassware that measures by a factor of 2 units, you can only measure to ½ of the last unit giving you a value of 0 or 5 as your last significant digit.



Experiment 1: Procedures and Data Sheet
(Submit as part of your informal report.)

Name: ______________________________________ Date: _______ Section: _______

TA Signature: ________________________________

All data must be written in pen at the time it is collected. **Pencil is not allowed!!**
Record all measurements with the correct number of significant figures and units.
TA signature & TA initials on any changes made to the data are required or the data is invalid.

Part 1: Determining density using mathematical dimensions

1. Tare the balance.
1. Place the copper cylinder on the balance and record the mass of the copper cylinder.
2. Using your ruler, measure the length and diameter of the cylinder.
3. Record both measurements with units of centimeters in your data sheet.

Mass of copper cylinder __________

Length of copper cylinder __________

Diameter of copper cylinder __________

Part 2: Determining density by water displacement

1. **Gently** slide the copper cylinder into the 25mL graduated cylinder.
2. Fill your 10.0mL graduated cylinder with exactly 5.0mL of water using your wash bottle.
3. Pour the water into the 25mL graduated cylinder.
4. If the water does not cover the copper cylinder, add 2.0mL of additional water from your 10.0 mL graduated cylinder. Repeat until the copper cylinder is covered.
5. Record the total volume of water added.
6. Swirl the 25mL graduated cylinder to remove any bubbles that may be present.
7. Record the volume of the 25mL graduated cylinder contents. Make sure to use the bottom of the meniscus and estimate to the correct number of significant figures.
8. Pour the water into the sink and dry the copper cylinder with a paper towel.

Total volume of water added to the graduated cylinder __________

Volume of the contents of the graduated cylinder __________

Part 3: Determining the density of a test sample

1. Record the code on the bottle of test solution.
2. Tare the balance.
3. Place a clean, dry 10 mL graduated cylinder on the balance and record the mass.
4. Remove the graduated cylinder from the balance and dispense exactly 10.0 mL of the test solution into the graduated cylinder.
5. Record the mass of the graduated cylinder and test solution.
6. When finished, pour the test solution into the waste container.
7. Repeat these measurements twice more. You will be graded on the accuracy of your measurements, **so do not rush these measurements.**

Test solution Code # __________

	Trial 1	**Trial 2**	**Trial 3**
Mass of graduated cylinder	__________	__________	__________
Mass of graduated cylinder and test solution	__________	__________	__________

Experiment 1: Data Rubric (20pts)

Points

Data are neat and legible	5pts	__________pts
Significant figures (>80% correct)	3pts	__________pts
Units (>80% correct)	2pts	__________pts
All data are present and make sense	10pts	__________pts

Deductions (sliding scale based on TA discretion)

Lab area left unclean	–20pts	__________pts
Improper waste disposal	–20pts	__________pts
Disruptive behavior	–20pts	__________pts
Lab coat or safety glasses removed while in lab	–20pts	__________pts
Data sheet is missing TA signature	–20pts	__________pts
Other: ____		__________pts

Comments: ____

Grade for Data Sheet **__________pts**

Experiment 1: Results Table
(Submit as part of your informal report.)

Name: ____ Date: ________ Section: ________

All results must be written in pen. **Pencil is not allowed!!**
Record all results with the correct number of significant figures and units.
No marks or notes should be present on this page. Only the tabulated results are allowed.

Part 1: Determining density using mathematical dimensions

Calculation	Values
Volume of the copper cylinder	
Density of copper	
Percent Error	

Part 2: Determining density by water displacement

Calculation	Values
Volume of the copper cylinder	
Density of copper	
Percent Error	

Part 3: Determining the density of a test sample

Calculation	Trial 1 Values	Trial 2 Values	Trial 3 Values
Mass of Test Solution			
Density of Test Solution			
Average Density of Test Solution			
Percent Difference			

Experiment 1: Results Table Rubric (20pts)

Points

Tables are neat and legible	5pts	__________pts
Significant figures (>80% correct)	3pts	__________pts
Units (>80% correct)	2pts	__________pts
All results are present and make sense	10pts	__________pts

Deductions (sliding based on TA discretion)

Results to not match data	-20pts	__________pts
Plagiarism!!! Results are identical to another student	**-100pts**	**__________pts**

Other: ____ __________pts

Comments: ____

Grade for Results Table **__________pts**

Experiment 1: Calculations

Perform the following calculations & submit as part of your informal report.

Part 1: Determining density using mathematical dimensions

Volume of the copper cylinder

Use this relationship to calculate the volume of the cylinder: $V = \pi r^2 h$

r = Radius of copper cylinder = ½ diameter of copper cylinder in centimeters.
h = Height of cylinder = length of cylinder in centimeters
$\pi = 3.1416$

Volume of copper cylinder ____________

Density of copper cylinder

Use the formula for density to calculate the density of the cylinder. Since the cylinder is made of copper, the density of the cylinder should be the same as the density of copper.

Density of copper cylinder ____________

Percent Error

Calculate the percent error in your mathematical determination of the density of the cylinder. Use the result of your density calculation for your experimental value and the accepted density of copper, 8.96g/mL as the actual value.

