Chapter Objectives:
• Learn to use and manipulate units, and convert from one unit to another.
• Learn to use the appropriate number of significant figures in measurements and calculations.
• Learn how to classify matter by state and composition.

What Is Chemistry?

• Chemistry is the science that seeks to understand the composition, properties, and transformations of matter by studying the behavior of atoms and molecules.

• Chemistry is subdivided into different specialized fields: organic chemistry, inorganic chemistry, physical chemistry, biochemistry, analytical chemistry, environmental chemistry, etc.

• We study chemistry to provide ourselves with a better understanding of the underlying workings of nature, to learn how to make new materials with useful properties that satisfy particular needs. Chemistry intersects with other important fields, such as biology, molecular biology and genetics, medicine, physics, etc.
Units of Measure

... there is poetry in science, but also a lot of bookkeeping.

Peter Medawar

Units

• Observations in science may be:
  – qualitative — a description which does not involve a number (e.g., “this coin is heavy,” “the sky is blue”).
  – quantitative — a measurement, which contains both a number and a unit (e.g., “this coin weighs 2.35 grams,” “the frequency of the light from the sky has a wavelength of 421 nm”).

• For a measurement to be meaningful, both the number and the unit must be present (usually).
Chapter 1: Essential Ideas

The SI System

- In science, the most commonly used set of units are those of the International System of Units (the SI System, for Système International d’Unités).

- There are seven fundamental units in the SI system. The units for all other quantities (e.g., area, volume, energy) are derived from these base units.

<table>
<thead>
<tr>
<th>Physical Quantity</th>
<th>Unit</th>
<th>Abbreviation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass</td>
<td>kilogram</td>
<td>kg</td>
</tr>
<tr>
<td>Length</td>
<td>meter</td>
<td>m</td>
</tr>
<tr>
<td>Time</td>
<td>second</td>
<td>s</td>
</tr>
<tr>
<td>Amount of substance</td>
<td>mole</td>
<td>mol</td>
</tr>
<tr>
<td>Temperature</td>
<td>kelvin</td>
<td>K</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>

The SI System — Mass, Length, and Time

- The **meter** (m) is the SI unit of length.
  - Defined as the distance light travels in a vacuum in 1/2999,792,458 seconds.
  - 1 m = 39.37 in = 1.094 yards

- The **kilogram** (kg) is the SI unit of mass.
  - 1 kg = 2.205 pounds (Avoirdupois)

- The **second** (s) is the SI unit of time.

- The **mole** (mol) is the SI unit for the amount of a substance (we’ll get into this much more later on).
The SI System — Temperature
(Is It Hot In Here, Or Is It Me?)

- In the SI system, temperature is measured in **kelvins (K)**, but often the **Celsius degree, °C**, is used instead.

- A kelvin is the same size as a Celsius degree, but with the zero point set at the coldest possible temperature, **absolute zero (−273.15°C)**.

- In most mathematical formulas, K must be used instead of °C.

\[
K = °C + 273.15
\]

\[
°F = \frac{9}{5}°C + 32
\]

\[
°C = \frac{5}{9}(°F - 32)
\]

Derived Units

- From the SI base units, we can derive other units, such as those for area, volume, density, force, etc.

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Definition</th>
<th>Units</th>
</tr>
</thead>
<tbody>
<tr>
<td>Area</td>
<td>Length × width</td>
<td>m²</td>
</tr>
<tr>
<td>Volume</td>
<td>Length × width × height</td>
<td>m³</td>
</tr>
<tr>
<td>Density</td>
<td>Mass / volume</td>
<td>kg/m³, g/cm³, g/mL</td>
</tr>
<tr>
<td>Speed</td>
<td>Distance / time</td>
<td>m s⁻¹</td>
</tr>
<tr>
<td>Acceleration</td>
<td>Change in speed / time</td>
<td>m s⁻²</td>
</tr>
<tr>
<td>Frequency</td>
<td>Event / time</td>
<td>s⁻¹</td>
</tr>
<tr>
<td>Force</td>
<td>Mass × acceleration</td>
<td>kg m s⁻² (newton, N)</td>
</tr>
<tr>
<td>Pressure</td>
<td>Force / area</td>
<td>kg m⁻¹ s⁻² (pascal, Pa)</td>
</tr>
<tr>
<td>Energy</td>
<td>Force × distance</td>
<td>kg m² s⁻² (joule, J)</td>
</tr>
</tbody>
</table>
Chapter 1: Essential Ideas

**Derived Units — Volume**

- **Volume** — the amount of space occupied by an object (length \( \times \) width \( \times \) height). In the SI system, volume has units of m\(^3\) (cubic meters).
  
