UTA-801
MASS AND VOLUME MEASUREMENTS

Objective: Determine the accuracy and precision of mass and volume measurements with common laboratory equipment.

INTRODUCTION

Chemistry is a quantitative science. Most experiments you will do this semester will involve measurements of one sort or another and subsequent mathematical manipulation of your data to produce a final answer. For instance, in the experiments to follow you will quantitatively determine the ΔH for a chemical reaction, the equilibrium constant for a reaction, and the chemical composition of a substance. It is therefore necessary for you to be familiar with the proper methods for obtaining mass and volume measurements, the accuracy and precision of these measurements, the experimental errors associated with your data, and how they affect your final results. Finally, you must be able to perform the mathematical operations needed for the treatment of your data.

Measurement of Mass

No matter how precisely an experimenter collects data; the accuracy of the experimentation is greatly dependent on the equipment used and on the proper use of that equipment. For accurate weighing, it is necessary to use a balance or scale that can take good readings at the level desired by the user. Most often, cheap inexpensive scales give less reliable data than more expensive, well-built scales. Consider, for example, weighing yourself and let's assume your true mass is 66.000000 kg. Do you believe a bathroom scale and a doctor's scale will give exactly the same mass? If so, will they agree to the kilogram, gram, milligram, and so on? Not likely! In fact, most bathroom scales only show readings to the nearest kilogram, thus you may obtain a mass reading of 65.9 or 66.0 or 66.1 kg by estimating the last digit. In total, only 3 significant figures may be read from these scales (unless you weigh 100 kg or more then 4 sig figs are obtained). A good doctor's scale, on the other hand, may allow accurate readings to the nearest decagram, thus giving your mass as 66.004 kg or 65.998 kg. As you can see, this 'better' scale gives 5 significant figures (6 sig figs if 100 kg or more) so you can estimate your mass to the nearest gram.

Note the application dictates the degree of accuracy and precision that we require from our instrument. Most people don't care to know their weight to the nearest gram and any scale that gives a reading ± 0.5 kg (~1-2 lbs) is sufficient for them – so why spend more for the better scale! However, if you are pharmacist weighing a medicinal compound to make an IV solution, a scale that reads to the nearest gram or better may be necessary! In this case, the extra money spent on the balance is well-worth it.

In a chemistry lab, the mass of a chemical or reagent can be determined to a milligram and approximated to a tenth of a milligram (6-7 sig. figs) by using an **analytical balance**. Such accuracy and precision is rarely called for in a general chemistry laboratory but is routine in analytical laboratories whether industrial, governmental, or academic. A **top-loading balance** is not as accurate (4-5 sig. figs), and a **triple-beam balance** may be even less accurate (3-4 sig. figs).

Technique also matters, especially as the measurements require greater and greater accuracy. For example, when weighing to a tenth of a milligram with an analytical balance, both the object to be weighed and the weights must be handled with tweezers because a fingerprint has a weight of the order of a tenth of a milligram. When less accuracy is called for, it is permissible to handle the object with clean, dry fingers and to use a top-loading or a triple-beam balance.

What if we do not know the quality of the instrument, in this case a mass scale that we are using? How do we evaluate the scale's reliability and quality of data? The *accuracy* and *precision* of any instrument may be determined by experiment, but first we need to define these terms.

Accuracy and Precision

The **accuracy** of a measurement describes the difference between an observed value and the "true" value of the quantity being measured. In some cases, the "true" value is given by a definition, but in many cases, "true" values are not known. If the same value for a measured quantity is independently obtained by several reliable investigators, then that value is assumed to be the "true" value. The accuracy of any measurement is reported as the **percent error**, which is defined as:

$$\%\ \text{error} = \frac{|V_n - V_t|}{V_t} \times 100$$

where V_t is the true value and V_n is the value determined by measurement.

Precision is concerned with the reproducibility of a result by the same method of measurement. If the difference in the numerical results of a series of measurements is small, the measurements have a high precision. Conversely, if the difference is large, the measurements have a low precision.

Precision is not the same as accuracy. A defective instrument may give an inaccurate result with a high precision.

The precision of some measurements can be reported in several different ways, one of which is the **average absolute deviation** from the mean, defined as:

$$\text{Average absolute deviation} = \frac{1}{n}\sum|v_i - \bar{v}|$$

where v_i is the ith measurement, v bar is the **average** of all the measurements, and n is the number of measurements.

