Numbers, Scientific Method, Conversions and Density

CHEM 121 Lab Experiment

Student Name	CHEM 121 Section/Date	Lab Instructor Sign Off

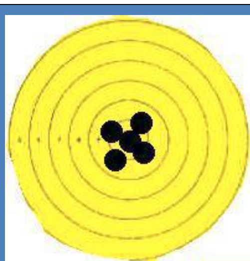
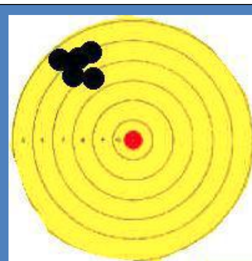
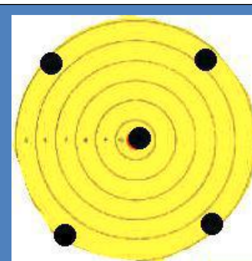
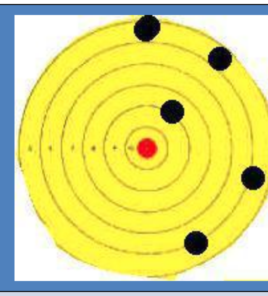
Conversions and Density

Introduction

Perhaps the most difficult portion of chemistry is that of learning to think in terms of both old and new manners in which to measure and calculate different things.

Accuracy and Precision

Accuracy is how close a value is to known values. Precision is within test agreement. The table, below, illustrates the concepts of accuracy, precision and deviations from both:

			
Precise, Accurate -- great	Precise, Inaccurate -- repair instrument; avoidable	Imprecise, Accurate -- Train person; accident	Imprecise, Inaccurate -- Hire new person; sad

Significant Figures

No science can proceed very far without resorting to quantitative observation[s] that are, of necessity, dependent upon accuracy and precision.

This is a big part of **<u>the</u>** Scientific Method.

Step 1: Observation

Step 2: Data Gathering: Generate a law or laws from the data – means of prediction

Step 3: Hypothesize

Step 4: Test Hypothesis [es] by Experiment

Step 5: Prove Hypothesis [es] to make Theory

That means that scientists must make measurements. To fulfill the requirements of the Scientific Method, there are an unlimited number of experimental techniques or experimental methodologies that may be used.

Quantifying data through measurements is one such methodology. Quantification usually means reading numbers from some sort of measuring device. There must be some way to limit the number of meaningful digits (significant figures) that may be obtained in an experimentally determined quantity. Significant figures are critical when reporting scientific data because they give the reader an idea of how well you … actually measure[d] and/or report[ed] your data.

Significant Figures' Rules

http://www.usca.edu/chemistry/genchem/sigfig.htm

1) ALL non-zero numbers (1,2,3,4,5,6,7,8,9) are ALWAYS significant.

2) ALL zeroes between non-zero numbers are ALWAYS significant.

3) ALL zeroes which are SIMULTANEOUSLY to the right of the decimal point AND at the end of the number are ALWAYS significant.

4) ALL zeroes which are to the left of a written decimal point and are in a number >= 10 are ALWAYS significant.

A helpful way to check rules 3 and 4 is to write the number in scientific notation. If you can/must get rid of the zeroes, then they are NOT significant.

Examples: How many significant figures are present in the following numbers?

Number	# Significant Figures	Rule(s)
54,356	5	1
3.142	4	1
750.04	5	1,2,4
0.00006 (= $6 * 10^{-5}$)	1	1,4
6.6000	5	1,3
606.060	6	1,2,3,4
8,000,000 (= $8 * 10^{+6}$)	1	1
30.0 (= $3.00 * 10^{+1}$)	3	1,3,4

Mathematical Operations
http://www.usca.edu/chemistry/genchem/sigfig2.htm

ADDITION AND SUBTRACTION

When adding or subtracting numbers, count the number of decimal places to determine the number of significant figures. The answer cannot contain more places after the decimal point than the smallest number of decimal places in the numbers being added or subtracted.

Example: in table, below, how many significant figures are there in the final answer?

Note: There are 4 significant figures in the answer.

