

LAB #1: INTRODUCTION TO THE CHEMICAL LABORATORY

PURPOSE:

- To become familiar with basic lab safety principles.
- To become acquainted with metric measurements, the use of common lab apparatus, and the basic parts of the lab write-up.

INTRODUCTION:

The chemistry lab is not a safe environment. The appreciation of what constitutes safe laboratory practices is paramount for any student working in the chemistry lab. Safety practices serve not only to protect you from harm but your fellow colleagues as well. Moreover, conscientious lab protocol will also serve to minimize problems associated with chemical cross contamination, which can ruin the results of an otherwise successful experiment.

In the lab, the primary goal is to make proper observations of chemical and physical properties and processes. In addition to qualitative observations, scientific measurements represent a vital component of your observation repertoire. Scientific measurements will be made using the metric system, including measurements of the quantities of mass (in grams), volume (in L or mL) and length (m or cm). These quantities along with temperature represent the basic measurement building blocks of quantitative observation and can be combined to establish more elaborate quantities of measure, such as density (grams/mL).

How well a measurement is taken determines both its precision and its accuracy. *Precision* is related to the reproducibility to the measurement. It is a comparison of several measured values obtained in the same way. For example, a student measures the volume of a liquid three times, obtaining in the same way. For example, a student measures the volume of a liquid three times, obtaining the values 3.66 mL, 3.62mL, and 3.64mL. These measurements have high precision; there is only a 0.040 mL difference between the highest and lowest values.

Accuracy is a comparison of a measured value to the accepted, or true, value. Suppose the volume of the liquid had actually been 5.44 mL. Then the accuracy of the three measurements taken would be low because a difference of 1.80 mL between the average measured value and the true value is relatively large.

A good scientist is always trying to achieve both high precision and high accuracy in the laboratory.

Measuring Length:

All measurements are estimated values. They are obtained by using a measuring device that is marked with a calibrated scale. The scale contains marks and printed numeric values at regular intervals. There is a degree of uncertainty in all measured values. The degree of uncertainty is indicated by the number of significant figures given in a measurement (See textbook). For example, of the two rulers shown in figure 1, is one more uncertain than the other?

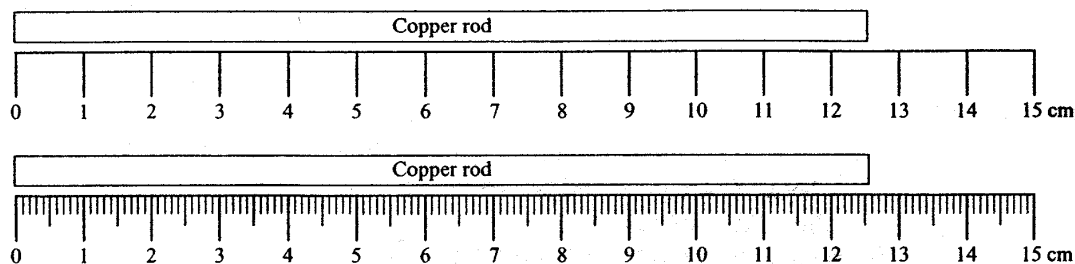


Figure 1

In the top ruler, each graduation represents 1 cm. It is certain that the right end of the copper rod lies between the 12th and 13th graduations. By estimating only one digit, the reading can be taken as 12.5 cm. In the bottom ruler, the larger graduations again represent 1 cm, and again the end of the rod lies between the 12th and 13th graduations. This ruler also has smaller graduations, however, representing 0.1 cm, and the rod lies between the 5th and 6th smaller graduations. By estimating only one digit, the reading can be taken as 12.55 cm.

