CAROLINA
DISTANCE LEARNING

Dry Lab

Chemistry

Measurement and Uncertainty

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MEASUREMENT AND UNCERTAINTY

Table of Contents

Overview

In this investigation, students will use basic scientific measuring equipment to determine the accuracy and uncertainty associated with measurements using common laboratory glassware.

Outcomes

- Determine the uncertainty of measurements with standard glassware and equipment.
- Determine the accuracy of measurements with standard glassware.
- Apply the mathematical concept of significant figures to measurements.

Time Requirements

Preparation .... 5 minutes

Activity 1: Determination of Uncertainty in Lab Balance .... 20 minutes

Activity 2: Determination of Uncertainty in Common Glassware .... 60 minutes

Activity 3: Determination of Accuracy in Common Glassware .... 40 minutes

Activity 4: Measurement and Significant Figures .... 20 minutes

Activity 5: Everyday Measurements .... 20 minutes

Activity 5 Key

Background

Measurements can come in many forms—such as length, weight (mass), volume, and temperature—but there are many other forms you may encounter in the future. This investigation will focus on two: how to measure weight and volume, and some common equipment used to measure both.

Using Your Balance

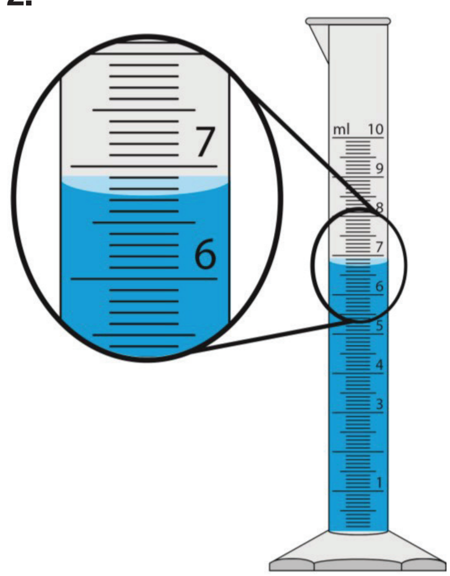
To measure the weight of an object (more scientifically referred to as the mass), a balance or scale is used. Figure 1 shows a small balance that was used to weigh all substances in this investigation. This balance has complete instructions on its use on the inside of its lid. When looking at the top of the balance, there are five points of interest. First, is the pan; this is the large, flat surface above the LCD screen. Objects you wish to weigh are placed on the pan. Below the pan is the LCD screen. This is where the mass of an object on the pan is displayed. Below the screen are three buttons reading, from left to right, ON/OFF, MODE, and TARE. When you first turn on the balance by pressing the ON/OFF button, the screen should read "0.00 g" (*g* stands for *grams* in this instance). This indicates there is no mass on the balance. If the letter at the end is not *g*, you would press the MODE button until "g" is listed as the units. If the screen is not indicating a mass of 0.00, pressing the TARE button will re-zero the scale. The scale should reset to 0.00 g. The scale has a maximum capacity of 100 g; if a mass greater than 100 g is placed on the pan, the screen will read "0_Ld," indicating too large a mass has been placed on the pan.

Figure 1.

Measuring Liquids

To measure a volume of liquid, typically a piece of glassware, such as a beaker or graduated cylinder, is used. The equipment is placed on a flat countertop or table, and liquid is poured into it. The bottom of the **meniscus** (the concave layer of water at the top) is where the volume is measured against the scale (Figure 2). As you will see in Activities 2 and 4, the volume you read from a particular piece of glassware may be at best an estimate.

Figure 2.



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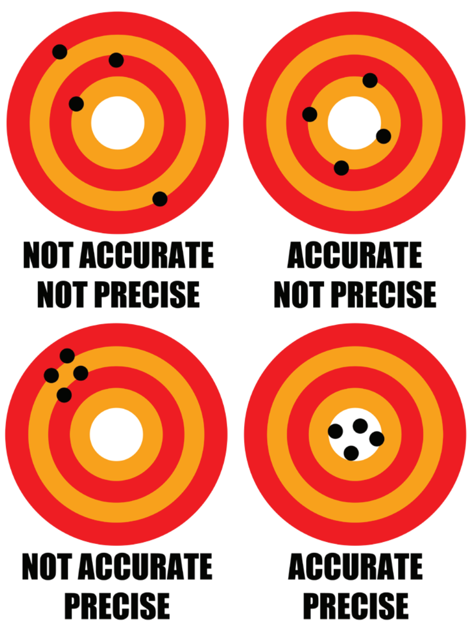
Background continued

Having measurable results is an integral part of the scientific method. Scientists must contend with two main factors while taking measurements: the **accuracy** of the measurement and the **precision** of the measurement. Accuracy is how close a set of data is to the actual value. It is gauged by comparing the measured value of a known standard to its true value.

