<table>
<thead>
<tr>
<th>Appendix A</th>
<th>Chemistry Skill Handbook</th>
<th>785</th>
</tr>
</thead>
<tbody>
<tr>
<td>Measurement in Science</td>
<td>785</td>
<td></td>
</tr>
<tr>
<td>The International System of Units</td>
<td>785</td>
<td></td>
</tr>
<tr>
<td>Other Useful Measurements</td>
<td>786</td>
<td></td>
</tr>
<tr>
<td>SI Prefixes</td>
<td>786</td>
<td></td>
</tr>
<tr>
<td>Relating SI, Metric, and English Measurements</td>
<td>788</td>
<td></td>
</tr>
<tr>
<td>Making and Interpreting Measurements</td>
<td>791</td>
<td></td>
</tr>
<tr>
<td>Expressing the Accuracy of Measurements</td>
<td>793</td>
<td></td>
</tr>
<tr>
<td>Expressing Quantities with Scientific Notation</td>
<td>795</td>
<td></td>
</tr>
<tr>
<td>Computations with the Calculator</td>
<td>799</td>
<td></td>
</tr>
<tr>
<td>Using the Factor Label Method</td>
<td>801</td>
<td></td>
</tr>
<tr>
<td>Organizing Information</td>
<td>804</td>
<td></td>
</tr>
<tr>
<td>Making and Using Tables</td>
<td>804</td>
<td></td>
</tr>
<tr>
<td>Making and Using Graphs</td>
<td>805</td>
<td></td>
</tr>
</tbody>
</table>

| Appendix B | Supplemental Practice Problems | 809 |

| Appendix C | Safety Handbook | 839 |
| Safety Guidelines in the Chemistry Laboratory | 839 |
| First Aid in the Laboratory | 839 |
| Safety Symbols | 840 |

| Appendix D | Chemistry Data Handbook | 841 |
| Table D.1 Symbols and Abbreviations | 841 |
| Table D.2 The Modern Periodic Table | 842 |
| Table D.3 Alphabetical Table of the Elements | 844 |
| Table D.4 Properties of Elements | 845 |
| Table D.5 Electron Configurations of the Elements | 848 |
| Table D.6 Useful Physical Constants | 850 |
| Table D.7 Names and Charges of Polyatomic Ions | 850 |
| Table D.8 Solubility Guidelines | 851 |
| Table D.9 Solubility Product Constants | 851 |
| Table D.10 Acid-Base Indicators | 852 |

| Appendix E | Answers to In-Chapter Practice Problems | 853 |
Measurement in Science

It’s easier to determine if a runner wins a race than to determine if the runner broke a world’s record for the race. The first determination requires that you sequence the runners passing the finish line—first, second, third. . . . The second determination requires that you carefully measure and compare the amount of time that passed between the start and finish of the race for each contestant. Because time can be expressed as an amount made by measuring, it is called a quantity. One second, three minutes, and two hours are quantities of time. Other familiar quantities include length, volume, and mass.

The International System of Units

In 1960, the metric system was standardized in the form of Le Système International d’Unités (SI), which is French for the “International System of Units.” These SI units were accepted by the international scientific community as the system for measuring all quantities.

SI Base Units  The foundation of SI is seven independent quantities and their SI base units, which are listed in Table A.1.

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Unit</th>
<th>Unit Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Length</td>
<td>meter</td>
<td>m</td>
</tr>
<tr>
<td>Mass</td>
<td>kilogram</td>
<td>kg</td>
</tr>
<tr>
<td>Time</td>
<td>second</td>
<td>s</td>
</tr>
<tr>
<td>Temperature</td>
<td>kelvin</td>
<td>K</td>
</tr>
<tr>
<td>Amount of substance</td>
<td>mole</td>
<td>mol</td>
</tr>
<tr>
<td>Electric current</td>
<td>ampere</td>
<td>A</td>
</tr>
<tr>
<td>Luminous intensity</td>
<td>candela</td>
<td>cd</td>
</tr>
</tbody>
</table>

SI Derived Units  You can see that quantities such as area and volume are missing from the table. The quantities are omitted because they are derived—that is, computed—from one or more of the SI base units. For example, the unit of area is computed from the product of two perpendicular length units. Because the SI base unit of length is the meter, the SI derived unit of area is the square meter, m². Similarly, the unit of volume is derived from three mutually perpendicular length units, each represented by the meter. Therefore, the SI derived unit of volume is the cubic meter, m³. The SI derived units used in this text are listed in Table A.2.
Table A.2 SI Derived Units

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Unit</th>
<th>Unit Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Area</td>
<td>square meter</td>
<td>m$^2$</td>
</tr>
<tr>
<td>Volume</td>
<td>cubic meter</td>
<td>m$^3$</td>
</tr>
<tr>
<td>Mass density</td>
<td>kilogram per cubic meter</td>
<td>kg/m$^3$</td>
</tr>
<tr>
<td>Energy</td>
<td>joule</td>
<td>J</td>
</tr>
<tr>
<td>Heat of fusion</td>
<td>joule per kilogram</td>
<td>J/kg</td>
</tr>
<tr>
<td>Heat of vaporization</td>
<td>joule per kilogram</td>
<td>J/kg</td>
</tr>
<tr>
<td>Specific heat</td>
<td>joule per kilogram-kelvin</td>
<td>J/kg•K</td>
</tr>
<tr>
<td>Pressure</td>
<td>pascal</td>
<td>Pa</td>
</tr>
<tr>
<td>Electric potential</td>
<td>volt</td>
<td>V</td>
</tr>
<tr>
<td>Amount of radiation</td>
<td>gray</td>
<td>Gy</td>
</tr>
<tr>
<td>Absorbed dose of radiation</td>
<td>sievert</td>
<td>Sv</td>
</tr>
</tbody>
</table>

Other Useful Measurements

Metric Units As previously noted, the metric system is a forerunner of SI. In the metric system, as in SI, units of the same quantity are related to each other by orders of magnitude. However, some derived quantities in the metric system have units that differ from those in SI. Because these units are familiar and equipment is often calibrated in these units, they are still used today. Table A.3 lists several metric units that you might use.

