

Significant Figures and Measurement of Density

Objectives: To investigate the concepts of accuracy and precision, and to review the use of significant figures in measurements and calculations. These concepts will be applied in the determination of the density of solids and solutions.

Materials: Solid metallic objects (tin, lead or copper shot or beads); 50- or 100-mL graduated cylinder; 125-mL Erlenmeyer flask with rubber stopper; top-loading balance; solution or solvent for liquid density measurement

Safety: Take care when inserting the rubber stopper into the fully filled Erlenmeyer flask—excessive force and increased hydraulic pressure may cause the neck of the flask to break. Safety goggles should be worn at all times.

Waste Disposal: All solutions should be flushed down the drain with plenty of tap water; solid metal shot/beads can be dried and placed in a collecting container as directed by your instructor.

Review: Rules for significant figures

INTRODUCTION

All scientific investigations involve making measurements. A measured value, however, is only as good as the equipment or tools used to obtain and make the measurement. It is important, therefore, to follow certain guidelines when making measurements or using measured values in calculations.

Consider measuring the mass of an object using a top-loading balance that can be read to the nearest 0.1 grams. The display on the balance indicates that the mass of the object is 42.5 grams. We would record the mass as 42.5 ± 0.1 g, which means we are fairly confident that the actual mass is between 42.4 g and 42.6 g. The uncertainty in our recorded mass would be ± 0.1 g. If we measured the mass of the same object with an analytical balance we might obtain a value of $42.467 \text{ g} \pm 0.001$, which implies a mass between 42.466 g and 42.468 g. The uncertainty in any measurement is usually implied as plus or minus 1 in the last recorded unit. Clearly, the uncertainty in the mass obtained using the analytical balance is much less than the uncertainty in the top-loading balance. The uncertainty of a measurement depends on the sensitivity of the instrument, and determines the number of significant figures used when recording the measured value.

Ideally, the measured values obtained in the laboratory reflect the true value we are trying to measure. The accuracy of our measurements is reflected in how close they are to the correct value. In an effort to ensure accurate results, scientists often make several measurements and then average them so that the error in any given measurement will be minimized. Agreement between multiple measurements is known as precision. Good precision does not necessarily ensure accuracy, however. Consider the following data obtained for the mass of an object on two different balances.

	Balance #1	Balance #2
Measurement #1	27.4 g	27.8 g
Measurement #2	26.9 g	26.1 g
Measurement #3	27.1 g	26.7 g
Average =	27.1 g	26.9 g

The range of measurements for Balance #1 is from 26.9 to 27.4, or only 0.5 g, while the range for Balance #2 is from 26.1 to 27.8, or 1.7 g. The precision of measurements for Balance #1 is better (i.e., better agreement between measurements), but is it more accurate? If the true mass of the object was 26.9 g then the value obtained using Balance #2 would be more accurate, although less precise.

We indicate the precision of a measured value by the number of significant figures we use to record it. Typically, the appropriate number of significant figures will depend on the sensitivity of the instruments we used to obtain the value. If these measured values are then used in a calculation, the precision of the final calculated answer will depend on the precision of the measured values used in the calculation. The calculated answer CANNOT be more precise than the values used in the calculation. It may be worthwhile to review the section in your textbook that discusses the rules for significant figures in calculations before beginning this lab exercise.

In this exercise we will use various approaches to determine the mass and volume of both solid objects and solutions, and use these measured values to calculate **density**. Density, defined as the mass per unit volume, is an intrinsic property of matter which is often used to identify unknown substances. It is important to record measured results to the appropriate number of significant figures, based on the precision of the equipment or instrument used. Mass is measured using an **analytical balance**, as illustrated in Figure 1.

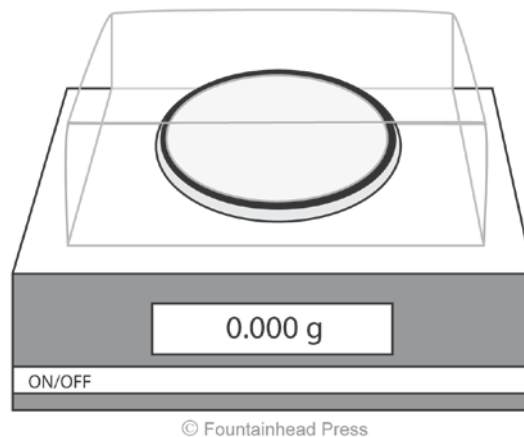
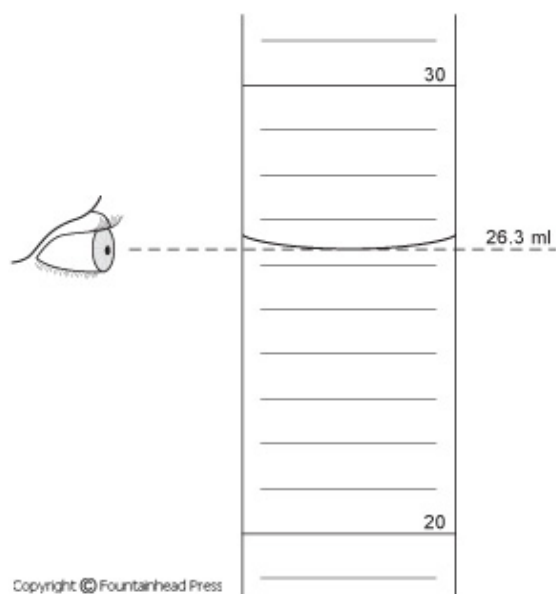


Figure 1. Analytical balance with precision of ± 0.001 g (1 mg).