Precision refers to how close a data point is to other measurements in a data set.

A data set that is accurate is not necessarily precise, whereas a very precise data set could be highly inaccurate (see Figure 3).

Figure 3.



Forces that affect the accuracy and precision in measurements are **error**. In scientific settings, error is defined as the difference between the measured value and the actual value, where the actual value is a known value, sometimes referred to as a standard. Two main types of error exist: **systematic error** and **random error**.

Systematic error causes measurements to be inaccurate by a certain value in a particular direction. Systematic error can be further divided into absolute and relative error. Absolute error has both magnitude and direction and is represented as a discrete value. For example, if your alarm clock is slow by five minutes, it has a systematic, absolute error. Each morning, you will be getting up five minutes later than planned and dealing with the potential repercussions. Absolute error can be calculated as follows:

$$\text{absolute error} = |\text{measurement} - \text{actual value}|$$
$$\text{absolute error} = |6{:}35 - 6{:}40| = 5 \text{ minutes}$$

The | | brackets indicate that you take the absolute value of a calculation. An absolute value means the value in the brackets will always be positive.

There is a second type of systematic error called **relative error**, or percent error, which is expressed as a percentage. One of the more common measuring devices with built-in percent error is the speedometer of a car. Most automobile manufacturers have a tolerance of ±2% in their speedometers. This means that any given speedometer could read between 2% too slow or 2% too fast. If your speedometer reads 61 mph while actually traveling 60 mph, the percent error is calculated using the equation below.

$$\text{relative error} = \frac{|\text{measurement} - \text{actual value}|}{\text{actual value}} \times 100\%$$

$$\text{relative error} = \frac{|61 \text{ mph} - 60 \text{ mph}|}{60 \text{ mph}} \times 100\% = 1.6\%$$

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Related to relative error is the concept of percent error. **Percent error** is calculated by comparing a measurement against an accepted value. Typically an accepted value is measured with a high level of precision and accuracy, but it is still a measured value—no matter what, there is always some form of error associated with a measured value.

$$\text{percent error} = \frac{|\text{measurement} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

An important characteristic of systematic error, both absolute and relative, is that it can be either corrected or accounted for in future measurements since it has both direction and magnitude. With your alarm clock, you could change the time so that it is no longer 5 minutes fast; with the speedometer, you could mathematically correct for the relative error in future readings.

Although systematic error can be corrected for if discovered, random error will be present in all measurements. Through improved experimental design and best lab practices, random error can be reduced but can never be eliminated. The most common form of random error in a lab setting comes from the equipment. This type of random error is most commonly referred to as **uncertainty**. Uncertainty is the limit of quantifiable measurement with confidence using measuring equipment.

One method for determining the uncertainty of an analog measuring device is to utilize the scale provided on the equipment. For example, on a 10-mL graduated cylinder, there are graduations (lines) every 0.1 mL.

In Figure 2, the bottom of the meniscus is between the graduations of 6.7 and 6.8. Most people would read the volume as 6.75 mL. You can say with certainty that the water is between 6.70 and 6.80 mL, but many people would have difficulty determining a finer range of certainty.

A simple method for determining the measured value and the uncertainty is as follows:

$$\text{measured value} = \frac{\text{high interval} + \text{low interval}}{2}$$

$$6.75 = \frac{6.80 + 6.70}{2}$$

$$\text{uncertainty} = \frac{\text{high interval} - \text{low interval}}{2}$$

$$0.05 = \frac{6.80 - 6.70}{2}$$

The measured value in this example would be 6.75 mL ± 0.05 mL. The ±0.05 mL indicates confidence that the actual value for this measurement is between 6.80 mL and 6.70 mL.

