CHM115 Lab 1

Precision, Accuracy and Density

Review the SAFETY POLICY before coming to lab. You need to know all the safety rules. Dress appropriately; bring a notebook, safety glasses and the hand-written procedure.

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 using different lab instruments, then draw a graph and carry out density calculations.

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. Discarding outlying data points is necessary to obtain meaningful results. But, accidentally discarding valid data points can lead to incorrect conclusions (like thinking a drug is safe when really it has dangerous side effects). This class will not delve into the intricacies of statistical analysis, but a basic understanding of accuracy, precision and deviations is necessary.

Precision tells us how consistent measurements are. It can be written by expressing the average value and then the uncertainty, either with the ± symbol or parentheses. If we know a measurement is between 1.18 and 1.28, we can express it as 1.23±0.05 or alternatively 1.23(5).

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 (eq 1, n = the number of data points you have). This is mathematically simple, but not statistically meaningful.

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

A much more useful concept for estimating precision is *standard deviation*. There are 2 forms of standard deviation: *population standard deviation* is used when data has been collected for every single member of a category. If you wanted to know the standard deviation of grades on an exam, you could use the population standard deviation for every student who took the exam. However, if you want to know the density of water, it would be a bit onerous to measure every single drop of water in the universe. Chemists rarely use population standard deviation, so we won't get into it (though it's not hard).

The other kind is *sample standard deviation* (s), used when you have measured a sample set of data points out of the population of interest. The formula for finding sample standard deviation is similar to average deviation, but contains a square, a square root, and you divide by n-1 instead of n.

$$s = \sqrt{\frac{\sum_{i=1}^{n}(|x_i - \bar{x}|)^2}{n-1}} \quad \text{(eq 2)}$$

You need to be careful when asked to find a deviation and always think about which formula, average or standard, you want to use. Why do we want to go to the bother of making the arithmetic more complicated? Because the standard deviation IS statistically significant. If your measurements contain only *random errors* (more on this below), then there are lots of fancy math equations we can use with the standard deviation to determine the probability of data being within a certain range.

Accuracy tells us how close a value is to the “true” 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 more reliable than what we can measure in General Chemistry. 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 or aardvark spit. 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. The top-loading balances are located on the end of each work bench, and the analytical balances are located in the balance room at the back corner of the lab. 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?

Volume – A buret is used to measure volumes of liquid. The numbers on most burets 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. Determining the volume of a solid is a bit trickier, as we will see in part B of the procedure.

In part A of the procedure, you will measure the mass and volume of water to determine the density.

In part B, you will determine the mass and volume of an unknown metal and determine the identity from among the following choices:

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

Procedure: 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, eye wash, fire blankets, and where the stockroom is (first aid kit).

Part A: Density of water

1. Measure the mass of a clean, dry 100 mL beaker on both the top-loading and the analytical balance. Be sure to tare first, and record the data in your notebook. (If you don't know what tare means, ask your instructor or TA).
2. Use your squirt bottle to fill your buret with deionized water (from the plastic taps at the end of the lab bench). Measure the volume precisely, to two decimal places. If all of your volumes end in .00, you are not reading the buret accurately. Transfer ~1 mL of water to the beaker, read the volume, and weigh it on both the top loading and analytical balances. Record the weight of the beaker + water and the exact volume in your notebook. Write down all of the significant figures the balances give you.
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 volume to (you may need to reload the buret and re-record the new values).
4. Density is temperature dependant, so find the lab thermometer and record the temperature. ___________ ^{o}C

Table 1: Raw data

	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)			

Part B: Unknown metal

For this part use just the top-loading balance, not the analytical.

1. Write down the number of the unknown metal you got. Always write down the numbers on any unknown.
2. Weigh the empty, dry 25 mL Erlenmeyer and the glass stopper.
3. Fill the flask about halfway with the unknown metal. Replace the stopper. Weigh to determine the mass of the metal.
4. Fill the rest of the flask with deionized water. Fill it all the way to the stopper, 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 pieces go down the drain. Dry the metal with a paper towel and put 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 water required to completely fill the flask.

Table 2

Unknown # _________

Mass of empty flask (g) _________

Mass of flask + metal (g) _________

Mass of flask + metal + water (g) _________

Mass of flask + water (g) _________

Make sure all 3 of your data tables are completely filled in, clean up your workspace, and return everything to the place it belongs. Yay, you're done!