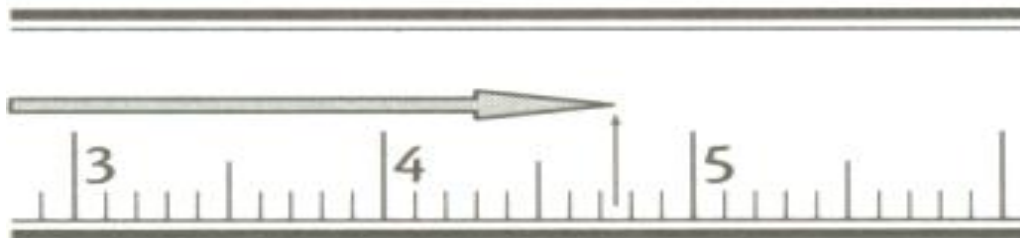


SIGNIFICANT FIGURES

As discussed earlier, measurements are an integral part of most chemical experimentation. However, the numerical measurements that result have some inherent uncertainty. This uncertainty is a result of the measurement device as well as the fact that a human being makes the measurement. No measurement is absolutely exact. When you use a piece of laboratory equipment, read and record the measurement to one decimal place beyond the smallest marking on the piece of equipment.



The length of the arrow placed along the centimeter stick is 4.75 cm long. There are no graduation markings to help you read the last measurement as 5. This is an estimate. As a result this digit is uncertain. Another person may read this as 4.76 cm. This is acceptable since it is an estimation. There is error (uncertainty) built into each measurement and cannot be avoided.

If the measurement is reported as 4.75 cm, scientists accept the principle that the last digit has an uncertainty of ± 0.01 cm. In other words the length might be as small as 4.74 cm or as large as 4.76 cm. It is understood by scientists that the last digit recorded is an estimation and is uncertain. It is important to follow this convention.

There are three processes involving significant figures that we will mention here. The first process is estimation in measurement. You will encounter this most frequently in the laboratory when measuring volumes with a graduated cylinder, a buret or perhaps a pipette. Estimation will also be important when measuring length with a ruler. The process of estimating the last digit will be very similar to the process described at the beginning of this section.

Another process is examining measurements or data already taken by you or someone else and determining the number of significant figures in the measurement. The next section provides you with six guidelines for determining the number of significant figures in a recorded measurement.

The third process is reporting a calculated answer to the proper number of significant figures. The rules for assigning the appropriate number of significant figures to an answer are covered in the section that begins on page 16.

Significant figures will be of concern any time you perform a mathematical computation based on data. It is important to present an answer that reflects the precision of the instruments used to collect the data.

Guidelines for Determining Significant Digits

1. **All digits recorded from a laboratory measurement are called significant figures (or digits).**

The measurement of 4.75 cm has three (3) significant figures.

- Note:** If you use an electronic piece of equipment, such as a balance, you should record the measurement exactly as it appears on the display.

2. **All non-zero digits are considered significant.**

There are special rules for zeros. Zeros in a measurement fall into three types: leading zeros, trailing zeros, and middle zeros (zeroes between non-zero digits).

3. **A middle zero is always significant.**

303 mm has 3 significant figures: a middle zero—always significant

4. **A leading zero is never significant.** It is only a placeholder; not a part of the actual measurement.

0.0123 kg has 3 significant figures: a leading zero—never significant

5. **A trailing zero is significant when it is to the right of a decimal point.** This is not a placeholder. It is a part of the actual measurement.

23.20 mL has 4 significant figures: a trailing zero—significant to the right of a decimal point

220 mL has only 2 significant figures: a trailing zero—there is no decimal point so this zero is NOT significant

6. **All significant figures include units since they are a result of a measurement.** A number without units has little significance.