Table A.3 Metric Units

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Unit</th>
<th>Unit Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume</td>
<td>liter (0.001 m$^3$)</td>
<td>L</td>
</tr>
<tr>
<td>Temperature</td>
<td>Celsius degree</td>
<td>°C</td>
</tr>
<tr>
<td>Specific heat</td>
<td>joule per kilogram-degree</td>
<td>J/kg•°C</td>
</tr>
<tr>
<td></td>
<td>degree Celsius</td>
<td></td>
</tr>
<tr>
<td>Pressure</td>
<td>millimeter of mercury</td>
<td>mm Hg</td>
</tr>
<tr>
<td>Energy</td>
<td>calorie</td>
<td>cal</td>
</tr>
</tbody>
</table>

SI Prefixes

When you express a quantity, such as ten meters, you are comparing the distance to the length of one meter. Ten meters indicates that the distance is a length ten times as great as the length of one meter. Even though you can express any quantity in terms of the base unit, it may not be convenient. For example, the distance between two towns might be 25 000 m. Here, the meter seems too small to describe that distance. Just as you would use 16 miles, not 82 000 feet, to express that distance, you would use a larger unit of length, the kilometer, km. Because the kilometer represents a length of 1000 m, the distance between the towns is 25 km.
In SI, units that represent the same quantity are related to each other by some factor of ten such as 10, 100, 1000, 1/10, 1/100, and 1/1000. In the example above, the kilometer is related to the meter by a factor of 1000; namely, 1 km = 1000 m. As you see, 25 000 m and 25 km differ only in zeros and the units.

To change the size of the SI unit used to measure a quantity, you add a prefix to the base unit or derived unit of that quantity. For example, the prefix *centi-* designates one one-hundredth (0.01). Therefore, a centimeter, cm, is a unit one one-hundredth the length of a meter and a centijoule, cJ, is a unit of energy one one-hundredth that of a joule. The exception to the rule is in the measurement of mass in which the base unit, kg, already has a prefix. To express a different size mass unit, you replace the prefix to the gram unit. Thus, a centigram, cg, represents a unit having one one-hundredth the mass of a gram. Table A.4 lists the most commonly used SI prefixes.

### Table A.4  SI Prefixes

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Meaning</th>
<th>Numerical value</th>
<th>Multiplier</th>
<th>Expressed as scientific notation</th>
</tr>
</thead>
</table>
| Greater than 1
| giga-  | G      | billion      | 1 000 000 000   | 1 × 10⁹    |
| mega-  | M      | million      | 1 000 000      | 1 × 10⁶    |
| kilo-  | k      | thousand     | 1 000          | 1 × 10³    |
| Less than 1
| deci-  | d      | tenth        | 0.1            | 1 × 10⁻¹   |
| centi- | c      | hundredth    | 0.01           | 1 × 10⁻²   |
| milli- | m      | thousandth   | 0.001          | 1 × 10⁻³   |
| micro- | µ      | millionth    | 0.000 001      | 1 × 10⁻⁶   |
| nano-  | n      | billionth    | 0.000 000 001  | 1 × 10⁻⁹   |
| pico-  | p      | trillionth   | 0.000 000 000 001 | 1 × 10⁻¹² |

### Practice Problems

**Use Tables A.1, A.2, and A.3 to answer the following questions.**

1. Name the following quantities using SI prefixes. Then write the symbols for each.
   a) 0.1 m
d) 0.000 000 001 m  
b) 1 000 000 000 J
e) 10⁻³ g  
c) 10⁻¹² m
   f) 10⁶ J

2. For each of the following, identify the quantity being expressed and rank the units in increasing order of size.
   a) cm, µm, dm
d) pg, cg, mg  
b) Pa, MPa, kPa
e) mA, MA, µA  
c) kV, eV, V
   f) dGy, mGy, nGy
Relating SI, Metric, and English Measurements

Any measurement tells you how much because it is a statement of quantity. However, how you express the measurement depends on the purpose for which you are going to use the quantity. For example, if you look at a room and say its dimensions are 9 ft by 12 ft, you are estimating these measurements from past experience. As you become more familiar with SI, you will be able to estimate the room size as 3 m by 4 m. On the other hand, if you are going to buy carpeting for that room, you are going to make sure of its dimensions by measuring it with a tape measure no matter which system you are using.

Estimating and using any system of measurement requires familiarity with the names and sizes of the units and practice. As you will see in Figure 1, you are already familiar with the names and sizes of some common units, which will help you become familiar with SI and metric units.

**Figure 1**
Relating Measurements

**Length**
A paper clip and your hand are useful approximations for SI units. For approximating, a meter and a yard are similar lengths. To convert length measurements from one system to another, you can use the relationships shown here.

2.54 cm = 1.00 in.
1.00 m = 39.37 in.
Temperature
Having a body temperature of 37° or 310 sounds unhealthy if you don’t include the proper units. In fact, 37°C and 310 K are normal body temperatures in the metric system and SI, respectively.

Volume
For approximating, a liter and a quart are similar volumes. Kitchen measuring spoons are used in estimating small volumes.

1.00 dm³ = 1.0 liter = 1.06 qt. = 1000 mL
Mass and Weight  

One of the most useful ways of describing an amount of stuff is to state its mass. You measure the mass of an object on a balance. Even though you are measuring mass, most people still refer to it as weighing. You might think that mass and weight are the same quantity. They are not. The mass of an object is a measure of its inertia; that is, its resistance to changes in motion. The inertia of an object is determined by its quantity of mass. A pint of sand has more matter than a pint of water; therefore, it has more mass.

