In scientific work, we recognize two kinds of numbers: exact numbers (those whose values are known exactly) and inexact numbers (those whose values have some uncertainty). Exact numbers are those that have defined values or integers that result from counting. For example, there are exactly 3 ft in a yard, exactly 1000 g in a kilogram, and exactly 12 eggs in a dozen. By definition, there are exactly 2.54 cm in an inch. The number 1 in any conversion factor between units, as in 1 m = 1.0936 yd, is also an exact number.

Numbers obtained by measurement are always inexact. There are always built-in errors in the equipment used to measure quantities (equipment errors), and there are differences in how different people make the same measurement (human errors). Suppose that ten students with ten different balances are given the same dime to weigh. The ten measurements will vary slightly. The balances might be calibrated slightly differently, and there might be differences in how each student reads the mass from the balance. Remember: Uncertainties always exist in measured quantities.

**Precision and Accuracy**

Two terms are usually used in discussing the uncertainties in measured values: precision and accuracy. Precision is a measure of how closely individual measurements agree with each other. Accuracy refers to how closely individual measurements agree with the correct, or “true,” value. In general, the more precise a measurement, the more accurate it will be. We gain confidence in the accuracy of a measurement if we obtain nearly the same value in many different experiments. Thus, in the laboratory, you will often perform several different “trials” of the same experiment. It is possible, however, for a precise value to be inaccurate. If a very sensitive balance is poorly calibrated, for example, the masses measured will be precise but inaccurate.

**Significant Figures**

Suppose you weigh a dime on a balance capable of measuring to the nearest 0.0001 g. You could report the mass as 2.2405 ± 0.0001 g. The ± notation (read “plus or minus 0.0001”) is a useful way to express the uncertainty of a measurement. In much scientific work, we drop the ± notation with the understanding that an uncertainty of at least one unit exists in the last digit of the measured quantity. That is, measured quantities are generally reported in such a way that only the last digit is uncertain. All the digits, including the uncertain one, are called significant figures. The number 2.2405 has five significant figures. The number of significant figures indicates the preciseness of a measurement.

**SAMPLE EXERCISE 1.6**

What is the difference between 4.0 g and 4.00 g?

**Solution:** Many people would say there is no difference, but a scientist would note the difference in the number of significant figures in the two measurements. The value 4.0 has two significant figures, while 4.00 has three. This implies that the second measurement is more precise. A mass of 4.0 g indicates that the mass is between 3.9 and 4.1 g; the mass is 4.0 ± 0.1 g. A measurement of 4.00 g implies that the mass is between 3.99 and 4.01 g; the mass is 4.00 ± 0.01 g.

**PRACTICE EXERCISE**

A balance has a precision of ±0.001 g. A sample that weighs about 25 g is weighed on this balance. How many significant figures should be reported for this measurement? **Answer:** 5

**Learning Goal 8:** Determine the number of significant figures in a measured quantity.
The following rules apply to determining the number of significant figures in a measured quantity:

1. All nonzero digits are significant—457 cm (three significant figures); 0.25 g (two significant figures).
2. Zeros between nonzero digits are significant—1005 kg (four significant figures); 1.03 cm (three significant figures).
3. Zeros to the left of the first nonzero digit in a number are not significant; they merely indicate the position of the decimal point—0.02 g (one significant figure); 0.0026 cm (two significant figures).
4. Zeros that fall both at the end of a number and to the right of the decimal point are significant—0.0200 g (three significant figures); 3.0 cm (two significant figures).
5. When a number ends in zeros that are not to the right of a decimal point, the zeros are not necessarily significant—130 cm (two or three significant figures); 10,300 g (three, four, or five significant figures). We describe how to remove this ambiguity below.

