Density: Accuracy and Precision

Introduction

Density is a measure of a substance's mass-to-volume ratio. For liquids and solids, density is usually expressed in units of g/mL or g/cm^3; these are equivalent since 1 mL is exactly equal to 1 cm^3. Density is an **intensive property** meaning that a substance's density will be the same regardless of the size of a sample. For example, steel has a density of 7.85 g/cm^3 whether you have a tiny steel ball bearing or a large steel beam. The mass and volume of each piece will be different, but the **mass-to-volume ratio** is the same!

In general, different substances have different densities even if they occupy the same volume or have the same mass. For example, in the opening sequence of *Raiders of the Lost Ark*, Indiana Jones tries to replace a gold statue with a bag of sand of about the same volume. Even though the volumes may be the same, gold is much more dense than sand. As Indiana Jones found out, ignoring density and only considering volume can have dire consequences.

Similarly, two substances may have the same mass but different densities. A classic riddle asks, "Which weighs more: a pound of feathers or a pound of gold?" The answer seems easy. Since a pound is a pound, they should weigh the same. In this case, both samples weigh the same, but feathers are definitely much less dense than gold.

In this lab, you will determine the density of brass as measured by two different techniques: 1) volume calculation based on water displacement and 2) volume calculation using measurements of the radius and length of the cylinder using calipers.

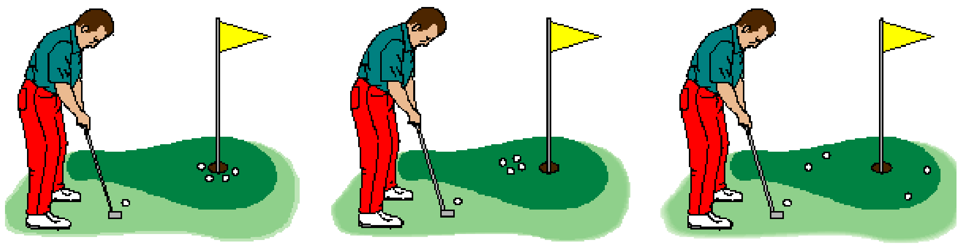
Mass and linear measurements are basic units in the metric system. In chemistry, mass is often measured in grams and length is often measured in centimeters. Volume, on the other hand, can be considered a derived unit. The volume of a solid object can be determined by linear measurements and calculated using volume formulas. The volume of a solid can also be determined by displacement. A solid that is more dense than water will sink and displace a volume of liquid equal to the volume of the solid object. Such a volume measurement is typically made in mL. (Remember, 1 mL = 1 cm^3.)

Accuracy versus Precision

Accuracy is a measure of how close your measured value is to the correct value. For example, if a substance has a density of 1.23 g/mL and you measure its density to be 1.24 g/mL, then you are accurate. The difference between the experimentally measured value and the accepted value is very small.

Precision is a measure of how close repeated measurements are to each other. For example, if you measure a substance with an actual density of 1.23 g/mL four times and calculate densities of 1.430 g/mL, 1.431 g/mL, 1.431 g/mL, and 1.429 g/mL, your measurements are very precise since the difference between the highest and lowest measurement (the range) is small.

However, these measurements are not accurate since they differ greatly from the accepted value of 1.23 g/mL. Ideally, your measurements should be both accurate and precise. The figure below provides a pictorial representation of accuracy and precision in terms of golf.



Good accuracy, good precision. | Poor accuracy, good precision. | Poor accuracy, poor precision.

Significant Figures

Every measurement has a degree of uncertainty associated with it. The uncertainty derives from the measuring device and from the skill of the person doing the measuring.
(This section adapted from http://chemistry.about.com/library/weekly/aa082701a.htm)

If you need to measure 7 mL of water, you could use a beaker that is marked in 5 mL increments. With this beaker, you could estimate a volume between 5 and 10 mL, and probably measure an amount close to 7 mL (within 1 mL). If you used a buret marked to 0.1 mL increments, you could measure a volume between 6.99 and 7.01 mL more reliably. The precision of your measurement is expressed by using the appropriate number of significant figures. The significant digits include the digits you can clearly read, and the last digit which is estimated.