  - More commonly, we use the liter (L), which is a cubic decimeter (dm\(^3\)).
    - 1 L = 1 dm\(^3\) = 1000 cm\(^3\) = 0.001 m\(^3\)
    - 1 m\(^3\) = 1000L
    - 1 L = 1.0567 quarts
  
  - In chemistry, we frequently use the milliliter (mL) or cubic centimeter (cm\(^3\)):
    - 1 mL = 1 cm\(^3\) (cubic centimeter, cc)
    - 1 L = 1000 mL

**Derived Units — Density**

- **Density**, the ratio of an object’s mass \((m)\) to its volume \((V)\), is given by the formula:
  
  \[ d = \frac{m}{V} \]
  
  - Density has units of mass over volume: g/mL, g/L, lb/gal, kg/m\(^3\), lb/ft\(^3\), etc.
  
  - Because volume changes with temperature, density is *temperature-dependent*.

### Densities of Various Substances at 20ºC

<table>
<thead>
<tr>
<th>Substance</th>
<th>State</th>
<th>Density (g/cm(^3))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxygen</td>
<td>gas</td>
<td>0.00133</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>gas</td>
<td>0.000084</td>
</tr>
<tr>
<td>Ethanol</td>
<td>liquid</td>
<td>0.789</td>
</tr>
<tr>
<td>Benzene</td>
<td>liquid</td>
<td>0.886</td>
</tr>
<tr>
<td>Water</td>
<td>liquid</td>
<td>0.99982</td>
</tr>
<tr>
<td>Magnesium</td>
<td>solid</td>
<td>1.74</td>
</tr>
<tr>
<td>Sodium chloride</td>
<td>solid</td>
<td>2.16</td>
</tr>
<tr>
<td>Aluminum</td>
<td>solid</td>
<td>2.70</td>
</tr>
<tr>
<td>Iron</td>
<td>solid</td>
<td>7.87</td>
</tr>
<tr>
<td>Copper</td>
<td>solid</td>
<td>8.96</td>
</tr>
<tr>
<td>Silver</td>
<td>solid</td>
<td>10.5</td>
</tr>
<tr>
<td>Lead</td>
<td>solid</td>
<td>11.34</td>
</tr>
<tr>
<td>Mercury</td>
<td>liquid</td>
<td>13.6</td>
</tr>
<tr>
<td>Gold</td>
<td>solid</td>
<td>19.32</td>
</tr>
</tbody>
</table>
Extensive and Intensive Properties

• **Extensive properties** depend on the size of the sample (mass, volume, length, etc.).

• **Intensive properties** are independent of the size of the sample (color, melting / boiling point, odor, etc.)
  
  – Despite the fact that the mass and volume of a sample are extensive properties, the density of a pure substance is an intensive property.

Larger and Smaller Units

• In many instances, decimal multipliers are added to the units in cases where numbers are inconveniently large or small:
  
  – the diameter of a sodium atom:
    • long-hand: 0.000 000 000 372 m
    • scientific notation: $3.72 \times 10^{-10}$ m
    • prefix units: 0.372 nm or 372 pm
  
  – the distance from the earth to the sun:
    • long-hand: 150,000,000,000 m
    • scientific notation: $1.5 \times 10^{11}$ m
    • prefix units: 150 Gm
### SI Prefix Multipliers — Large Units

<table>
<thead>
<tr>
<th>Factor</th>
<th>Prefix</th>
<th>Symbol</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>$10^{24}$</td>
<td>yotta</td>
<td>Y</td>
<td></td>
</tr>
<tr>
<td>$10^{21}$</td>
<td>zetta</td>
<td>Z</td>
<td></td>
</tr>
<tr>
<td>$10^{18}$</td>
<td>exa</td>
<td>E</td>
<td></td>
</tr>
<tr>
<td>$10^{15}$</td>
<td>peta</td>
<td>P</td>
<td></td>
</tr>
<tr>
<td>$10^{12}$</td>
<td>tera</td>
<td>T</td>
<td></td>
</tr>
<tr>
<td>$10^{9}$</td>
<td>giga</td>
<td>G</td>
<td>$1 \text{ Gm} = 1 \times 10^9 \text{ m}$</td>
</tr>
<tr>
<td>$10^{6}$</td>
<td>mega</td>
<td>M</td>
<td>$1 \text{ Mm} = 1 \times 10^6 \text{ m}$</td>
</tr>
<tr>
<td>$10^{3}$</td>
<td>kilo</td>
<td>k</td>
<td>$1 \text{ kg} = 1 \times 10^3 \text{ g} = 1000 \text{ g}$</td>
</tr>
<tr>
<td>$10^{2}$</td>
<td>hecto</td>
<td>h</td>
<td>$1 \text{ hm} = 100 \text{ m}$</td>
</tr>
<tr>
<td>$10^{1}$</td>
<td>deka</td>
<td>da</td>
<td>$1 \text{ dag} = 10 \text{ g}$</td>
</tr>
<tr>
<td>$10^{0}$</td>
<td>—</td>
<td>—</td>
<td>—</td>
</tr>
</tbody>
</table>