With a digital device, such as a balance, uncertainty is generally limited to the last significant figure. For example, a balance that can read to tenths of a gram would have an uncertainty in the tenths place, typically of ±0.1 or ±0.2 grams. Uncertainty is generally calculated using a **standard** and a high number of measurements. A standard is a chemical or piece of equipment that has a known quantity associated with it, in this case a mass. For this activity, plastic cups are used as your standard for determining the uncertainty in your balance.

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Background continued

Error, uncertainty, and equipment segue into the mathematical concept of **significant figures**. Significant figures are digits relating to the precision of measurement. There are some general rules for determining if a digit is significant:

- All nonzero digits are considered significant.
- Zeros appearing anywhere between two nonzero digits are significant (0.1003 has four significant figures).
- Leading zeros are not significant (0.0076 has two significant figures).
- Trailing zeros in a number containing a decimal point are significant. For example, 35.000 has five significant figures.

Uncertainty limits the precision and the number of significant figures in a measurement. In the example discussed, the 6.75 mL of water in the graduated cylinder has three significant figures. The "6" before the decimal and the "7" and "5" after the decimal are all considered significant. This is confirmed with the uncertainty of 0.05 mL. In this instance, the uncertainty indicates that there are no additional significant figures beyond the hundredths place. However, if the graduated cylinder was measured at 6.75 but the uncertainty was determined to be 0.20 mL, the number of significant figures would be limited to two and the measurement would be reported as 6.8 mL ± 0.2 mL.

In the next example, let's assume that the volume measurement above had a relative error of 1%.

$$6.75 \text{ mL} \times \frac{1\%}{100\%} = 0.0675 \text{ mL}$$

This would equate to an absolute error of 0.0675 mL in the measurement. Like 6.75 mL, 0.0675 mL has three significant figures. However, the process of multiplication and division has added a false precision to the result. The measurement 6.75 mL ± 0.0675 mL is incorrect because the calculated error has additional precision that the original measurement can contain. In this instance, the proper measured value would be written as 6.75 mL ± 0.07 mL. In general, you cannot gain significant figures and you cannot gain precision in a measurement through mathematical functions.

Safety

There are no safety concerns for this investigation.

Preparation

Read all the instructions for this laboratory activity before beginning. Follow the instructions closely.

ACTIVITY

ACTIVITY 1

A Determination of Uncertainty in Lab Balance

Observe the procedure in the figures shown, and note your observations in the table indicated.

1. Assume your balance is turned on and the reading is stabilized at 0.00 (see Figure 4).
2. Two cups are labeled “#1” and “#2” (see Figure 5).

Figure 4.

Figure 5.

3. Assume Cup #1 is placed on the balance. The mass is shown in Figure 6. Record this in **Data Table 1** (see page 15) below “Mass of Cup #1” as Trial 1.

Figure 6.

4. Assume Cup #1 is removed from the balance and the balance is restabilized at 0.00.
5. Repeat Steps 3 and 4 for four additional trials, recording the readings shown in Figures 7–10 in **Data Table 1** as Trials 2–5.

Figure 7.

Figure 8.

Figure 9.

Figure 10.

6. Assume Cup #1 has been removed, the balance as been restabilized at 0.00, and Cup #2 is on the balance. The mass is shown in Figure 11. Record this in **Data Table 1** below “Mass of Cup #2” as Trial 1.

Figure 11.

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ACTIVITY 1 continued

7. Assume Cup #2 is removed from the balance and the balance is restabilized at 0.00.
8. Repeat Steps 6 and 7 for four additional trials, recording the readings shown in Figures 12–15 in **Data Table 1** as Trials 2–5.
9. Determine the average mass of each cup, and record the value in **Data Table 1**.
10. For each trial, perform the following calculation:

    deviation from average =
    |average mass of cup – mass of cup in trial|

11. Determine the average "deviation from average" for each cup. This is the uncertainty of your measurement with the balance.

Figure 12.

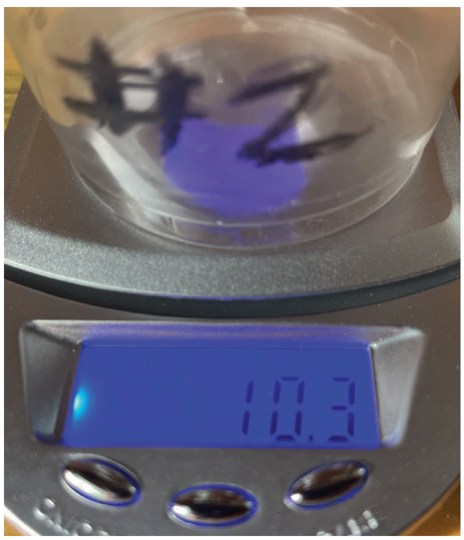

Figure 13.