Measurement	Number of Significant Figures
123 g	3
46.54 mL	4
0.33 cm	2
726.8 mmHg	4
0.0336	3

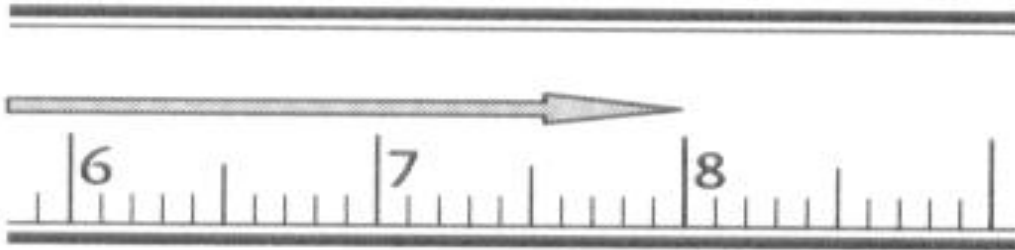
Figure 5 Significant Figures

The most common errors concerning significant figures are (1) recording all digits on the calculator readout, (2) failing to include significant trailing zeroes (e.g., 14.150 g), and (3) considering leading zeroes to be significant (e.g., 0.002 g: only 1 significant figure, not three).

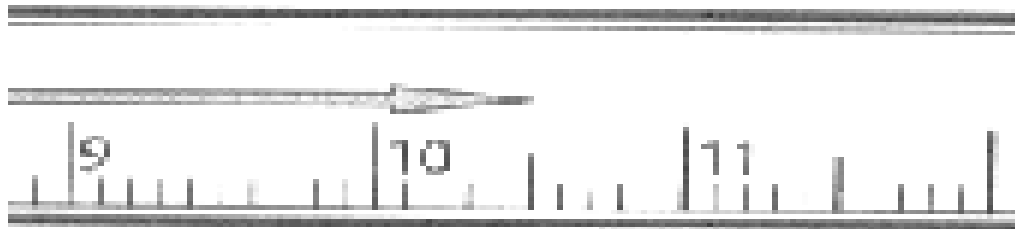
Practice Problems

- How many significant figures are in each of the following?
 - 451 000 m
 - $6.626 \times 10^{-34} \text{ J} \cdot \text{s}$
 - 0.0065 g
 - 4056 V
 - 0.0540 mL
- For the centimeter rulers below record the length of the arrow shown.

a.



b.



USING SIGNIFICANT FIGURES IN CALCULATIONS

When performing calculations in chemistry you will take measurements and apply some mathematical operation. You might be adding masses or finding a difference in masses. You might be solving a problem through multiplication or division. It is easy to plug the numbers into the calculator and get a display that may have up to ten digits. Not all of those digits can be kept in the final answer! Chemists, and chemistry students, have to be honest at the end of every calculation and reflect the level of precision in the data. Here are some rules to help you determine how many digits to keep at the end of a mathematical operation.

Addition and Subtraction

The number of decimal places in the answer should be the same as in the measured quantity with the **least precision**. The easiest way to determine this is by examining the number of places, e.g., hundreds, tens, ones, tenths, hundredths, thousandths, in the measurement. The answer can have as many places in the answer as the term with the fewest places.

$$\begin{array}{r} 1259.1 \text{ grams} \\ 2.365 \text{ grams} \\ \hline 15.34 \text{ grams} \\ 1277.075 \text{ grams} = 1277.1 \text{ grams} \end{array} \qquad \begin{array}{r} 1875 \text{ J} \\ 950 \text{ J} \\ 86 \text{ J} \\ 2911 \text{ J} = 2910 \text{ J} \end{array}$$

± 0.1 gram is least precise measurement
i.e., you can only go to the tenths place

± 10 J is the least precise measurement
i.e., you can only go to the tens place

Multiplication and Division

The number of significant figures in the answer should be the same as in the measured quantity with the **smallest number of significant figures**.

$$\frac{13.356 \text{ g}}{10.42 \text{ mL}} = 1.2817658 \text{ g mL}^{-1} = 1.282 \text{ g mL}^{-1}$$

13.345 g has 5 significant figures, 10.42 mL has 4 significant figures. The answer can only have 4 significant figures.