### SI Prefix Multipliers — Small Units

<table>
<thead>
<tr>
<th>Factor</th>
<th>Prefix</th>
<th>Symbol</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>$10^{0}$</td>
<td>—</td>
<td>—</td>
<td>—</td>
</tr>
<tr>
<td>$10^{-1}$</td>
<td>deci</td>
<td>d</td>
<td>$1 \text{ dm} = 1 \times 10^{-1} \text{ m} = 0.1 \text{ m}$</td>
</tr>
<tr>
<td>$10^{-2}$</td>
<td>centi</td>
<td>c</td>
<td>$1 \text{ cm} = 1 \times 10^{-2} \text{ m} = 0.01 \text{ m}$ or $1 \text{ m} = 100 \text{ cm}$</td>
</tr>
<tr>
<td>$10^{-3}$</td>
<td>milli</td>
<td>m</td>
<td>$1 \text{ mg} = 0.001 \text{ g}$</td>
</tr>
<tr>
<td>$10^{-6}$</td>
<td>micro</td>
<td>μ</td>
<td>$1 \text{ μm} = 1 \times 10^{-6} \text{ m}$</td>
</tr>
<tr>
<td>$10^{-9}$</td>
<td>nano</td>
<td>n</td>
<td>$1 \text{ ns} = 1 \times 10^{-9} \text{ s}$</td>
</tr>
<tr>
<td>$10^{-12}$</td>
<td>pico</td>
<td>p</td>
<td>$1 \text{ pm} = 1 \times 10^{-12} \text{ m}$</td>
</tr>
<tr>
<td>$10^{-15}$</td>
<td>femto</td>
<td>f</td>
<td></td>
</tr>
<tr>
<td>$10^{-18}$</td>
<td>atto</td>
<td>a</td>
<td></td>
</tr>
<tr>
<td>$10^{-21}$</td>
<td>zepto</td>
<td>z</td>
<td></td>
</tr>
<tr>
<td>$10^{-24}$</td>
<td>yocto</td>
<td>y</td>
<td></td>
</tr>
</tbody>
</table>
Scientific Notation

• In science, we have the opportunity to work with numbers that are extremely large:

\[ 602,000,000,000,000,000,000,000 \]

and numbers that are extremely small:

\[ 0.000000000000000000000000000911 \]

• Numbers like this may be written more compactly using scientific notation:

\[ 6.02 \times 10^{23} \quad 9.11 \times 10^{-28} \]

Scientific Notation

• To put a number into scientific notation, we move the decimal point behind the first significant figure, and multiply it by the appropriate power of ten.

• For a number larger than 1, move the decimal point to the left, behind the first nonzero digit, and use a positive power of ten to indicate how many places the decimal point was moved:

\[ 6.02 = 6.02 \times 10^0 \]
\[ 60.2 = 6.02 \times 10^1 \]
\[ 602 = 6.02 \times 10^2 \]
\[ 6,020 = 6.02 \times 10^3 \]
\[ 6,020,000 = 6.02 \times 10^6 \]


**Scientific Notation**

- For a decimal number (smaller than 1), move the decimal point to the right, behind the first nonzero digit, and use a negative power of ten to indicate how many places the decimal point was moved:
  
  \[
  \begin{align*}
  0.602 &= 6.02 \times 10^{-1} \\
  0.0602 &= 6.02 \times 10^{-2} \\
  0.00602 &= 6.02 \times 10^{-3} \\
  0.0000602 &= 6.02 \times 10^{-6}
  \end{align*}
  \]

---

**Arithmetic with Scientific Notation**

\[
(6.71 \times 10^2)(2.21 \times 10^3) = (6.71 \times 2.21)(10^{2+3}) = 14.8 \times 10^5 = 1.48 \times 10^6
\]

\[
\frac{7.143 \times 10^8}{2.55 \times 10^3} = (7.143 \div 2.55)(10^{8-3}) = 2.80 \times 10^5
\]

\[
(1.6 \times 10^5)^3 = (1.6)^3 \times (10^{5 \times 3}) = 4.1 \times 10^{15}
\]

\[
\sqrt{6.32 \times 10^8} = \sqrt{6.32 \times (10^{8/2})} = 2.51 \times 10^4
\]
**Examples: SI Prefixes**