Figure 14.

Figure 15.

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ACTIVITY 2

B Determination of Uncertainty in Common Glassware

1. Assume your balance is turned on and the reading is stabilized at 0.00 (see Figure 16).
2. Assume the 10-mL graduated cylinder is placed on the balance. The mass is shown in Figure 17. Record this in **Data Table 2**.

Figure 16.

Figure 17.

3. Assume the graduated cylinder has been removed from the balance and approximately 5 mL of water has been added to it (see Table 1).
4. Record the volume of water in **Data Table 2** based on the meniscus of the water as shown in Figure 18. Realize that, for some pieces of glassware, this may be an estimate.

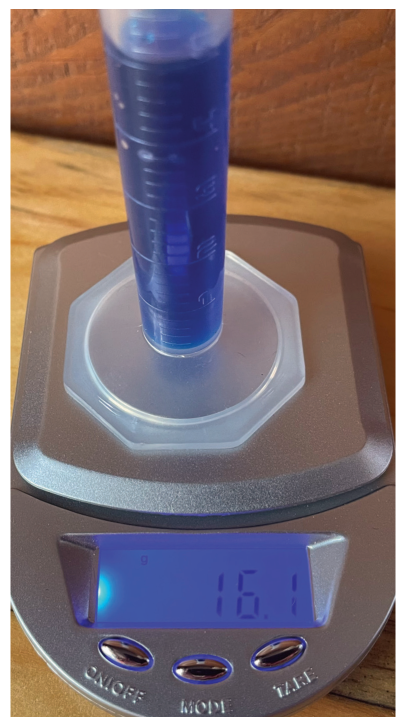
Figure 18.

5. Record the high and low volume intervals in **Data Table 2**. These should be volumes that you are certain the actual volume is between. Use the graduations (lines) on the glassware to help determine the high and low intervals.
6. Calculate the uncertainty of your measurement, and record it in **Data Table 2**.

   Uncertainty =
   (high volume interval – low volume interval)/2

7. Assume the balance reading is zeroed. The mass of the 10-mL graduated cylinder with water is shown in Figure 19. Record the mass in **Data Table 2**.

Figure 19.



8. Repeat Steps 1–7 with each remaining glassware shown in Figures 20–25. Replace the volumes in Step 3 with the target volumes listed in Table 1. Be sure to record the data in **Data Table 2**.
9. Calculate the mass of water in each piece of glassware, and record the mass in **Data Tables 2** and **3**.

   Mass of water = Mass of glassware with water – Mass of empty glassware

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ACTIVITY 2 continued

Table 1.

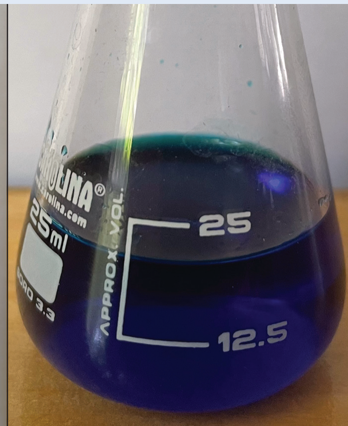
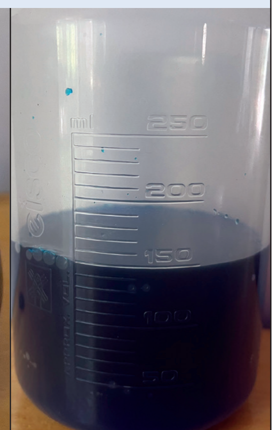
Glassware			
			
Graduated cylinder, 10 mL	Graduated cylinder, 100 mL	Erlenmeyer flask, 25 mL	Beaker, 250 mL
Target Volume of Water (mL)			
5	60	21	150

Figure 20.

Figure 21.

Figure 22.

Figure 23.

Figure 24.

Figure 25.

