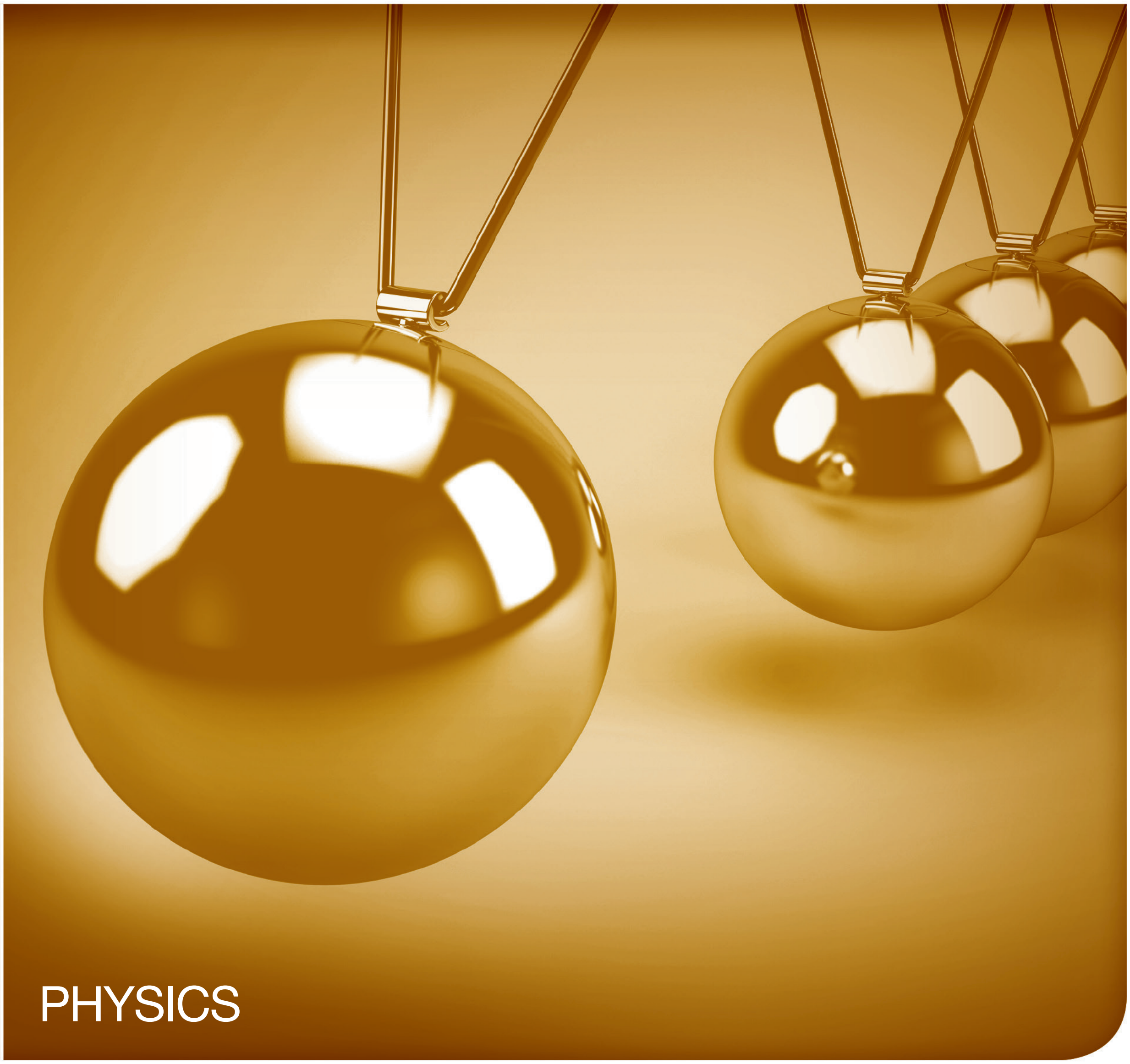

PHYSICS

Measurement and Uncertainty

Investigation Manual

CAROLINA DISTANCE LEARNING

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.
- Define accuracy and precision as they pertain to measurements.
- Identify different types of errors that may affect accuracy and precision of measurements.
- 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	30 minutes
Activity 3: Determination of Accuracy in Common Glassware	20 minutes