1. Complete the following conversion factors:

   a. \(1 \text{ km} = \underline{1000} \text{ m}\)
   
   b. \(1 \text{ cm} = \underline{\text{\_\_\_}} \text{ m}\)
   
   c. \(1.2 \times 10^{-9} \text{ g} = 1.2 \underline{\text{\_\_\_}} \text{ g}\)
   
   d. \(1 \text{ cm} = \underline{\text{\_\_\_}} \text{ mm}\)
   
   e. \(3.5 \mu\text{g} = 3.5 \times 10^{\underline{\text{\_\_\_}}} \text{ g}\)
   
   f. \(1,000,000 \text{ g} = 1 \underline{\text{\_\_\_}} \text{ g}\)
   
   g. \(0.001 \text{ g} = 1 \underline{\text{\_\_\_}} \text{ g}\)

---

**Examples: Conversions with SI Prefixes**

2. A bar of aluminum has a mass of 1210 g. What is its mass in kilograms (kg)?

\[
1 \text{ kg} = \underline{\text{\_\_\_}} \text{ g}
\]

\[
1 \text{ kg} = 1000 \text{ g}
\]

\[
1210 \text{ g} \times \frac{1000 \text{ g}}{1 \text{ kg}} =
\]

\[
1210 \text{ g} \times \frac{1 \text{ kg}}{1000 \text{ g}} =
\]
Examples: Metric Conversions

3. Convert 0.123 cm to mm.

Answer: 1.23 mm

Examples: Scientific Notation

4. Round off each of the following numbers to the indicated number of significant digits, and write the answer in both long-hand and scientific notation:

a. 0.07565 to two significant figures

long-hand: ______________
scientific notation: ______________

b. 325801 to three significant figures

long-hand: ______________
scientific notation: ______________
### Examples: Scientific Notation

5. Write the following in scientific notation:

| a. 360.0 | 3.600×10^2 |
| b. 0.000000208 | 2.08×10^-7 |
| c. 5.61 | 5.61×10^0 |
| d. 60200 | 6.02×10^4 |
| e. 224.5×10^-14 | 2.245×10^-12 |
| f. 0.02245×10^8 | 2.245×10^6 |

---

### Examples: Scientific Notation

6. Write the following in long hand notation:

| a. 1.046×10^-4 | 0.001046 |
| b. 2.50×10^2 | 250. |
| c. 1638×10^-3 | 1.638 |
| d. 0.00224×10^-4 | 0.000000224 |
| e. 0.0224×10^7 | 224000 |
| f. 500×10^0 | 500 |
Measurement and Significant Figures

Uncertainty in Measurement

• For instance, on the scale below, the pointer is pointing between “13” and “14”:

• We can “guess” that the pointer is about a third of the way to the 14, so we can estimate the reading as 13.3. Since the last digit is an estimation, anything further than that would be a wild guess.

• All measurements have some degree of uncertainty to them.
Chapter 1: Essential Ideas

**Significant Figures**

- The total number of digits in a measurement is called the number of *significant figures*.

- **When reading a scale, the value you record should use all of the digits you are sure of, plus one additional digit that you estimate.** This last estimated digit is the last significant figure in your reading. (On a digital readout, the last number on the screen is usually the last significant figure.)

- The greater the number of significant figures, the greater the certainty of the measurement.

---

**Accuracy and Precision**

- Whenever possible, a measurement must be performed more than once in order to improve the confidence with which the result is reported.

  - The *precision* of a series of measurements is a measure of how close all of the reported numbers are to each other (reproducibility).

  - The *accuracy* of the measurements is a measure of how close they are to the actual value.

---

(a) inaccurate imprecise  
(b) inaccurate precise  
(c) accurate precise
Random and Systematic Error

• In the following example, a lead block with a mass of 10.00 g has been measured by three students:

<table>
<thead>
<tr>
<th></th>
<th>Student 1</th>
<th>Student 2</th>
<th>Student 3</th>
</tr>
</thead>
<tbody>
<tr>
<td>10.49 g</td>
<td>9.78 g</td>
<td>10.03 g</td>
<td></td>
</tr>
<tr>
<td>9.79 g</td>
<td>9.82 g</td>
<td>9.99 g</td>
<td></td>
</tr>
<tr>
<td>9.92 g</td>
<td>9.75 g</td>
<td>10.03 g</td>
<td></td>
</tr>
<tr>
<td>10.31 g</td>
<td>9.80 g</td>
<td>9.98 g</td>
<td></td>
</tr>
<tr>
<td><strong>Avg: 10.13 g</strong></td>
<td><strong>Avg: 9.79 g</strong></td>
<td><strong>Avg: 10.01 g</strong></td>
<td></td>
</tr>
</tbody>
</table>

inaccurate    inaccurate    accurate
imprecise      precise

• The results in (1) are due to random error. Random errors often average out in repeated trials.

• In (2), there is a systematic error — all of the measurements are off in one direction. Systematic errors do not average out in repeated trials, because the same error is being made every time.