Determining the number of significant figures in a reported measurement:

- Non-zero digits are always significant (e.g., 549 cm).
- Zeroes between non-zero digits are significant (e.g., 1025 mL).
- Zeroes at the end of a number that contain decimal point **are** significant (43.21000 g).
- Zeroes at the beginning of a number **are not** significant (e.g., 0.00482 m). You can always determine which digits are significant by rewriting this number in scientific notation. Leading zeroes do not appear when numbers are written in scientific notation.
- Zeroes at the end of a number and before the decimal point (5300) are assumed to not be significant. If they are significant, the number should be reported in scientific notation so that the significant zeroes are after the decimal place. 5300 would be written 5.300×10^3.

Uncertainty in Calculations:
Measured quantities are often used in calculations in lab. The precision of a calculation is limited by the precision of the measurements on which it is based.

- **Addition and Subtraction:**
  When measured quantities are used in addition or subtraction, the final answer is limited by the **last decimal place in which all digits are significant.** In this example, The tenths place (one after the decimal place) is the last significant digit.

  Example: 32.01 m + 5.325 m + 12.1 m = 49.435 m

  The sum of the numbers is 49.435, but the sum should be reported as 49.4 meters because 12.1 m is only known to the tenths place.

- **Multiplication and Division:**
  When experimental quantities are multiplied or divided, the answer is limited by the number in the calculation with the fewest **number of significant figures** than any of the original numbers.

  Example: 25.624 g / 25 mL = 1.02616 g/mL

  Your calculator will read 1.02616, but the final answer should be reported as 1.0 g/mL because 25 mL only has 2 significant digits.

Using Lab Equipment

Measuring mass using an analytical balance:
When using an analytical balance to measure mass, there are two ways to determine the mass of your sample. In both cases, you should record **all** stable digits displayed on the balance readout. Each digit is considered significant.

1) Measure the mass of an empty container and then the mass of the container plus the sample. Subtracting these two masses will give you mass of just the sample.
2) Analytical balances have a tare button that "zeroes" the mass of the container. If you place the container on the scale and tare the balance (press the O/T bar), then the weight of the container will automatically be subtracted out. You can then remove the container, add the sample, replace the container on the scale (**without** pressing the O/T bar again), and the display will show the mass of just the sample.

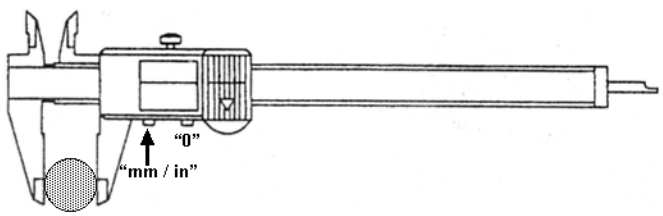
<u>Notes on using analytical balances:</u>

- Before weighing, make sure the doors are closed and that the scale reads 0.0000 g. If it doesn't read all zeroes, **gently press** the O/T bar in front to reset the balance to zero.
- To weigh an item, open the door and carefully place the item on the center of the pan. Close the door and wait for the digital readout to stabilize (the small circle to the left of the readout will go off). Generally, the readout will continue changing over time, but the mass shown when the circle disappears will be accurate.

- **Never add samples to containers that are *inside* the balance. Always remove the container from the balance pan before adding or removing sample. Placing chemicals directly on the balance can damage the balance.**
- Keep the doors closed except when loading or unloading samples.
- **Do not lean on the counter** – it will change the weight displayed.
- Record all digits that are displayed on the balance – they are all significant as long as they are steady and do not change.
- Close the door and tare the balance in between **each** sample to be measured.