Percent error ____________

Part 2: Determining density by water displacement

Volume of copper cylinder

Calculate the volume of water displaced by the copper cylinder when you placed the copper in the graduated cylinder. This volume is also the volume of the copper cylinder.

Volume of copper cylinder ____________

Density of copper

Calculate the density of the cylinder using the water displacement volume. This density should be very similar to (ideally the same as) the density calculated in Part 1.

Density of copper cylinder: ____________

Percent error

Calculate the percent error in your determination of the density of the cylinder using water displacement. Use the result of your density calculation for your experimental value and the accepted density of copper, 8.96g/mL as the actual value.

Percent error: ____________

Part 3: Determining the density of a test sample

Mass of test solution

Subtract the mass of the graduated cylinder from the mass of the graduated cylinder and the test solution to get the mass of just the test solution.

Mass of test solution Trial 1: ____________

Mass of test solution Trial 2: ____________

Mass of test solution Trial 3: ____________

Density of test solution

Calculate the density of your test solution based on the data from each of the three trials. The volume should be 10.0 mL, but if you did not put exactly 10.0 mL in your graduated cylinder, make sure to use the volume that you actually measured out in your calculation.

Density of test solution Trial 1: ____________

Density of test solution Trial 2: ____________

Density of test solution Trial 3: ____________

Average Density of test solution

Calculate the average of the three values for the density of the test solution.

Average density of test solution: ____________

Standard Deviation

Calculate the standard deviation in the measurement of your test solution. Use the result of each density calculation for the experimental values (x) and the average density of your test solution as μ.

$$\sigma = \sqrt{\frac{\Sigma|x_1 - \mu|}{N}}$$

Standard Deviation of Test Solution Measurements: ____________

Experiment 1: Questions

Answer the following questions & submit as part of your informal report.

1. Consider the two methods that you used to calculate the density of the copper cylinder.
    a. Describe two potential sources of error for each of these methods.

    b. Which method did you feel was more accurate? Briefly explain your answer.

2. Suppose you had added 11.0 mL of test solution to your graduated cylinder rather than 10.0 mL.
    a. Would this error cause your calculation of the density to be too high or too low? Briefly explain your answer.

    b. Is this an example of a random error or a systematic error? Briefly explain your answer.

3. A student measures the density of a block of wood, and obtains the following values: 4.27g/mL, 4.33g/mL, 4.29g/mL, and 4.31g/mL. The correct density for that wood isi 2.25g/mL.
    a. Are their measurements precise? Briefly explain your answer.

    b. Are their measurements accurate? Briefly explain your answer

    c. What type of error (random or systematic) would be represented by this data? Briefly explain your answer

4. A plastic square has a mass of 0.1476kg and measures 44.3mm on each side. Will the plastic square float? (density of water ~ 1g/mL at room temperature) Show all work and explain your answer.

Experiment 1: Prelab Worksheet

(Submit via Brightspace BEFORE the start of your lab session.)

Name: ______________________________ Date: ________ Section: ________ Grade: ________

All information needed to complete this worksheet can be found in the pre-lab information and calculations sections of the lab manual. Read this introductory material first!

- Record all values with the correct number of significant figures and units.
- Place all answers on the line when provided.
- Show calculations for any numerical answers; **work must be shown to receive credit**.
- See any 102 TA in the help office before your prelab is due if you have any questions.
- Each question is worth 2 points.

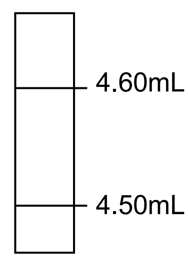
You collect the following data for a liquid:

Mass of graduated cylinder:	51.80g
Mass of graduated cylinder and liquid:	60.04g
Volume of unknown liquid:	10.48mL

1. Calculate the density of the liquid.

Answer: ___________

2. Draw the meniscus in the picture below to correspond to a reading of 4.52mL.



3. Special weights, with known masses, are used to calibrate balances. In your lab, a 5.000g weight was placed on a tared balance and a mass of 5.022g was recorded. Calculate the percent error in this measurement.

Answer: ___________

4. Your balance reads 3.024g before it is tared. What value should the balance read immediately after you push the tare button? (Include the correct number of significant figures)

Answer: ___________

5. You continually read the top of the meniscus rather than the bottom when using a graduated cylinder. Is this an example of a systematic or random error?

Answer: ___________

6. Calculate the **DEVIATION** in the **FIRST** measurement if you have the following 3 measurements. You do not know the actual value of the mass used. Include units with your answer.

3.450g 3.255g 3.601g

Answer: ___________

7. Calculate the **standard deviation** for the set of measurements in question 6.

Answer: ___________

8. Is the deviation used in the previous question an estimation of the accuracy or precision of the measurement? Briefly explain your answer.

Answer: ___________

Brief explanation: ____

9. What is the mass of a liquid if you have 42.2mL and the density of the liquid is 0.7800g/mL?

Answer: ___________

10. Can you pass in the prelab for an experiment once you have started the experiment?

Answer: ___________