Measuring volume:

When taking volume measurements, you will notice that the surface of the liquid is curved. This curved surface is called a *meniscus*. You should always take our reading from eye level and report the position of the *bottom* of the meniscus (See figure 2)

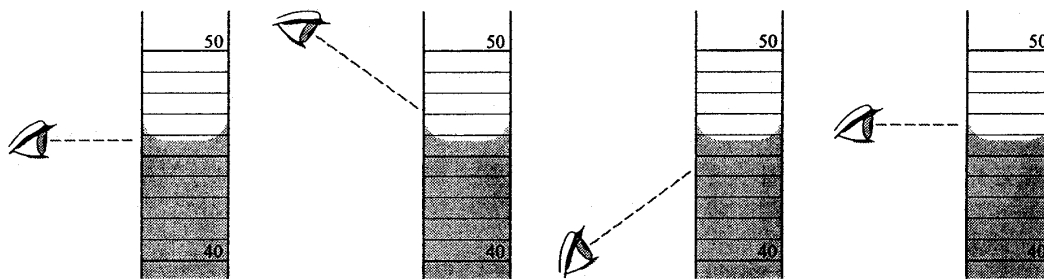


Figure 2

Once you measure a volume (or any other physical quantity) experimentally, you should consider the precision of your measurements. To quantify the uncertainty of a measurement we shall use a calculation called *percent range* to express precision. Percent range can be calculated for an individual measurement or a set of measurements.

For an individual measurement:

$$\% \text{ Range} = \left(\frac{\text{smallest measureable increment}}{\text{measured value}} \right) \times 100\%$$

Note: The smallest measurable increment is smallest discernible value you can make with a particular measurement device.

For a set of measurements:

$$\% \text{ Range} = \left(\frac{\text{highest value} - \text{lowest value}}{\text{average experimental value}} \right) \times 100\%$$

When you have more than one data point, the precision of the measurements is related to the span, or range, of the values (i.e. highest value minus lowest value) and the average of the values.

Part A: LAB SAFETY

A. Review the lab safety video

B. Identify the locations of the following safety items in the labs:

1. First Aid Kit:
2. Safety shower:
3. Eye Wash Fountain:
4. Fire Extinguisher:
5. Fire Blanket:
6. Fume Hoods:
7. Emergency Phone
8. Fire Alarm (s):

Part B: Volume Measurements

1. Measure out approximately 8.10 mL of water using a 10 mL pipette and pump.
2. Read the water volume from the marks on the cylinder and record your reading on the report sheet, using the correct number of significant figures.

Volume of water in 10 mL pipette = _____ mL

3. Carefully dispense this water into a 100 mL graduated cylinder.
4. Read the water volume from the marks on the cylinder and record your reading on the report sheet, using the correct number of significant figures.

Volume of water in 100 mL graduated cylinder = _____ mL

5. Carefully pour all the water into a 50 mL beaker. Be sure the beaker has measurements
6. Read and record the water volume, using the correct number of significant figures.

Volume of water in 50 mL beaker = _____ mL

Question:

1. Which type of glassware – the 10 mL pipette, the 100-mL graduated cylinder, or beaker- has the least amount of uncertainty? Explain.
2. What is the uncertainty of the volume measurement performed with the 10 mL pipette?
3. What is the uncertainty of the volume measurement performed with the 100 mL graduated cylinder?

Part C: Mass Measurements

1. Obtain a digital gram scale and turn it on.
2. Zero the digital gram scale.
3. Set an empty 50 mL beaker onto a digital gram scale and record the mass.

Mass of 50 mL beaker = _____ g

4. Place approximately 10 mL of water into the beaker and measure the mass.

Mass of water and 50 mL beaker = _____ g

5. Determine the mass of the water by subtracting step 3 from step 4.

Mass of 10 mL water = _____ g

6. A more direct way to measure the mass of the water would have been to zero the digital gram scale with the empty 50 mL beaker on it. Then when you add the 10 mL water sample, the scale will yield the mass of the water alone. Try it.

Mass of 10 mL water = _____ g

Questions:

1. How do the mass measurements of the 10 mL water sample compare for the techniques above?
2. Given the limits of the measurement itself, what is the precision (percent range) of your 10 mL water mass measurements above? Explain.
3. How many digits should the measured mass value be expressed with? Explain.