Chemistry: The Study of Change

Chapter 1
Chemistry: A Science for the 21st Century

- Health and Medicine
  - Sanitation systems
  - Surgery with anesthesia
  - Vaccines and antibiotics
  - Gene therapy

- Energy and the Environment
  - Fossil fuels
  - Solar energy
  - Nuclear energy
Chemistry: A Science for the 21st Century

• Materials and Technology
  • Polymers, ceramics, liquid crystals
  • Room-temperature superconductors?
  • Molecular computing?

• Food and Agriculture
  • Genetically modified crops
  • “Natural” pesticides
  • Specialized fertilizers
The Study of Chemistry

Macroscopic

Microscopic

O₂

Fe₂O₃

Fe
The **scientific method** is a systematic approach to research.

A **hypothesis** is a tentative explanation for a set of observations.
A **law** is a concise statement of a relationship between phenomena that is always the same under the same conditions.

\[ \text{Force} = \text{mass} \times \text{acceleration} \]

A **theory** is a unifying principle that explains a body of facts and/or those laws that are based on them.

Atomic Theory
Chemistry In Action: Primordial Helium and the Big Bang Theory

In 1940 George Gamow hypothesized that the universe began with a gigantic explosion or big bang.

Experimental Support

- expanding universe
- cosmic background radiation
- primordial helium
Chemistry is the study of matter and the changes it undergoes.

Matter is anything that occupies space and has mass.

A substance is a form of matter that has a definite composition and distinct properties.
A *mixture* is a combination of two or more substances in which the substances retain their distinct identities.

1. **Homogenous mixture** – composition of the mixture is the same throughout
   - soft drink, milk, solder

2. **Heterogeneous mixture** – composition is not uniform throughout
   - cement, iron filings in sand
Physical means can be used to separate a mixture into its pure components.
An element is a substance that cannot be separated into simpler substances by chemical means.

- 114 elements have been identified
- 82 elements occur naturally on Earth
  - gold, aluminum, lead, oxygen, carbon, sulfur
- 32 elements have been created by scientists
  - technetium, americium, seaborgium
<table>
<thead>
<tr>
<th>Name</th>
<th>Symbol</th>
<th>Name</th>
<th>Symbol</th>
<th>Name</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminum</td>
<td>Al</td>
<td>Fluorine</td>
<td>F</td>
<td>Oxygen</td>
<td>O</td>
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<td>Au</td>
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<td>Iodine</td>
<td>I</td>
<td>Potassium</td>
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<td>Br</td>
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<td>Fe</td>
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<td>Pb</td>
<td>Silver</td>
<td>Ag</td>
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<td>Carbon</td>
<td>C</td>
<td>Magnesium</td>
<td>Mg</td>
<td>Sodium</td>
<td>Na</td>
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<td>Chlorine</td>
<td>Cl</td>
<td>Manganese</td>
<td>Mn</td>
<td>Sulfur</td>
<td>S</td>
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<tr>
<td>Chromium</td>
<td>Cr</td>
<td>Mercury</td>
<td>Hg</td>
<td>Tin</td>
<td>Sn</td>
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<tr>
<td>Cobalt</td>
<td>Co</td>
<td>Nickel</td>
<td>Ni</td>
<td>Tungsten</td>
<td>W</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu</td>
<td>Nitrogen</td>
<td>N</td>
<td>Zinc</td>
<td>Zn</td>
</tr>
</tbody>
</table>
A **compound** is a substance composed of atoms of two or more elements chemically united in fixed proportions.

Compounds can only be separated into their pure components (elements) by **chemical** means.

- lithium fluoride
- quartz
- dry ice – carbon dioxide
Classifications of Matter

Matter

Mixtures
- Homogeneous mixtures
- Heterogeneous mixtures

Substances

Compounds

Elements

Separation by physical methods

Separation by chemical methods
A Comparison: The Three States of Matter

Solid

Liquid

Gas
The Three States of Matter: Effect of a Hot Poker on a Block of Ice
Types of Changes

A *physical change* does not alter the composition or identity of a substance.