Measuring volume with calipers:

- With the calipers tightly closed, be sure the display reads 0.00 (press the "zero" button if it doesn't). Also be sure you are measuring lengths in **millimeters** (mm) and not inches (in).
- In order to measure length, slowly roll the wheel (on the bottom right of the digital reader) until the jaws grip the solid object. **Do not** just slide the jaws open. To ensure a snug fit, you should be able to suspend the object <u>momentarily</u> between the lower portion of the jaws – as shown below.
- Return the jaws to the closed position in between **each reading** to make sure the display returns to zero before measuring the next length.



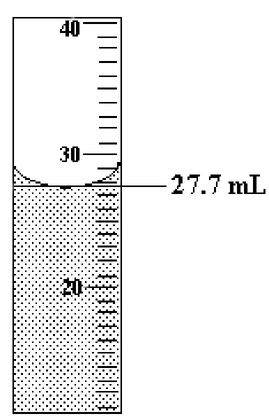
Measuring volume with glassware:

To correctly record volume measurements, you must determine how many significant digits should be recorded for each type of glassware. When water is placed in a glass cylinder, a concave surface forms; this curve is called the ***meniscus***. Glassware is manufactured so that the line at the bottom of the meniscus gives the most accurate reading. In order to read glassware accurately, it must be level (sitting on the counter, **NOT** hand-held) and your eye must be at the same level as the meniscus.

<u>**Reading a 100-mL graduated cylinder:**</u>

Graduated cylinders are designed to measure the amount of liquid contained. They are available in many sizes. **100 mL graduated cylinders can only be read to <u>1</u> decimal place (e.g. 27.7 mL or 50.5 mL)**.

- Note that 100 mL graduated cylinders have markings for each mL, so measurements are reported to one decimal place (to the nearest 0.1 mL), as in the example to the right.
- In the picture, you can see the water volume is definitely between 27 and 28 mL so you should **estimate** the distance between them to get a measurement of the tenths place.



Using your lab Notebook

Print out the "Keeping a Lab Notebook" link from the CHM151 Lab website and keep it inside the back cover of your lab notebook for reference.

You will record **ALL** lab data and calculations in your lab notebook for this course. To help you learn correct set-ups of data tables and calculations, the lab handouts will include example tables. Since the notebook pages use "carbonless" copies, you **must** use a pen. Do not use a pencil since mistakes cannot be erased from the copy.

ALL data and calculations will be recorded **only** in your notebook for every experiment. DO NOT fill in the data tables in this handout. Record all the information directly in your lab notebook. Leave sufficient space between tables and calculations lines to show your work. It is better to overestimate and use plenty of room rather than underestimate and run out of space.

Showing Calculations

ALL calculations must be shown in your lab notebook during the semester. To receive full credit for your calculations, the following should always be followed

- Write down any formulas that you use in your calculations.
- Each step in the calculation should be shown clearly.
- Report the result of the calculation with units (unrounded)
- Report the final answer, with units, to the correct number of significant figures.

Example:

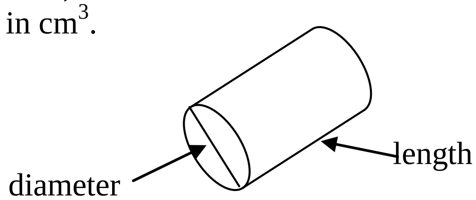
A = l x w
A = 2.34 cm x 1.2 cm
A = 2.808 cm^2
A = 2.8 cm^2

Procedure

1. Use an analytical balance to weigh the 4 brass slugs.

2. Use calipers to measure the length and diameter of the four brass slugs. Be sure that the calipers are reading in **millimeters** (rather than inches). Convert the caliper readings from mm to cm. Calculate the radius and volume of the cylinder. Calculate the density of the cylinder.

From the linear measurements, use the formula for the volume of a cylinder to calculate the volume of each slug in cm^3.



The volume of a cylindrical object (depicted above) is calculated by:
$\mathbf{V = \pi r^2 l}$ where **r** is the radius of the cylinder and **l** is the length of the cylinder.

