DENSITY OF WATER AND METAL

Review the SAFETY POLICY before coming to lab. You need to know all the safety rules. The link is at the top of the website; be sure to read them. Dress appropriately, and bring a lab notebook (with your hand-written procedure in it) and safety glasses. You don't need to record the background parts of the lab instructions, just the procedures.

Purpose: This first experiment is a simple set of exercises to get you used to working with laboratory apparatus and thinking about data analysis. You are going to measure the mass and volume of several samples of water using different lab instruments, then draw a graph and carry out density calculations. You will determine the identity of an unknown metal using the density of water you calculated.

Background: There is no such thing as a perfectly exact measurement. All measurements have some uncertainty associated with them. Understanding this uncertainty and its implications is one of the fundamental skills needed for good science. Sometimes discarding outlying data points is necessary to obtain meaningful results, but accidentally discarding valid data points can lead to incorrect conclusions. This course will not delve into the intricacies of statistical analysis, but a basic understanding of accuracy, precision, and deviations is necessary. Please review the definitions of **precision**, **accuracy**, **significant figures**, and **density** in our textbook.

Precision tells us how consistent measurements are. It can be expressed by writing the average measurement and then the associated uncertainty. For example, if we know (from a set of measurements) that a particular measurement comes out between 1.18 and 1.28, we can express it as 1.23±0.05.

One way of estimating precision is to use the *average deviation*, finding how far each individual measurement (x) is from the average value ($\bar{x}$) and then taking the average (n = the number of data points you have). This is mathematically simple, but not statistically meaningful.

$$\text{Avg Deviation} = \frac{\sum_{i=1}^{n}|x_i - \bar{x}|}{n}$$

A much more useful concept for estimating precision is *standard deviation*, but we will not be collecting enough data points to get reliable statistics. So we will not be using standard deviations in this class, only average deviations. Outside this class, standard deviation is more common, so keep that in mind.

Accuracy tells us how close a measured value is to the actual value. The standard deviation and confidence interval tell us nothing about the accuracy of a measurement. Values in reference books do change over time. They aren't true in any absolute sense, but they are accepted by the scientific community and are more reliable than what we can measure in this course. If a value is inaccurate or imprecise, there are several different ways of classifying the reason. *Systematic errors* are always off by a fixed amount. A poorly calibrated balance might always give a mass that is too heavy, or reading at the top of a meniscus (rather than the correct way, at the bottom) might always give too large a volume. Systematic errors generally result in high precision but poor accuracy. *Random errors* affect data in irreproducible ways. These can include balance readings that fluctuate, impurities in materials, and glassware that is not perfectly clean. Random errors usually give poor precision, but don't necessarily affect accuracy. *Erratic errors* or *blunders* are when you make a mistake in the procedure, for example writing down 12 when you mean 21… or spilling stuff. If you make a blunder, you usually have to start over to get good data.

We are going to measure the accuracy and precision of measurements related to **density.** An old children's riddle asks "Which weighs more a pound of feathers or a pound of gold?" Obviously, a pound weighs one pound, whether its feathers or gold. However, we think of gold as being heavier because it is denser. Density is mass per unit volume. We don't have one instrument to easily measure density, but we can easily measure mass and volume, and then divide them.

Mass: In General Chemistry, we have two types of balances available for measuring mass: top-loading and analytical. Both types are located along the bench at the back of the lab. They can be distinguished by the glass-sided cage around the balance tray of the analytical balances. A good top loading balance costs several thousand dollars and reads 2 places past the decimal, while a good analytical balance costs tens of thousands of dollars and reads 4 places past the decimal. **To think about: Is the analytical more precise or more accurate than the top-loading?** Please read and think about the document "Using an analytical balance" before coming to lab.

Volume: A buret is used to measure volumes of liquid. The numbers on most burets might initially appear backwards: when it is completely full it reads 0.00 mL and when it is empty it reads 50.00 mL. This is because the readings are supposed to indicate how much liquid has been delivered, rather than how much remains. Directions for using a buret are in the "How to use a buret" document, which you should also read and think about before coming to lab.

