

INTRODUCTION TO MEASUREMENTS

Introduction: In the first three experiments you will learn how to use equipment that to measure length, mass, and volume. You will apply the rules about significant figures every time you make a measurement or a calculation involving measured quantities.

Significant Figures

A measurement consists of a numerical value and a unit. Both have to be specified or the measurement is meaningless. For example if you say that you live 5.6 from campus, we don't know if that means 5.6 miles, 5.6 blocks, 5.6 minutes by car etc

The concept of **significant figures** applies to measured quantities because of the uncertainty associated with every measuring device. It does not apply to exact numbers, such as counting numbers or numbers that are defined. **The number of significant figures in a measurement is the number of figures that are known with confidence plus one figure that is approximate.** The uncertain digit is the last of the number. One of the essential tasks you must do when you are asked to make a measurement is to investigate the measuring tool. How many digits can you read and in what decimal place is the doubtful digit, i.e. where is the uncertainty? Data is *never rounded off*.

Mass measurements

Example: When weighing a sample of salt, of appx 3 grams, to the nearest 0.01g, the reading should be recorded to 2 decimal places – i.e. 3.01 g or 2.99 g or 3.00 g.

Any weighing's made with an electronic balance should be recorded with **all** the digits shown on the screen. Because it is a delicate instrument that can be damaged easily, certain precautions should be taken.

1. Never place a chemical directly on the balance pan. Use a **dry** container
2. Never pour liquids into a container while it is on the balance!
3. All measurements need to be within the capacity of the balance.
4. Should a spill occur, clean it up immediately.
5. Never place any item that is hot onto the balance pan.
6. To weigh substances in containers or vials, two common procedures are used.

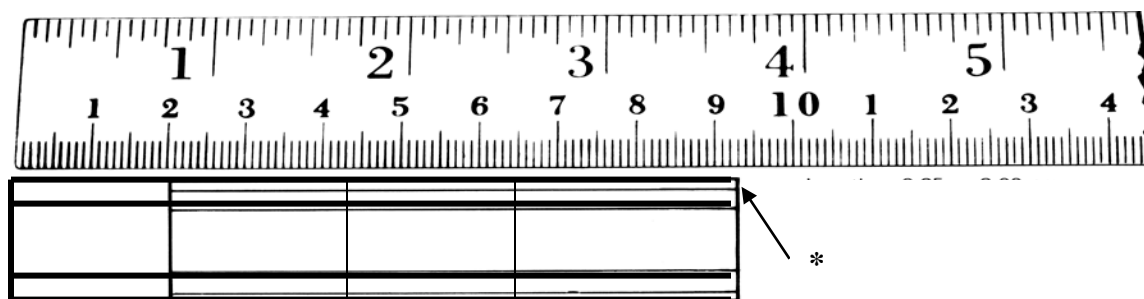
Weighing by difference: Weigh the container and record its mass. The substance is added to the container and both are weighed together. The mass of the substance is found by subtracting the mass of the container from the mass of the container with substance.

Weighing by taring the container: The weighing container (vial, watch glass, beaker, etc) is placed on the balance so the mass is displayed on the screen. Next the **TARE** function is pressed so the screen on the balance then reads **zero**. Add the chemical to the container from the stock bottle and observe and record the mass of the substance displayed directly on the screen.

Length Measurement: Length is measured with a centimeter ruler.

The numbered calibration marks are in centimeters, and the small calibration marks are 0.1 centimeters (1 millimeter) apart. Do not use the side of the ruler which is marked in inches and fractions thereof.

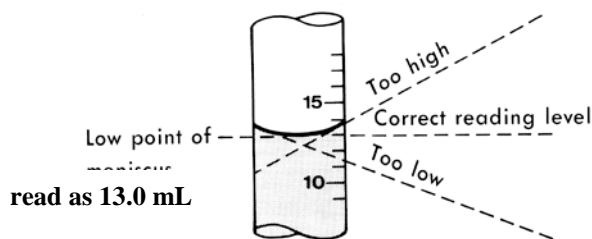
The centimeter ruler is read by placing one edge of the object being measured on a calibration mark and estimating the reading on the other edge. We need to be careful here, to estimate one digit past the calibration mark. Note the estimated length below. Is this the same reading you would make?



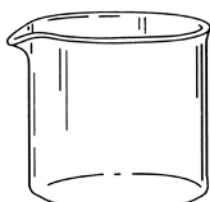
Because each calibration mark is 0.1 cm, it is possible to estimate to a fraction of that – ideally to ± 0.01 cm. If the object being measured in the figure above aligns a little more than half way between 9.20 and 9.30, you can estimate the reading as 9.27 cm.

Volume Measurement: Several tools can be used to measure volume.