- ice melting
- sugar dissolving in water

A *chemical change* alters the composition or identity of the substance(s) involved.

- hydrogen burns in air to form water
Extensive and Intensive Properties

An **extensive property** of a material depends upon how much matter is being considered.

- mass
- length
- volume

An **intensive property** of a material **does not** depend upon how much matter is being considered.

- density
- temperature
- color
Matter - anything that occupies space and has \textit{mass} \\

\textit{mass} – measure of the quantity of matter \\
SI unit of mass is the \textit{kilogram} (kg) \\
1 kg = 1000 g = 1 \times 10^3 g \\

\textit{weight} – force that gravity exerts on an object \\
weight = c \times \text{mass} \\
on earth, \ c = 1.0 \\
on moon, \ c \sim 0.1 \\
A 1 \text{ kg bar will weigh} \\
1 \text{ kg on earth} \\
0.1 \text{ kg on moon}
# International System of Units (SI)

## Table 1.2: SI Base Units

<table>
<thead>
<tr>
<th>Base Quantity</th>
<th>Name of Unit</th>
<th>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>Electrical current</td>
<td>ampere</td>
<td>A</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>Luminous intensity</td>
<td>candela</td>
<td>cd</td>
</tr>
<tr>
<td>Prefix</td>
<td>Symbol</td>
<td>Meaning</td>
</tr>
<tr>
<td>--------</td>
<td>--------</td>
<td>---------</td>
</tr>
<tr>
<td>tera-</td>
<td>T</td>
<td>1,000,000,000,000, or $10^{12}$</td>
</tr>
<tr>
<td>giga-</td>
<td>G</td>
<td>1,000,000,000, or $10^{9}$</td>
</tr>
<tr>
<td>mega-</td>
<td>M</td>
<td>1,000,000, or $10^{6}$</td>
</tr>
<tr>
<td>kilo-</td>
<td>k</td>
<td>1,000, or $10^{3}$</td>
</tr>
<tr>
<td>deci-</td>
<td>d</td>
<td>1/10, or $10^{-1}$</td>
</tr>
<tr>
<td>centi-</td>
<td>c</td>
<td>1/100, or $10^{-2}$</td>
</tr>
<tr>
<td>milli-</td>
<td>m</td>
<td>1/1,000, or $10^{-3}$</td>
</tr>
<tr>
<td>micro-</td>
<td>μ</td>
<td>1/1,000,000, or $10^{-6}$</td>
</tr>
<tr>
<td>nano-</td>
<td>n</td>
<td>1/1,000,000,000, or $10^{-9}$</td>
</tr>
<tr>
<td>pico-</td>
<td>p</td>
<td>1/1,000,000,000,000, or $10^{-12}$</td>
</tr>
</tbody>
</table>
**Volume** – SI derived unit for volume is cubic meter \((m^3)\)

- \(1\, cm^3 = (1 \times 10^{-2}\, m)^3 = 1 \times 10^{-6}\, m^3\)
- \(1\, dm^3 = (1 \times 10^{-1}\, m)^3 = 1 \times 10^{-3}\, m^3\)
- \(1\, L = 1000\, mL = 1000\, cm^3 = 1\, dm^3\)

\[
1\, mL = 1\, cm^3
\]
Density – SI derived unit for density is kg/m$^3$

1 g/cm$^3$ = 1 g/mL = 1000 kg/m$^3$

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

\[
d = \frac{m}{V}
\]
<table>
<thead>
<tr>
<th>Substance</th>
<th>Density (g/cm³)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Air*</td>
<td>0.001</td>
</tr>
<tr>
<td>Ethanol</td>
<td>0.79</td>
</tr>
<tr>
<td>Water</td>
<td>1.00</td>
</tr>
<tr>
<td>Graphite</td>
<td>2.2</td>
</tr>
<tr>
<td>Table salt</td>
<td>2.2</td>
</tr>
<tr>
<td>Aluminum</td>
<td>2.70</td>
</tr>
<tr>
<td>Diamond</td>
<td>3.5</td>
</tr>
<tr>
<td>Iron</td>
<td>7.9</td>
</tr>
<tr>
<td>Mercury</td>
<td>13.6</td>
</tr>
<tr>
<td>Gold</td>
<td>19.3</td>
</tr>
<tr>
<td>Osmium†</td>
<td>22.6</td>
</tr>
</tbody>
</table>