23.112233	(6 places after the decimal point
1.3324	(4 places after the decimal point)
+ 0.25	(2 places after the decimal point)
24.694633	(on calculator
24.69	(rounded to 2 places in the answer)

Mathematical Operations
http://www.usca.edu/chemistry/genchem/sigfig2.htm

MULTIPLICATION AND DIVISION

When multiplying or dividing numbers, count the number of significant figures. The answer cannot contain more significant figures than the number being multiplied or divided with the least number of significant figures.

Example: in table, below, how many significant figures are there in the final answer?

Note: There are 5 significant figures in the answer.

23.123123	(8 significant figures)
x 1.3344	(5 significant figures)
30.855495	(on calculator)
30.855	(rounded to 5 significant figures; and 3 places)

Significant Figures in the Laboratory

Masses should always be recorded to as many places after the decimal point as are read off the balance. NEVER round data! Calculation of mass by difference using a tare should be reported to this same number of places.

10 mL Graduate cylinders should be read to the nearest 0.01 mL. 25 mL and 100 mL graduate cylinders should be read to the nearest 0.1 mL.

Rules for Rounding Whole Numbers
http://mathatube.com/rounding-rules-for-rounding.html

If the number you are rounding is followed by 5, 6, 7, 8, or 9, round the number up. Example: 38 rounded to the nearest ten is 40.

If the number you are rounding is followed by 0, 1, 2, 3, or 4, round the number down. Example: 33 rounded to the nearest ten is 30.

To round a number to a particular place, look at the digit to its right. If it is 5 or more, round up. If it is 4 or less, round down.

Number Rounding Examples

Example: Round 53 to the nearest 10. Look at the one's position. It's a "3". Round down to 50

Round 368 to the nearest 100. Look at the ten's position. It's a "6". Round up to 400.

Note that all of the numbers to the right of the place you are rounding to become zeros.

When rounding a number, you first need to ask: what are you rounding it to?

Numbers can be rounded to the nearest ten, the nearest hundred, the nearest thousand, and so on.

Example: look at the number 2,827.

2,827 rounded to the nearest ten is 2,830; 2,827 rounded to the nearest hundred is 2,800; 2,827 rounded to the nearest thousand is 3,000.

Dimensional Analysis (Factor Labeling or Unit Factor) Method of Problem Solving
http://www.chem.tamu.edu/class/fyp/mathrev/mr-da.html

Dimensional Analysis is a problem-solving method that uses the fact that any number or expression can be multiplied by one without changing its value. It is a useful technique. The only danger is that you may end up thinking that chemistry is simply a math problem - which it definitely is not.

Unit factors may be made from any two terms that describe the same or equivalent "amounts" of what we are interested in, e.g.,

1” = 2.54 cm	39.37” = 1 m	1000 m = 1 km
1 lb = 454 g	1 mi = 1760 yds	0.001 m = 1 mm
1 oz = 30 mL	1 qt = 946 mL	1000 mL = 1 L

We can make two unit factors from this information:

$$\frac{1\,\text{qt}}{946\,\text{mL}}\text{OR}\frac{946\,\text{mL}}{1\,\text{qt}} \qquad \frac{1000\,\text{mL}}{1\,\text{L}}\text{OR}\frac{1\,\text{L}}{1000\text{mL}} \qquad \frac{1\,\text{inch}}{2.54\,\text{cm}}\text{OR}\frac{2.54\,\text{cm}}{1\,\text{inch}}$$

$$\frac{39.37\,\text{inches}}{1\,\text{m}}\text{OR}\frac{1\,\text{m}}{39.37\,\text{inches}} \qquad \frac{1\,\text{pt}}{16\,\text{oz}}\text{OR}\frac{16\,\text{ounces}}{1\,\text{pint}} \qquad \frac{1\,\text{mi}}{5280\,\text{ft}}\text{OR}\frac{5280\,\text{ft}}{1\,\text{mi}}$$

As you may have surmised from the choices, above, the dominant measuring method in science is the metric system. Base quantities are fundamental quantities of measurement. They number 4:

– Mass,

The base unit of MASS is the kilogram (kg). The most commonly used units are the gram (g), milligram (mg) and microgram (μg – be aware that clinicians have stepped back 25 years and are re-using mcg).

– Length,

Length is a physical quantity that describes how far an object extends in some direction. It is the distance between two points. The basic unit of length is the meter.

Volume is derived from length. It is defined as the space an object occupies and is described by length units. Remember that for a cube, the volume is equal to the cube of the length of one side.