Practice Problems

1. Answer the following problems using the correct number of significant figures.
 - a. $16.27 + 0.463 + 32.1$
 - b. $42.05 - 3.6$
 - c. 15.1×0.032
 - d. $13.36 \div 0.0468$
 - e.
$$\frac{(13.36 + 0.046) \times 12.6}{1.424}$$
2. In the laboratory a group of students was assigned to determine the density of an unknown liquid. They used a buret to measure the liquid and found a volume of 2.04 mL. The mass was determined on an analytical balance to be 2.260 g. How should they report the density of the liquid?
3. In the first laboratory activity of the year, students were assigned to find the total area of three tabletops in the room. To save time, each of the three students grabbed a ruler and measured the dimensions. They then calculated the area for each tabletop and added them together. Figure 6 presents the students' measurements. What is the total area of the three tabletops?

Student	Length	Width
A	127 cm	74 cm
B	1.30 m	0.80 m
C	50.0 in	29.5 in

Figure 6 Tabletop Dimensions

Note: Only numbers resulting from measurements made using instruments have significant figures. Exact numbers have an infinite number of significant figures. Exact numbers include numbers derived from counting or definition.

Examples: 25 desks in a room or 100 cm = 1 meter

SCIENTIFIC NOTATION

In chemistry we deal with very small and very large numbers. It is awkward to use many zeros to express very large or very small numbers, so scientific notation is used. The number is rewritten as the product of a number between 1 and 10 and an exponential term— 10^n , where n is a whole number.

Examples

1. **The distance between New York City and San Francisco = 4,741,000 meters:**

$$4,741,000 \text{ m} = (4.741 \times 1,000,000) \text{ m, or } 4.741 \times 10^6 \text{ m}$$

2. **The mass of 25 water molecules = 0.000 000 000 000 000 000 000 748 grams:**

$$0.000 \ 000 \ 000 \ 000 \ 000 \ 000 \ 000 \ 748 \text{ grams} = 7.48 \times 0.000 \ 000 \ 000 \ 000 \ 000 \ 000 \ 1 \text{ g} = 7.48 \times 10^{-22} \text{ grams}$$

It is easier to assess the magnitude and to perform operations with numbers written in scientific notation. It is also easier to identify the proper number of significant figures.

Addition/Subtraction Using Scientific Notation

1. Convert the numbers to the same power of ten.
2. Add (subtract) the nonexponential portion of the numbers.
3. The power of ten remains the same.

$$\text{Example: } 1.00 \times 10^4 + 2.30 \times 10^5$$

A good rule to follow is to express all numbers in the problem to the highest power of ten.

$$\text{Convert } 1.00 \times 10^4 \text{ to } 0.100 \times 10^5.$$

$$0.100 \times 10^5 + 2.30 \times 10^5 = 2.40 \times 10^5$$

Multiplication Using Scientific Notation

1. The numbers (including decimals) are multiplied.
2. The exponents are added.
3. The answer is converted to proper scientific notation—the product of a number between 1 and 10 and an exponential term.

$$\text{Example: } (4.24 \times 10^2) \times (5.78 \times 10^4)$$

$$(4.24 \times 5.78) \times (10^{2+4}) = 24.5 \times 10^6$$

$$\text{Convert to proper scientific notation} = 2.45 \times 10^7$$

Division Using Scientific Notation

1. Divide the decimal parts of the number.
2. Subtract the exponents.
3. Express the answer in scientific notation.

Example: $(3.78 \times 10^5) \div (6.2 \times 10^8)$

$$(3.78 \div 6.2) \times (10^{5-8}) = 0.61 \times 10^{-3}$$

$$\text{Convert to proper scientific notation} = 6.1 \times 10^{-4}$$

Practice Problems

1. Convert the following numbers to exponential notation.
 - a. 0.0000369
 - b. $\frac{36}{1000}$
 - c. 0.0452
 - d. 4 520 000
 - e. 365 000
2. Carry out the following operations:
 - a. $(1.62 \times 10^3) + (3.4 \times 10^2)$
 - b. $(1.75 \times 10^{-1}) - (4.6 \times 10^{-2})$
 - c. $(15.1 \times 10^2) \times (3.2 \times 10^{-2})$
 - d. $(6.02 \times 10^{23}) \times (2.0 \times 10^2)$
 - e. $(6.02 \times 10^{23}) \div (12.0)$