The weight of an object is the amount of gravitational force acting on the mass of the object. On Earth, you sense the weight of an object by holding it and feeling the pull of gravity on it.

An important aspect of weight is that it is directly proportional to the mass of the object. This relationship means that the weight of a 2-kg object is twice the weight of a 1-kg object. Therefore, the pull of gravity on a 2-kg object is twice as great as the pull on a 1-kg object. Because it’s easier to measure the effect of the pull of gravity rather than the resistance of an object to changes in its motion, mass can be determined by a weighing process.

Two instruments used to determine mass are the double-pan balance and a triple-beam balance. How each functions is illustrated in Figure 2. You can read how an electronic balance functions in How it Works in Chapter 12.
Making and Interpreting Measurements

Using measurements in science is different from manipulating numbers in math class. The important difference is that numbers in science are almost always measurements, which are made with instruments of varying accuracy. As you will see, the degree of accuracy of measured quantities must always be taken into account when expressing, multiplying, dividing, adding, or subtracting them. In making or interpreting a measurement, you should consider two points. The first is how well the instrument measures the quantity you’re interested in. This point is illustrated in Figure 4.
Because the bottom ruler in Figure 4 is calibrated to smaller divisions than the top ruler, the lower ruler has more precision than the upper. Any measurement you make with it will be more precise than one made on the top ruler because it will contain a smaller estimated value.

The second point to consider in making or interpreting a measurement is how well the measurement represents the quantity you’re interested in. Looking at Figure 4, you can see how well the edge of the paper strip aligns with the value 4.27 cm. From sight, you know that 4.27 cm is a better representation of the strip’s length than 4.3 cm. Because 4.27 cm better represents the length of the strip (the quantity you’re interested in) than does 4.3 cm, it is a more accurate measurement of the strip’s length.
In Figure 5, you will see that a more precise measurement might not be a more accurate measurement of a quantity.

**Figure 5**

**Precise and Accurate Measurements**

Describing the width of the index card as 10.16 cm indicates that the ruler has 0.1-cm calibrations and the 0.06 cm is an estimation. Similarly, the uniform alignment of the edge of the index card with the ruler indicates that the measurement 10.16 cm is also an accurate measurement of width.

Describing the width of a brick as 10.16 cm indicates that this measurement is as precise as the measurement of the index card. However, 10.16 cm isn’t a good representation of the width of the ragged- and jagged-edged brick.

A better representation of the width of the brick is made using a ruler with less precision. As you can see, 10.2 cm is a better representation, and therefore a more accurate measurement, of the brick’s width than 10.16 cm.

**Expressing the Accuracy of Measurements**

In measuring the length of the strip as 4.3 cm and 4.27 cm in Figure 4, you were aware of the difference in the calibration of the two rulers. This difference appeared in the way each measurement was recorded. In one measurement, the digits 4 and 3 were meaningful. In the second, the digits 4, 2, and 7 were meaningful. In any measurement, meaningful digits are called *significant digits*. The significant digits in a measured quantity include all the digits you know for sure, plus the final estimated digit.
**Significant Digits**  Several rules can help you express or interpret which digits of a measurement are significant. Notice how those rules apply to readings from a digital balance (left) and a graduated cylinder (right).

<table>
<thead>
<tr>
<th>Digital balance readout</th>
<th>Graduated cylinder</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>2</strong> 100.00 g</td>
<td><strong>3</strong> 10.0 mL</td>
</tr>
<tr>
<td><strong>8</strong> 10.00 g</td>
<td><strong>2</strong> (±1) 0.1 mL</td>
</tr>
<tr>
<td><strong>3</strong> 1.00 g</td>
<td><strong>0</strong></td>
</tr>
<tr>
<td><strong>4</strong> 0.10 g</td>
<td>32.2 mL</td>
</tr>
<tr>
<td><strong>7 (±1)</strong> 0.01 g</td>
<td></td>
</tr>
<tr>
<td>283.47 g</td>
<td>3 significant digits</td>
</tr>
<tr>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>5</strong> 10.00 g</td>
<td><strong>1</strong> 10.0 mL</td>
</tr>
<tr>
<td><strong>6</strong> 1.00 g</td>
<td><strong>0</strong></td>
</tr>
<tr>
<td><strong>0</strong> 0.10 g</td>
<td><strong>7 (±1)</strong> 0.1 mL</td>
</tr>
<tr>
<td><strong>6 (±1)</strong> 0.01 g</td>
<td><strong>10.7 mL</strong></td>
</tr>
<tr>
<td>56.06 g</td>
<td>3 significant digits</td>
</tr>
<tr>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>3</strong> 10.00 g</td>
<td><strong>2</strong> 10.0 mL</td>
</tr>
<tr>
<td><strong>3</strong> 1.00 g</td>
<td><strong>0</strong></td>
</tr>
<tr>
<td><strong>0</strong> 0.10 g</td>
<td><strong>0 (±1)</strong> 0.1 mL</td>
</tr>
<tr>
<td><strong>0 (±1)</strong> 0.01 g</td>
<td><strong>20.0 mL</strong></td>
</tr>
<tr>
<td>73.00 g</td>
<td>3 significant digits</td>
</tr>
<tr>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>1</strong> 0.09 g</td>
<td><strong>0</strong> 1.0 mL</td>
</tr>
<tr>
<td><strong>0</strong> 0.10 g</td>
<td><strong>7 (±1)</strong> 0.1 mL</td>
</tr>
<tr>
<td><strong>0 (±1)</strong> 0.01 g</td>
<td><strong>0.7 mL</strong></td>
</tr>
<tr>
<td>0.09 g</td>
<td>1 significant digit</td>
</tr>
</tbody>
</table>