*Measured at 1 atmosphere.
†Osmium (Os) is the densest element known.
Example 1.1

Gold is a precious metal that is chemically unreactive. It is used mainly in jewelry, dentistry, and electronic devices.

A piece of gold ingot with a mass of 301 g has a volume of 15.6 cm$^3$. Calculate the density of gold.
**Solution**  We are given the mass and volume and asked to calculate the density. Therefore, from Equation (1.1), we write

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

\[
= \frac{301 \text{ g}}{15.6 \text{ cm}^3}
\]

\[
= 19.3 \text{ g/cm}^3
\]
Example 1.2

The density of mercury, the only metal that is a liquid at room temperature, is 13.6 g/mL. Calculate the mass of 5.50 mL of the liquid.
Solution  We are given the density and volume of a liquid and asked to calculate the mass of the liquid.

We rearrange Equation (1.1) to give

\[ m = d \times V \]

\[ = 13.6 \ \frac{g}{mL} \times 5.50 \ mL \]

\[ = 74.8 \ g \]
A Comparison of Temperature Scales

\[ K = ^0C + 273.15 \]

\[ 273.15 \text{ K} = 0 ^0C \]

\[ 373.15 \text{ K} = 100 ^0C \]

\[ ^0F = \frac{9}{5} \times ^0C + 32 \]

\[ 32 ^0F = 0 ^0C \]

\[ 212 ^0F = 100 ^0C \]
(a) Solder is an alloy made of tin and lead that is used in electronic circuits. A certain solder has a melting point of 224°C. What is its melting point in degrees Fahrenheit?

(b) Helium has the lowest boiling point of all the elements at 2452°F. Convert this temperature to degrees Celsius.

(c) Mercury, the only metal that exists as a liquid at room temperature, melts at 238.9°C. Convert its melting point to kelvins.
Example 1.3

**Solution** These three parts require that we carry out temperature conversions, so we need Equations (1.2), (1.3), and (1.4). Keep in mind that the lowest temperature on the Kelvin scale is zero (0 K); therefore, it can never be negative.

(a) This conversion is carried out by writing

\[
\frac{9 \, ^\circ\text{F}}{5 \, ^\circ\text{C}} \times (224 \, ^\circ\text{C}) + 32 \, ^\circ\text{F} = 435 \, ^\circ\text{F}
\]

(b) Here we have

\[
(-452 \, ^\circ\text{F} - 32 \, ^\circ\text{F}) \times \frac{5 \, ^\circ\text{C}}{9 \, ^\circ\text{F}} = -269 \, ^\circ\text{C}
\]

(c) The melting point of mercury in kelvins is given by

\[
(-38.9 \, ^\circ\text{C} + 273.15 \, ^\circ\text{C}) \times \frac{1 \, \text{K}}{1 \, ^\circ\text{C}} = 234.3 \, \text{K}
\]
Chemistry In Action

On 9/23/99, $125,000,000 Mars Climate Orbiter entered Mars’ atmosphere 100 km (62 miles) lower than planned and was destroyed by heat.