The basic unit of volume is the cubic meter (m^3). 1 m^3 is equal to 1,000,000 cm^3. Since the cubic meter is very large and cumbersome, volume is based upon the density of water: 1 gram of water occupies a volume of 1 milliliter (mL). Given that 1 cubic centimeter (cc) of water is very close to 1 mL, the mL is commonly used.

– Time

Time is a measure of how long events last. The SI unit of time is the second.

– Temperature

Temperature (degrees in the F and C scales) is used to describe the hotness or coldness of an object. The SI unit of temperature is the Kelvin -- NOT "degrees" Kelvin!

In its simplest form, applications of the metric system to science may be used by learning the following four (4) equivalencies:

American	Metric
1 inch	2.54 cm
2.2 lbs	1 kg
1 qt	946 mL
1 lb	454 g

The rest of the American units of measurements are known to us, and, hence, may be derived with great ease.

The metric system, however, is not as well known in this country and can create a few problems. If one remembers the prefixes and what they stand for, it is not difficult as one might first believe:

Prefix	Abbreviation	Fraction	Scientific notation
Mega	M	1,000,000	10^{6}
Kilo	k	1,000	10^{3}
Deci	d	1/10	10^{-1}
Centi	c	1/100	10^{-2}
Milli	m	1/1000	10^{-3}
Micro	µ	1/1,000,000	10^{-6}
Nano	n	1/1,000,000,000	10^{-9}
Pico	p	1/1,000,000,000,000	10^{-12}

The prefixes are simply another way in which to say a number, e.g., one might say "I have 1000 grams.", another might say, "I have 1 kilogram." And both are saying the same thing. The ease at which one may manipulate his or her way through the metric system is best shown through the factor label (dimensional analysis) method. This method takes into account equivalencies as in the table, above.

Example 1: 1 mile is how many cm?

$$(1 mile)\left(\frac{5280\ ft}{1 mile}\right)\left(\frac{12\,inches}{1\ ft}\right)\left(\frac{2.54\,cm}{1 inch}\right)=160{,}934.4\,cm$$

Notice how the units in the numerator cancel out with units in the denominator leaving the answer in the units we wanted in the first place: cm. The point is that if you label your numbers and use equivalencies, you will get the correct answer every time.

Example 2: 1 oz is how many liters?

$$(1 oz)\left(\frac{1 qt}{32\,oz}\right)\left(\frac{946\,mL}{1 qt}\right)\left(\frac{1 L}{1000\,mL}\right)=0.0296\,L$$

Once again, the answer was already set up by using this method of "converting" from American to metric measure.

Example 3: 200 g is how many oz?

$$(200\,g)\left(\frac{1\,lb}{454\,g}\right)\left(\frac{16\,oz}{1\,lb}\right)=7.048\,oz$$

Again, the answer was set up by using this method of "converting" to American measure from metric.

Example 4: The density of mercury (Hg) is 13.6 g/mL. To two decimal places, what is the density of Hg in ounces per gallon?

$$\text{ALWAYS start with what you know!}$$
$$13.6\,\frac{\text{g}}{\text{mL}}$$
$$\text{And THEN look at your units for the final result}$$
$$\text{X}\,\frac{\text{ounces}}{\text{gallon}}$$
$$\left(\frac{13.6\,\text{g}}{\text{mL}}\right)\left(\frac{1\,\text{lb}}{454\,\text{g}}\right)\left(\frac{16\,\text{oz}}{1\,\text{lb}}\right)\left(\frac{946\,\text{mL}}{1\,\text{qt}}\right)\left(\frac{4\,\text{qt}}{1\,\text{gal}}\right)=1813.652863\,\frac{\text{oz}}{\text{gal}}\approx 1813.65\,\frac{\text{oz}}{\text{gal}}$$

Example 5: There are 5280 ft in one mile. Using scientific notation, to 3 decimal places, how many mm are there in 1 mile?

$$\left(\frac{5280\,\text{ft}}{1\,\text{mile}}\right)\left(\frac{12\,\text{inches}}{1\,\text{ft}}\right)\left(\frac{2.54\,\text{cm}}{1\,\text{inch}}\right)\left(\frac{10\,\text{mm}}{1\,\text{cm}}\right)=1{,}609{,}344\,\frac{\text{mm}}{\text{mile}}\approx 1.609*10^{6}\,\frac{\text{mm}}{\text{mile}}$$