The fourth rule sometimes causes difficulty in expressing such measurements as 20 L. Because the zero is a placeholder in the measurement, it is not significant and 20 L has one significant digit. You should interpret the measurement 20 L as 20 L plus or minus the value of the least significant digit, which is 10 L. Thus, a volume measurement of 20 L indicates 20 L ± 10 L, a range of 10-30 L. However, suppose you made the measurement with a device that is accurate to the nearest 1 L. You would want to indicate that both the 2 and the 0 are significant. How would you do this? You can’t just add a decimal point after the 20 because it could be mistaken for a
period. Adding a decimal point followed by a zero would indicate that the final zero past the decimal is significant (Rule 3), and the measurement would then be 20.0 ± 0.1 L. To solve the dilemma, you have to express 20 L as 2.0 × 10^1 L. Now the zero is a significant digit because it is a final zero past the decimal (Rule 3). The measurement has two significant digits and signifies (2.0 ± 0.1) × 10^1 L.

**Expressing Quantities with Scientific Notation**

The most common use of scientific notation is in expressing measurements of very large and very small dimensions. Using scientific notation is sometimes referred to as using powers of ten because it expresses quantities by using a small number between one and ten, which is then multiplied by ten to give the quantity its proper magnitude. Suppose you went on a trip of 9000 km. You know that 10^3 = 1000, so you could express the distance of your trip as 9 × 10^3 km. In this example, it may seem that scientific notation wouldn’t be terribly useful. However, consider that an often-used quantity in chemistry is 602 000 000 000 000 000 000, the number of atoms or molecules in a mole of a substance. Recall that the mole is the SI unit of amount of a substance. Rather than writing out this huge number every time it is used, it’s much easier to express it in scientific notation.

**Determining Powers of 10** To determine the exponent of ten, count as you move the decimal point left until it falls just after the first nonzero digit—in this case, 6. If you try this on the number above, you’ll find that you’ve moved the decimal point 23 places. Therefore, the number expressed in scientific notation is 6.02 × 10^{23}.

Expressing small measurements in scientific notation is done in a similar way. The diameter of a carbon atom is 0.000 000 000 000 154 m. In this case, you move the decimal point right until it is just past the first nonzero digit—in this case, 1. The number of places you move the decimal point right is expressed as a negative exponent of ten. The diameter of a carbon atom is 1.54 × 10^{-13} m. You always move the decimal point until the coefficient of ten is between one and less than ten. Thus, scientific notation always has the form, M × 10^n where 1 ≤ M < 10.

Notice how the following examples are converted to scientific notation.

**Quantities greater than 1**

17.16 g \[ \rightarrow \ 1.716 \times 10^1 \text{ g} \]  
Decimal point moved 1 place left.

152.6 L \[ \rightarrow \ 1.526 \times 10^2 \text{ L} \]  
Decimal point moved 2 places left.

73 621 kg \[ \rightarrow \ 7.3621 \times 10^4 \text{ kg} \]  
Decimal point moved 4 places left.

**Quantities between 0 and 1**

0.29 mL \[ \rightarrow \ 2.9 \times 10^{-1} \text{ mL} \]  
Decimal point moved 1 place right.

0.0672 m \[ \rightarrow \ 6.72 \times 10^{-2} \text{ m} \]  
Decimal point moved 2 places right.

0.0008 g \[ \rightarrow \ 8 \times 10^{-4} \text{ g} \]  
Decimal point moved 4 places right.
Calculations with Measurements  You often must use measurements to calculate other quantities. Remember that the accuracy of measurement depends on the instrument used and that accuracy is expressed as a certain number of significant digits. Therefore, you must indicate which digits in the result of any mathematical operation with measurements are significant. The rule of thumb is that no result can be more accurate than the least accurate quantity used to calculate that result. Therefore, the quantity with the least number of significant digits determines the number of significant digits in the result.

The method used to indicate significant digits depends on the mathematical operation.

### Addition and Subtraction
The answer has only as many decimal places as the measurement having the least number of decimal places.

<table>
<thead>
<tr>
<th>Mass (g)</th>
<th>Calculation</th>
</tr>
</thead>
<tbody>
<tr>
<td>190.2 g</td>
<td>Because the masses were measured on balances differing in accuracy, the least accurate measurement limits the number of digits past the decimal point.</td>
</tr>
<tr>
<td>65.291 g</td>
<td></td>
</tr>
<tr>
<td>12.38 g</td>
<td></td>
</tr>
</tbody>
</table>

The answer is rounded to the nearest tenth, which is the accuracy of the least accurate measurement:

\[
267.9 \text{ g}
\]

### Multiplication and Division
The answer has only as many significant digits as the measurement with the least number of significant digits.

\[
density = \frac{\text{mass}}{\text{volume}}
\]

\[
D = \frac{m}{V} = \frac{13.78 \text{ g}}{11.3 \text{ mL}} = 1.219469 \text{ g/mL}
\]

The answer is rounded to three significant digits, 1.22 g/mL, because the least accurate measurement, 11.3 mL, has three significant digits.