Volumes of liquids are typically measured using **graduated** glassware, or equipment that is marked with lines to indicate the volume of the liquid. When reading volumes from graduated glassware it is important to read the liquid level at the bottom of the meniscus, or curved surface, while viewing the meniscus at eye level, as illustrated in Figure 2. In this case, the first two significant figures are easily determined, but the last significant figure is estimated.

Figure 2. Reading volumes in graduated glassware.

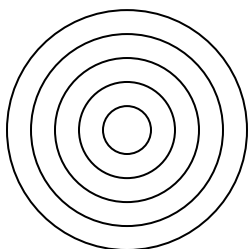
Volumes of liquids can be measured directly using appropriate glassware, but the volumes of irregularly shaped solids must be determined by the volume of liquid displaced by that solid. For example, suppose you wanted to measure the volume of a spherical object. One way to do this would be to partially fill a graduated cylinder with water. Then, place the spherical object in the graduated cylinder. The water level will rise due to the added object. The volume of the solid can be calculated as the difference between the initial and final liquid levels in the graduated cylinder.

In this lab you will determine the density of both liquids and solids. The density of solid substances is typically reported in units of g/cm^3 , while the density of liquids is typically reported in units of g/mL . Since $1 \text{ cm}^3 = 1 \text{ mL}$, these units are often used interchangeably.

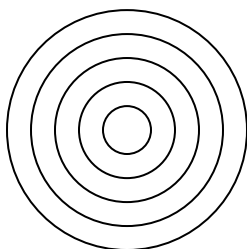
Pre-Lab Questions

1. Use the three targets below to indicate the accuracy and precision of the following sets of measurements:

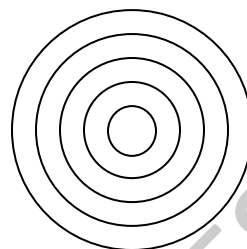
- (a) Place four X's to represent data points with good accuracy but poor precision.
- (b) Place four X's to represent data points with good precision but poor accuracy.
- (c) Place four X's to represent data points with good accuracy and good precision.



(a)



(b)



(c)

2. Write the implied range for a temperature recorded as 38.9°C.

3. Define each of the following terms with regard to scientific measurements.

(a) Accuracy :

(b) Precision :

(c) Sensitivity :

(d) Uncertainty :

4. Indicate the number of significant figures in each of the following:

(a) 20.05 _____

(b) 2.37×10^{-2} _____

(c) 1.460 _____

(d) 0.0462 _____

(e) 3040 _____

(f) 3.040×10^3 _____

5. Perform the following calculations and report the answer to the appropriate number of significant figures:

(a) $48.3 \text{ mL} - 9.27 \text{ mL} =$ _____ (b) $(17.36 \text{ g}) / (22.0 \text{ mL}) =$ _____

6. Explain the rules used to determine the number of significant figures in your answers to Question 5.

7. A student determines the density of a solid object using the procedures described in Part B of this exercise. The following data is obtained:

Mass of object(s): 48.65 grams

Volume of water: 37.6 mL

Volume of water + objects(s): 41.9 mL

Volume of object(s): _____ mL

Density of unknown solid: _____ g/mL

Complete the calculations to find the density of the unknown solid object. Show your calculations below.

PROCEDURE

Part A : Density of a Liquid

1. Clean and dry a 50- or 100-mL graduated cylinder.
2. Weigh the graduated cylinder on the balance provided and record the mass on the data sheet.
3. Obtain a sample of the assigned unknown solution and record the identity. Pour the solution into the graduated cylinder until it is about half full.
4. Record the volume of solution to the nearest 0.1 mL. Be sure to read the volume at the bottom of the meniscus and at eye level. (Note: You may have to estimate the last significant figure.)
5. Place the graduated cylinder with the solution on the balance and record the mass. Work quickly to avoid loss of solution by evaporation.

Part B : Density of a Solid by Displacement

1. Obtain a sample of the assigned unknown solid and record its identity on the Data Sheet provided. Use the same unknown for both procedures B and C.
2. Weight out approximately 50 grams of the unknown solid, and record the mass on the Data Sheet to the appropriate number of significant figures.
3. Fill your graduated cylinder about halfway with water and record the volume to the nearest 0.1 mL.
4. Carefully add all of the unknown solid to the graduated cylinder. Shake gently to release any air bubbles. Record the volume of water + solid in the graduated cylinder. The difference between this new volume and the volume recorded in step 3 is the volume displaced by the unknown solid.