Measured Numbers vs. Exact Numbers

• Exact numbers are relationships that are arrived at by counting discrete objects (3 eggs = 3.00000… eggs) or that are true by definition (12 inches = 1 foot, 60 s = 1 min, 5280 feet = 1 mile, 100 cm = 1 m, 2.54 cm = 1 inch, etc.).

  – There is no uncertainty in these numbers, and they have an infinite number of significant figures (i.e., they do not affect the number of significant figures in the result of a calculation).

• All measured numbers will have some limit to how precisely they are known, and there is a limit to the number of significant digits in the number.

  – This must be taken into account when doing calculations with those numbers!
Chapter 1: Essential Ideas

Counting Significant Figures

Rules for Counting Significant Figures:

a. All nonzero digits are significant. (42 has 2 sf’s.)
b. Leading zeros are not significant; they are there to locate the decimal point. (0.00123 g has three sf’s.)
c. Zeros in the middle of a number \((\text{middle zeros or captive zeros})\) are significant. (4.803 cm has 4 sf’s.)
d. Trailing zeros are significant if the number contains a decimal point. (55.220 K has five sf’s; 50.0 mg has three sf’s, 5.100\times10^{-3} \text{ has four sf’s.})
e. Trailing zeros are not significant if the number does not contain a decimal. (34,200 m has three sf’s.)
   – Because trailing zeros can be ambiguous, it is a good practice to avoid potential errors by reporting the number in scientific notation.

Counting Significant Figures

a. All nonzero digits are significant.
   
   \[164.87\] 5 sf’s
   \[395\] 3 sf’s

b. Leading zeros are not significant.
   
   \[0.766\] 3 sf’s
   \[0.000033\] 2 sf’s
   \[0.00591.3\] 4 sf’s

c. Middle zeros are always significant.
   
   \[2.028\] 4 sf’s
   \[5107\] 4 sf’s
   \[0.00304\] 3 sf’s
Chapter 1: Essential Ideas

**Counting Significant Figures**

d. Trailing zeros are significant if the number contains a decimal point.

- 14.30 4 sf’s
- 0.0030 2 sf’s
- 500. 3 sf’s
- 500.0 4 sf’s

e. Trailing zeros are ambiguous if the number contains no decimal point. We usually assume that they are not significant.

- 2500 2 sf’s
- 60 1 sf’s

---

**Manipulating Significant Figures**

- The results of calculations are only as reliable as the least precise measurement.

\[
\text{Mileage} = \frac{\text{Miles}}{\text{Gallons}} = \frac{278 \text{ mi}}{11.70 \text{ gal}} = 23.760683760684 \ldots \text{ mi/gal}
\]

This is far more significant figures than either of the measurements

**Rules for Calculating Numbers involving sig. figs.:**

- During multiplication or division, the result has the same number of sf’s as the factor with the fewest sf’s.
- During addition or subtraction, the result has the same number of decimal places as the quantity with the fewest decimal places.
- The final answer is then rounded off appropriately.
Chapter 1: Essential Ideas

**Manipulating Significant Figures**

3 sf's $\rightarrow$ 278 mi
4 sf's $\rightarrow$ 11.70 gal

$\frac{278 \text{ mi}}{11.70 \text{ gal}} = 23.8 \text{ mi/gal} \rightarrow$ 3 sf's

3.18 $\rightarrow$ 2 decimal places
+ 0.01315 $\rightarrow$ 5 decimal places

3.19315
3.19 $\rightarrow$ 2 decimal places

**Adding and Subtracting Measured Numbers**

- During addition or subtraction, the result has the same number of decimal places as the quantity with the fewest decimal places.

  - It is possible to either lose or gain significant figures in addition or subtraction.

  $6.14 + 0.0375 - 0.0014278 + 63.4$
  $\underline{6.1775} - 0.0012032 + 5263.4$
  $\underline{6.18} - 0.001203 + 5260$

  $12.6198 - 12.5202 + 3.48$
  $\underline{0.0996} + 100.00$
**Multiplying and Dividing Measured Numbers**

- **During multiplication or division**, the result has the same number of sf’s as the factor with the fewest sf’s.

\[
\frac{5.77}{1.9} = 3.03684210526\ldots \text{rounds off to 3.0}
\]

\[
\frac{(28.71)(0.0626)(128.54)}{(5.0)} = 46.203600168\ldots \text{rounds off to 46}
\]

---

**Combined Operations**

- **Combined Operations.** When operations involving both addition/subtraction and multiplication/division are performed, the order of operations is important when determining the number of significant figures in the final answer:

\[
\frac{500.00}{275.0 - 225.00} = \frac{500.00}{50.0} = 10.0
\]

\[
10.00 \times 10.00 - \frac{200.0}{20.00} = 90.0
\]
Chapter 1: Essential Ideas

Examples: Significant Figures

1. (a) Report the volume on the graduated cylinder shown below (which is read at the bottom of the meniscus) to the correct number of significant figures. (b) Report the temperature on the thermometer shown below to the correct number of significant figures.