Multiplying or dividing measured quantities results in a derived quantity. For example, the mass of a substance divided by its volume is its density. But mass and volume are measured with different tools, which may have different accuracies. Therefore, the derived quantity can have no more significant digits than the least accurate measurement used to compute it.
Practice Problems

3. Determine the number of significant digits in each of the following measurements.
   a) 64 mL
   b) 0.650 g
   c) 30 cg
   d) 724.56 mm
   e) 47 080 km
   f) 0.072 040 g
   g) 1.03 mm
   h) 0.001 mm

4. Write each of the following measurements in scientific notation.
   a) 76.0°C
   b) 212 mm
   c) 56.021 g
   d) 0.78 L
   e) 0.076 12 m
   f) 763.01 g
   g) 10 301 980 nm
   h) 0.001 mm

5. Write each of the following measurements in scientific notation.
   a) 73 000 ± 1 mL
   b) 4000 ± 1000 kg
   c) 100 ± 10 cm
   d) 100 000 ± 1000 km

6. Solve the following problems and express the answer in the correct number of significant digits.
   a) 45.761 g
      \[ - 42.65 \text{ g} \]
      \[ = 1.1018 \text{ g} \]
   b) \[ 1.6 \text{ km} + 0.62 \text{ km} \]
   c) \[ 0.340 \text{ cg} + 1.20 \text{ cg} \]
      \[ = 1.54 \text{ cg} \]
   d) \[ 6000 \mu\text{m} - 202 \mu\text{m} \]

7. Solve the following problems and express the answer in the correct number of significant digits.
   a) \[ 5.761 \text{ cm} \times 6.20 \text{ cm} \]
   b) \[ \frac{23.5 \text{ kg}}{4.615 \text{ m}^3} \]
   c) \[ \frac{0.2 \text{ km}}{5.4 \text{ s}} \]
   d) \[ 11.00 \text{ m} \times 12.10 \text{ m} \times 3.53 \text{ m} \]
   e) \[ \frac{4.500 \text{ kg}}{1.500 \text{ m}^2} \]
   f) \[ \frac{18.21 \text{ g}}{4.4 \text{ cm}^3} \]

Adding and Subtracting Measurements in Scientific Notation

Adding and subtracting measurements in scientific notation requires that for any problem, the measurements must be expressed as the same power of ten. For example, in the following problem, the three length measurements must be expressed in the same power of ten.

\[
1.1012 \times 10^4 \text{ mm} \\
2.31 \times 10^3 \text{ mm} \\
+ 4.573 \times 10^2 \text{ mm}
\]
In adding and subtracting measurements in scientific notation, all measurements are expressed in the same order of magnitude as the measurement with the greatest power of ten. When converting a quantity, the decimal point is moved one place to the left for each increase in power of ten.

\[
\begin{align*}
2.31 \times 10^3 & \quad 3 \times 10^3 \rightarrow 0.231 \times 10^4 \\
4.573 \times 10^2 & \quad 5.73 \times 10^2 \rightarrow 0.04573 \times 10^4
\end{align*}
\]

\[
\begin{align*}
1.1012 \times 10^4 \text{ mm} & \\
0.231 \times 10^4 \text{ mm} & + 0.04573 \times 10^4 \text{ mm} \\
\hline
1.37793 \times 10^4 \text{ mm} = 1.378 \times 10^4 \text{ mm (rounded)}
\end{align*}
\]

**Multiplying and Dividing Measurements in Scientific Notation** Multiplying and dividing measurements in scientific notation requires that similar operations are done to the numerical values, the powers of ten, and the units of the measurements.

a) The numerical coefficients are multiplied or divided and the resulting value is expressed in the same number of significant digits as the measurement with the least number of significant digits.

b) The exponents of ten are algebraically added in multiplication and subtracted in division.

c) The units are multiplied or divided.

The following problems illustrate these procedures.

**Sample Problem 1**

\[
(3.6 \times 10^3 \text{ m})(9.4 \times 10^3 \text{ m})(5.35 \times 10^{-1} \text{ m})
\]

\[
= (3.6 \times 9.4 \times 5.35) \times (10^3 \times 10^3 \times 10^{-1}) \text{ (m} \times \text{m} \times \text{m})
\]

\[
= (3.6 \times 9.4 \times 5.35) \times 10^{3+3+(-1)} \text{ (m} \times \text{m} \times \text{m})
\]

\[
= (181.044) \times 10^5 \text{ m}^3
\]

\[
= 1.8 \times 10^2 \times 10^5 \text{ m}^3
\]

\[
= 1.8 \times 10^7 \text{ m}^3
\]

**Sample Problem 2**

\[
\frac{6.762 \times 10^2 \text{ m}^3}{(1.231 \times 10^1 \text{ m})(2.80 \times 10^{-2} \text{ m})}
\]

\[
= \frac{6.762}{1.231 \times 2.80} \times \frac{10^2}{10^1 \times 10^{-2}} \times \frac{\text{m}^3}{\text{m} \times \text{m}}
\]

\[
= 1.961819659 \times 10^{(2-(-1-2))} \text{ m}^{(3-(+2))}
\]

\[
= 1.96 \times 10^{(2-(-1))} \text{ m}^{(1)}
\]

\[
= 1.96 \times 10^3 \text{ m}
\]
Practice Problems
8. Solve the following addition and subtraction problems.
   a) $1.013 \times 10^3 \text{ g} + 8.62 \times 10^2 \text{ g} + 1.1 \times 10^1 \text{ g}$
   b) $2.82 \times 10^6 \text{ m} - 4.9 \times 10^4 \text{ m}$
9. Solve the following multiplication and division problems.
   a) $1.18 \times 10^{-3} \text{ m} \times 4.00 \times 10^2 \text{ m} \times 6.22 \times 10^2 \text{ m}$
   b) $3.2 \times 10^2 \text{ g} \div 1.04 \times 10^2 \text{ cm}^2 \div 6.22 \times 10^{-1} \text{ cm}$

Computations with the Calculator
Working problems in chemistry will require you to have a good understanding of some of the advanced functions of your calculator. When using a calculator to solve a problem involving measured quantities, you should remember that the calculator does not take significant digits into account. It is up to you to round off the answer to the correct number of significant digits at the end of a calculation. In a multistep calculation, you should not round off after each step. Instead, you should complete the calculation and then round off. Figure 6 shows how to use a calculator to solve a subtraction problem involving quantities in scientific notation.