The use of exponential notation (Appendix A) avoids the potential ambiguity of whether the zeros at the end of a number are significant (rule 5). For example, a mass of 10,300 g can be written in exponential notation showing three, four, or five significant figures:

\[ 1.03 \times 10^4 \text{ g} \quad (\text{three significant figures}) \]
\[ 1.030 \times 10^4 \text{ g} \quad (\text{four significant figures}) \]
\[ 1.0300 \times 10^4 \text{ g} \quad (\text{five significant figures}) \]

In these numbers all the zeros to the right of the decimal point are significant (rules 2 and 4). (All significant figures come before the exponent; the exponential term does not add to the number of significant figures.)

Exact numbers can be treated as if they have an infinite number of significant figures. This rule applies to many definitions between units. Thus, when we say, “There are 12 inches in 1 foot,” the number 12 is exact, and we need not worry about the number of significant figures in it.

**SAMPLE EXERCISE 1.7**

How many significant figures are in each of the following numbers (assume that each number is a measured quantity): (a) 4.003; (b) 6.023 \times 10^{13}; (c) 5000?

**Solution:** (a) Four; the zeros are significant figures. (b) Four; the exponential term does not add to the number of significant figures. (c) One, two, three, or four. In this case, the ambiguity could have been avoided by using standard exponential notation. Thus \( 5 \times 10^3 \) has only one significant figure; \( 5.00 \times 10^3 \) has three.

**PRACTICE EXERCISE**

How many significant figures are in each of the following measurements:
(a) 3.549 g \hspace{1cm} (b) 2.3 \times 10^4 \text{ cm} \hspace{1cm} (c) 0.00134 \text{ m}^3?

**Answers:** (a) four \hspace{1cm} (b) two \hspace{1cm} (c) three

**Significant Figures in Calculations**

In carrying measured quantities through calculations, observe this point: The precision of the result is limited by the least precise measurement. *In multiplica-
tion and division the result must be reported as having no more significant figures than the measurement with the fewest significant figures. When the result contains more than the correct number of significant figures, it must be rounded off.

For example, the area of a rectangle whose edge lengths are 6.221 cm and 5.2 cm should be reported as 32 cm²:

\[
\text{Area} = (6.221 \text{ cm})(5.2 \text{ cm}) = 32.3492 \text{ cm}^2 \rightarrow \text{round off to 32 cm}^2
\]

We round off to two significant figures because 5.2 cm has only two significant figures.

In rounding off numbers, the following rules are followed (each example is rounded to two digits):

1. If the leftmost digit to be removed is more than 5, the preceding number is increased by 1; 2.376 rounds to 2.4.
2. If the leftmost digit to be removed is less than 5, the preceding number is left unchanged; 7.248 rounds to 7.2.
3. If the leftmost digit to be removed is 5, the preceding number is not changed if it is even and is increased by 1 if it is odd; 2.25 rounds to 2.2; 4.35 rounds to 4.4.

The rule used to determine the number of significant figures in multiplication and division cannot be used for addition and subtraction. For these operations, the result should be reported to the same number of decimal places as that of the term with the least number of decimal places. In the following example the uncertain digits appear in color:

\[
\begin{align*}
\text{This number limits} & \quad 20.4 \quad \leftarrow \text{one decimal place} \\
\text{the number of significant} & \quad 1.322 \quad \leftarrow \text{three decimal places} \\
\text{figures in the result} & \quad 8.3 \quad \leftarrow \text{zero decimal places} \\
& \quad 104.722 \quad \rightarrow \text{round off to 105} \\
& \quad \text{(one uncertain digit)}
\end{align*}
\]

**SAMPLE EXERCISE 1.8**

A person’s height is measured to be 67.50 in. What is this height in centimeters?

**Solution:** There are 2.54 cm in an inch; this is an exact number and can be treated as if it had an infinite number of significant figures. The precision of the answer is thus limited by the measurement in inches and should be reported to four significant figures (67.50 has four significant figures). The answer is

\[
(67.50 \text{ in.})(\frac{2.54 \text{ cm}}{\text{in.}}) = 171.45 \text{ cm, which is rounded to 171.4 cm}
\]

**PRACTICE EXERCISE**

There are 1609.4 m in a mile. How many meters are in a distance of 1.35 mi?