Determining the Accuracy and Precision of a Top-Loading Balance

The accuracy of a balance can be determined by measuring the mass of an object whose 'true' mass is known. Standard weights, e.g. a 50.0000 g piece of metal, are made with respect to the standard kilogram, which is by definition a platinum-iridium bar in the custody of the International Bureau of Weights and Measures (BIPM) near Paris, France. Copies of this bar are kept by the standard agencies of all major industrial nations, including the U.S. National Institute of Standards and Technology (NIST). These standard weights can be used to examine the accuracy of the mass balance. By taking multiple measurements of this mass, we can determine the precision of this balance.

CALCULATIONS

Suppose we analyze a standard mass that is 50.0000 g. The three mass measurements from the same balance were 50.00, 50.01, 49.99 g. We can calculate the accuracy of each measurement by calculating the percent error, as defined as:

$$\%\ \text{error} = \frac{|V_n - V_t|}{V_t} \times 100$$

where V_t is the true value and V_n is the value determined by measurement.

Measurement 1: % error = 0.00 %

Measurement 2: % error = 0.02 %

Measurement 3: % error = 0.02 %

Average % error 0.013% = 0.01%

Note that if we average the three mass measurements, we obtain an average mass of 50.00g. Before we round to 50.00 g (to avoid round-off error), we calculate the % error for the average of three measurements to be 0.013% or 0.01% after rounding. Thus we see, that the accuracy can be improved by taking multiple measurements and averaging them, but if only a *single* measurement is taken, the percent error is larger. In this example, it's up to 0.02%.

The precision of our measurements can be measured in several different ways, one of which is the average absolute deviation from the mean, defined as:

$$\text{Average absolute deviation} = \frac{1}{n}\sum|v_i - \bar{v}|$$

where v_i is the ith measurement, v bar is the **average** of all the measurements, and n is the number of measurements. The average absolute deviation for our three mass determinations would be calculated as:

$$\frac{1}{3}\left(|\ 50.00\text{g} - 50.00\text{g}\ | + |\ 50.01\text{g} - 50.00\text{g}\ | + |\ 49.99\text{g} - 50.00\text{g}\ | \right)$$

$$= \frac{1}{3}\left(|\ 0.00\text{g}\ | + |\ 0.01\text{g}\ | + |\ -0.01\text{g}\ | \right)$$

$$= \frac{1}{3}(0.00\text{g} + 0.01\text{g} + 0.01\text{g})$$

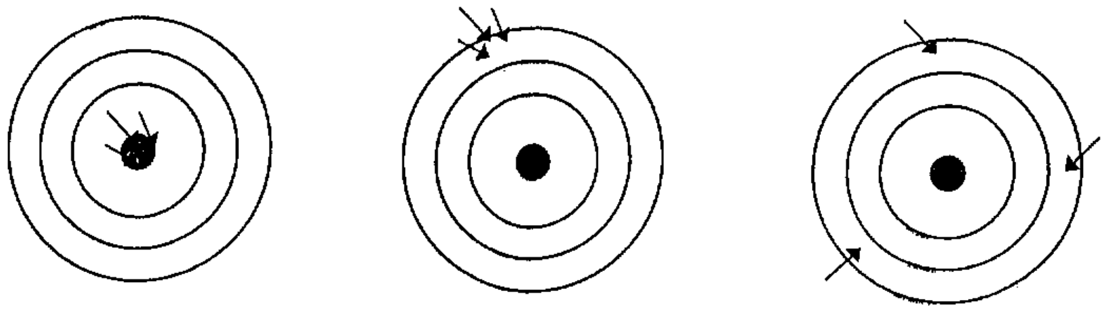
$$= \frac{1}{3}(0.02\text{g}) = 0.0066\text{g} = 0.007\text{g}$$

Our final result would then be reported as 50.00 g ± 0.01 g. This balance can be relied on to weigh objects (with masses around 50 g) to the hundredths of a gram with an accuracy that is within 0.02 % of the true value for a single measurement or 0.01% for three averaged measurements. The reading is precise to ± 0.01 g (one one hundredth of a gram)

Note the units of average absolute deviation are *grams* in this case not %. If the precision of a volume reading was being measured then the units would be mL or L, etc.

Consider the drawing in Figure 1. If the center of the target represents the true value and the arrows indicate an individual measurement, we can observe that the target on the left shows an instrument with good accuracy and precision. The center target represents an instrument with good precision (gives very similar readings each time) but poor accuracy (much like a bathroom scale set 10 lb too light to soothe the ego of the user). Finally the target on the right represents an instrument that gives widely scattered readings (poor precision) and *individually* gives poor accuracy.