Example 6: The volume of a wooden block is 6.66 in^3. This is equivalent to how many cubic centimeters?

$$\left(6.66\,\text{in}^{3}\right)\left(\frac{2.54\,\text{cm}}{1\,\text{inch}}\right)\left(\frac{2.54\,\text{cm}}{1\,\text{inch}}\right)\left(\frac{2.54\,\text{cm}}{1\,\text{inch}}\right)=109.1378\,\text{cm}^{3}$$

Example 7: A person’s heart beats on average 70 times a minute. Each heart beat pumps about 85 mL of blood to the body. If a person lives to be 85 years old, how many gallons of blood has the person’s heart pumped in his/her lifetime?

$$\left(\frac{70\,\text{beats}}{\text{minute}}\right)\left(\frac{85\,\text{mL}}{\text{beat}}\right)\left(\frac{1\,\text{qt}}{946\,\text{mL}}\right)\left(\frac{1\,\text{gallon}}{4\,\text{qts}}\right)\left(\frac{60\,\text{min}}{1\,\text{hr}}\right)\left(\frac{24\,\text{hr}}{1\,\text{day}}\right)\left(\frac{365\,\text{days}}{1\,\text{year}}\right)(85\,\text{years}) = 826458.7738\,\text{gallons}$$

$$\approx 826459\,\text{gallons of blood}$$

Example 8: The Mosteller formula for body surface area is the standard formula to calculate a patient's body surface area when a nomogram is not easily accessible:

BSA (m^2) = ([Height(cm) x Weight(kg)]/ 3600)^{½}

Using both 2.5 cm = 1" and 2.54 cm = 1", determine the difference in the BSA of a patient who is 5'10" tall and weighs 175 lbs.

$$\text{BSA} = \sqrt{\left(\frac{(70'')\left(\frac{2.5\,\text{cm}}{1\,\text{inch}}\right)\left(175\,\text{lbs}\,\frac{1\,\text{kg}}{2.2\,\text{lbs}}\right)}{3600}\right)} = 1.9664\,\text{m}^2$$

$$\text{BSA} = \sqrt{\left(\frac{(70'')\left(\frac{2.54\,\text{cm}}{1\,\text{inch}}\right)\left(175\,\text{lbs}\,\frac{1\,\text{kg}}{2.2\,\text{lbs}}\right)}{3600}\right)} = 1.9821\,\text{m}^2$$

Extending the previous example, what's the difference in drug delivery to the patient if Methotrexate (a cancer chemotherapy drug), 30 mg/m^2 is given twice weekly, intravenously?

$$\frac{30\,\text{mg}}{\text{m}^2}(1.9664\,\text{m}^2) = 58.992\,\text{mg (at 2.5 cm} = 1\,\text{inch)}$$

$$\frac{30\,\text{mg}}{\text{m}^2}(1.9821\,\text{m}^2) = 59.463\,\text{mg (at 2.54 cm} = 1\,\text{inch)}$$

To two sig figs, at the nearest ten, the dosage difference is minimal : both patients would get 60 mg.

A 235 pound patient was admitted onto your floor the other day. You are to give the patient Drug Y at a rate of 125 mg/kg body weight/hour in a 1 liter bag of normal saline that is running at 250 mL per hour. How much of Drug Y will you add to the bag to start the medication off?

$$(235\,\text{lbs})\left(\frac{1\,\text{kg}}{2.2\,\text{lbs}}\right)\left(\frac{125\,\text{mg}}{\text{kg body weight} - \text{hour}}\right)\left(\frac{1\,\text{hour}}{250\,\text{mL}}\right)\left(\frac{1000\,\text{mL}}{\text{bag}}\right) = 53409.0909\,\text{mg} \approx 53.4\,\text{grams in the bag}$$

Experiment

Obtain a metric ruler and complete the following exercises in the lab:

NOTE: record your measurements in the correct units in the table below, using decimals -- NOT fractions! And record data (what you measure or observe in the lab) in INK! Calculations and final results may be written in pencil.

	Inches (in)	Centimeters (cm)	Millimeters (mm)
The length of this piece of paper			
The width of this piece of paper			
The thickness of the lab bench			

Using the data you collected (NOT what you've memorized in class or in this write-up) in your data table, above, calculate the following (show your work):

1) How many cm are in one (1) inch (HINT: cm/in)?