**Answer:** \(2.17 \times 10^3\) m

**SAMPLE EXERCISE 1.9**

A gas at 25°C exactly fills a container previously determined to have a volume of \(1.05 \times 10^4\) cm³. The container plus gas are weighed and found to have a mass of 837.6 g. The container, when emptied of all gas, has a mass of 836.2 g. What is the density of the gas at 25°C?
Solution: The mass of the gas is just the difference in the two masses: \((837.6 - 836.2) \text{ g} = 1.4 \text{ g}\). Notice that 1.4 g has only two significant figures, even though the masses from which it is obtained have four.

From the definition of density we have

\[
\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{1.4 \text{ g}}{1.05 \times 10^3 \text{ cm}^3} = 0.0013 \text{ g/cm}^3
\]

There are two significant figures in this quantity, corresponding to the smaller number of significant figures in the two numbers that form the ratio.

**PRACTICE EXERCISE**

To how many significant figures should the container be weighed (with and without the gas) in Sample Exercise 1.9 in order for the density to be calculated to three significant figures? 

*Answer: 5*

It is important to have a feeling for significant figures when you use a calculator, because calculators ordinarily display more digits than are significant. For example, a typical calculator would give \(1.333333 \times 10^{-3}\) as the answer to the calculation in Sample Exercise 1.9. This result must be rounded off because of the uncertainties in the measured quantities used in the calculation.

When a calculation involves two or more steps, retain at least one additional digit—past the number of significant figures—for intermediate answers. This procedure ensures that small errors from rounding at each step do not combine to affect the final result. In using a calculator, you may enter the numbers one after another, rounding only the final answer. Accumulated round-off errors often account for small differences between results you obtain and answers given in the text for numerical problems.

Before we continue, perhaps a word of caution is in order. To have a real knowledge of chemistry, you must understand both the basic concepts and how to use these concepts to solve problems. Your instructor and this text will show you how to solve different types of problems, many of which involve numbers. But is that enough? Is watching someone else solve problems an effective way for you to learn how to solve problems yourself (as you will have to do on exams)? The answer is no.

One of our goals in this text (and, we hope, one of your goals in studying chemistry) is to make you comfortable solving chemical problems. Achieving this goal requires active practice in problem solving through the use of homework problems and the like. Throughout the book, we have provided sample exercises in which the solutions are shown in detail. A practice exercise, for which only the answer is given, accompanies each sample exercise; it is important that you use these exercises as learning aids. End-of-chapter exercises provide additional questions to help you understand the material in the chapter. Colored numbers indicate exercises whose answers are given at the back of the book. A review of basic mathematics is given in Appendix A.

Throughout the text, we use an approach called **dimensional analysis** as an aid in problem solving. In dimensional analysis, we carry units through all calculations. Units are multiplied together, divided into each other, or "canceled." Dimensional analysis will help ensure that the solutions to problems yield the proper units. More important, though, dimensional analysis provides a means by which problems can be solved.
The key to using dimensional analysis is the correct use of conversion factors to change one unit into another. A conversion factor is a fraction whose numerator and denominator are different units of the same quantity. Consider the statement, "There are 2.54 centimeters per inch," or the equivalent equation, \(2.54 \text{ cm} = 1 \text{ in.}\). This relationship allows us to write two conversion factors:

\[
\begin{align*}
\frac{2.54 \text{ cm}}{1 \text{ in.}} & \quad \frac{1 \text{ in.}}{2.54 \text{ cm}}
\end{align*}
\]

The first of these factors is used when we want to convert inches to centimeters. For example, the length in centimeters of an object that is 8.50 in. long is given by

\[
\text{Number of centimeters} = (8.50 \text{ in.}) \left(\frac{2.54 \text{ cm}}{1 \text{ in.}}\right) = 21.6 \text{ cm}
\]

Note that the units of inches in the denominator of the conversion factor cancel the units of inches in the data we were given (8.50 inches). The centimeters in the numerator of the conversion factor become the units of the final answer. In general, the units multiply and divide as follows:

\[
\text{Given unit} \times \frac{\text{desired unit}}{\text{given unit}} = \text{desired unit}
\]

**SAMPLE EXERCISE 1.10**

A man weighs 185 lb. What is his mass in grams?