Figure 6
Subtracting Numbers in Scientific Notation with a Calculator
A quantity in scientific notation is entered by keying in the coefficient and then striking the [EXP] or [EE] key followed by the value of the exponent of ten. At the end of the calculation, the calculator readout must be corrected to the appropriate number of decimal places. The answer would be rounded to the second decimal place and expressed as $2.56 \times 10^4 \text{ kg}$.

To solve the problem
\[
\begin{array}{c}
2.61 \times 10^4 \\
- 5.2 \times 10^2
\end{array}
\]

Keystrokes

[=]

Calculator display

\[
\begin{array}{c}
2.61 \text{ 04} \\
- 5.2 \text{ 02} \\
2.558 \text{ 04}
\end{array}
\]

\[
\begin{array}{c}
2.61 \times 10^4 \text{ kg} \\
- 0.052 \times 10^4 \text{ kg} \\
2.56 \times 10^4 \text{ kg}
\end{array}
\]
Look at Figure 7 to see how to solve multiplication and division problems involving scientific notation. The problems are the same as the two previous sample problems.

**Figure 7**

**Multiplying and Dividing Measurements Expressed in Scientific Notation**

A negative power of ten is usually entered by striking the [EXP] or [EE], entering the positive value of the exponent, and then striking the [±] key.

<table>
<thead>
<tr>
<th>Keystrokes</th>
<th>Calculator display</th>
</tr>
</thead>
<tbody>
<tr>
<td>3 x 6 EXP 3</td>
<td>3.6 03</td>
</tr>
<tr>
<td>9 x 4 EXP 3</td>
<td>9.4 03</td>
</tr>
<tr>
<td>5 x 3 5 EXP 1</td>
<td>5.35 -01</td>
</tr>
<tr>
<td>1.8104 07</td>
<td>1.8104 07</td>
</tr>
</tbody>
</table>

Rounded off to 2 significant digits: $1.8 \times 10^7$

The numerical value of the answer can have no more significant digits than the measurement that has the least number of significant digits.

<table>
<thead>
<tr>
<th>Keystrokes</th>
<th>Calculator display</th>
</tr>
</thead>
<tbody>
<tr>
<td>6 x 7 6 2 EXP 2</td>
<td>6.762 02</td>
</tr>
<tr>
<td>1 ÷ 1 2 3</td>
<td>1.231 01</td>
</tr>
<tr>
<td>1 EXP 1</td>
<td>1.9618 03</td>
</tr>
<tr>
<td>2 ÷ 8 0</td>
<td>2.80 -02</td>
</tr>
</tbody>
</table>

Rounded off to 3 significant digits: $1.96 \times 10^3$
Practice Problems
10. Solve the following problems and express the answers in scientific notation with the proper number of significant digits.
   a) \(2.01 \times 10^2 \text{ mL}\)
      \(3.1 \times 10^1 \text{ mL}\)
      \(+ 2.712 \times 10^3 \text{ mL}\)
   b) \(7.40 \times 10^2 \text{ mm}\)
      \(- 4.0 \times 10^1 \text{ mm}\)
   c) \(2.10 \times 10^1 \text{ g}\)
      \(- 1.6 \times 10^{-1} \text{ g}\)
   d) \(5.131 \times 10^2 \text{ J}\)
      \(2.341 \times 10^1 \text{ J}\)
      \(+ 3.781 \times 10^3 \text{ J}\)

11. Solve the following problems and express the answers in scientific notation with the proper number of significant digits.
   a) \((2.00 \times 10^1 \text{ cm})(2.05 \times 10^1 \text{ cm})\)
   b) \(\frac{5.6 \times 10^3 \text{ kg}}{1.20 \times 10^4 \text{ m}^3}\)
   c) \((2.51 \times 10^1 \text{ m})(3.52 \times 10^1 \text{ m})(1.2 \times 10^{-1} \text{ m})\)
   d) \(\frac{1.692 \times 10^4 \text{ dm}^3}{(2.7 \times 10^{-2} \text{ dm})(4.201 \times 10^1 \text{ dm})}\)

Using the Factor Label Method

The factor label method is used to express a physical quantity such as the length of a pen in any other unit that measures that quantity. For example, you can measure the pen with a metric ruler calibrated in centimeters and then express the length in meters.

If the length of a pen is measured as 14.90 cm, you can express that length in meters by using the numerical relationship between a centimeter and a meter. This relationship is given by the following equation.

\[100 \text{ cm} = 1 \text{ m}\]

If both sides of the equation are divided by 100 cm, the following relationship is obtained.

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

To express 14.90 cm as a measurement in meters, you multiply the quantity by the relationship, which eliminates the cm unit.

\[14.90 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} = \frac{14.90}{1} \times \frac{1 \text{ m}}{100} = \frac{14.90}{100} \text{ m} = 0.1490 \text{ m}\]
The factor label method doesn’t change the value of the physical quantity because you are multiplying that value by a factor that equals 1. You choose the factor so that when the unit you want to eliminate is multiplied by the factor, that unit and the similar unit in the factor cancel. If the unit you want to eliminate is in the numerator, choose the factor which has that unit in the denominator. Conversely, if the unit you want to eliminate is in the denominator, choose the factor which has that unit in the numerator. For example, in a chemistry lab activity, a student measured the mass and volume of a chunk of copper and calculated its density as 8.80 g/cm³. Knowing that 1000 g = 1 kg and 100 cm = 1 m, the student could then use the following factor label method to express this value in the SI unit of density, kg/m³.