2) How many mm are in one inch (HINT: mm/in)?

3) How many inches are in one centimeter (HINT: in/cm)?

4) How many inches are in one millimeter (HINT: in/mm)?

Obtain a piece of string, a metric ruler and 5 or 6 beakers. Complete the following exercises: measure the diameter and circumference of your beakers. Record your measurements in ink in the following table:

Volume of Beaker (written ON the beaker's side)	Diameter of Beaker (cm)	Circumference of Beaker (cm)

Graph your data as follows using either QuattroPro, Lotus, or Excel or the like: plot the diameter of your beakers on the "x" (horizontal) axis and the circumference of your beakers in the "Y" (vertical) axis. Attach it to your exercises, below, for turn-in with the lab.

Your graph ought to be a straight line (or at least the best straight line you can get fitted to the data points – do NOT dot-to-dot the data pairs) and ought to fit the equation of a straight line: y = mx + b, where "y" is a "y" data point that matches up with a specific "x" data point, "b" is the point at which the line intersects the "y" axis and "m" is the slope of your line. By manipulating the equation to read per the image at right:

$$\frac{y-b}{x}=m$$

Determine the slope of the line. In this case, when you determine the slope of the line from your data, compare it to the value for "pi" (π) in your calculator. How close does your slope come to pi? An alternative method for determining the slope of the line is to use the "rise over run" method you learned in high school geometry class.

Exercises

Complete these exercises on separate paper before you leave the lab (you are welcome to complete them before you come to lab) and attach to this experiment for turn-in.

1) Convert 30 mL to oz, qt, L, gal.
2) Convert 16 oz to kg, lb, tons, g.
3) Convert 25.4 cm to inches, m, km, Mm.
4) Convert 39.37 inches to nm, pm, cm, m.

Introduction

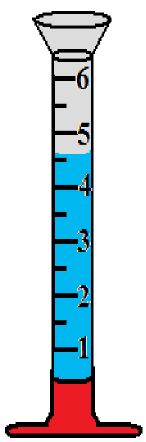
The density of an object relates to how much matter (mass) is present per unit volume. This may be expressed as the formula at right for our purposes.

$$Density = \rho = \frac{mass}{volume} = \frac{grams}{mL}$$

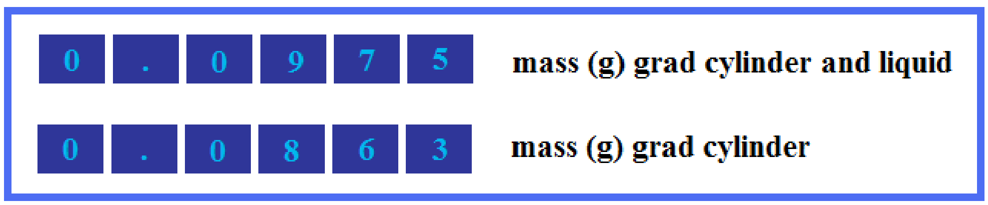
Example at right and graphics below:

$$\rho_{\text{liquid}} = \frac{\text{mass}}{\text{volume}} = \frac{\text{g}}{\text{mL}}$$

$$\rho_{\text{liquid}} = \frac{0.0112\,\text{g}}{4.5\,\text{mL}} = 0.00249\,\frac{\text{g}}{\text{mL}}$$

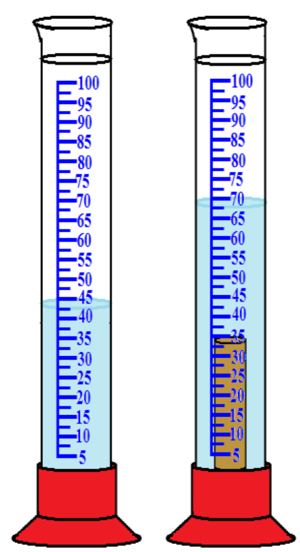


Keep in mind that, although the definition of density is given here as g/mL, it is possible to convert between units as defined and others (as you did in the previous experiment).



Aluminum (Al) cylinders have a density of about 2.701 g/mL, while copper (cu) cylinders have a density of about 8.9 g/mL. In other words, different substances occupying the same volume may have more (or less) matter present than others.