1 lb ≠ 1 N
1 lb = 4.45 N

“This is going to be the cautionary tale that will be embedded into introduction to the metric system in elementary school, high school, and college science courses till the end of time.”
Scientific Notation

The number of atoms in 12 g of carbon:

\[ 602,200,000,000,000,000,000,000 \]

\[ 6.022 \times 10^{23} \]

The mass of a single carbon atom in grams:

\[ 0.0000000000000000000000000000199 \]

\[ 1.99 \times 10^{-23} \]

\[ N \times 10^n \]

N is a number between 1 and 10

\[ n \] is a positive or negative integer
Scientific Notation

568.762

\[ \text{move decimal left} \quad n > 0 \]

\[ 568.762 = 5.68762 \times 10^2 \]

0.00000772

\[ \text{move decimal right} \quad n < 0 \]

\[ 0.00000772 = 7.72 \times 10^{-6} \]

Addition or Subtraction

1. Write each quantity with the same exponent \( n \)
2. Combine \( N_1 \) and \( N_2 \)
3. The exponent, \( n \), remains the same

\[ 4.31 \times 10^4 + 3.9 \times 10^3 = 4.70 \times 10^4 \]
Scientific Notation

Multiplication
1. Multiply $N_1$ and $N_2$
2. Add exponents $n_1$ and $n_2$

$$\begin{align*}
(4.0 \times 10^{-5}) \times (7.0 \times 10^3) &= (4.0 \times 7.0) \times (10^{-5+3}) = \\
&= 28 \times 10^{-2} = 2.8 \times 10^{-1} 
\end{align*}$$

Division
1. Divide $N_1$ and $N_2$
2. Subtract exponents $n_1$ and $n_2$

$$\begin{align*}
8.5 \times 10^4 \div 5.0 \times 10^9 &= (8.5 \div 5.0) \times 10^{4-9} = \\
&= 1.7 \times 10^{-5} 
\end{align*}$$
Significant Figures

• Any digit that is not zero is significant
  1.234 kg  4 significant figures

• Zeros between nonzero digits are significant
  606 m  3 significant figures

• Zeros to the left of the first nonzero digit are **not** significant
  0.08 L  1 significant figure

• If a number is greater than 1, then all zeros to the right of the decimal point are significant
  2.0 mg  2 significant figures

• If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant
  0.00420 g  3 significant figures
Determine the number of significant figures in the following measurements:

(a) 478 cm
(b) 6.01 g
(c) 0.825 m
(d) 0.043 kg
(e) $1.310 \times 10^{22}$ atoms
(f) 7000 mL
**Example 1.4**

**Solution**

(a) 478 cm -- Three, because each digit is a nonzero digit.

(b) 6.01 g -- Three, because zeros between nonzero digits are significant.

(c) 0.825 m -- Three, because zeros to the left of the first nonzero digit do not count as significant figures.

(d) 0.043 kg -- Two. Same reason as in (c).

(e) 1.310 × 10^{22} atoms -- Four, because the number is greater than one so all the zeros written to the right of the decimal point count as significant figures.
(f) 7000 mL -- This is an ambiguous case. The number of significant figures may be four \((7.000 \times 10^3)\), three \((7.00 \times 10^3)\), two \((7.0 \times 10^3)\), or one \((7 \times 10^3)\).

This example illustrates why scientific notation must be used to show the proper number of significant figures.
Significant Figures

Addition or Subtraction

The answer cannot have more digits to the right of the decimal point than any of the original numbers.

\[
\begin{align*}
89.332 & \quad +1.1 & \quad \text{one significant figure after decimal point} \\
\hline
90.432 & \quad \text{round off to 90.4} \\
\end{align*}
\]

\[
\begin{align*}
3.70 & \quad -2.9133 & \quad \text{two significant figures after decimal point} \\
\hline
0.7867 & \quad \text{round off to 0.79} \\
\end{align*}
\]
Significant Figures

Multiplication or Division

The number of significant figures in the result is set by the original number that has the smallest number of significant figures.

\[
4.51 \times 3.6666 = 16.536366 = 16.5
\]

3 sig figs

round to

3 sig figs

\[
6.8 \div 112.04 = 0.0606926 = 0.061
\]

2 sig figs

round to

2 sig figs
Significant Figures

Exact Numbers

Numbers from definitions or numbers of objects are considered to have an infinite number of significant figures.