If you were using a rectangular object, the volume would be given by the formula
V = l x w x h where **l** is the length, **w** is the width and **h** is the height of the rectangle.

$$\text{density} = \frac{\text{mass (in g)}}{\text{volume (in cm}^3\text{ or mL)}}$$

In your lab notebook, create a table of mass and caliper measurements (both millimeter and centimeter) for the four brass slugs. These EXAMPLE tables show the format you should use in your notebook. Data is not to be recorded here! **Leave space between the data tables for your calculations.**

Density Using Calipers Data Table:
Brass (4 cylindrical samples)

Mass (g)	Length (mm)	Length (cm)	Diameter (mm)	Diameter (cm)	Radius (cm)	Volume (cm^3)	Density (g/cm^3)

3. Place a rubber stopper on the bottom of the a 100 mL graduated cylinder to serve as a cushion and prevent breakage. Add about 50 mL of water to the graduated cylinder (this should be sufficient to cover the largest slug when it is placed in the cylinder). Measure and record the volume of the water and the stopper to the maximum precision possible as shown above.

4. Carefully insert one **brass slug** into the graduated cylinder by tipping the cylinder and sliding the sample down the side. Measure the volume of the water, stopper and slug to the proper number of significant figures.

5. Repeat Steps 3 and 4 for each slug. Calculate the volume and density of each slug

$$\mathbf{V_{slug} = (volume\ of\ water + stopper + slug) - (volume\ of\ water + stopper)}$$

Density Using Graduated Cylinder Data Table:
Brass (4 cylindrical samples)

Mass (g)	Volume water + stopper (mL)	Volume water + stopper + slug (mL)	Volume (mL)	Density (g/mL)

6. The accepted value of density is 8.470 g/cm^3 for brass. Calculate the average density value for the caliper method and for the graduated cylinder method.

Data Summary Table:

	Average density from calculation (g/cm^3) Experimental value	Average density from displacement (g/mL) Experimental value
Brass		

7. Determine the ***accuracy*** of your average measured values by calculating the % error for each method. **Percent error values are always reported as positive numbers because the absolute value of the difference is used.** We do not report negative % error values.

$$\textbf{\% error} = \left|\frac{\text{experimental value - accepted value}}{\text{accepted value}}\right| \times 100$$

Accuracy by percent error:

Metal (Accepted Density)	**Percent Error: Caliper Density**	**Percent Error: Graduated Cylinder Density**
Brass (8.470 g/cm^3)		

Name: ______________________________ Section: ____________________

Group members:____

Post-Lab Questions: Write your answers on these pages and attach them to your report.

1. *According to your experimental results* which method (circle one) of determining volume gave the more accurate density?

caliper graduated cylinder

Discuss your results briefly.

2. List one advantage for the water displacement method.

3. List one advantage for the caliper method.

4. List one disadvantage for the water displacement method.

5. List one disadvantage for the caliper method.

6. An object is known to have a density of 4.9435 g/cm^3. A freshman lab class measures objects made of the same material to have the following densities: 4.57, 4.58, 4.55, 4.57, 4.55, 4.56, 4.58, 4.57, and 4.56 g/cm^3.

   Describe the students' data in terms of **accuracy and precision**.

   Explain your answer.

7. Two objects have the same mass but different volumes. Which will be more dense, the one with the larger volume or smaller volume?

Explain your reasoning.

8. A solid cylinder of plastic has a density of 1.6 g/cm^3. It is then cut exactly in half. What is the density of each of the pieces now?

d = _______________

Explain.

9. Liquid A has a density of 0.90 g/cm^3, liquid B has a density of 1.15 g/cm^3, and liquid C has a density of 0.65 g/cm^3. A 10-mL sample of each liquid is poured into a beaker and allowed to sit overnight. Assuming the liquids do not mix into one another, explain which liquid will be on the bottom, in the middle, and on the top of the beaker.

Explanation:

Draw a picture of the cylinder containing the liquids and identify each liquid in your illustration.