There are times that we may calculate the density using the calculated volume from simple geometry. From MATH pre-requisite courses, as well as the **Math Primer**, you already know how to calculate the volume of a cylinder and a sphere (do not calculate cylindrical volume for this experiment – experimentally obtain it).



Density of an Irregularly- or Regularly Shaped Solid

Determine the density of a metal cylinder of mass 8.735 grams using the image at right.

$$\rho = \frac{\text{mass (g)}}{\text{volume (mL)}} = \frac{8.735\,\text{g}}{(67.5\,\text{mL} - 42.5\,\text{mL})} = 0.3494\,\frac{\text{g}}{\text{mL}}$$

There are, however, simpler methods in which to determine density -- particularly the density of irregularly shaped objects. The simplest method requires that one place a sample of the substance (with known mass) into a pre-massed graduated cylinder and measure the difference in water levels. All one has to do, then, is simply calculate the density of the substance by our first, and definitive, equation. That is the aim of the experiment you are about to perform.

One final note is that even solutions have different densities. Many of these solutions are based in water. The density of water at room temperature is taken to be 1 g/mL. This density is easily obtained by dividing a known mass of water by the volume it occupies. The same may be done for other solutions.

Specific Gravity – Sp.G.

Many times, health care professionals and occupational health professionals wish to know how much more dense a solution is than water. This is done by dividing the density of the other solution by the density of water.

$$\text{Specific Gravity} = \text{Sp.G.} = \frac{\rho \text{ of one liquid}}{\rho \text{ density of a second liquid}} = \frac{g/_{mL}}{g/_{mL}} = \text{UNITLESS}$$

In simplest terms, the Sp.G. is equal to the ratio of the density (ρ) of one liquid to the density (ρ) of water, for our purposes, e.g., the density of mercury (Hg) is 13.6 g/mL and that of water is 1 g/mL. The Sp.G. of Hg is 13.6 and is UNITLESS. This means that Hg is 13.6 times more dense than water.

Clinically, the normal range of urinary Sp.G. is 1.005 to 1.040. If the Sp.G. is less than this range, the urine is very dilute and more like water. Conversely, if the Sp.G. is greater than this range, then the urine is really concentrated and unlike water.

Experimental

Obtain a dry 10 mL graduated cylinder and tap water. Perform the experiment in triplicate and complete the following data table once you have obtained your supplies and data:

	Trial 1	Trial 2	Trial 3
Mass of Graduated Cylinder WITH water (**g**)			
Mass of DRY Graduated Cylinder (**g**)			
Mass of water (**g**)			
Volume of Water in Graduated Cylinder (**mL**)			
Density (**g/mL**)			
Average Density (g/mL)			

Use the above space for calculations.

Obtain an Al cylinder and a Cu cylinder with a 100-mL graduated cylinder containing ca 60 mL water and determine the densities of each of the cylinders by completing the following tables (remember to slide the metal cylinders in at an angle so you don't knock out the bottom of the graduated cylinder!):

	Trial 1	Trial 2	Trial 3
Mass of Al cylinder (DRY) (g)			
Volume of water in graduated cylinder WITH Al cylinder in it (mL)			
Volume of water in graduated cylinder WITHOUT Al cylinder in it (mL)			
Total volume of water displaced by the Al cylinder (mL)			
Density of Al cylinder (g/mL)			
Average Density of Al Cylinder (g/mL)			

Use above space for calculations.

	Trial 1	Trial 2	Trial 3
Mass of Cu cylinder (DRY) (g)			
Volume of water in graduated cylinder WITH Cu cylinder in it (mL)			
Volume of water in graduated cylinder WITHOUT Cu cylinder in it (mL)			
Total volume of water displaced by the Cu cylinder (mL)			
Density of Cu cylinder (g/mL)			
Average density of Cu Cylinder (g/mL)			

Use above space for calculations

Exercises

Complete the exercises, in the space, below, before you come to lab for turn-in at the end of the lab period.

1) A sphere of diameter 5 cm has a mass of 10 g. Determine the density of the sphere.

2) A cylinder of height 12 cm, diameter of 12 cm has a mass of 5 g. Determine the density of the cylinder.

This file upgraded 26 Dec 2013, 1442 hours PST.