\[
8.80 \text{ g/cm}^3 \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \left( \frac{100 \text{ cm}}{1 \text{ m}} \right)^3 =
\]

\[
8.80 \times \frac{\text{g}}{\text{cm}^3} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \frac{100 \text{ cm}}{1 \text{ m}} \times \frac{100 \text{ cm}}{1 \text{ m}} \times \frac{100 \text{ cm}}{1 \text{ m}} =
\]

\[
\frac{8.80 \times (10^2 \times 10^2 \times 10^2)}{1000} \text{ kg/m}^3 = \frac{8.80 \times 10^{(2+2+2)+1}}{10^3} \text{ kg/m}^3 =
\]

\[
8.80 \times 10^{(6-3)} \text{ kg/m}^3 = 8.80 \times 10^3 \text{ kg/m}^3
\]

The factor label method can be extended to other types of calculations in chemistry. To use this method, you first examine the data that you have. Next, you determine the quantity you want to find and look at the units you will need. Finally, you apply a series of factors to the data in order to convert it to the units you need.
Sample Problem 1
The density of silver sulfide (Ag₂S) is 7.234 g/mL. What is the volume of a lump of silver sulfide that has a mass of 6.84 kg?

First, you must apply a factor that will convert kg of Ag₂S to g Ag₂S.

\[
\frac{6.84 \text{ kg Ag}_2\text{S}}{1 \text{ kg Ag}_2\text{S}} \times \frac{1000 \text{ g Ag}_2\text{S}}{1 \text{ kg Ag}_2\text{S}} \ldots
\]

Next, you use the density of Ag₂S to convert mass to volume.

\[
\frac{6.84 \text{ kg Ag}_2\text{S}}{1 \text{ kg Ag}_2\text{S}} \times \frac{1000 \text{ g Ag}_2\text{S}}{1 \text{ mL Ag}_2\text{S}} \times \frac{1 \text{ mL Ag}_2\text{S}}{7.234 \text{ g Ag}_2\text{S}} = 946 \text{ mL Ag}_2\text{S}
\]

Notice that the new factor must have grams in the denominator so that grams will cancel out, leaving mL.

Sample Problem 2
What mass of lead can be obtained from 47.2 g of Pb(NO₃)₂?

Because mass is involved, you will need to know the molar mass of Pb(NO₃)₂.

\[
Pb = 207.2 \text{ g}
\]
\[
2N = 28.014 \text{ g}
\]
\[
6O = 95.994 \text{ g}
\]

Molar mass of Pb(NO₃)₂ = 331.208 g.

Rounding off according to the rules for significant digits, the molar mass of Pb(NO₃)₂ = 331.2 g.

You can see that the mass of lead in Pb(NO₃)₂ is 207.2/331.2 of the total mass of Pb(NO₃)₂.

Now you can set up a relationship to determine the mass of lead in the 47.2-g sample.

\[
\frac{47.2 \text{ g Pb(NO}_3\text{)}_2}{331.2 \text{ g Pb(NO}_3\text{)}_2} \times \frac{207.2 \text{ g Pb}}{331.2 \text{ g Pb(NO}_3\text{)}_2} = 29.52850242 \text{ g Pb}
\]

= 29.5 g Pb rounded to 3 significant digits.

Notice that the equation is arranged so that the unit g Pb(NO₃)₂ cancels out, leaving only g Pb, the quantity asked for in the problem.

Practice Problems in the Factor Label Method
12. Express each quantity in the unit listed to its right.
   a) 3.01 g cg
e) 0.2 L dm³
   b) 6200 m km
   f) 0.13 cal/g J/kg
c) 6.24 \times 10^{-7} \text{ g } \mu \text{g}
g) 5 \text{ ft}, 1 \text{ in. } m
   d) 3.21 L mL
   h) 1.2 qt L
Organizing Information

It is often necessary to compare and sequence observations and measurements. Two of the most useful ways are to organize the observations and measurements as tables and graphs. If you browse through your textbook, you’ll see many tables and graphs. They arrange information in a way that makes it easier to understand.

Making and Using Tables

Most tables have a title telling you what information is being presented. The table itself is divided into columns and rows. The column titles list items to be compared. The row headings list the specific characteristics being compared among those items. Within the grid of the table, the information is recorded. Any table you prepare to organize data taken in a laboratory activity should have these characteristics. Consider, for example, that in a laboratory experiment, you are going to perform a flame test on various solutions. In the test, you place drops of the solution containing a metal ion in a flame, and the color of the flame is observed, as shown in Figure 8. Before doing the experiment, you might set up a data table like the one below.

![Flame Test](image)

**Figure 8**

A drop of solution containing the potassium ion, K⁺, causes the flame to burn with violet color.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Metal ion</th>
<th>Color of Flame</th>
</tr>
</thead>
<tbody>
<tr>
<td>KNO₃</td>
<td>K⁺</td>
<td>violet-pink</td>
</tr>
</tbody>
</table>

While performing the experiment, you would record the name of the solution and then the observation of the flame color. If you weren’t sure of the metal ion as you were doing the experiment, you could enter it into the table by checking oxidation numbers afterward. Not only does the table organize your observations, it could also be used as a reference to determine whether a solution of some unknown composition contains one of the metal ions listed in the table.
Making and Using Graphs

After organizing data in tables, scientists usually want to display the data in a more visual way. Using graphs is a common way to accomplish that. There are three common types of graphs—bar graphs, pie graphs, and line graphs.

**Bar Graphs**  Bar graphs are useful when you want to compare or display data that do not continuously change. Suppose you measure the rate of electrolysis of water by determining the volume of hydrogen gas formed. In addition, you decide to test how the number of batteries affects the rate of electrolysis. You could graph the results using a bar graph as shown in Figure 9. Note that you could construct a line graph, but the bar graph is better because there is no way you could use 0.4 or 2.6 batteries.