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ACTIVITY 3

C Determination of Accuracy in Common Glassware

1. Using the thermometer shown in Figure 26, record the room temperature in degrees centigrade in **Data Table 3** as the current water temperature. (For simplicity, assume that the water used in Activity 2 was at this room temperature.)
2. Using Table 2, record the density of water at the current room temperature in **Data Table 3.**
3. Calculate the volume of water from Activity 2 for each piece of glassware, and record it in **Data Table 3.**

Reminder: volume = mass/density

Figure 26.

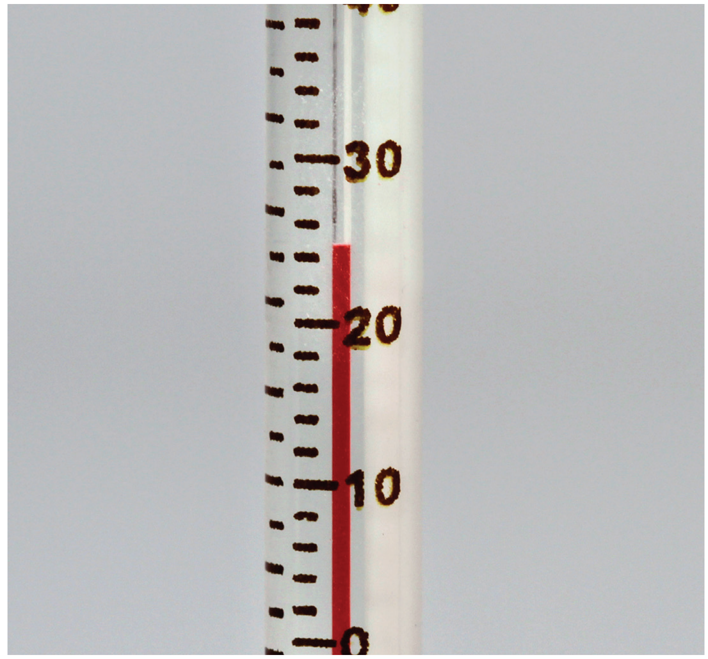

Table 2.

Temperature ° C	Density (g/mL)	g/mL + 0.1° C	g/mL + 0.2° C	g/mL + 0.3° C	g/mL + 0.4° C	g/mL + 0.5° C	g/mL + 0.6° C	g/mL + 0.7° C	g/mL + 0.8° C	g/mL + 0.9° C
18	0.9986	0.9986	0.9986	0.9985	0.9985	0.9985	0.9985	0.9985	0.9984	0.9984
19	0.9984	0.9984	0.9984	0.9983	0.9983	0.9983	0.9983	0.9983	0.9982	0.9982
20	0.9982	0.9982	0.9982	0.9981	0.9981	0.9981	0.9981	0.9981	0.9980	0.9980
21	0.9980	0.9980	0.9979	0.9979	0.9979	0.9979	0.9979	0.9978	0.9978	0.9978
22	0.9978	0.9977	0.9977	0.9977	0.9977	0.9977	0.9976	0.9976	0.9976	0.9976
23	0.9975	0.9975	0.9975	0.9975	0.9974	0.9974	0.9974	0.9974	0.9973	0.9973
24	0.9973	0.9973	0.9972	0.9972	0.9972	0.9972	0.9971	0.9971	0.9971	0.9971
25	0.9970	0.9970	0.9970	0.9970	0.9970	0.9969	0.9969	0.9969	0.9968	0.9968
26	0.9968	0.9968	0.9967	0.9967	0.9967	0.9966	0.9966	0.9966	0.9966	0.9965
27	0.9965	0.9965	0.9965	0.9964	0.9964	0.9964	0.9963	0.9963	0.9963	0.9963

How to use this table: If the water temperature is 23.4° C: Start at the 23° C row and go over to the “g/mL + 0.4° C” column. The density at 23.4° C would be 0.9974 g/mL.

ACTIVITY 4

A Measurement and Significant Figures

1. Observe the measurements in Figures 27–39, and note your observations in **Data Table 4**.
2. Calculate the level of uncertainty for each Figure 27–39, and record the calculation in **Data Table 4**.
3. Record the number of significant figures that are present for each measurement in Figures 27–39. Record the calculation in **Data Table 4**.

Recall: A simple method for determining the measured value and the uncertainty is

$$\text{measured value} = \frac{\text{high interval} + \text{low interval}}{2}$$

$$\text{uncertainty} = \frac{\text{high interval} - \text{low interval}}{2}$$

Recall: There are some general rules for determining if a digit is significant.