Key

Personal protective equipment (PPE)


goggles


gloves


apron


follow link to video


photograph results and submit


stopwatch required


warning


corrosion


flammable


toxic


environment


health hazard

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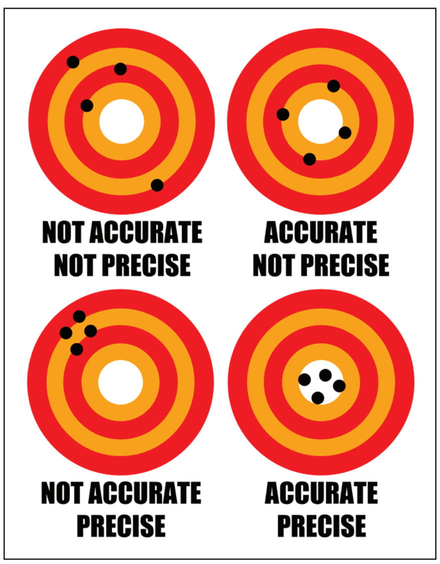
Background

Accuracy is gauged by comparing the measured value of a known standard to its true value. 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. Precision refers to how close datum 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 1). 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 and random.

Figure 1.



Systematic error is a type of error that 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. It is important to note that **systematic errors are consistent. They produce either consistently large or consistently small errors in a fixed proportion. If you repeat the experiment with the same equipment, you will see the same 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 vertical bars shown above indicate that you take the absolute value of a calculation. An absolute value means the value inside the bars 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 following equation.

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

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Background continued

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

Percent error (relative error) is used when comparing a measured value to a known standard. For example, if an experiment required a calculation of the speed of sound at a given temperature, you would use the calculation for percent error to compare your result to the accepted value for the speed of sound at the given temperature, since multiple experiments have confirmed this accepted value. If, however, your experiment requires you to measure or calculate a result using two different methods, you would calculate the percent difference. For example, you could measure the height of a building by dropping a golf ball from the roof and measuring the time, or you could stand on the ground a known distance from the base of the building and calculate the height using trigonometry. To see how close the values are to each other, you would calculate the percentage difference.

$$\text{percent difference} = \left| \frac{\text{first value} - \text{second value}}{\frac{(\text{first value} + \text{second value})}{2}} \right| \times 100\%$$

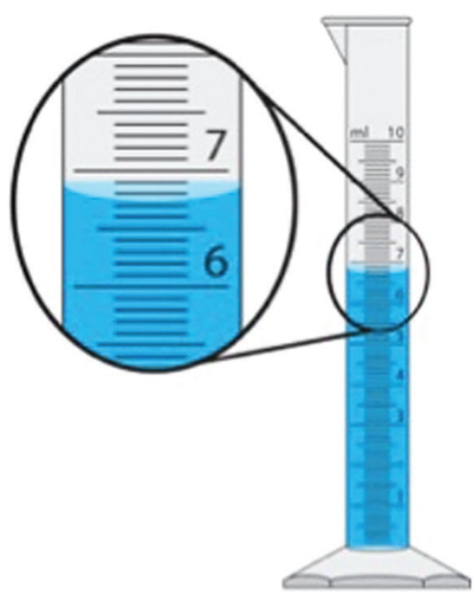
An important characteristic of systematic error, both absolute and relative, is that it can be either corrected or accounted for in future measurements as it has both direction and magnitude. With your alarm clock, you could change the time so that it was no longer 5 minutes fast; with the speedometer,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

Although systemic error can be corrected for if discovered, random error has no pattern and is usually unavoidable. Through improved experimental design and best lab practices, random error can be reduced, but it can never be eliminated. The most common form of random error in a lab setting comes from the limitations of reading 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, in Figure 2, there are graduations (lines) every 0.1 mL on the 10 mL graduated cylinder.

In Figure 2, the bottom of the meniscus is

Figure 2.



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between the graduations of 6.7 and 6.8. In one case the volume is low, in another case it is high. This is how we know an instrument's level of precision. 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 with confidence that the actual value for this measurement is between 6.80 mL and 6.70 mL.

With a digital device, such as the balance supplied in your equipment kit, 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.

Figure 3.

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 non-zero digits are considered significant.
- Zeros appearing anywhere between two non-zero digits are significant; for example, 1,003 has 4 significant figures.
- Leading zeros are not significant; for example, 0.0076 has 2 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 above, 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

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Background continued

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 a volume of 6.75 mL, 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 (6.75 mL) 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. 6.75 mL ±0.0675 mL is incorrect because the calculated error has additional precision than the original measurement can contain. In this instance, the proper measured value would be written as 6.75 mL ± 0.06 mL. In general, you cannot gain significant figures and you cannot gain precision in a measurement through mathematical functions.

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 your 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.

Other examples of everyday measurements are picking out the clothes you wear, estimating what time you need to leave your house to make it to work on time, and how much fuel you need to put in your car to make it to your destination.