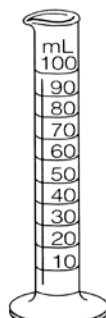
1. If the object is a rectangular solid, you can measure the 3 dimensions and calculate volume (length x width x height). The unit of volume will be cm^3 .
2. The graduated cylinder, volumetric flask, pipet, and burets are used to measure the volume of a liquid. The 50 or 100-mL grad cylinder can be read to the nearest 0.1 mL since the smallest marks are every 1 mL. The volumetric flasks, burets, and pipets are valid ± 0.01 mL. A flask or beaker should only be used if good precision is not required.
3. Water in a glass container forms a **meniscus**, a curved surface that is lower in the middle than at the edge. Volumetric laboratory equipment is calibrated to measure volume by sighting to the **bottom** of the meniscus and read the value at eye level). In the example below, record the value to the nearest 0.1 mL; (you would read **13.0** mL).



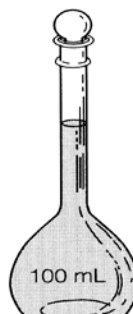
Erlenmeyer flask



Beaker



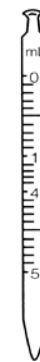
Graduated cylinder



Volumetric flask



Buret



Pipet

The precision attainable with the above glassware ranges from ± 10 mL with the beaker and Erlenmeyer flask, to ± 0.1 mL with the grad cylinder and ± 0.01 mL with the volumetric flasks, burets and pipets.

Density: an intensive physical property of matter.

Density is a physical property of matter that can be used to characterize a pure substance (an element or compound does not depend on how much of the material you have, but, rather, is a ratio of the mass over the volume ($d = m / V$). We know that 1.00 g water occupies a volume of 1.00 mL. Thus the density is $1.00 \text{ g} / 1.00 \text{ mL} = 1.00 \text{ g/mL}$. The (read this as one gram **per** milliliter, where “**per**” means **divided by**).

. Metallic elements have densities ranging from about 0.5 to 22 g/mL. (or g/cm³)

A bar of pure gold that weighs 10.0 g occupies a volume of 0.518 mL. The density is calculated as follows:

$$1 \text{ mL} = 1 \text{ cm}^3 \quad \frac{10.0 \text{ g}}{0.518 \text{ mL}} = 19.3 \text{ g/mL} \text{ or } 19.3 \text{ g/cm}^3$$

Other units of mass and volume can be used to express density, such as g/L (used to express densities of gases), lbs/ft³ in the English system, kg/m³ and so on.

The term “specific gravity” is a unitless quantity that is the ratio of the density of a substance over the density of a reference material (usually water with density approximately 1.00 g/mL). Thus, the specific gravity of gold is calculated as follows:

$$\frac{19.3 \text{ g/mL}}{1.00 \text{ g/mL}} = 19.3 \text{ (no units)}$$

Sometimes this is a useful quantity, since densities of materials can change at different temperatures, while specific gravities are relatively constant.

Densities of elements are summarized in the table below.

Densities of Elements

H 0.084	Densities of solid and liquid elements at 20°C are given in g / cm ³ . Densities of gases at 20°C are given in g/L (italic numbers).																He 0.17
Li Be 0.53 1.85																	B C N O F Ne 2.46 3.51 <i>1.17</i> <i>1.33</i> <i>1.58</i> <i>0.84</i>
Na Mg 0.97 1.74	Al Si P S Cl Ar 2.70 2.33 1.82 2.06 2.95 <i>1.66</i>																
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br Kr 0.86 1.54 2.99 4.51 6.09 7.14 7.44 7.87 8.89 8.91 8.92 7.14 5.91 5.32 5.72 4.82 3.14 <i>3.48</i>																	
Rb Sr Y Zr Nb Mo Tc Ru Rh Pd Ag Cd In Sn Sb Te I Xe 1.53 2.63 4.47 6.51 8.58 10.28 11.49 12.45 12.41 12.02 10.49 8.64 7.31 7.29 6.69 6.25 4.94 <i>4.49</i>																	
Cs Ba La Hf Ta W Re Os Ir Pt Au Hg Tl Pb Bi Po At Rn 1.90 3.65 6.16 13.31 16.68 19.26 21.03 22.61 22.65 21.45 19.32 13.55 11.85 11.34 9.80 9.20 -- <i>9.23</i>																	
Fr Ra Ac Rf -- 5.50 10.07 --																	
Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu 6.77 6.48 7.00 7.22 7.54 5.25 7.89 8.25 8.56 8.78 9.05 9.32 6.97 9.84																	
Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr 11.72 15.37 18.97 20.48 19.74 13.67 13.51 13.25 15.1 -- -- -- -- --																	

Data analysis and calculations

1. Precision

When gathering, the measuring device (balance, graduated cylinder, ruler, etc) is used to the maximum degree of *precision*, or degree of reproducibility, of which it is capable.