- All nonzero digits are considered significant.
- Zeros appearing anywhere between two nonzero digits are significant. For example, 0.1003 has four significant figures.
- Leading zeros are not significant. For example, 0.0076 has two significant figures.
- Trailing zeros in a number containing a decimal point are significant. For example, 35.000 has five significant figures.

Figure 27.

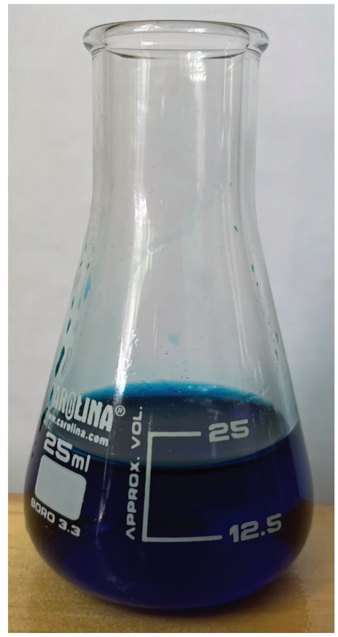

Figure 28.

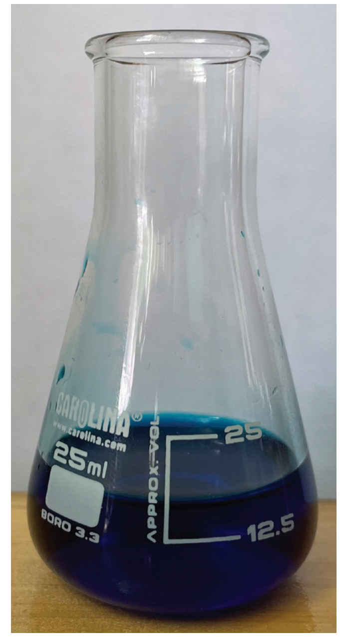

Figure 29.

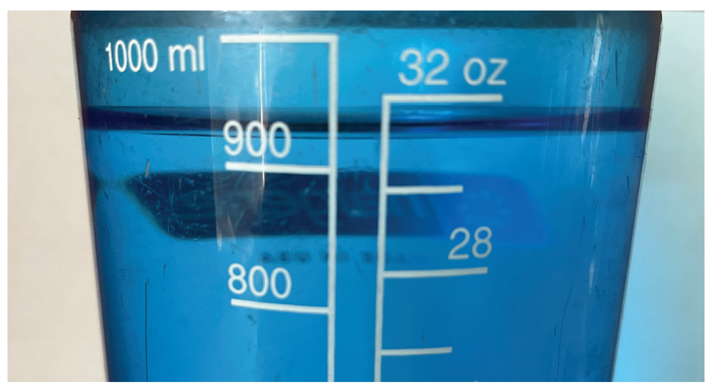

Figure 30.

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Figure 31.

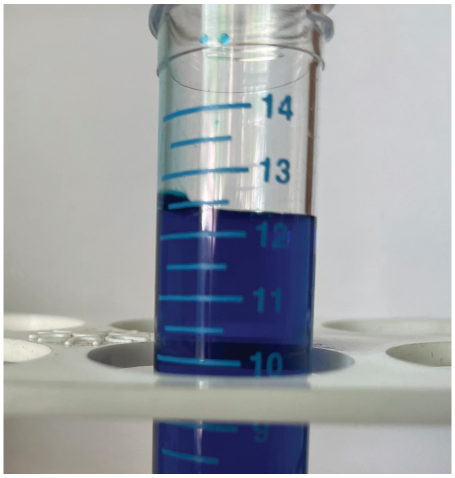

Figure 32.

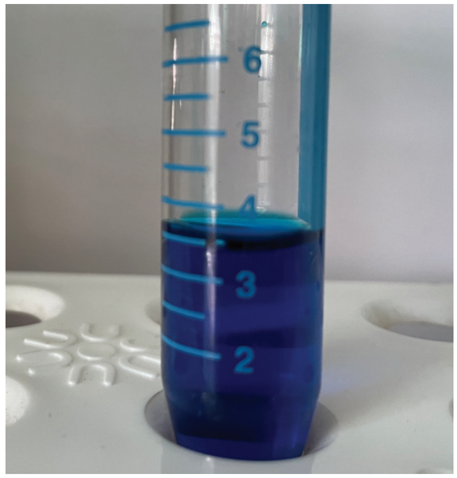

Figure 33.