**Solution:** From the back inside cover we have 1 lb = 453.6 g. In order to cancel pounds and leave grams, we will use the conversion factor with grams in the numerator and pounds in the denominator:

\[
\text{Mass in grams} = (185 \text{ lb}) \left(\frac{453.6 \text{ g}}{1 \text{ lb}}\right) = 8.39 \times 10^4 \text{ g}
\]

Note that the answer can be given only to three significant figures.

**PRACTICE EXERCISE**

By using a conversion factor from the back inside cover, determine the length in kilometers of a 500.0-mi automobile race. **Answer:** 804.7 km

We can use more than one conversion factor in the solution of a problem. For example, suppose we want to know the length in inches of an 8.00-m rod. The table at the end of the text doesn't give the relationship between meters and inches. It does give the relationship between centimeters and inches, though, and we know that 1 m = 100 cm. Thus, we can convert first from meters to centimeters, and then from centimeters to inches:

\[
= (\text{length in m})(\text{factor converting m} \to \text{cm})(\text{factor converting cm} \to \text{in.})
\]
The relationship between meters and centimeters gives us the first conversion factor. Because we need to cancel meters, we write meters in the denominator:

\[
\frac{100\, \text{cm}}{1\, \text{m}}
\]

We then use the relationship \(2.54\, \text{cm} = 1\, \text{in.}\) to write the second conversion factor with the desired units, inches, in the numerator:

\[
\frac{1\, \text{in.}}{2.54\, \text{cm}}
\]

Thus, we have

\[
\text{Number of inches} = (8.00\, \text{m}) \left(\frac{100\, \text{cm}^3}{1\, \text{m}^3}\right) \left(\frac{1\, \text{in.}}{2.54\, \text{cm}}\right) = 315\, \text{in.}
\]

The conversion factors above convert from length to length or, more generally, from one unit of a given measure to another unit of the same measure. We also have conversion factors that convert from one measure to a different one. The density of a substance, for example, is a conversion factor between mass and volume. Suppose that we want to know the mass in grams of a cubic inch of gold, which has a density of 19.3 g/cm\(^3\). The density gives us the following conversion factors:

\[
\frac{19.3\, \text{g}}{1\, \text{cm}^3} = \frac{1\, \text{cm}^3}{19.3\, \text{g}}
\]

Because the answer we want is a mass in grams, we can see that we will use the first of these factors, which has mass in grams in the numerator. To use this factor, however, we must first relate cubic centimeters to cubic inches. We know the conversion factor from inches to centimeters, and the cube of this gives us the desired conversion factor:

\[
\left(\frac{2.54\, \text{cm}}{1\, \text{in.}}\right)^3 = \frac{(2.54)^3\, \text{cm}^3}{(1\, \text{in.})^3} = \frac{16.4\, \text{cm}^3}{1\, \text{in.}^3}
\]

Notice that both the numbers and the units are cubed. Applying our conversion factors, we can now solve the problem:

\[
\text{Mass in grams} = (1\, \text{in.}^3) \left(\frac{16.4\, \text{cm}^3}{1\, \text{in.}^3}\right) \left(\frac{19.3\, \text{g}}{1\, \text{cm}^3}\right) = 317\, \text{g}
\]

**Summary of Dimensional Analysis**

In using dimensional analysis to solve problems, we will always ask three questions:

1. **What data are we given in the problem?**
2. **What quantity do we wish to obtain in the problem?**
3. **What conversion factors do we have available to take us from the given quantity to the desired one?**
When solving numerical problems, always ask yourself whether your answer makes sense. \textit{(Calvin and Hobbes. Copyright 1986 Universal Press Syndicate. Reprinted with permission. All rights reserved.)}

If you carry the units through during your calculations, you will always know whether you are using the correct conversion factors. Finally, whenever you finish a calculation, look at the numerical value of your answer (as well as the units) and ask yourself whether your answer makes any sense.