Answer: (a) 4.58 mL, (b) 103.4ºF

Examples: Significant Figures

2. How many significant figures are in each of the following?
   a. 0.04450 m
   b. 1000 m = 1 km
   c. 0.00002 g
   d. 5.0003 km
   e. 1.000×10⁻³ mL
   f. 10,000 m
   g. 4080 kg
   h. 1.4500 L
   i. 5280 feet = 1 mile
   j. 2.54 cm = 1 inch
   k. 0.000304 s
Examples: Significant Figures

3. Perform the following calculations to the correct number of significant figures.

   a. \(1.10 \times 0.5120 \times 4.0015 \div 3.4555 = 0.652\)

   b. \(0.355 + 105.1 - 100.5820 = 4.9\)

   c. \(4.562 \times 3.99870 \div (452.6755 - 452.33) = 53\)

   d. \((14.84 \times 0.55) - 8.02\) or 0.1 depending on when you round off

Examples: Significant Figures

4. Do the following calculations, rounding the answers off to the correct number of significant digits.

   a. \(12.43 \text{ miles} \left(\frac{5280 \text{ ft}}{1 \text{ mile}}\right) \left(\frac{12 \text{ in}}{1 \text{ ft}}\right) \left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right) = 2000414.592 \text{ cm}\)

   b. \(\frac{90.00 \text{ g}}{(1.20 \text{ cm})(0.27 \text{ cm})(1.05 \text{ cm})} = 264.5502646 \text{ g/cm}^3\)

   c. \(8.233 \text{ g} + 35.5 \text{ g} + 75.8060 \text{ g} = 119.539 \text{ g}\)

   d. \(\frac{165.75 \text{ g}}{(12.20 \text{ mL} - 1.15 \text{ mL})} = 15 \text{ g/mL}\)
Units

• Every number that you measure in the laboratory and most of the numbers that you calculate have not only a numerical value, but also a set of units associated with them. A number without units is (usually) meaningless; for a calculation to be regarded as correct, the correct units must be included in the final answer.

• Units can be manipulated just like numbers:

\[
\begin{align*}
6 \text{ cm} & \quad 8.0 \text{ ft} \times 2.0 \text{ ft} = 16 \text{ ft}^2 \\
-2 \text{ cm} & \quad \frac{24 \text{ mL}}{4.0 \text{ mL}} = 6.0 \text{ (no units)} \\
4 \text{ cm} & \quad \frac{24 \text{ g}}{6.0 \text{ mL}} = 4.0 \text{ g/mL} \\
& \quad 5.0 \text{ ft} \times 3.0 \text{ lb} = 15 \text{ ft} \cdot \text{lb} \\
& \quad 0.00165 s^{-1} = 0.00165 /s = 0.00165 \text{ Hz}
\end{align*}
\]
Manipulating Units

- **Dimensional analysis** (also known as the **factor-label method**) is a way of analyzing the setup of a problem by manipulating the units in the same way you would manipulate the numbers in the calculations.
  
  - If your final units are correct, there is a good chance the problem has been set up correctly.
  
  - If you end up with incorrect units (for instance, units of time when you’re measuring distance), or units which are clearly nonsense (for instance, cm•inches), the problem has been set up incorrectly.

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Examples: Unit Conversions

1. **Convert 5.0 lb to kg (kilograms)**
   - The conversion factor is 2.2 lb = 1 kg.
   - Do you multiply 5.0 by 2.2, divide 5.0 by 2.2, or divide 2.2 by 5.0?

$$5.0 \text{ lb} \left( \frac{2.2 \text{ lb}}{1 \text{ kg}} \right) = 11 \text{ lb}^2/\text{kg} \quad \times$$

$$2.2 \text{ lb} \left( \frac{1}{5.0 \text{ lb}} \right) = 0.44 \text{ kg}^{-1} \quad \times$$

$$5.0 \text{ lb} \left( \frac{1 \text{ kg}}{2.2 \text{ lb}} \right) = 2.3 \text{ kg} \quad \checkmark$$
### Examples: Unit Conversions

2. Convert 58 cm³ to gallons.

   \[ 1 \text{ cm}^3 = 1 \text{ mL} \]
   \[ 1 \text{ L} = 1000 \text{ mL} \]
   \[ 1 \text{ L} = 1.057 \text{ qt} \text{ (not exact)} \]
   \[ 1 \text{ gal} = 4 \text{ qt} \]

*Answer:* 0.015 gal

### Examples: Density

3. What is the volume in liters of a sample of acetone having a mass of 925 g? (The density of acetone is 0.788 g/mL)

   \[ d = \frac{m}{V} \]

   \[ V = \frac{d}{m} \quad V = \frac{\text{g/mL}}{\text{g}} = \]

   \[ V = \frac{m}{d} \quad V = \frac{\text{g}}{\text{g/mL}} = \]

*Answer:* 1.17 L
4. What is the mass in grams of a sample of acetone that has a volume of 1.180 L? The density of acetone is 0.788 g/mL.