In Part 2, we will be measuring a solid, not a liquid. We cannot directly use a buret to measure the volume of a solid. So instead, we're going to measure several masses. You will use the density value you determined for water (in Part 1) to convert the mass of water into volume. You will apply your analytical reasoning skills to identify which of the following materials is your unknown metal.

metal	**Density (g/cm^3)**
aluminum	2.70
zinc	7.14
chromium	7.19
nickel	8.90
manganese	7.43
iron	7.87
copper	8.96
silver	10.49
lead	11.36
mercury	13.55
gold	19.32

<u>Procedure</u>: This is the part that needs to be HAND WRITTEN in your notebook prior to lab.

Safety Orientation: Your instructor will show you the safety showers, the eyewash, the fire extinguishers, and the location of the first aid kit (prep room).

Work with a partner. Make sure you have read the background, including the documents “Using an analytical balance” and “Using a buret”.

Part 1: Density of water

1. Measure the mass of one clean, dry 100 mL beaker on both the top-loading and analytical balances. Be sure to tare first, and record the data correctly in your notebook. Check with your instructor to make sure you have the right number of sig figs recorded if you are not sure.
2. Use your squirt bottle to fill your buret with deionized water (from the DI tap at the end of the lab bench). Measure the volume precisely, and record it to two decimal places. If all of your volumes end in .00 or .05, you are not reading the buret accurately. Transfer at least 1.0 mL of water to the beaker and read the volume. Record the mass on both the top loading and analytical balances. Record the weight of the beaker + water and the exact volume in your notebook.
3. Dispense more water to get a total of ~10 mL in the beaker (don’t add 10 mL, make the total 10 mL). Record the weight of the beaker + water and the exact volume in your notebook. Repeat for totals of 25 and 50 mL, being sure to record the actual volumes (you may need to reload the buret and re-record the new values).
4. Density is temperature dependent, so find the lab thermometer and record the temperature. __________ ^{o}C

Example Data Table 1:

	Volume reading (mL)	Mass (g)	
		Top loading	Analytical
Empty beaker	------------------		
Initial buret		------------------	------------------
Buret reading 1 (~1.00 mL)			
Buret reading 2 (~10.00 mL)			
Buret reading 3 (~25.00 mL)			
Buret reading 4 (~50.00 mL)			

Exchange data with two other groups. Make sure your lab area is neat and your buret is put away. You’re done with Part 1! Go on to Part 2.

Part 2: Density of Metal

This time we will use just the top-loading balance, not the analytical.

1. Write down the number of the unknown metal you get from the designated area. Always write down the numbers on any unknown.
2. Weigh the empty, dry 25 mL Erlenmeyer and the glass stopper you get from the designated area.
3. Fill the flask no more than halfway with the unknown metal. Replace the stopper. Weigh to determine the mass of the flask + metal.
4. Fill the rest of the flask with deionized water. Fill it all the way to the top, NOT to the 25 mL mark. Gently shake to make sure there are no bubbles. Insert the stopper, and make sure there are no bubbles under the stopper. Dry off any drops on the outside. Weigh the flask + metal + water.
5. Carefully pour out the water; don't let the metal pieces go down the drain. Dry the metal with a paper towel, and put it back in the numbered container.
6. Fill the flask completely with water, and insert the stopper. Again, no bubbles allowed. Weigh to determine the mass of the flask + water. This will allow you to calculate the mass of water required to completely fill the flask.
7. Return your materials to the designated area and clean up your lab area, as always.

Example Data Table 2:

Unknown # _________
Mass of empty flask (g) _________
Mass of flask + metal (g) _________
Mass of flask + metal + water (g) _________
Mass of flask + water (g) _________

After completing the procedure but before leaving lab, write in your notebook a brief statement (two to three sentences) on the quality and reasonableness of the data you collected. Note what you might do differently if you performed the lab again.