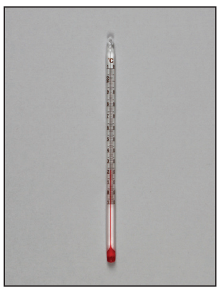
Materials

Needed from the equipment kit:

Electronic balance

Graduated cylinder, 250 mL

Beaker, 250 mL



Thermometer

2 Large metal washers

Needed but not supplied:
- Permanent marker

Reorder Information: Replacement supplies for the Measurement and Uncertainty Investigation (item 580090) can be ordered from Carolina Biological Supply Company.

Call: 800.334.5551 to order.

Preparation

1. Read through the activities.
2. Clear the work space.
3. Obtain all materials and set up for the lab activity.

Safety

Safety goggles should be worn during this investigation. There are no additional safety concerns.

Read all of the instructions for this laboratory activity before starting the activity. Follow the instructions closely and observe established laboratory safety practices, including the use of appropriate personal protective equipment.

Do not eat, drink, or chew gum while performing this activity. Wash your hands with soap and water before and after performing the activity. Clean up the work area with soap and water after completing the investigation. Keep pets and children away from lab materials and equipment.

ACTIVITY

ACTIVITY 1

A Determination of Uncertainty in Lab Balance

1. Turn on your balance and allow the reading to stabilize at 0.0. If the balance does not read 0.0, press the tare button.
2. Label two large metal washers as “1” and “2.”
3. Place Washer 1 on the balance and record the mass in **Data Table 1**.
4. Remove the washer from the balance and allow the balance to restabilize at 0.0.
5. Repeat Steps 3 and 4 for four additional readings.
6. Place Washer 2 on the balance and record the mass in **Data Table 1**.
7. Remove the washer from the balance and allow the balance to restabilize at 0.0.
8. Repeat Steps 6 and 7 for four additional readings.
9. Determine the average mass of each washer and record the value in **Data Table 1**. (Remember, the average is the sum of the mass of all trials, divided by the number of trials.)
10. For each trial, perform the following calculation:

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

11. Calculate the average **deviation from average** for each washer and record this value in **Data Table 1**.

Data Table 1. Determination of Uncertainty in Lab Balance

Mass of Washer 1			Mass of Washer 2	
Trial	**Mass (g)**	**Deviation from average (g)**	**Mass (g)**	**Deviation from average (g)**
1				
2				
3				
4				
5				
Average				

continued on next page

ACTIVITY 2

A Determination of Uncertainty in Common Glassware

1. Turn on your balance and allow the reading to stabilize at 0.0. If the balance does not read 0.0, press the tare button.
2. Place the 250 mL graduated cylinder on the balance and record the mass in **Data Table 2**.
3. Remove the graduated cylinder from the balance.
4. Fill the graduated cylinder with approximately 50 mL of water.
5. Record the volume of water in **Data Table 2**.

> **Note:** The volumes in Step 5 are high confidence estimates; i.e., you are sure that the true value is between the high volume and low volume.

6. Record the highest and lowest volume interval 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 higher and lower interval.
7. Calculate the uncertainty of your measurement.

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

8. Zero the balance, and record the mass of the graduated cylinder with water in **Data Table 2**.
9. Repeat Steps 1–8 with the graduated cylinder for one additional trial.
10. Repeat Steps 1–8 with the 250 mL beaker.
11. 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

Data Table 2. Determination of Uncertainty in Common Glassware

	250 mL Graduated Cylinder	250 mL Beaker
Mass of empty glassware		
Estimated volume of water	50 mL	50 mL
High volume		
Low volume		
Average volume		
Uncertainty		
Mass of glassware with water		
Mass of water		

continued on next page

ACTIVITY 3

A Determination of Accuracy in Common Glassware

1. With the thermometer, record the current water temperature in **Data Table 3.**
2. Using Table 1 on page 11, record the density of water at the current water temperature in **Data Table 3.**
3. Calculate the volume of water using the density for the current water temperature from **Data Table 3.**

Remember: density = mass/volume

4. Calculate the percent difference between the volume of water in Activity 2 (50 mL) and the value calculated for the volume of water in Activity 3.

$$\text{percent difference} = \left| \frac{\text{first value} - \text{second value}}{\frac{(\text{first value} + \text{second value})}{2}} \right| \times 100\%$$

5. Repeat Step 4 with the 250 mL beaker using the target volume in **Data Table 2.**

Disposal and Cleanup

Rinse and dry the lab equipment and return the materials to your equipment kit.

Data Table 3. Determination of Accuracy in Common Glassware

	250 mL Graduated Cylinder	250 mL Beaker
Mass of water in Activity 2		
Current water temperature		
Density of water at current temperature (Data Table 1)		
Calculated volume of water		

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Table 1. Density vs. Temperature

Temp. °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.

PHYSICS
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