**SAMPLE EXERCISE 1.11**

A certain printed page has an average of 25 words per square inch of paper. The average length of the words is 5.3 letters. What is the average number of letters per square centimeter of paper?

**Solution:** We are given the “print density” in the units of words per square inch. We want to convert this to a print density in the units of letters per square centimeter. We are given the relation 1 word = 5.3 letters and, from the back inside cover, 1 in. = 2.54 cm. The solution is given by

\[
\text{Number of letters/cm}^2 = \left(\frac{25 \text{ words}}{1 \text{ in.}^2}\right) \left(\frac{5.3 \text{ letters}}{1 \text{ word}}\right) \left(\frac{1 \text{ in.}^2}{2.54 \text{ cm}}\right)^2
\]

\[
= 21 \text{ letters/cm}^2
\]

**PRACTICE EXERCISE**

In a certain part of the country, there is an average of 710 people per square mile and 0.72 telephones per person. What is the average number of telephones in an area of 5.0 km$^2$? \textit{Answer:} 990 telephones

**SAMPLE EXERCISE 1.12**

What is the mass in grams of 1.00 gal of water? The density of water is 1.00 g/mL.

**Solution:** Following the procedure summarized above, we note the following:

1. We are given 1.00 gal of water.
2. We wish to obtain the mass in grams.
3. We have the following conversion factors either given, commonly known, or available on the back inside cover of the text:

<table>
<thead>
<tr>
<th>1.00 g water</th>
<th>1 L</th>
<th>1 L</th>
<th>1 gal</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mL water</td>
<td>1000 mL</td>
<td>1.057 qt</td>
<td>4 qt</td>
</tr>
</tbody>
</table>
We realize that the first of these conversion factors must be used as written (with grams in the numerator) to give the desired result. We also realize that the last conversion factor must be “turned over” in order to cancel gallons. The solution is given by

\[
\text{Mass in grams} = (1.00 \text{ gal of water})(4 \text{ qt})(1 \text{ L})(1000 \text{ mL})(1.00 \text{ g})
\]

\[
= 378 \times 10^3 \text{ g water}
\]

**PRACTICE EXERCISE**

A car travels 28 mi/gal of gasoline. How many kilometers per liter will it go?

*Answer:* 12 km/L

---

**SUMMARY**

Chemistry is the study of the properties, composition, and changes of matter. Matter exists in three states: gas, liquid, and solid. Most matter consists of a mixture of substances. Mixtures can be either homogeneous or heterogeneous; homogeneous mixtures are called solutions. Mixtures can be separated into two types of pure substances: elements and compounds. Each substance has a unique set of physical and chemical properties that can be used to identify it. Matter can undergo physical changes and chemical changes (chemical reactions).

Measurements in chemistry are made using the metric system. Special emphasis is placed on a particular set of metric units called SI units, which are based on the meter, kilogram, and second as the basic units of length, mass, and time, respectively. The metric system employs a set of prefixes to indicate decimal fractions or multiples of the base units.

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**KEY TERMS**

- matter (Sec. 1.1)
- gas (Sec. 1.1)
- liquid (Sec. 1.1)
- solid (Sec. 1.1)
- pure substance (Sec. 1.1)
- physical properties (Sec. 1.1)
- chemical properties (Sec. 1.1)
- physical changes (Sec. 1.1)
- changes of state (Sec. 1.1)
- chemical changes (Sec. 1.1)
- chemical reactions (Sec. 1.1)
- mixtures (Sec. 1.1)
- solution (Sec. 1.1)
- elements (Sec. 1.2)
- compounds (Sec. 1.2)
- law of constant composition (Sec. 1.2)
- law of definite proportions (Sec. 1.2)
- metric system (Sec. 1.3)
- SI units (Sec. 1.3)
- mass (Sec. 1.3)
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