However, note that by averaging the *multiple* measurements we can dramatically improve the accuracy of this last instrument's readings.



Good Accuracy Good Precision	Poor Accuracy Good Precision	Good Accuracy (average in bullseye) Poor Precision

Figure 1. Pictorial description of precision and accuracy.

Generally, poor precision means poor accuracy but as the right target shows this is not always true if multiple measurements are made. Conversely, as the middle target shows, good precision does not always mean good accuracy as systemic errors can appear and skew the accuracy of a good instrument (e.g. imagine the scope on a good rifle being poorly calibrated).

Determining the Accuracy and Precision of a Volume Measurement

For liquids and specifically for water, there exist numerous ways to measure the volume to varying degrees of accuracy and precision. **Volumetric glassware** is one of the most accurate and precise measuring devices. Such glassware includes **volumetric flasks**, **volumetric pipettes** and **burets**. **Graduated cylinders** and **graduated pipettes** are also commonly used but generally are less accurate and less precise. **Beakers** and **Erlenmeyer flasks** are generally only useful for very crude measurements of volume. Representative pictures of these items are shown in Figure 2.

In this experiment, you will determine the accuracy and precision of some common laboratory instruments including a top-loading balance, beaker, graduated cylinder and buret. You will also measure a simple linear relationship and appropriately plot the data. As you will see during this and next semester, we often use graphical representations of data to extract information (equilibrium constants, stoichiometry, rate laws) from experiments.

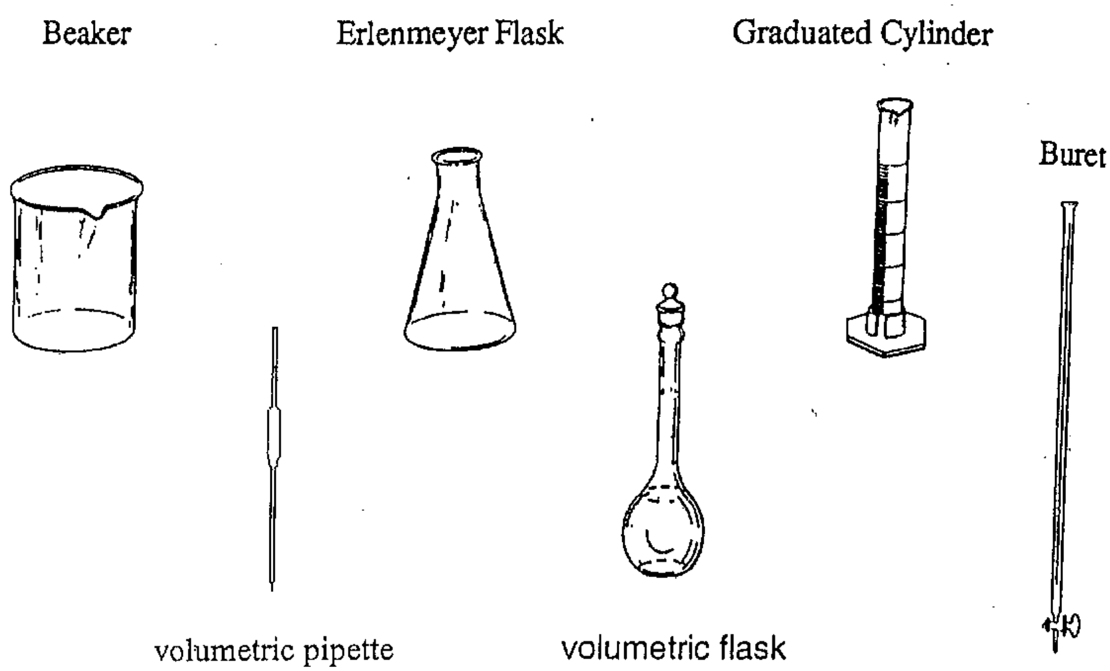


Figure 2. Pictures of common laboratory glassware.

Densities of some common metals:

Metal	Density (g/mL)
Magnesium	1.74
Aluminum	2.70
Zinc	7.13
Tin	7.28
Iron	7.87
Brass (alloy)	8.4-8.7 (depending on exact composition)
Copper	8.96
Silver	10.5
Lead	11.3
Mercury	13.6
Gold	19.3

Hazardous waste disposal: There is no hazardous waste in this experiment. All DI water may be poured down the drain.