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If a thermometer is inaccurate, you cannot accurately determine your temperature with it. If the thermometer always reads 1.0 °F low and you measure your temperature as 98.6 °F, then your temperature is actually 99.6 °F. The measured value differs from the true value by 1.0 °F and is inaccurate by that amount. Note that the accuracy of the measured value cannot be improved by repeated measurements. Repeated measurements can only increase the precision of the measured value.

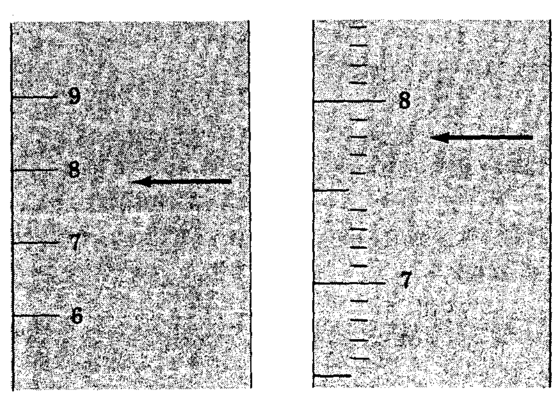
2.3 Significant Figures Give the Uncertainty in Measurements

If you count the number of students in a class you may determine that there are 26 students. There is no uncertainty in the number 26; it is an exact number. There also is no uncertainty in the number of feet in 1 yd; it is exactly 3 ft. **Exact numbers** arise either from a direct count or from a defined equivalence such as 1 yd = 3 ft.

Scientists use measured numbers, which are never exact. Perfect measurements are not possible; there is always an uncertainty due to limitations either in the measuring tool or in how the measurement is made. When measured quantities are reported, a scientist must indicate their reliability. The number of digits reported in a measured number that give reliable information is the number of significant figures. The **significant figures** in a measured number include all the numbers known with certainty plus the next digit to the right, which is estimated.

What parts of a measurement are certain, and what parts are estimated? Consider the measurement of the mass of an object obtained by using a balance. The indicator arrow is between 7 and 8 grams (g) but is closer to 8 g (Figure 2.1a). Divide the space between the marked values into 10 parts. Then estimate that the indicator arrow is $\frac{8}{10}$ of the way toward 8 g. The mass is 7.8 g; the quantity has one certain figure and one estimated figure, or a total of two significant figures. If the object is placed on a more accurate balance (Figure 2.1b), the mass is more easily seen to be 7.8 g. However, the mass can now be estimated to the nearest hundredth of a gram as 7.82 g. In this case, there are three significant figures and the 2 is the estimated figure.

The reliability of any measurement is given by the number of significant figures. Whenever you make a measurement, it is necessary to make sure that the recorded quantity has the number of significant figures that reflects the reliability of the measurement.



(a) (b)

Figure 2.1 Significant Figures in Determining a Mass
In case (a) the scale of the balance gives the mass in grams only to the nearest gram. The mass is somewhat less than 8 g and is clearly more than 7 g. You should report the mass as 7.8 g by estimating to the nearest 0.1 g. This value of the mass has two significant figures. In case (b) the balance gives greater accuracy. The mass is slightly more than 7.8 g and clearly less than 7.9 g. You should report the mass as 7.82 g, a quantity with three significant figures.

Is Zero a Significant Figure?

With the exception of zero, all digits in a number that is exact or in a number properly expressing the accuracy of a measurement are significant. However, zeros in a quantity may or may not be significant figures. Zeros in measured quantities are divided into three classes.

1. **Leading zeros** are zeros that precede all of the nonzero digits.
2. **Captive zeros** are zeros between nonzero digits.
3. **Trailing zeros** are zeros that are to the right of the last nonzero digit.

Suppose that different samples are weighed on a balance that gives the mass accurately to 0.001 g. If one sample has a mass of 0.015 g, the leading zeros before and after the decimal point are not significant figures. The sample has a mass between 0.014 and 0.016 g, and the number 0.015 contains only two significant figures.

Leading zeros are not significant figures.

0.015 g

If the mass of a second sample is 2.004 g, the captive zeros that are "sandwiched" between the 2 and the 4 are known with certainty and are significant figures. There are four significant figures in the number, which represents a mass known to be between 2.003 and 2.005 g.

Captive zeros are significant figures.

2.004 g

If the mass of a third sample is found to be 5.000 g, the sample is between 4.999 and 5.001 g. Thus, the three trailing zeros following the decimal point are significant figures, and the number of significant figures is four.

Trailing zeros are significant figures.

5.000 g

Example 2.1 What is the number of significant figures in each of the following numbers?

(a) 5041 (b) 0.05401 (c) 0.5401 (d) 0.5410

SOLUTION There are four significant figures in each case. The zeros that are not significant are underlined below.

(a) 5041 (b) 0.05401 (c) 0.5401 (d) 0.5410

The captive zeros in (a), (b), and (c) are significant figures. The trailing zero in (d) is a significant figure.

Problem 2.1 What is the number of significant figures in each of the numbers 0.04010, 0.00410, and 0.000401?

Significant Figures in Calculated Quantities

What happens when you use measured quantities in calculations? How many significant figures can be retained in the calculated quantity? The number of significant figures depends on the number of significant figures in each measured quantity and the type of math operation used. The rules governing the use of significant figures in mathematical operations as well as the procedures for rounding off numbers are discussed in Appendix A. Sample calculations are also provided. Reference to this appendix will be given in later chapters when it will be necessary for you to use these rules to ensure that calculated quantities are not expressed with any more reliability than the least reliable quantity used. Simply stated, the rules are

1. When numbers are added or subtracted, the answer must not contain any significant figures beyond the place value common to all of the numbers.
2. In multiplication or division of measured numbers, the answer must not contain more significant figures than the least number of significant figures in the measurements.

2.4 Expressing Large and Small Numbers

You often read about large numbers such as the number of starving people in the world or the size of the national debt, but in everyday life you seldom encounter very small numbers. Both very large and very small numbers are used in chemistry. The large numbers have a long string of trailing zeros that are not significant figures. Similarly, the small numbers have a string of leading zeros that are not significant figures. Scientists have found it convenient to express such numbers using powers of 10.

In **scientific notation,** a number is expressed as a product of a coefficient multiplied by a power of 10. The **coefficient** is a number equal to or greater than 1 but less than 10. The power of 10 is an **exponent.**

exponent

$$a \times 10^b$$

coefficient

The coefficient retains only the significant figures of the original number; the exponent gives the location of the decimal point. For example, 18.0160 g of water contains 6.02217×10^{23} molecules of water. The mass of a water molecule is 2.99161×10^{-23} g. Each number has six significant figures.

$$6.02217 \times 10^{23} \text{ means } 602{,}217{,}000{,}000{,}000{,}000{,}000{,}000$$

$$2.99161 \times 10^{-23} \text{ means } 0.000\ 000\ 000\ 000\ 000\ 000\ 000\ 029\ 9161$$

The positive exponent for the number of molecules indicates that the quantity is very large; the negative exponent for the weight of one molecule indicates that the quantity is very small. Because each quantity is expressed to six significant figures, the zeros that are not significant figures are eliminated by using scientific notation.

If you are not comfortable using scientific notation, Appendix A provides some examples for you to work. However, most calculators express small and large numbers using scientific notation and you will not have to "move" decimal points to write the proper coefficient and exponent.