Answer: 930. g

5. A common way to measure the density of an irregular solid is by the displacement of water. An irregularly-shaped piece of a shiny yellowish material having a mass of 51.842 g is submerged in a graduated cylinder containing 17.1 mL of water. The volume rose to 19.8 mL.
(a) What is the density of this material?
(b) Use the table of densities given earlier in this chapter to identify the material.

Answer: (a) 19 g/cm$^3$
Examples: Unit Conversions

6. The radius of a copper atom is 0.1280 nanometers (nm). What is its radius in picometers (pm)?

Answer: 128.0 pm

Examples: Unit Conversions

7. How many square centimeters (cm²) are there in 2.00 square meters (m²)?

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

\[ 2.00 \text{ m}^2 \times \frac{100 \text{ cm}}{1 \text{ m}} = \]

\[ 2.00 \text{ m}^2 \times \left( \frac{100 \text{ cm}}{1 \text{ m}} \right)^2 = \]
Examples: Unit Conversions

8. Convert 75.0 mi/hr to ft/s.

In a case like this, since you’re converting two sets of units, it’s helpful to split it up into a numerator and denominator:

\[ 75.0 \text{ mi/hr} = \frac{75.0 \text{ mi}}{1 \text{ hr}} \]

Answer: 110 ft/s

Examples: Unit Conversions

9. To leave the surface of the Earth, an object must attain an escape velocity of \(11,200 \text{ m/s}\). What is this speed in units of miles per hour?

Answer: 25,000 \text{ mi/hr} (or \(2.50 \times 10^4 \text{ mi/hr}\))
Examples: Unit Conversions

10. Convert 37°C to °F and Kelvin.

Answer: 98.6°C (what’s wrong with this number?)

Examples: Unit Conversions

11. A small hole in the heat shield of a space capsule requires a 32.70 cm² patch. If the patching material costs NASA $2.75/in², what is the cost of the patch? (1 in = 2.54 cm [exact])

Answer: $13.94
Examples: Unit Conversions

12. A runner wants to run 10.0 km. She knows that her running pace is 7.5 miles per hour. How many minutes must she run? (1 mi = 1.6093 km)

Answer: 50. min.

Examples: Unit Conversions

13. A 4.00 quart sample of antifreeze weighs 9.26 pounds. What is the density of the antifreeze in units of g/mL? (1 lb = 453.59 g; 1 L = 1.0567 qt)

Answer: 111 g/mL
Matter and Energy

The States of Matter

- **Matter** is anything that occupies space and has mass. Matter is classified by its *state* and by its *composition*:

  - **Solid**: Has fixed shape and volume
  - **Liquid**: Takes shape of container, forms horizontal surface, has fixed volume
  - **Gas**: Expands to fill container
**The States of Matter — Solids**

- **Solids** have a fixed shape and volume that does not conform to the container shape.
  - The atoms or molecules vibrate, but don’t move past each other, making solids *rigid* (more or less) and *incompressible*.
  - In *crystalline solids*, the atoms and molecules are arranged with some kind of long-range, repeating order, as in diamonds, ice, or salt.
  - In *amorphous solids*, there is no long-range order, as in charcoal, glass, plastics.

**The States of Matter — Liquids**

- **Liquids** have fixed volumes that conform to the container shape, but forms an upper surface.
  - The particles are packed closely, as in solids, but are free to move past each other, making them *fluid* and *incompressible* (more or less).
  - e.g., liquid water
Chapter 1: Essential Ideas

**The States of Matter — Gases**

- **Gases** have no fixed shape or volume; they conform to the container shape, but fill the entire volume (i.e., there is no surface).
  - Gas particles are widely separated, making them both *fluid* and *compressible*.
  - e.g., steam

**The Composition of Matter**

- Matter can be classified as either **pure substances**, which have *fixed (constant) compositions*, or **mixtures**, which have *variable compositions*.
  - **Pure substances** (*elements* and *compounds*) are unique materials with their own chemical and physical properties, and are composed of only one type of atom or molecule.
  - **Mixtures** are random combinations of two or more different types of atoms or molecules, and *retain the properties of the individual substances*.
  - Mixtures can be separated (sometimes with difficulty) by *physical means* (such as boiling, distillation, melting, crystallizing, magnetism, etc.).
Chapter 1: Essential Ideas

**The Composition of Matter — Pure Substances**

- **Pure Substances:**
  - An *element* is the simplest type of matter with unique physical and chemical properties.
    - Elements consist of only one kind of atom.
  - A *compound* is a pure substance that is composed of atoms of two or more different elements.
    - Chemical compounds can be divided into *ionic compounds* and *molecular compounds*.
    - Compounds cannot be broken down by physical means into simpler substances, but can be broken down (although sometimes with difficulty) by *chemical reactions*.