Figure 34.

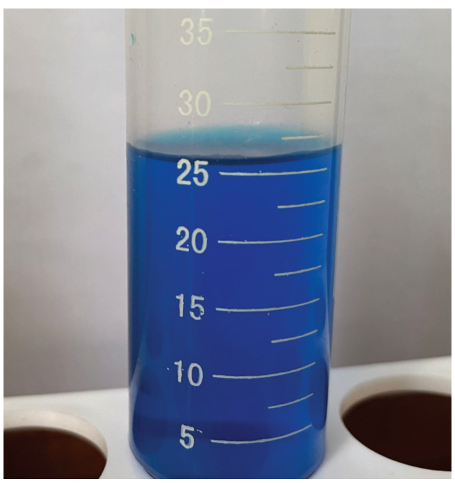

Figure 35.

Figure 36.

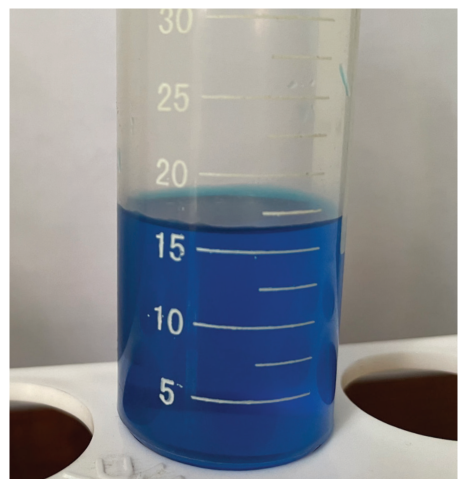

Figure 37.

Figure 38.

Figure 39.

ACTIVITY 5

A Everyday Measurements

Measurements can be found all around us, and they are often taken for granted. If you feel sick, for example, you may take your temperature using a thermometer. The thermometer is the tool used to measure your temperature. If you want to take some liquid medicine to feel better, you may measure the amount in a small measuring cup. Pouring the amount accurately is essential to feeling better and your well-being.

If you like to bake, proper measurement is essential to get the best results. If you do not mix with the correct ratios of ingredients, you will get flat, thin, crispy cookies instead of fluffy, thick, scrumptious ones.

1. Find 3 items/objects in your location. For each one, estimate what you think the length, height, weight, or volume is. Record each item's name and your estimate in **Data Table 5.**
2. Use an appropriate measuring device (i.e., ruler, scale, measuring cup) to determine each item's actual length, height, weight, or volume. Record your data in **Data Table 5.**

Observations

Data Table 1.

Determination of Uncertainty in Lab Balance				
	Mass of Cup #1		Mass of Cup #2	
Trial	**Mass (g)**	**Deviation from Average (g)**	**Mass (g)**	**Deviation from Average (g)**
1				
2				
3				
4				
5				
Average				

Data Table 2.

Determination of Uncertainty in Common Glassware				
	10-mL Graduated Cylinder	**100-mL Graduated Cylinder**	**25-mL Erlenmeyer Flask**	**250-mL Beaker**
Mass of empty glassware (g)				
Estimated volume of water (mL)				
High volume interval (mL)				
Low volume interval (mL)				
Uncertainty (mL)				
Mass of glassware with water (g)				
Mass of water (g)				

ACTIVITY

Data Table 3.

Determination of Accuracy in Common Glassware				
	10-mL Graduated Cylinder	**100-mL Graduated Cylinder**	**25-mL Erlenmeyer Flask**	**250-mL Beaker**
Mass of water in Activity 2 (g)				
Current water temperature (˚ C)				
Density of water at room temperature (g/mL)				
Calculated volume of water (mL)				

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Data Table 4.

Determining Measurement and Significant Figures				
Figure	**Measurement Device**	**Observed Measurement with Units**	**Level of Uncertainty**	**Significant Figures**
27				
28				
29				
30				
31				
32				
33				
34				
35				
36				
37				
38				
39				

Data Table 5.

Estimating and Determining Everyday Measurements		
Item/Object	**Estimate with Units**	**Actual with Units**

NOTES

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Measurement and Uncertainty
Investigation Manual

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