The average of three measured lengths: 6.64, 6.68 and 6.70?

\[
\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67 = 7
\]

Because 3 is an exact number
Example 1.5

Carry out the following arithmetic operations to the correct number of significant figures:

(a) 11,254.1 g + 0.1983 g

(b) 66.59 L − 3.113 L

(c) 8.16 m × 5.1355

(d) 0.0154 kg ÷ 88.3 mL

(e) $2.64 \times 10^3$ cm + $3.27 \times 10^2$ cm
Example 1.5

Solution  In addition and subtraction, the number of decimal places in the answer is determined by the number having the lowest number of decimal places. In multiplication and division, the significant number of the answer is determined by the number having the smallest number of significant figures.

(a) 11,254.1 g
    + 0.1983 g
    11,254.2983 g ← round off to 11,254.3 g

(b) 66.59 L
    - 3.113 L
    63.477 L ← round off to 63.48 L
Example 1.5

(c) \[8.16 \text{ m} \times 5.1355 = 41.90568 \text{ m} \leftarrow \text{round off to 41.9 m}\]

(d) \[\frac{0.0154 \text{ kg}}{88.3 \text{ mL}} = 0.000174405436 \text{ kg/mL} \leftarrow \text{round off to}
\]
\[0.000174 \text{ kg/mL}
\]
\[\text{or } 1.74 \times 10^{-4} \text{ kg/mL}\]

(e) First we change \(3.27 \times 10^2 \text{ cm}\) to \(0.327 \times 10^3 \text{ cm}\) and then carry out the addition \((2.64 \text{ cm} + 0.327 \text{ cm}) \times 10^3\). Following the procedure in (a), we find the answer is \(2.97 \times 10^3 \text{ cm}\).
**Accuracy** – how close a measurement is to the *true* value

**Precision** – how close a set of measurements are to each other

---

(a) accurate 
& 
precise

(b) precise 
but 
not accurate

(c) not accurate 
& 
not precise
Dimensional Analysis Method of Solving Problems

1. Determine which unit conversion factor(s) are needed
2. Carry units through calculation
3. If all units cancel except for the *desired unit(s)*, then the problem was solved correctly.

\[
given \text{ quantity} \times \text{ conversion factor} = desired \text{ quantity}
\]

\[
given \text{ unit} \times \frac{\text{desired unit}}{\text{given unit}} = \text{desired unit}
\]
Example 1.6

A person’s average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g.)
Example 1.6

**Strategy** The problem can be stated as

\[ \text{? mg} = 0.0833 \text{ lb} \]

The relationship between pounds and grams is given in the problem. This relationship will enable conversion from pounds to grams.

A metric conversion is then needed to convert grams to milligrams (1 mg = 1 \times 10^{-3} g).

Arrange the appropriate conversion factors so that pounds and grams cancel and the unit milligrams is obtained in your answer.
Solution  The sequence of conversions is

\[
\text{pounds} \rightarrow \text{grams} \rightarrow \text{milligrams}
\]

Using the following conversion factors

\[
\frac{453.6 \text{ g}}{1 \text{ lb}} \quad \text{and} \quad \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}}
\]

we obtain the answer in one step:

\[
? \text{ mg} = 0.0833 \text{ lb} \times \frac{453.6 \text{ g}}{1 \text{ lb}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}} = 3.78 \times 10^4 \text{ mg}
\]
Example 1.6

**Check** As an estimate, we note that 1 lb is roughly 500 g and that \(1 \text{ g} = 1000 \text{ mg}\). Therefore, 1 lb is roughly \(5 \times 10^5 \text{ mg}\).

Rounding off 0.0833 lb to 0.1 lb, we get \(5 \times 10^4 \text{ mg}\), which is close to the preceding quantity.
An average adult has 5.2 L of blood. What is the volume of blood in m$^3$?
Strategy

The problem can be stated as

\[ ? \text{ m}^3 = 5.2 \text{ L} \]

How many conversion factors are needed for this problem?