Part C : Density of a Solid by Displaced Mass of Water

1. Recover your unknown sample used in Part B and dry it thoroughly.
2. Take a clean, dry 125-mL Erlenmeyer flask and obtain a rubber stopper that fits. Insert the stopper carefully and mark the level of the bottom of the stopper on the neck of the flask with a marker. Weigh the flask with the stopper and record the mass on the data sheet.

3. Add the unknown solid to the flask and replace the stopper. Record the mass of the flask + solid + stopper to the nearest 0.01 g on the data sheet. The difference between this mass and the mass recorded in step 2 is the mass of unknown solid.
4. Remove the solid from the flask, and fill the flask with water to the line marked in step 2. Insert the stopper carefully so that no air bubbles are left. Dry the outside of the flask thoroughly. Weigh the flask and record the mass on your Data Sheet.
5. Pour out some of the water and add the unknown solid to the flask. Re-fill the flask with water to the line marked in step 2. Dry the outside of the flask, and re-weigh. Record the mass of the flask + stopper + water + unknown on your Data Sheet.
6. Measure the temperature of the water and record it on your Data Sheet.

CALCULATIONS

Part A

1. Subtract the mass of the empty graduated cylinder from the mass of the cylinder + solution. This difference is equal to the mass of the unknown solution.
2. Density is defined as $D = \text{mass} / \text{volume}$. Calculate the density of unknown solution and record it on your Data Sheet to the appropriate number of significant figures.

Part B

1. Subtract the initial volume of water from the volume of water + solid to obtain the volume of the unknown solid by difference.
2. Calculate the density of the solid and record it on your Data Sheet to the appropriate number of significant figures.

Part C

1. Subtract the mass of the flask + stopper from the mass of flask + stopper + solid to obtain the mass of the solid by difference.
2. Subtract the mass of the flask + stopper from the mass of flask + stopper + water to obtain the mass of water by difference.
3. Use the density of water at the measured temperature (Table 1) to calculate the volume of water in the filled flask as follows:

$$\text{Volume of water (mL)} = (\text{mass of water, g}) / (\text{density of water, g/mL})$$

You may need to interpolate between temperatures to obtain the correct density for your temperature. Record this calculated volume of water as the volume of the flask on your Data Sheet.

- Next we must find the volume of the unknown solid. In step 5 of Part C you weighed the flask + stopper + solid + water. In step 3 you weighed the flask + stopper + solid. You can use the difference between these two masses to obtain the weight of water added to the flask. Using the density of water you can obtain the volume of water added to the flask. The volume of the solid can now be obtained as follows:

$$\text{Volume of solid} = \text{volume of flask} - \text{volume of water added}$$

- Use the mass and volume of the solid to calculate the density. Be sure to record the density to the appropriate number of significant figures.

Table 1. Density of Water vs Temperature

<u>Temp (°C)</u>	<u>Density (g/mL)</u>	<u>Temp (°C)</u>	<u>Density (g/mL)</u>
14.0	0.99999	20.0	0.99823
15.0	0.99913	21.0	0.99800
16.0	0.99896	22.0	0.99777
17.0	0.99879	23.0	0.99754
18.0	0.99862	24.0	0.99732
19.0	0.99845	25.0	0.99707

Significant Figures and Measurement of Density Data Sheet

PART A : Solution Density

Unknown: _____

Mass of graduated cylinder _____ g

Volume of solution _____ mL

Mass of grad. cylinder + solution _____ g

Mass of solution _____ g

Density of solution _____ g/mL

Show calculations:

PART B : Density of Solid by Volume Difference

Unknown: _____

Mass of solid _____ g

Volume of water _____ mL

Volume of water + solid _____ mL

Volume of solid _____ mL

Density of solid _____ g/mL

Show calculations:

PART C : Density of Solid by Mass Difference

Unknown: _____

Mass of flask + stopper _____ g

Mass of flask + stopper + solid _____ g

Mass of solid _____ g

Mass of flask + stopper + water _____ g

Mass of water _____ g

Temperature of water _____ °C

Density of water (Table 1) _____ g/mL

Volume of water = volume of flask _____ mL

Mass of flask + stopper + solid + water _____ g

Mass of water added _____ g

Volume of water added _____ mL

Volume of solid _____ cm³

Density of solid _____ g/mL

Show calculations:

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Post-Lab Questions

1. Compare the sensitivity of Parts B and C. Which calculated density has less uncertainty? Explain.
2. If you repeated the measurements from Parts B and C several times, which procedure do you think would give greater precision? Why?
3. How many significant figures did you report for Parts A, B, and C. Which measured value was the determining factor in each case? Explain.
4. There is a very old riddle which asks “Which weighs more--a pound of gold or a pound of feathers?” The most common answer (gold) is not correct. Explain the problem in reasoning that might cause someone to give a wrong answer.