If several runs are performed, the average value or **mean** value is calculated by adding the numbers and dividing the sum by the number of measurements performed. Example:

Three values: 2.1, 2.2 and 2.3 yield a mean value of 2.2 with a deviation of ± 0.1

Three values: 1.9, 2.2 and 2.5 yield the same mean value but with a deviation of ± 0.3 . Thus the first set of values is more reliable.

Your lab instructor looks for the consistency of results, not just how close the mean value is to the correct (or accepted) value.

2. Accuracy

The term *accuracy* refers to how close an experimental result comes to the “correct” value (also referred to as the true, or known or accepted or theoretical value). Typically students are assigned “unknowns”, where the instructor has the correct value and the student is left to determine a value. The difference between the known value and the experimental value can be calculated. This is called the *error*:

$$\text{error} = [\text{experimental value} - \text{true value}]$$

The **percent error** is more often used and is calculated by dividing the error by the actual value, and multiplying by 100.

$$\% \text{ error} = \left[\frac{\text{exp. value} - \text{true value}}{\text{true value}} \right] \times 100$$

Note: the error or % error will have a positive or negative sign associated with it, which means that the experimental value is higher or lower than the actual value. Do report the + or – sign with your percent error when writing your report.

3. Significant Figures.

The number of “*sig figs*” or significant figures in a number which is measured or calculated from measured numbers, is defined as:

the number of digits known with confidence plus one digit with uncertainty.

If you know the precision with which your instrument is capable, you can assign the sig figs. For example, if you measure 25.5 mL with a graduated cylinder, you have three sig figs. The last 5 in the number has some uncertainty. Below are some rules about assigning significant figures (more extensive discussion can be found in your text)

1. All integers (non-zero numbers) are sig figs. 254 & 2.54 & 25.4 all have 3 sig figs

2. 0's between integers are sig figs. 2002, 2.002, 20.02 all have 4 sig figs

3. 0's preceding (to the left of) the first integer are never sig figs. 0.0035 has 2 sig figs.
scientific notation removes the doubt about the zeroes, thus $0.0035 = 3.5 \times 10^{-3}$

4. 0's to the right of the last integer are sig figs if they are to the right of the decimal point. So 2.500 or 25.00 or 250.0 all have 4 sig figs

5. Large numbers with many zeroes are problematic, so unless you know how the number was generated, you cannot assume that zeroes behind the last integer are sig figs if they are to the left of the decimal point. So 3,600,000 has 2 sig figs and is better expressed as 3.6×10^6 .

6. Defined numbers (not measured) are exact, and contain an infinite number of sig figs.

When doing calculations, to find the number of sig figs in the answer, follow these rules:

1. For multiplications and divisions, the number of sig figs in the answer is the least number of sig figs in the operation. Ex: $126.3 \times 5.0 = 631.5$, which when rounded to 2 sig figs becomes 6.3×10^2

2. For additions & subtractions – don't count sig figs. Ex 123.1576

Instead, round to the first uncertain digit
which in this case is the 2nd decimal place.

$$\begin{array}{r} 123.1576 \\ + \quad 5.27 \\ \hline 128.4276 \end{array}$$

(round to 128.43)

The following is a table of some commonly used units and conversion factors. You should commit much of this material to memory

The Metric System, (SI System) prefixes, units, and numerical values

Prefix	Symbol	Number	Power of 10	Conversion of units
<i>Mega</i>	M	1,000,000	1×10^6	1 meter (m) = 100 cm = 1000 mm
<i>kilo</i>	k	1,000	1×10^3	2.54 cm = 1 inch
<i>hecto</i>	h	100	1×10^2	453.6 g = 1 lb or 1 kg = 2.20 lb
<i>deca</i>	dk	10	1×10^1	1 mL = 1 cm ³ 1 L = 1 dm ³
<i>deci</i>	d	0.1	1×10^{-1}	1 Å (Ångström) = 10^{-10} m (a non-SI unit used to measure atomic size)
<i>centi</i>	c	0.01	1×10^{-2}	
<i>milli</i>	m	0.001	1×10^{-3}	1 oz = 28.35 g, 1 troy oz = 31.1 g (troy ounces are used to measure Ag, Au)
<i>micro</i> (Greek letter mu)	μ	0.000001	1×10^{-6}	
<i>nano</i>	n	0.000000001	1×10^{-9}	1 quart = 946 mL 1 mile = 1.61 km

Temperature Conversions

$$^{\circ}\text{C} = \frac{(\text{F} - 32)}{1.8} \quad ^{\circ}\text{F} = 1.8 \times ^{\circ}\text{C} + 32 \quad \text{K} = ^{\circ}\text{C} + 273.15$$

Celsius Fahrenheit Kelvin (note – no degree symbol on the K)
More precisely, 0 K = – 273.15 °C

Most commonly used units in chemistry:

Dimensions: m, mm, cm, nm, Å, km Mass: mg, g, kg

Volume: mL = cm³, L, μ L Energy: J (Joule), cal (calorie), V (volts)
1 calorie = 4.184 Joules