Recall that 1 L = 1000 cm³ and 1 cm = 1 \times 10^{-2} \text{ m}.
Example 1.7

Solution  We need two conversion factors here: one to convert liters to cm\(^3\) and one to convert centimeters to meters:

\[
\frac{1000 \text{ cm}^3}{1 \text{ L}} \quad \text{and} \quad \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}
\]

Because the second conversion factor deals with length (cm and m) and we want volume here, it must therefore be cubed to give

\[
\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} = \left(\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}\right)^3
\]

This means that \(1 \text{ cm}^3 = 1 \times 10^{-6} \text{ m}^3\).
Now we can write

\[ ? \text{ m}^3 = 5.2 \text{ L} \times \frac{1000 \text{ cm}^3}{1 \text{ L}} \times \left( \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \right)^3 = 5.2 \times 10^{-3} \text{ m}^3 \]

**Check** From the preceding conversion factors you can show that 1 L = 1 × 10⁻³ m³. Therefore, 5 L of blood would be equal to 5 × 10⁻³ m³, which is close to the answer.
Liquid nitrogen is obtained from liquefied air and is used to prepare frozen goods and in low-temperature research.

The density of the liquid at its boiling point (−196°C or 77 K) is 0.808 g/cm³. Convert the density to units of kg/m³.
Example 1.8

**Strategy**  The problem can be stated as

\[ ? \text{ kg/m}^3 = 0.808 \text{ g/cm}^3 \]

Two separate conversions are required for this problem:

\[ \text{g} \rightarrow \text{kg} \quad \text{and} \quad \text{cm}^3 \rightarrow \text{m}^3 \]

Recall that 1 kg = 1000 g and 1 cm = 1 \( \times \) \( 10^{-2} \) m.
Example 1.8

**Solution** In Example 1.7 we saw that $1 \text{ cm}^3 = 1 \times 10^{-6} \text{ m}^3$. The conversion factors are

$$\frac{1 \text{ kg}}{1000 \text{ g}} \quad \text{and} \quad \frac{1 \text{ cm}^3}{1 \times 10^{-6} \text{ m}^3}$$

Finally

$$? \text{ kg/m}^3 = \frac{0.808 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \frac{1 \text{ cm}^3}{1 \times 10^{-6} \text{ m}^3} = 808 \text{ kg/m}^3$$

**Check** Because $1 \text{ m}^3 = 1 \times 10^6 \text{ cm}^3$, we would expect much more mass in $1 \text{ m}^3$ than in $1 \text{ cm}^3$. Therefore, the answer is reasonable.
A modern pencil “lead” is actually composed primarily of graphite, a form of carbon.

Estimate the mass of the graphite core in a standard No. 2 pencil before it is sharpened.
Strategy  Assume that the pencil lead can be approximated as a cylinder.

Measurement of a typical unsharpened pencil gives a length of about 18 cm (subtracting the length of the eraser head) and a diameter of roughly 2 mm for the lead.

The volume of a cylinder $V$ is given by

$$V = \pi r^2 l$$

where $r$ is the radius and $l$ is the length.

Assuming that the lead is pure graphite, you can calculate the mass of the lead from the volume using the density of graphite given in Table 1.4.
Example 1.9

Solution

Converting the diameter of the lead to units of cm gives

\[
2 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} = 0.2 \text{ cm}
\]

which, along with the length of the lead, gives

\[
V = \pi \left( \frac{0.2 \text{ cm}}{2} \right)^2 \times 18 \text{ cm}
\]

\[
= 0.57 \text{ cm}^3
\]
Example 1.9

Rearranging Equation (1.1) gives

\[ m = d \times V \]

\[ = 2.2 \frac{g}{cm^3} \times 0.57 \text{ cm}^3 \]

\[ = 1 \text{ g} \]

**Check** Rounding off the values used to calculate the volume of the lead gives

\[ 3 \times (0.1 \text{ cm})^2 \times 20 \text{ cm} = 0.6 \text{ cm}^3. \]

Multiplying that volume by roughly 2 g/cm\(^3\) gives around 1 g, which agrees with the value just calculated.