![Figure 9 A Sample Bar Graph](image)

**Pie Graphs**  Pie graphs are especially useful in comparing the parts of a whole. You could use a pie graph to display the percent composition of a compound such as sodium dihydrogen phosphate, \( \text{NaH}_2\text{PO}_4 \), as shown in Figure 10.

In constructing a pie graph, recall that a circle has 360°. Therefore, each fraction of the whole is that fraction of 360°. Suppose you did a census of your school and determined that 252 students out of a total of 845 were 17 years old. You would compute the angle of that section of the graph by multiplying \( \frac{252}{845} \times 360° = 107° \).

![Figure 10 A Sample Pie Graph](image)
Line Graphs  Line graphs have the ability to show a trend in one variable as another one changes. In addition, they can suggest possible mathematical relationships between variables.

Table A.5 shows the data collected during an experiment to determine whether temperature affects the mass of potassium bromide that dissolves in 100 g of water. If you read the data as you slowly run your fingers down both columns of the table, you will see that solubility increases as the temperature increases. This is the first clue that the two quantities may be related.

To see that the two quantities are related, you should construct a line graph as shown in Figure 11.

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>Solubility (g of KBr/100 g H₂O)</th>
</tr>
</thead>
<tbody>
<tr>
<td>10.0</td>
<td>60.2</td>
</tr>
<tr>
<td>20.0</td>
<td>64.3</td>
</tr>
<tr>
<td>30.0</td>
<td>67.7</td>
</tr>
<tr>
<td>40.0</td>
<td>71.6</td>
</tr>
<tr>
<td>50.0</td>
<td>75.3</td>
</tr>
<tr>
<td>60.0</td>
<td>80.1</td>
</tr>
<tr>
<td>70.0</td>
<td>82.6</td>
</tr>
<tr>
<td>80.0</td>
<td>86.8</td>
</tr>
<tr>
<td>90.0</td>
<td>90.2</td>
</tr>
</tbody>
</table>

Figure 11  Constructing a Line Graph

1. Plot the independent variable on the x-axis (horizontal axis) and the dependent variable on the y-axis (vertical). The independent variable is the quantity changed or controlled by the experimenter. The temperature data in Table A.5 were controlled by the experimenter, who chose to measure the solubility at 10°C intervals.

2. Scale each axis so that the smallest and largest data values of each quantity can be plotted. Use divisions such as ones, fives, or tens or decimal values such as hundredths or thousandths.

3. Label each axis with the appropriate quantity and unit.

4. Plot each pair of data from the table as follows.
   - Place a straightedge vertically at the value of the independent variable on the x-axis.
   - Place a straightedge horizontally at the value of the dependent variable on the y-axis.
   - Mark the point at which the straight-edges intersect.

5. Fit the best straight line or curved line through the data points.
One use of a line graph is to predict values of the independent or dependent variables. For example, from Figure 11 you can predict the solubility of KBr at a temperature of 65°C by the following method:

- Place a straightedge vertically at the approximate value of 65°C on the x-axis.
- Mark the point at which the straightedge intersects the line of the graph.
- Place a straightedge horizontally at this point and approximate the value of the dependent variable on the y-axis as 82 g/100 g H₂O.

To predict the temperature for a given solubility, you would reverse the above procedure.

**Practice Problems**

*Use Figure 11 to answer these questions.*

13. Predict the solubility of KBr at each of the following temperatures.
   a) 25.0°C  
   b) 52.0°C  
   c) 6.0°C  
   d) 96.0°C

14. Predict the temperature at which KBr has each of the following solubilities.
   a) 70.0 g/100 g H₂O  
   b) 88.0 g/100 g H₂O

**Graphs of Direct and Inverse Relationships**  
Graphs can be used to determine quantitative relationships between the independent and dependent variables. Two of the most useful relationships are relationships in which the two quantities are directly proportional or inversely proportional.

When two quantities are directly proportional, an increase in one quantity produces a proportionate increase in the other. A graph of two quantities that are directly proportional is a straight line, as shown in Figure 12.

**Figure 12**

*Graph of Quantities That Are Directly Proportional*

As you can see, doubling the mass of the carbon burned from 2.00 g to 4.00 g doubles the amount of energy released from 66 kJ to 132 kJ. Such a relationship indicates that mass of carbon burned and the amount of energy released are directly proportional.
When two quantities are inversely proportional, an increase in one quantity produces a proportionate decrease in the other. A graph of two quantities that are inversely proportional is shown in Figure 13.

**Figure 13**
**Graph of Quantities That Are Inversely Proportional**
As you can see, doubling the pressure of the gas reduces the volume of the gas by one-half. Such a relationship indicates that the volume and pressure of a gas are inversely proportional.

**Practice Problems**
15. Plot the data in the following table and determine whether the two quantities are directly proportional.

<table>
<thead>
<tr>
<th>Temperature (K)</th>
<th>Pressure (kPa)</th>
</tr>
</thead>
<tbody>
<tr>
<td>300.0</td>
<td>195</td>
</tr>
<tr>
<td>320.0</td>
<td>208</td>
</tr>
<tr>
<td>340.0</td>
<td>221</td>
</tr>
<tr>
<td>360.0</td>
<td>234</td>
</tr>
<tr>
<td>380.0</td>
<td>247</td>
</tr>
<tr>
<td>400.0</td>
<td>261</td>
</tr>
</tbody>
</table>

16. Plot the data in the following table and determine whether the two quantities are inversely proportional.

<table>
<thead>
<tr>
<th>Number of mini-lightbulbs</th>
<th>Current (mA)</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>3.94</td>
</tr>
<tr>
<td>4</td>
<td>1.98</td>
</tr>
<tr>
<td>6</td>
<td>1.31</td>
</tr>
<tr>
<td>9</td>
<td>0.88</td>
</tr>
</tbody>
</table>