**The Composition of Matter — Mixtures**

- **Mixtures:**
  - *Heterogeneous mixtures* — the mixing is not uniform, and there are regions of different compositions — i.e., there are observable boundaries between the components.
    - e.g., ice-water, salad dressing, milk, dust in air.
  - *Homogeneous mixtures* (or *solutions*) — the mixing is uniform and there is a constant composition throughout — i.e., there are no observable boundaries because the substances are intermingled on the molecular level.
    - e.g., salt water, sugar water, metal alloys, air.
Chapter 1: Essential Ideas

**The Composition of Matter**

Matter

- Pure Substances
  - Elements
  - Ionic Compounds
- Mixtures
  - Compounds
    - Molecular Compounds
  - Heterogeneous Mixtures
  - Homogeneous Mixtures

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**Separation of Mixtures**

- Since the components of a mixture are different substances, with at least some physical properties that are unique to each compound, mixtures can be separated by physical means into their components by physical means, such as filtration, distillation, chromatography, etc.
Chapter 1: Essential Ideas

Separation of Mixtures

Chromatography

Filtration

Atoms and Molecules

- **Atoms** are the smallest particles of an element that has the properties of that element.

- **Molecules** consist of two or more atoms joined by some kind of chemical bond (*more later*).
Chapter 1: Essential Ideas

Physical Changes and Physical Properties

- A **physical change** occurs when a substance alters its physical form, but not its composition — the atoms and molecules in the sample retain their identities during a physical change (e.g., ice melting into liquid water, liquid water boiling to steam.)

  \[ \text{H}_2\text{O(s)} \rightarrow \text{H}_2\text{O(l)} \]

- **Physical properties** are properties that do not involve a change in a substance’s chemical makeup (e.g., melting and boiling points, color, density, odor, solubility, etc.).

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Chemical Changes and Chemical Properties

- A **chemical change** is one that alters the composition of matter — the atoms in the sample rearrange their connections in a chemical reaction, transforming the substance into a different substance (e.g., the rusting of iron, the formation of water from hydrogen and oxygen, etc.)

  \[
  \begin{align*}
  4\text{Fe(s)} + 3\text{O}_2(g) & \rightarrow 2\text{Fe}_2\text{O}_3(s) \\
  2\text{H}_2(g) + \text{O}_2(g) & \rightarrow 2\text{H}_2\text{O(l)}
  \end{align*}
  \]

- **Chemical properties** are properties that do involve a change in chemical makeup (e.g., flammability, corrosiveness, reactivity with acids, etc.).
Chapter 1: Essential Ideas

Chemical Changes

- In a chemical change, the products have a different composition, and therefore different properties, from the starting materials.

Physical and Chemical Changes
Chapter 1: Essential Ideas

**Physical and Chemical Properties of Copper**

- **Copper — Physical Properties**
  - reddish brown, metallic luster
  - it is malleable (easily formed into thin sheets) and ductile (easily drawn into wires)
  - good conductor of heat and electricity
  - can be mixed with zinc to form brass or with tin to form bronze
  - density = 8.95 g/cm³
  - melting point = 1083°C
  - boiling point = 2570°C

- **Copper — Chemical Properties**
  - slowly forms a green carbonate in moist air
  - reacts with nitric acid and sulfuric acid
  - forms a deep blue solution in aqueous ammonia

**Examples: Physical and Chemical Changes**

1. Which of the following processes are physical changes, and which are chemical changes?
   a. the evaporation of rubbing alcohol
   b. the burning of lamp oil
   c. the bleaching of hair with hydrogen peroxide
   d. the forming of frost on a cold night
   e. the beating of a copper wire into a sheet
   f. a nickel dissolving in acid to produce H₂ gas
   g. dry ice evaporating without melting
   h. the burning of a log in a fireplace
Energy

- **Energy** is defined as the ability to do work or produce heat.
  - **Work** is done when a force is exerted through a distance \( (w = Fd) \).
  - **Heat** is the energy that flows from one object to another because of a temperature difference.

- Energy may be converted from one form to another, but it is neither created nor destroyed (the law of conservation of energy).

- The total energy possessed by an object is the sum of its **kinetic energy** (energy of motion) and **potential energy** (stored energy due to position).

- Energy is measured in **Joules** (J) \((\text{kg m}^2 \text{s}^{-2})\) or **calories** (cal).

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Chemical Energy

- The **chemical potential energy** of a substance results from the relative positions and the attractions and repulsions among all its particles. Under some circumstances, this energy can be released, and can be used to do work: