Fundamentals of General, Organic, and Biological
CHEMISTRY
Eighth Edition

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with contributions by
Sara Madsen
# Instructor and Student Resources

<table>
<thead>
<tr>
<th>Name of Supplement</th>
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| Instructor Resource Manual  
0134283171 / 9780134283173 |  | ✓ | Instructor | The IRM features lecture outlines with presentation suggestions, teaching tips, suggested in-class demonstrations, topics for classroom discussion, and answers to group problems. |
| TestGen Test Bank  
013426147X / 9780134261478 |  | ✓ | Instructor | The test bank has been updated to reflect revisions in this text, and contains more than 2,000 multiple choice, true/false, matching, and short answer questions. |
| Instructor's Resource Materials  
0134261267 / 9780134261263 |  | ✓ | Instructor | The Instructor Resource area provides the following downloadable resources: All illustrations, tables and photos from the text in JPEG format, and pre-built PowerPoint® Presentations (lecture—including Worked Examples, images). |
| Study Guide and Full Solutions Manual  
0134261372 / 9780134261379 | ✓ |  | Student | This manual, prepared by Susan McMurry, provides solutions to all problems in the text. It explains in detail how the answers to the in-text and end-of-chapter problems are obtained. It also contains chapter summaries, study hints, and self-tests for each chapter. |
According to the U.S. Centers for Disease Control and Prevention, the U.S. population is suffering from a fat epidemic, with more than one-third (34.9% or 78.6 million) of U.S. adults characterized as obese. But how do we define obesity, and how is it measured? Obesity is defined as an excessive amount of body fat. But some body fat is important for good health, so how much body fat is healthy and how much is too much? What is fat, and how do we measure it? Body fat can be estimated using a Body Mass Index (BMI) as discussed later in the chapter, or can be measured directly using underwater immersion, or buoyancy testing, as illustrated in the photo above. The immersion tank uses buoyancy—a property related to the differences in density—to determine the percentage of body fat.
of body fat. Checking the observed buoyancy on a standard table then gives an estimation of body fat. Density is just one of the concepts we will explore in this chapter, as we learn about the properties of matter and the various forms that matter can take.

The ancient philosophers believed that all matter was composed of four fundamental substances—earth, air, fire, and water. We now know that matter is much more complex, made up of 91 naturally occurring fundamental substances, or elements, in millions of unique combinations. Everything you see, touch, taste, and smell is made of chemicals formed from these elements. Many chemicals occur naturally, but others are synthetic, including the plastics, fibers, and medicines that are so critical to modern life. Just as everything you see is made of chemicals, many of the natural changes you observe taking place around you are the result of chemical reactions—the change of one chemical into another. The crackling fire of a log burning in the fireplace, the color change of a leaf in the fall, and the changes that a human body undergoes as it grows and ages are all results of chemical reactions. To understand these and other natural processes, you must have a basic understanding of chemistry.

As you might expect, the chemistry of living organisms is complex, and it is not possible to understand all concepts without a proper foundation. Thus, we will gradually learn to connect the basic concepts, beginning in the first 11 chapters with a grounding in the scientific fundamentals that govern all of chemistry. Next, in the following six chapters, we look at the nature of the carbon-containing substances, or organic chemicals, that compose all living things. In the final 12 chapters, we apply what we have learned in the first part of the book to the study of biological chemistry.

We begin in Chapter 1 with an examination of the states and properties of matter. Since our knowledge of chemistry is based on observations and measurements, we include an introduction to the systems of measurement that are essential to our understanding of matter and its behavior.

1.1 Chemistry: The Central Science

Learning Objective:

- Identify properties of matter and differentiate between chemical and physical changes.

Chemistry is often referred to as "the central science" because it is essential to nearly all other sciences. In fact, as more and more is learned, the historical dividing lines between chemistry, biology, and physics are fading, and current research is more interdisciplinary. Figure 1.1 diagrams the relationship of chemistry and biological chemistry to other fields of scientific study.

Chemistry is the study of matter—its nature, properties, and transformations. Matter, in turn, is an all-encompassing word used to describe anything physically real—anything you can see, touch, taste, or smell. In more scientific terms, matter is anything that has mass and volume. Like all the other sciences, our knowledge of chemistry has developed by application of a process called the scientific method. The discovery of aspirin, for example, is a combination of serendipity and the scientific method: observation, evaluation of data, formation of a hypothesis, and the design of experiments to test the hypothesis and further our understanding (see the Chemistry in Action on p. 7). Advances in scientific knowledge are typically the result of this systematic approach; hypotheses can be tested by carefully designed experiments, modified based on the results of those experiments, and further tested to refine our understanding.

All of chemistry is based on the study of matter and the changes that matter undergoes. How might we describe different kinds of matter more specifically? Any characteristic used to describe or identify something is called a property; size, color, density, and more. Chemistry is the study of the nature, properties, and transformations of matter. Matter is the physical material that makes up the universe; anything that has mass and occupies space. Scientific method is the systematic process of observation, hypothesis, and experimentation used to expand and refine a body of knowledge. Property is a characteristic useful for identifying a substance or object.
and temperature are all familiar examples. Less familiar properties include chemical composition, which describes what matter is made of, and chemical reactivity, which describes how matter behaves. Rather than focusing on the properties themselves, it is often more useful to think about changes in properties. There are two types of changes: physical and chemical. A physical change is one that does not alter the identity of a substance, whereas a chemical change does alter a substance’s identity. For example, the melting of solid ice to give liquid water is a physical change because the water changes only in form but not in chemical makeup. However, the rusting of an iron bicycle left in the rain is a chemical change because iron combines with oxygen and moisture from the air to give a new substance, rust.

Table 1.1 lists some chemical and physical properties of several familiar substances—water, table sugar (sucrose, a carbohydrate), and baking soda (sodium bicarbonate). Note in Table 1.1 that changes occurring when sugar and baking soda are heated are chemical changes because new substances are produced.

Worked Example 1.1 Chemical vs. Physical Change

Identify each of the following as a chemical change or a physical change:

a) Sugar dissolving in water.
b) Sugar heated in a saucepan to make caramel.

ANALYSIS A physical change does not result in a change in the identity of the substance, whereas a chemical change results in the creation of a new substance with properties that are different than the original substance.
SOLUTION

a) Physical change: When sugar dissolves in water, the sugar and the water retain their identity. The water can be removed by evaporation, and the sugar can be recovered in its original form.

b) Chemical change: When sugar is heated in a saucepan, it melts and darkens and thickens into caramel. When cooled, the caramel clearly has significantly different properties (color, consistency) than the original sugar, indicating that a chemical change has occurred and a new substance has been formed.

Table 1.1 Some Properties of Water, Sugar, and Baking Soda

<table>
<thead>
<tr>
<th></th>
<th>Water</th>
<th>Sugar (Sucrose)</th>
<th>Baking Soda (Sodium Bicarbonate)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Physical properties</td>
<td>Colorless liquid</td>
<td>White crystals</td>
<td>White powder</td>
</tr>
<tr>
<td></td>
<td>Odorless</td>
<td>Odorless</td>
<td>Odorless</td>
</tr>
<tr>
<td>Melting point: 0 °C</td>
<td>Begins to decompose at 160 °C, turning black and giving off water.</td>
<td></td>
<td>Decomposes at 270 °C, giving off water and carbon dioxide.</td>
</tr>
<tr>
<td>Boiling point: 100 °C</td>
<td>—</td>
<td>—</td>
<td>—</td>
</tr>
<tr>
<td>Chemical properties</td>
<td>Composition:* 11.2% hydrogen</td>
<td>6.4% hydrogen</td>
<td>27.4% sodium</td>
</tr>
<tr>
<td></td>
<td>88.8% oxygen</td>
<td>42.1% carbon</td>
<td>1.2% hydrogen</td>
</tr>
<tr>
<td></td>
<td>51.5% oxygen</td>
<td>14.3% carbon</td>
<td>57.1% oxygen</td>
</tr>
<tr>
<td>Does not burn.</td>
<td>Burns in air.</td>
<td>—</td>
<td>Does not burn.</td>
</tr>
</tbody>
</table>

*Compositions are given by mass percent.

HANDS-ON CHEMISTRY 1.1

Look in the refrigerator or on the counter top in your home, apartment, or work place. If there is a bowl of fruit, onions, potatoes, etc., take a look at these items and compare what they would look like in the grocery store versus in their current location. Do you see mold? Is the flesh of the food soft, etc.? If so, would this be a physical change or a chemical change? What evidence can you cite to support your answer?
1.2 States of Matter

Learning Objective:
• Identify the three states of matter and describe their properties.

Matter exists in three forms: solid, liquid, and gas. A solid has a definite volume and a definite shape that does not change regardless of the container in which it is placed; for example, a wooden block, marbles, or a cube of ice all keep their volume and shape whether they are placed on a table or in a box. A liquid, by contrast, has a definite volume but an indefinite shape. The volume of a liquid, such as water, remains the same when it is poured into a different container, but its shape changes as it takes the shape of the container. A gas is different still, having neither a definite volume nor a definite shape. A gas expands to fill the volume and take the shape of any container it is placed in, such as the helium in a balloon or steam formed by boiling water (Figure 1.2).

Many substances, such as water, can exist in all three phases, or states of matter—the solid state \((s)\), the liquid state \((l)\), and the gaseous state \((g)\)—depending on the temperature. In general, a substance that is a solid can be converted to the liquid state if the temperature is increased sufficiently. Likewise, many liquids can be converted to the gaseous state by increasing the temperature even further. The conversion of a substance from one state to another is known as a change of state. The melting of a solid, the freezing or boiling of a liquid, and the condensing of a gas to a liquid are physical changes familiar to everyone.

Worked Example 1.2 Identifying States of Matter

Formaldehyde is a disinfectant, a preservative, and a raw material for the manufacturing of plastics. Its melting point is \(-92\ °C\), and its boiling point is \(-19.5\ °C\). Is formaldehyde a gas, a liquid, or a solid at room temperature \((25\ °C)\)? (Note: Room temperature in the Fahrenheit scale \((°F)\), with which you may be more familiar, is around 78 °F. These two scales will be compared in Section 1.11.)

**ANALYSIS** The state of matter of any substance depends on its temperature. How do the melting point and boiling point of formaldehyde compare with room temperature?

**SOLUTION**
Room temperature \((25\ °C)\) is above the boiling point of formaldehyde \((-19.5\ °C)\), and so the formaldehyde is a gas.

**Problem 1.1**
Pure acetic acid, which gives the sour taste to vinegar, has a melting point of \(16.7\ °C\) and a boiling point of \(118\ °C\). Predict the physical state of acetic acid when the ambient temperature is \(10\ °C\).
1.3 Classification of Matter

Learning Objective:

• Distinguish between mixtures and pure substances and classify pure substances as elements or compounds

The first question a chemist asks about an unknown substance is whether it is a pure substance or a mixture. Every sample of matter is one or the other. Separately, water and sugar are pure substances, but stirring some sugar into a glass of water creates a mixture.

What is the difference between a pure substance and a mixture? One difference is that a pure substance is uniform in its chemical composition and its properties all the way down to the microscopic level. Every sample of water, sugar, or baking soda, regardless of source, has the composition and properties listed in Table 1.1. A mixture, however, can vary in both composition and properties, depending on how it is made. A homogeneous mixture is a blend of two or more pure substances having a uniform composition at the microscopic level. Sugar dissolved in water is one example. You cannot always distinguish between a pure substance and a homogeneous mixture just by looking. The sugar–water mixture looks just like pure water but differs on a molecular level. The amount of sugar dissolved in a glass of water will determine the sweetness, boiling point, and other properties of the mixture. A heterogeneous mixture, by contrast, is a blend of two or more pure substances having nonuniform composition, such as a vegetable stew in which each spoonful is different. It is relatively easy to distinguish heterogeneous mixtures from pure substances.

Another difference between a pure substance and a mixture is that the components of a mixture can be separated without changing their chemical identities. For example, water can be separated from a sugar–water mixture by boiling the mixture to drive off the steam and then condensing the steam to recover the pure water. Pure sugar is left behind in the container.

Pure substances are classified into two groups: those that can undergo a chemical breakdown to yield simpler substances and those that cannot. A pure substance that cannot be broken down chemically into simpler substances is called an element. Examples include hydrogen, oxygen, aluminum, gold, and sulfur. At the time this book was printed, 118 elements had been identified, although only 91 of these occur naturally.

Any pure material that can be broken down into simpler substances by a chemical change is called a chemical compound. The term compound implies “more than one” (think “compound fracture”). A chemical compound, therefore, is formed by combining two or more elements to make a new substance. Water, for example, is a chemical compound consisting of hydrogen and oxygen; it can be chemically changed by passing an electric current through it to produce the elements hydrogen and oxygen). In Section 1.5, we will discuss chemical changes in more detail. Figure 1.3 summarizes the classification of matter into mixtures, pure compounds, and elements.

**Worked Example 1.3 Classifying Matter**

Classify each of the following as a mixture or a pure substance. If a mixture, classify it as heterogeneous or homogeneous. If a pure substance, identify it as an element or a compound.

(a) Vanilla ice cream  
(b) Sugar

**ANALYSIS** Refer to the definitions of pure substances and mixtures. Is the substance composed of more than one kind of matter? Is the composition uniform?

**SOLUTION**

(a) Vanilla ice cream is composed of more than one substance—cream, sugar, and vanilla flavoring. The composition appears to be uniform throughout, so this is a homogeneous mixture.

(b) Sugar is composed of only one kind of matter—pure sugar. This is a pure substance. It can be converted to some other substance by a chemical change (see Table 1.1), so it is not an element. It must be a compound.
CHAPTER 1 Matter and Measurements

**Figure 1.3**
A map for the classification of matter.

- **Physical change**
  - **Mixture:**
    - Seawater
    - Mayonnaise
    - Concrete
  - **Is the mixture uniform?**
    - **Yes**
      - Homogeneous mixtures:
        - Salt water
        - Coffee
        - Air
    - **No**
      - Heterogeneous mixtures:
        - Chocolate chip cookies
        - Pot pie
  - **Pure substance:**
    - Water
    - Gold
    - Sugar
  - **Is separation by chemical reaction into simpler substances possible?**
    - **Yes**
    - Chemical compound:
      - Water
      - Sugar
      - Table salt
    - **No**
    - Element:
      - Oxygen
      - Gold
      - Sulfur

**PROBLEM 1.2**
Classify each of the following as a mixture or a pure substance. If a mixture, classify it as heterogeneous or homogeneous. If a pure substance, identify it as an element or a compound.

(a) Concrete  (b) The helium in a balloon  
(c) A lead weight  (d) Wood

**PROBLEM 1.3**
Classify each of the following as a physical change or a chemical change:

(a) Dissolving sugar in water  
(b) Producing carbon dioxide gas and solid lime by heating limestone  
(c) Frying an egg  
(d) The conversion of salicylic acid to acetylsalicylic acid (see the Chemistry in Action feature on the next page)

**KEY CONCEPT PROBLEM 1.4**
In the next image, red spheres represent element A and blue spheres represent element B. Identify the process illustrated in the image as a chemical change or a physical change. Also, identify the substance(s) on the left and the substance(s) on the right as pure substances or mixtures. Explain your answer.
CHEMISTRY IN ACTION

Aspirin—A Case Study

Acetylsalicylic acid (ASA), more commonly known as aspirin, is perhaps the first true wonder drug. It is a common staple in today's medicine chest, but its discovery can be traced back to 400 B.C. The ancient Greek physician Hippocrates prescribed the bark and leaves of the willow tree to relieve pain and fever. His knowledge of the therapeutic properties of these substances was derived through trial and error. In 1828, scientists isolated a bitter-tasting yellow extract, called salicin, from willow bark and identified salicin as the active ingredient responsible for the observed medical effects. Salicin could be easily converted to salicylic acid (SA). SA, however, had an unpleasant taste and often caused stomach irritation and indigestion. Further experiments were performed to convert SA to a substance that retained the therapeutic activity of SA but without the unpleasant side effects. Bayer marketed the new drug, now called aspirin, in water-soluble tablets.

But how does aspirin work? Once again, experimental data provided insights into the therapeutic activity of aspirin. In 1971, the British pharmacologist John Vane discovered that aspirin suppresses the body's production of prostaglandins, which are responsible for the pain and swelling that accompany inflammation. The discovery of this mechanism led to the development of new analgesic drugs.

Aspirin is classified as a non-steroidal anti-inflammatory drug (NSAID), but its therapeutic value goes well beyond relieving aches and pains. Because aspirin also has anticoagulant activity, a daily, low-dose aspirin regimen (100 mg) is recommended by many physicians to reduce the risks associated with cardiovascular disease—heart attacks and strokes. Its anti-inflammatory activity is also believed to reduce the risk of developing certain types of cancer, especially in patients who suffer from chronic or persistent inflammation. For example, in a study of almost 20,000 women, the risk of ovarian cancer decreased by over 20% for women who followed a daily low-dose aspirin regimen, and that these benefits increased with long-term use. Individuals who followed the low-dose regimen for 5 years or more experienced lower incidence of colorectal cancers, and the 20-year risk of cancer death remained lower for a wide variety of other cancers, including stomach and esophageal cancers and adenocarcinomas—common malignant cancers that develop in the lungs, colon, and prostate.

CIA Problem 1.1 The active ingredient in aspirin, ASA, melts at 140 °C. Is it a solid or a liquid at room temperature?

CIA Problem 1.2 Do you think the conversion of SA to aspirin is a chemical change or a physical change? Give evidence to support your answer.

1.4 Chemical Elements and Symbols

Learning Objective:
• Identify the symbols and names of the common elements.

As of the date this book was printed, 118 chemical elements have been identified. Some are certainly familiar to you—for example, oxygen, helium, iron, aluminum, copper, and gold—but many others are probably unfamiliar—rhenium, niobium, thulium, and promethium. Rather than writing out the full names of elements, chemists use a shorthand notation in which elements are referred to by one- or two-letter symbols. The names and symbols of some common elements are listed in Table 1.2, and a complete alphabetical list is given inside the front cover of this book.

Note that all two-letter symbols have only their first letter capitalized, whereas the second letter is always lowercase. The symbols of most common elements are the first one or two letters of the elements' commonly used names, such as H (hydrogen) and Al (aluminum). Pay special attention, however, to the elements grouped in the last column to the right in Table 1.2. The symbols for these elements are derived from their original Latin names, such as Na for sodium, once known as natrium. The only way to learn these symbols is to memorize them; fortunately, they are few in number.

Prostaglandins can be synthesized from arachidonic acid and have many biological effects, which are discussed in the Chemistry in Action feature in Chapter 23, p. 735. Aspirin can inhibit the formation of prostaglandins.

We will discuss the creation of new elements by nuclear bombardment in Chapter 11. Many of these new substances are used as medical diagnostic tracers or therapeutic agents.
Table 1.2 Names and Symbols for Some Common Elements

<table>
<thead>
<tr>
<th>Elements with Symbols Based on Modern Names</th>
<th>Elements with Symbols Based on Latin Names</th>
</tr>
</thead>
<tbody>
<tr>
<td>Al</td>
<td>Aluminum</td>
</tr>
<tr>
<td>Ar</td>
<td>Argon</td>
</tr>
<tr>
<td>Ba</td>
<td>Barium</td>
</tr>
<tr>
<td>Bi</td>
<td>Bismuth</td>
</tr>
<tr>
<td>B</td>
<td>Boron</td>
</tr>
<tr>
<td>Br</td>
<td>Bromine</td>
</tr>
<tr>
<td>Ca</td>
<td>Calcium</td>
</tr>
<tr>
<td>C</td>
<td>Carbon</td>
</tr>
<tr>
<td>Cl</td>
<td>Chlorine</td>
</tr>
</tbody>
</table>

In Chapter 29, elements found in the human body will be discussed in greater detail along with body fluids.

Chemical formula A notation for a chemical compound using element symbols and subscripts to show how many atoms of each element are present.

We’ll learn more about the structure of atoms and how they form compounds in Chapter 2.

Only 91 elements occur naturally; the remaining elements have been produced artificially by chemists and physicists. Each element has its own distinctive properties, and just about all of the first 95 elements have been put to use in some way that takes advantage of those properties. As indicated in Table 1.3, which shows the approximate elemental composition of the earth's crust and the human body, the naturally occurring elements are not equally abundant. Oxygen and silicon together account for nearly 75% of the mass in the earth's crust; oxygen, carbon, and hydrogen account for nearly all the mass of a human body.

Table 1.3 Elemental Composition of the Earth's Crust and the Human Body*

<table>
<thead>
<tr>
<th>Earth's Crust</th>
<th>Human Body</th>
</tr>
</thead>
<tbody>
<tr>
<td>Oxygen</td>
<td>46.1%</td>
</tr>
<tr>
<td>Silicon</td>
<td>28.2%</td>
</tr>
<tr>
<td>Aluminum</td>
<td>8.2%</td>
</tr>
<tr>
<td>Iron</td>
<td>5.6%</td>
</tr>
<tr>
<td>Calcium</td>
<td>4.1%</td>
</tr>
<tr>
<td>Sodium</td>
<td>2.4%</td>
</tr>
<tr>
<td>Magnesium</td>
<td>2.3%</td>
</tr>
<tr>
<td>Potassium</td>
<td>2.1%</td>
</tr>
<tr>
<td>Titanium</td>
<td>0.57%</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>0.14%</td>
</tr>
</tbody>
</table>

*Mass percent values are given.

Just as elements combine to form chemical compounds, symbols are combined to produce **chemical formulas**, which use subscripts to identify how many atoms (the smallest fundamental units) of each element are in a given chemical compound. For example, the formula H₂O represents water, which contains two hydrogen atoms combined with one oxygen atom. Similarly, the formula CH₄ represents methane (natural gas), and the formula C₁₂H₂₂O₁₁ represents table sugar (sucrose). When no subscript is given for an element, as for carbon in the formula CH₄, a subscript of “1” is understood.

Those elements essential for human life are listed in Table 1.4. In addition to the well-known elements such as carbon, hydrogen, oxygen, and nitrogen, less familiar elements such as molybdenum and selenium are also important.
### Table 1.4 Elements Essential for Human Life*

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Function</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon</td>
<td>C</td>
<td>These four elements are present in all living organisms [Ch. 12-29].</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H</td>
<td>May affect cell growth and heart function.</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O</td>
<td>Aids in the use of Ca, P, and Mg.</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N</td>
<td>Necessary for growth of teeth and bones.</td>
</tr>
<tr>
<td>Arsenic</td>
<td>As</td>
<td>Necessary for maintaining salt balance in body fluids [Ch. 29].</td>
</tr>
<tr>
<td>Boron</td>
<td>B</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Calcium*</td>
<td>Ca</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Chlorine*</td>
<td>Cl</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Chromium</td>
<td>Cr</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Cobalt</td>
<td>Co</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Fluorine</td>
<td>F</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Iodine</td>
<td>I</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Magnesium*</td>
<td>Mg</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Manganese</td>
<td>Mn</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Molybdenum</td>
<td>Mo</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Nickel</td>
<td>Ni</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Phosphorus*</td>
<td>P</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Potassium*</td>
<td>K</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Selenium</td>
<td>Se</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Silicon</td>
<td>Si</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Sodium*</td>
<td>Na</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Sulfur*</td>
<td>S</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
<tr>
<td>Zinc</td>
<td>Zn</td>
<td>Component of vitamin B[12] [Ch. 19].</td>
</tr>
</tbody>
</table>

* C, H, O, and N are present in most foods. Other elements listed vary in their distribution in different foods. Those marked with an asterisk are macronutrients, essential in the diet at more than 100 mg/day; the rest, other than C, H, O, and N, are micronutrients, essential at 15 mg or less per day.

**PROBLEM 1.5**

Match the names of the elements described below (a–f) with their elemental symbols (1–6).

- (a) Sodium, a major component in table salt
- (b) Tungsten, a metal used in light bulb filaments
- (c) Strontium, used to produce brilliant red colors in fireworks
- (d) Titanium, used in artificial hips and knee-replacement joints
- (e) Fluorine, added to municipal water supplies to strengthen tooth enamel
- (f) Tin, a metal used in solder

(1) W        (2) Na        (3) Sn        (4) F        (5) Ti        (6) Sr

**PROBLEM 1.6**

Identify the elements represented in each of the following chemical formulas and tell the number of atoms of each element:

- (a) NH₃ (ammonia)
- (b) NaHCO₃ (sodium bicarbonate)
- (c) C₈H₁₈ (octane, a component of gasoline)
- (d) C₆H₈O₆ (vitamin C)
**Chemical reaction** A process in which the identity and composition of one or more substances are changed.

**Reactant** A starting substance that undergoes change during a chemical reaction.

**Product** A substance formed as the result of a chemical reaction.

We will discuss how reactions are classified in Chapter 5.

---

### 1.5 Chemical Reactions: Examples of Chemical Change

**Learning Objective:**
- Identify a chemical change as a chemical reaction.

Chemists represent chemical changes using a symbolic shorthand notation called a chemical reaction. In writing this chemical change, the initial substances, or reactants, are written on the left; the new substances, or products, are written on the right. An arrow connects the two parts to indicate the chemical change or the chemical reaction. The conditions necessary to bring about the reaction are written above and below the arrow. Consider again the example of a chemical change discussed previously, in which electric current was passed through the reactant water ($\text{H}_2\text{O}$) to break it down into the products, the elements hydrogen ($\text{H}_2$) and oxygen ($\text{O}_2$). This chemical reaction can be expressed in words as shown next.

$$\text{Reactant} \xrightarrow{\text{Water Electric current}} \text{Products}$$

Chemists, however, find it more convenient to use chemical symbols to represent the elements and compounds involved in the reaction. This chemical reaction would more commonly be expressed as

$$\text{H}_2\text{O}(l) \xrightarrow{\text{Electric current}} \text{H}_2(g) + \text{O}_2(g)$$

Note that the reactants and products are represented using their chemical formulas, but that the physical states of the reactants and products are also indicated as a liquid ($l$) or gas ($g$). The formation of gas bubbles is an indication that a chemical reaction has occurred.

If we take a quick look at another example of a chemical reaction in Figure 1.4, we can reinforce these ideas. The element nickel is a hard, shiny metal, and the compound hydrogen chloride is a colorless gas that dissolves in water to give a solution called hydrochloric acid. When pieces of nickel are added to hydrochloric acid in a test tube, the nickel slowly dissolves, the colorless solution turns green, and a gas bubbles out of the test tube. The change in color, the dissolving of the nickel, and the appearance of gas bubbles are indications that a chemical reaction is taking place.

Again, the overall reaction of nickel with hydrochloric acid can be written in words or represented in a shorthand notation using symbols to represent the elements or compounds involved as reactants and products, as shown below. The physical states of the reactants are indicated as solid ($s$) for Ni and the HCl as ($aq$), which means "aqueous," or "dissolved in water." The physical states of the products are ($aq$) for the nickel(II) chloride dissolved in water and ($g$) for the H$_2$ gas. If the water is evaporated away, the nickel(II) chloride product can be collected as a solid, also shown in Figure 1.4.

$$\text{Reactants} \xrightarrow{\text{Nickel (II) chloride + Hydrogen}} \text{Products}$$

$$\text{Ni} + 2\text{HCl} \xrightarrow{(s) (aq)} \text{NiCl}_2 + \text{H}_2 \xrightarrow{(aq) (g)}$$
Figure 1.4
Reactants and products of a chemical reaction.
(a) The reactants: Nickel (shown on the flat dish), an element that is a typical lustrous metal, and hydrochloric acid (in the bottle), a solution of the chemical compound hydrogen chloride in water. (b) The reaction: As the chemical reaction occurs, the colorless solution turns green when water-insoluble nickel metal slowly changes into the water-soluble chemical compound nickel(II) chloride. Hydrogen gas bubbles are produced and rise slowly through the green solution. (c) The product: Hydrogen gas can be collected as it bubbles from the solution and removal of water from the solution leaves behind the other product, a solid, green chemical compound known as nickel(II) chloride.

1.6 Physical Quantities: Units and Scientific Notation

Learning Objective:
• Write very large and very small numbers using scientific notation or units with appropriate numerical prefixes.

Our understanding of matter depends on our ability to measure the changes in physical properties associated with physical and chemical change. Mass, volume, temperature, density, and other physical properties that can be measured are called physical quantities and are described by both a number and a unit that defines the nature and magnitude of the number.

Units of Measurement

The number alone is not much good without a unit. If you ask how much blood an accident victim has lost, the answer “three” would not tell you much. Three drops? Three milliliters? Three pints? Three liters? By the way, an adult human has only 5–6 liters of blood.

Any physical quantity can be measured in many different units. For example, a person’s height might be measured in inches, feet, yards, centimeters, or many other units. To avoid confusion, scientists from around the world have agreed on a system of standard units, called by the French name Système International d’Unités (International System of Units), abbreviated SI. SI units for some common physical quantities

<table>
<thead>
<tr>
<th>Physical quantity</th>
<th>A physical property that can be measured.</th>
</tr>
</thead>
<tbody>
<tr>
<td>Unit</td>
<td>A defined quantity used as a standard of measurement.</td>
</tr>
</tbody>
</table>

| SI units | Units of measurement defined by the International System of Units. Examples include kilograms, meters, and kelvins. |
Mercury and Mercury Poisoning

Mercury, the only metallic element that is liquid at room temperature, has fascinated people for millennia. Much of the recent interest in mercury has concerned its toxicity, but mercury, in nontoxic forms, has a wide array of clinical uses. For example, the mercury compound $\text{Hg}_2\text{Cl}_2$ (called calomel) has a long history of medical use as a laxative, yet it is also used as a fungicide and rat poison. Dental amalgam, a solid alloy of elemental mercury, silver, tin, copper, and zinc, was used by dentists for many years to fill tooth cavities, with little or no adverse effects except in individuals with a hypersensitivity to mercury. Yet, exposure to elemental mercury vapor for long periods leads to mood swings, headaches, tremors, and loss of hair and teeth. The widespread use of mercuric nitrate, a mercury compound used to make the felt used in hats, exposed many hatters of the eighteenth and nineteenth centuries to toxic levels of mercury. The eccentric behavior displayed by hatters suffering from mercury poisoning led to the phrase “mad as a hatter.”

Why is mercury more toxic in some forms than in others? It turns out that the toxicity of mercury and its compounds is related to solubility. Only soluble mercury compounds are highly toxic because they can be transported through the bloodstream to all parts of the body, where they react with enzymes and interfere with various biological processes. Elemental mercury and insoluble mercury compounds become toxic only when converted into soluble compounds, reactions that are extremely slow in the body. Calomel, for example, is an insoluble mercury compound that passes through the body long before it is converted into any soluble compounds. Mercury alloys were considered safe for dental use because mercury does not evaporate readily from the alloys and it neither reacts with nor dissolves in saliva. Mercury vapor, however, remains in the lungs when breathed, until it is slowly converted into soluble compounds. Soluble organic forms of mercury can be particularly toxic. Trace amounts are found in nearly all seafood, but some larger species such as king mackerel and swordfish contain higher levels of mercury. Because mercury can affect the developing brain and nervous system of a fetus, pregnant women are often advised to avoid consuming them.

Recent events have raised new concerns regarding the safe use of mercury in some other applications. Perhaps the most controversial example is the use of thimerosal, an organic mercury compound, as a preservative in flu vaccines. Concerns about possible links between thimerosal and autism in children resulted in elimination of its use in 1999, although most scientific data seem to refute any connection. In response to these concerns, preservative-free versions of the influenza vaccine are available for use in infants, children, and pregnant women.

CIA Problem 1.3 Calomel ($\text{Hg}_2\text{Cl}_2$) is not toxic but methyl mercury chloride ($\text{CH}_3\text{HgCl}$) is highly toxic. What physical property explains this difference in toxicity?
are often used for density—the mass of substance in a given volume. We will see other such derived units in future chapters.

One problem with any system of measurement is that the sizes of the units often turn out to be inconveniently large or small for the problem at hand. A biologist describing the diameter of a red blood cell (0.000 006 m) would find the meter to be an inconveniently large unit, but an astronomer measuring the average distance from the earth to the sun (150,000,000,000 m) would find the meter to be inconveniently small. For this reason, metric and SI units can be modified by prefixes to refer to either smaller or larger quantities. For instance, the SI unit for mass—the kilogram—differs by the prefix kilo- from the metric unit gram. Kilo- indicates that a kilogram is 1000 times as large as a gram:

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

Small quantities of active ingredients in medications are often reported in milligrams (mg). The prefix milli- shows that the unit gram has been divided by 1000, which is the same as multiplying by 0.001:

\[ 1 \text{ mg} = \left( \frac{1}{1000} \right)(1 \text{ g}) = (0.001)(1 \text{ g}) = 0.001 \text{ g} \]

A list of prefixes is given in Table 1.6, with the most common ones displayed in color. Centi- is seen most often in the length unit centimeter (1 cm = 0.01 m), and deci- is used most often in clinical chemistry, where the concentrations of blood components are given in milligrams per deciliter (1 dL = 0.1 L). These prefixes allow us to compare the magnitudes of different numbers by noting how the prefixes modify a common unit.

### Table 1.6 Some Prefixes for Multiples of Metric and SI Units

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Base Unit Multiplied By</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>mega</td>
<td>M</td>
<td>1,000,000 = 10^6</td>
<td>1 megameter (Mm) = 10^6 m</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>1000 = 10^3</td>
<td>1 kilogram (kg) = 10^3 g</td>
</tr>
<tr>
<td>hecto</td>
<td>h</td>
<td>100 = 10^2</td>
<td>1 hectogram (hg) = 100 g</td>
</tr>
<tr>
<td>deka</td>
<td>da</td>
<td>10 = 10^1</td>
<td>1 dekaliter (dal.) = 10 L</td>
</tr>
<tr>
<td>deci</td>
<td>d</td>
<td>0.1 = 10^-1</td>
<td>1 deciliter (dL) = 0.1 L</td>
</tr>
<tr>
<td>centi</td>
<td>c</td>
<td>0.01 = 10^-2</td>
<td>1 centimeter (cm) = 0.01 m</td>
</tr>
<tr>
<td>milli</td>
<td>m</td>
<td>0.001 = 10^-3</td>
<td>1 milligram (mg) = 0.001 g</td>
</tr>
<tr>
<td>micro</td>
<td>µ</td>
<td>0.000 001 = 10^-6</td>
<td>1 micrometer (µm) = 10^-6 m</td>
</tr>
<tr>
<td>nano</td>
<td>n</td>
<td>0.000 000 001 = 10^-9</td>
<td>1 nanogram (ng) = 10^-9 g</td>
</tr>
<tr>
<td>pico</td>
<td>p</td>
<td>0.000 000 000 001 = 10^-12</td>
<td>1 picogram (pg) = 10^-12 g</td>
</tr>
<tr>
<td>femto</td>
<td>f</td>
<td>0.000 000 000 000 001 = 10^-15</td>
<td>1 femtogram (fg) = 10^-15 g</td>
</tr>
</tbody>
</table>
CHAPTER 1 Matter and Measurements

The HIV-1 virus particles (in green) budding from the surface of a lymphocyte have an approximate diameter of 0.000 000 120 m.

One approach for deactivating the HIV-1 virus is called inhibition and will be discussed in Chapter 19. How small would the inhibition agent have to be to fit in a specific location on the surface of the virus particle?

Scientific notation
A number expressed as the product of a number between 1 and 10, times the number 10 raised to a power.

For example,

1 meter = 10 dm = 100 cm = 1000 mm = 1,000,000 μm

Such comparisons will be useful when we start performing calculations involving units. It is worth noting that, as mentioned before, all the metric units displayed above are related by factors of 10. Note also in Table 1.6 that numbers having five or more digits to the right of the decimal point are shown with thin spaces every three digits for convenience—0.000 001, for example. This manner of writing numbers is becoming more common and will be used throughout this book.

Scientific Notation
Another way to solve the problem of representing very large or very small numbers is to use scientific notation. Rather than write very large or very small numbers in their entirety, it is more convenient to express them using scientific notation. A number is written in scientific notation as the product of a number between 1 and 10, times the number 10 raised to a power. Thus, 215 is written in scientific notation as $2.15 \times 10^2$:

$215 = 2.15 \times 100 = 2.15(10 \times 10) = 2.15 \times 10^2$

Notice that in this case, where the number is larger than 1, the decimal point has been moved to the left until it follows the first digit. The exponent on the 10 is positive and tells how many places we had to move the decimal point to position it just after the first digit:

$215. = 2.15 \times 10^2$

Decimal point is moved two places to the left, so exponent is 2.

To express a number smaller than one in scientific notation, we have to move the decimal point to the right until it follows the first digit. The number of places moved is the negative exponent of 10. For example, the number 0.002 15 can be rewritten as $2.15 \times 10^{-3}$:

$0.002\,15 = 2.15 \times \frac{1}{1000} = 2.15 \times \frac{1}{10 \times 10 \times 10} = 2.15 \times \frac{1}{10^3} = 2.15 \times 10^{-3}$

$0.00215 = 2.15 \times 10^{-3}$

Decimal point is moved three places to the right, so exponent is -3.

To convert a number written in scientific notation to standard notation, the process is reversed. For a number with a positive exponent, the decimal point is moved to the right a number of places equal to the exponent:

$3.7962 \times 10^4 = 37,962$

Positive exponent of 4, so decimal point is moved to the right four places.

For a number with a negative exponent, the decimal point is moved to the left a number of places equal to the exponent:

$1.56 \times 10^{-8} = 0.000\,000\,015\,6$

Negative exponent of -8, so decimal point is moved to the left eight places.
**Worked Example 1.4** Units and Scientific Notation

The HIV-1 virus particles seen in the margin photo on p. 14 are very small, on the order of 0.000 000 120 m in diameter. Express this value using scientific notation and using an appropriate numerical prefix to modify the basic unit.

**ANALYSIS** The number is significantly less than one, so when we convert to scientific notation we should have a number with a negative exponent. We can use the value of that exponent to identify the appropriate numerical prefix.

**SOLUTION**

To convert to scientific notation we have to move the decimal place to the right by seven places, so 0.000 000 120 m = $1.20 \times 10^{-7} m$. From Table 1.6, the closest numerical prefixes are micro ($10^{-6}$) or nano ($10^{-9}$). If we moved the decimal place six places to the right we would obtain:

$$0.000 \ 000 \ 120 \ m = 0.120 \times 10^{-6} m = 0.120 \text{ micrometers (\mu m)}$$

If we move the decimal place nine places to the right we obtain:

$$0.000 \ 000 \ 120 \ m = 120 \times 10^{-9} m = 120 \text{ nanometers (nm)}.$$
The two-pan balance is used to measure the mass of objects, such as the pennies on the left pan, by comparing them with the mass of standard objects, such as the brass weights on the right pan.

**Table 1.7 Units of Mass**

<table>
<thead>
<tr>
<th>Unit</th>
<th>Equivalent</th>
<th>Unit</th>
<th>Equivalent</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 kilogram (kg)</td>
<td>= 1000 grams</td>
<td>1 ton</td>
<td>= 2000 pounds</td>
</tr>
<tr>
<td></td>
<td>= 2.205 pounds</td>
<td></td>
<td>= 907.03 kilograms</td>
</tr>
<tr>
<td>1 gram (g)</td>
<td>= 0.001 kilogram</td>
<td>1 pound (lb)</td>
<td>= 16 ounces</td>
</tr>
<tr>
<td></td>
<td>= 1000 milligrams</td>
<td>= 0.454 kilogram</td>
<td>= 454 grams</td>
</tr>
<tr>
<td>1 milligram (mg)</td>
<td>= 0.001 gram</td>
<td>= 28.35 grams</td>
<td>= 28,350 milligrams</td>
</tr>
<tr>
<td></td>
<td>= 1000 micrograms</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1 microgram (μg)</td>
<td>= 0.000 001 gram</td>
<td>1 ounce (oz)</td>
<td>= 0.028 35 kilogram</td>
</tr>
<tr>
<td></td>
<td>= 0.001 milligram</td>
<td></td>
<td>= 28.35 grams</td>
</tr>
</tbody>
</table>

**Table 1.8 Units of Length**

<table>
<thead>
<tr>
<th>Unit</th>
<th>Equivalent</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 kilometer (km)</td>
<td>= 1000 meters</td>
</tr>
<tr>
<td></td>
<td>= 0.6214 mile</td>
</tr>
<tr>
<td>1 meter (m)</td>
<td>= 100 centimeters</td>
</tr>
<tr>
<td></td>
<td>= 1000 millimeters</td>
</tr>
<tr>
<td></td>
<td>= 1.0936 yards</td>
</tr>
<tr>
<td></td>
<td>= 39.37 inches</td>
</tr>
<tr>
<td>1 centimeter (cm)</td>
<td>= 0.01 meter</td>
</tr>
<tr>
<td></td>
<td>= 10 millimeters</td>
</tr>
<tr>
<td></td>
<td>= 0.3937 inch</td>
</tr>
<tr>
<td>1 millimeter (mm)</td>
<td>= 0.001 meter</td>
</tr>
<tr>
<td></td>
<td>= 0.1 centimeter</td>
</tr>
<tr>
<td>1 mile (mi)</td>
<td>= 1.609 kilometers</td>
</tr>
<tr>
<td></td>
<td>= 1609 meters</td>
</tr>
<tr>
<td>1 yard (yd)</td>
<td>= 0.9144 meter</td>
</tr>
<tr>
<td></td>
<td>= 91.44 centimeters</td>
</tr>
<tr>
<td>1 foot (ft)</td>
<td>= 0.3048 meter</td>
</tr>
<tr>
<td></td>
<td>= 30.48 centimeters</td>
</tr>
<tr>
<td>1 inch (in)</td>
<td>= 2.54 centimeters</td>
</tr>
<tr>
<td></td>
<td>= 25.4 millimeters</td>
</tr>
</tbody>
</table>

Commonly used in chemistry and medicine. One liter has the volume of a cube 10 cm (1 dm) on edge and is a bit larger than one U.S. quart. Each liter is further divided into 1000 milliliters (mL), with 1 mL being the size of a cube 1 cm on edge, or 1 cm³. In fact, the milliliter is often called a cubic centimeter (cm³ or cc) in medical work. Figure 1.6 shows the divisions of a cubic meter, and Table 1.9 shows the relationships among units of volume.
SECTION 1.7 Measuring Mass, Length, and Volume

Figure 1.6
A cubic meter is the volume of a cube 1 m on edge. Each cubic meter contains 1000 cubic decimeters (liters), and each cubic decimeter contains 1000 cubic centimeters (milliliters). Thus, there are 1000 mL in a liter and 1000 L in a cubic meter.

Table 1.9 Units of Volume

<table>
<thead>
<tr>
<th>Unit</th>
<th>Equivalent</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 cubic meter (m³)</td>
<td>= 1000 liters</td>
</tr>
<tr>
<td></td>
<td>= 264.2 gallons</td>
</tr>
<tr>
<td>1 liter (L)</td>
<td>= 0.001 cubic meter</td>
</tr>
<tr>
<td></td>
<td>= 1000 milliliters</td>
</tr>
<tr>
<td></td>
<td>= 1.057 quarts</td>
</tr>
<tr>
<td>1 deciliter (dL)</td>
<td>= 0.1 liter</td>
</tr>
<tr>
<td></td>
<td>= 100 milliliters</td>
</tr>
<tr>
<td>1 milliliter (mL)</td>
<td>= 0.001 liter</td>
</tr>
<tr>
<td></td>
<td>= 1000 microliters</td>
</tr>
<tr>
<td>1 microliter (μL)</td>
<td>= 0.001 milliliter</td>
</tr>
<tr>
<td>1 gallon (gal)</td>
<td>= 3.7854 liters</td>
</tr>
<tr>
<td>1 quart (qt)</td>
<td>= 0.9464 liter</td>
</tr>
<tr>
<td></td>
<td>= 946.4 milliliters</td>
</tr>
<tr>
<td>1 fluid ounce (fl oz)</td>
<td>= 29.57 milliliters</td>
</tr>
</tbody>
</table>

HANDS-ON CHEMISTRY 1.2

The mass of an object provides us with important information about its composition, that is, what elements it contains. But mass is not the only property that can be used to distinguish between objects. Consider the U.S. penny—its composition has changed significantly over time. For example, the penny was pure copper from 1793 to 1837, and then incorporated varying amounts of other metals [mainly zinc, nickel, and tin] from 1837 to 1982. Interestingly, in 1943, the penny was made mainly of zinc-coated steel because copper and zinc were needed for the war effort. From 1962 to 1982, the penny contained 95% copper and 5% zinc; after 1982, the composition changed to 2.5% copper and 97.5% zinc. The significant difference in composition can be seen in the different properties of pre- and post-1982 pennies.

To explore these differences, sort through your spare change jar and collect 10 pre-1982 and 10 post-1982 pennies and perform the following activities:

a. Collect two identical glasses or jars. Take the 10 pre-1982 pennies and drop them in the glass jar and listen to them as they hit the glass sides. Then, take the 10 post-1982 pennies and drop them in the glass jar and listen to them as they hit the glass sides. Do they sound different? [Note: If glasses or jars are not readily available, you can use a hard surface, like the kitchen counter.]

b. If a food scale is available, weigh 10 pre-1982 pennies and weigh 10 post-1982 pennies. Are their masses different? Which has more mass?

c. Find a pair of tin snips or heavy duty metal shears. Carefully cut a penny from each group in half. Is one penny easier to cut than the other? Which one is easier? Now look at the inside of the pennies. How are they different?
CHAPTER 1 Matter and Measurements

1.8 Measurement and Significant Figures

Learning Objective:
- Use significant figures and scientific notation to represent the precision of a measurement.

How much does a tennis ball weigh? If you put a tennis ball on an ordinary bathroom scale, the scale would probably register 0 lb (or 0 kg if you have a metric scale). If you placed the same tennis ball on a common laboratory balance, however, you might get a reading of 54.07 g. On an expensive analytical balance like those found in clinical and research laboratories, you might find the ball has a mass of 54.071 38 g. Clearly, the precision of your answer depends on the equipment used for the measurement.

Every experimental measurement, no matter how precise, has a degree of uncertainty to it because there is always a limit to the number of digits that can be determined. An analytical balance, for example, might reach its limit in measuring mass to the fifth decimal place, and weighing the tennis ball several times might produce slightly different readings, such as 54.071 39 g, 54.071 38 g, and 54.071 37 g. Also, different people making the same measurement might come up with slightly different answers. How, for instance, would you record the volume of the liquid shown in Figure 1.7? It is clear that the volume of liquid lies between 17.0 and 18.0 mL, but the exact value of the last digit must be estimated.

To indicate the precision of a measurement, the value recorded should use all the digits known with certainty, plus one additional estimated digit that is usually considered uncertain by plus or minus 1 (written as ± 1). The total number of digits used to express such a measurement is called the number of significant figures. Thus, the quantity 54.07 g has four significant figures (5, 4, 0, and 7), and the quantity 54.071 38 g has seven significant figures. Remember: All but one of the significant figures are known with certainty; the last significant figure is only an estimate accurate to ± 1.

Deciding the number of significant figures in a given measurement is usually simple, but it can be troublesome when zeroes are involved. Depending on the circumstances, a zero might be significant or might be just a space filler to locate the decimal point. For example, how many significant figures does each of the following measurements have?

- 94.072 g Five significant figures (9, 4, 0, 7, 2)
- 0.0834 cm Three significant figures (8, 3, 4)
- 0.029 07 mL Four significant figures (2, 9, 0, 7)
- 138.200 m Six significant figures (1, 3, 8, 2, 0, 0)
- 23,000 kg Anywhere from two (2, 3) to five (2, 3, 0, 0, 0) significant figures

The following rules are helpful for determining the number of significant figures when zeroes are present:

RULE 1: Zeros in the middle of a number are like any other digit; they are always significant. Thus, 94.072 g has five significant figures.

RULE 2: Zeros at the beginning of a number are not significant; they act only to locate the decimal point. Thus, 0.0834 cm has three significant figures, and 0.029 07 mL has four.
RULE 3: Zeros at the end of a number and after the decimal point are significant. It is assumed that these zeroes would not be shown unless they were significant. Thus, 138.200 m has six significant figures. If the value were known to only four significant figures, we would write 138.2 m.

RULE 4: Zeros at the end of a number and before an implied decimal point may or may not be significant. We cannot tell whether they are part of the measurement or whether they act only to locate the unwritten but implied decimal point. Thus, 23,000 kg may have two, three, four, or five significant figures. Adding a decimal point at the end would indicate that all five numbers are significant.

Often, however, a little common sense is useful. A temperature reading of 20 °C probably has two significant figures rather than one, because one significant figure would imply a temperature anywhere from 10 °C to 30 °C and would be of little use. Similarly, a volume given as 300 mL probably has three significant figures. On the other hand, a figure of 150,000,000 km for the distance between the earth and the sun has only two or three significant figures because the distance is variable. We will see a better way to deal with this problem in the next section.

One final point about significant figures: some numbers, such as those obtained when counting objects and those that are part of a definition, are exact and effectively have an unlimited number of significant figures. Exact numbers are not measured and do not affect the number of significant figures in a calculated answer. Thus, a class might have exactly 32 students (not 31.9, 32.0, or 32.1), and 1 foot is defined to have exactly 12 inches.

WORKED EXAMPLE 1.5 Significant Figures of Measurements

How many significant figures do the following measurements have?
(a) 2730.78 m (b) 0.0076 mL (c) 3400 kg (d) 3400.0 m²

ANALYSIS All nonzero numbers are significant; the number of significant figures will then depend on the status of the zeroes in each case. (Hint: Which rule applies in each case?)

SOLUTION
(a) Six (rule 1; Zeroses in the middle of a number are significant.)
(b) Two (rule 2; Leading zeroes after a decimal point are not significant.)
(c) Two, three, or four (rule 4; Trailing zeroes with no decimal point may or may not be significant.)
(d) Five (rule 3; Trailing zeroes are significant if a decimal point is included.)

PROBLEM 1.8

How many significant figures do the following measurements have?
(a) 3.45 m (b) 0.1400 kg (c) 10.003 L (d) 35 cents

KEY CONCEPT PROBLEM 1.9

How would you record the temperature reading on the following Celsius thermometer? How many significant figures do you have in your answer?
Scientific Notation and Significant Figures

Scientific notation is particularly helpful for indicating how many significant figures are present in a number that has zeroes at the end but to the left of a decimal point. If we read, for instance, that the distance from the earth to the sun is 150,000,000 km, we do not really know how many significant figures are indicated. Some of the zeroes might be significant, or they might merely act to locate the decimal point. Using scientific notation, however, we can indicate how many of the zeroes are significant. Rewriting 150,000,000 as $1.5 \times 10^8$ indicates two significant figures, whereas writing it as $1.500 \times 10^8$ indicates four significant figures. Scientific notation is not ordinarily used for numbers that are easily written, such as 10 or 175, although it is sometimes helpful in doing arithmetic.

**Worked Example 1.6** Significant Figures and Scientific Notation

There are 1,760,000,000,000,000,000,000 molecules of sucrose (table sugar) in 1 g. Use scientific notation to express this number with four significant figures.

**ANALYSIS** Because the number is larger than 1, the exponent will be positive. You will have to move the decimal point 21 places to the left.

**SOLUTION**

The first four digits—1, 7, 6, and 0—are significant, meaning that only the first of the 19 zeroes is significant. Because we have to move the decimal point 21 places to the left to put it after the first significant digit, the answer is $1.760 \times 10^{21}$.

**Worked Example 1.7** Scientific Notation

The rhinovirus responsible for the common cold has a diameter of 20 nm or 0.000 000 020 m. Express this number in scientific notation.

**ANALYSIS** The number is smaller than 1, and so the exponent will be negative. You will have to move the decimal point eight places to the right.

**SOLUTION**

There are only two significant figures because zeroes at the beginning of a number are not significant. We have to move the decimal point eight places to the right to place it after the first digit, so the answer is $2.0 \times 10^{-8}$ m.

**Worked Example 1.8** Scientific Notation and Unit Conversions

A clinical laboratory found that a blood sample contained 0.0026 g of phosphorus and 0.000 101 g of iron.

(a) Give these quantities in scientific notation.

(b) Give these quantities in the units normally used to report them—milligrams for phosphorus and micrograms for iron.

**ANALYSIS** Is the number larger or smaller than 1? How many places do you have to move the decimal point?

**SOLUTION**

(a) $0.0026 \text{ g phosphorus} = 2.6 \times 10^{-3} \text{ g phosphorus}$

$0.000 101 \text{ g iron} = 1.01 \times 10^{-4} \text{ g iron}$
(b) We know from Table 1.6 that 1 mg = $1 \times 10^{-3}$ g, where the exponent is $-3$. Expressing the amount of phosphorus in milligrams is straightforward because the amount in grams ($2.6 \times 10^{-3}$ g) already has an exponent of $-3$. Thus, $2.6 \times 10^{-3}$ g = 2.6 mg of phosphorus.

We know from Table 1.6 that 1 μg = $1 \times 10^{-6}$ g where the exponent is $-6$. Expressing the amount of iron in micrograms thus requires that we restate the amount in grams so that the exponent is $-6$. We can do this by moving the decimal point six places to the right:

$$0.000\,010\,1\,\text{g iron} = 101 \times 10^{-6}\,\text{g iron} = 101\,\mu\text{g iron}$$

PROBLEM 1.10
Convert the following values to scientific notation:
(a) 0.058 g  
(b) 46,792 m  
(c) 0.006 072 cm  
(d) 345.3 kg

PROBLEM 1.11
Convert the following values from scientific notation to standard notation:
(a) $4.885 \times 10^4$ mg  
(b) $8.3 \times 10^{-6}$ m  
(c) $4.00 \times 10^{-2}$ m

PROBLEM 1.12
Rewrite the following numbers in scientific notation as indicated:
(a) 630,000 with five significant figures  
(b) 1300 with three significant figures  
(c) 794,200,000,000 with four significant figures

1.9 Rounding Off Numbers

Learning Objective:
• Determine the appropriate number of significant figures in a calculated result and round off numbers in calculations involving measurements.

It often happens, particularly when doing arithmetic on a pocket calculator, that a quantity appears to have more significant figures than are really justified. For example, you might calculate the gas mileage of your car by finding that it takes 11.70 gallons of gasoline to drive 278 miles:

$$\text{Mileage} = \frac{\text{Miles}}{\text{Gallons}} = \frac{278 \text{ mi}}{11.70 \text{ gal}} = 23.760\,684 \text{ mi/gal (mpg)}$$

Although the answer on a calculator has eight digits, your calculated result is really not as precise as it appears. In fact, as we will see next, your answer is good to only three significant figures and should be rounded off to 23.8 mi/gal.

How do you decide how many digits to keep? The full answer to this question is a bit complex and involves a mathematical treatment called error analysis, but for our purposes, a simplified procedure using just two rules is sufficient:

RULE 1: In carrying out a multiplication or division, the answer cannot have more significant figures than either of the original numbers. After all, if you do not know the number of miles you drove to better than three significant figures.
(278 could mean 277, 278, or 279), you certainly cannot calculate your mileage to more than the same number of significant figures.

\[
\frac{278 \text{ mi}}{11.70 \text{ gal}} = 23.8 \text{ mi/gal}
\]

**RULE 2:** In carrying out an addition or subtraction, the answer cannot have more digits after the decimal point than either of the original numbers. For example, if you have 3.18 L of water and you add 0.013 15 L more, you now have 3.19 L.

If you do not know the volume you started with past the second decimal place (it could be 3.17, 3.18, or 3.19), you cannot know the total of the combined volumes past the same decimal place.

<table>
<thead>
<tr>
<th>Volume of water at start</th>
<th>Volume of water added</th>
<th>Total volume of water</th>
</tr>
</thead>
<tbody>
<tr>
<td>3.18?? L</td>
<td>+ 0.013 15 L</td>
<td>3.19?? L</td>
</tr>
</tbody>
</table>

Two digits after decimal point

Five digits after decimal point

Two digits after decimal point

If a calculation has several steps, it is generally best to round off at the end after all the steps have been carried out, keeping the number of significant figures determined by the least precise number in your calculations. Once you decide how many digits to retain for your answer, the rules for rounding off numbers are straightforward:

**RULE 1:** If the first digit you remove is four or less, drop it and all following digits.

Thus, 2.4271 becomes 2.4 when rounded off to two significant figures because the first of the dropped digits (2) is four or less.

**RULE 2:** If the first digit you remove is five or greater, round the number up by adding a 1 to the digit to the left of the one you drop. Thus, 4.5832 becomes 4.6 when rounded off to two significant figures because the first of the dropped digits (8) is five or greater.

---

**Worked Example 1.9 Significant Figures and Calculations: Addition/Subtraction**

Suppose that you weigh 124 lb before dinner. How much will you weigh after dinner if you eat 1.884 lb of food?

**ANALYSIS** When performing addition or subtraction, the number of significant figures you report in the final answer is determined by the number of digits in the least precise number in the calculation.

**SOLUTION**

Your after-dinner weight is found by adding your original weight to the weight of the food consumed:

\[
\begin{align*}
124 \text{ lb} + 1.884 \text{ lb} &= 125.884 \text{ lb (Unrounded)}
\end{align*}
\]

Because the value of your original weight has no significant figures after the decimal point, your after-dinner weight also must have no significant figures after the decimal point. Thus, 125.884 lb must be rounded off to 126 lb.
**Worked Example 1.10** Significant Figures and Calculations: Multiplication / Division

To make currant jelly, 13.75 cups of sugar was added to 18 cups of currant juice. How much sugar was added per cup of juice?

**ANALYSIS** For calculations involving multiplication or division, the final answer cannot have more significant figures than either of the original numbers.

**SOLUTION**

The quantity of sugar must be divided by the quantity of juice:

\[ \frac{13.75 \text{ cups sugar}}{18 \text{ cups juice}} = 0.763\ 888\ 89 \text{ cup sugar per cup juice} \]

The number of significant figures in the answer is limited to two by the quantity 18 cups in the calculation and must be rounded to 0.76 cup of sugar per cup of juice.

**PROBLEM 1.13**

Round off the following quantities to the indicated number of significant figures:

(a) 2.304 g (three significant figures)

(b) 188.3784 mL (five significant figures)

(c) 0.008 87 L (one significant figure)

(d) 1.000 39 kg (four significant figures)

**PROBLEM 1.14**

Carry out the following calculations, rounding each result to the correct number of significant figures:

(a) 4.87 mL + 46.0 mL

(b) 3.4 \times 0.023 g

(c) 19.333 m - 7.4 m

(d) 55 mg - 4.671 mg + 0.894 mg

(e) 62,911 ÷ 611

### 1.10 Problem Solving: Unit Conversions and Estimating Answers

**Learning Objective:**
- Use the factor-label method [conversion factors] to solve a problem and check the result to ensure that it makes sense chemically and physically.

Many activities in the laboratory and in medicine—measuring, weighing, preparing solutions, and so forth—require converting a quantity from one unit to another. For example: "These pills contain 1.3 grains of aspirin, but I need 200 mg. Is one pill enough?" Converting between units is not mysterious; we all do it every day. If you run nine laps around a 400 m track, for instance, you have to convert between the distance unit "lap" and the distance unit "meter" to find that you have run 3600 m (9 laps times 400 m/lap). If you want to find how many miles, you have to convert again to find that 3600 m = 2.237 mi.

The simplest way to carry out calculations involving different units is to use the **factor-label method.** In this method, a quantity in one unit is converted into an equivalent quantity in a different unit by using a **conversion factor** that expresses the relationship between units:

\[ \text{Starting quantity} \times \text{Conversion factor} = \text{Equivalent quantity} \]
As an example, we learned from Table 1.8 that $1 \text{ km} = 0.6214 \text{ mi}$. Writing this relationship as a fraction restates it in the form of a conversion factor, either kilometers per mile or miles per kilometer.

Since $1 \text{ km} = 0.6214 \text{ mi}$, then:

$$\frac{1 \text{ km}}{0.6214 \text{ mi}} = 1 \quad \text{or} \quad \frac{0.6214 \text{ mi}}{1 \text{ km}} = 1$$

Note that this and all other conversion factors are numerically equal to 1 because the value of the quantity above the division line (the numerator) is equal in value to the quantity below the division line (the denominator). Thus, multiplying by a conversion factor is equivalent to multiplying by 1 and so does not change the value of the quantity being multiplied.

The key to the factor-label method of problem solving is that units are treated like numbers and can thus be multiplied and divided (though not added or subtracted) just as numbers can. When solving a problem, the idea is to set up an equation so that all unwanted units cancel, leaving only the desired units. Usually, it is best to start by writing what you know and then manipulating that known quantity. For example, if you know there are 26.22 mi in a marathon and want to find how many kilometers that is, you could write the distance in miles and multiply by the conversion factor in kilometers per mile. The unit “mi” cancels because it appears both above and below the division line, leaving “km” as the only remaining unit.

$$26.22 \text{ mi} \times \frac{1 \text{ km}}{0.6214 \text{ mi}} = 42.20 \text{ km}$$

The factor-label method gives the right answer only if the equation is set up so that the unwanted unit (or units) cancels. If the equation is set up in any other way, the units will not cancel and you will not get the right answer. Thus, if you selected the incorrect conversion factor (miles per kilometer) for the above problem, you would end up with an incorrect answer expressed in meaningless units:

$$26.22 \text{ mi} \times \frac{0.6214 \text{ mi}}{1 \text{ km}} = 16.29 \text{ mi}^2 \text{ km}^{-1} \text{ Incorrect}$$

The main drawback to using the factor-label method is that it is possible to get an answer without really understanding what you are doing. It is therefore best when solving a problem to first think through a rough estimate, or ballpark estimate, as a check on your work. If your ballpark estimate is not close to the final calculated solution, there is a misunderstanding somewhere and you should think the problem through again. If, for example, you came up with the answer 5.3 cm$^3$ when calculating the volume of a human cell, you should realize that such an answer could not possibly be right. Cells are too tiny to be distinguished with the naked eye, but a volume of 5.3 cm$^3$ is about the size of a walnut. The Worked Examples 1.11, 1.12, and 1.13 at the end of this section show how to estimate the answers to simple unit-conversion problems.

The factor-label method and the use of ballpark estimates are techniques that will help you solve problems of many kinds, not just unit conversions. Problems sometimes seem complicated, but you can usually sort out the complications by analyzing the problem properly:

**STEP 1:** Identify the information given, including units.

**STEP 2:** Identify the information needed in the answer, including units.
STEP 3: Find the relationships between the known information and unknown answer, and plan a series of steps, including conversion factors, for getting from one to the other.

STEP 4: Solve the problem.

BALLPARK CHECK Make a ballpark estimate at the beginning and check it against your final answer to be sure the value and the units of your calculated answer are reasonable.

**Worked Example 1.11 Factor Labels: Unit Conversions**

Write conversion factors for the following pairs of units (use Tables 1.7–1.9):

(a) Deciliters and milliliters
(b) Pounds and grams

**ANALYSIS** Start with the appropriate equivalency relationship and rearrange to form conversion factors.

**SOLUTION**

(a) Since $1 \text{ dL} = 0.1 \text{ L}$ and $1 \text{ mL} = 0.001 \text{ L}$, then $1 \text{ dL} = \left( \frac{1 \text{ mL}}{0.001 \text{ L}} \right) = 100 \text{ mL}$. The conversion factors are

\[
\frac{1 \text{ dL}}{100 \text{ mL}} \quad \text{and} \quad \frac{100 \text{ mL}}{1 \text{ dL}}
\]

(b) $\frac{1 \text{ lb}}{454 \text{ g}}$ and $\frac{454 \text{ g}}{1 \text{ lb}}$

**Worked Example 1.12 Factor Labels: Unit Conversions**

(a) Convert 0.75 lb to grams.
(b) Convert 0.50 qt to deciliters.

**ANALYSIS** Start with conversion factors and set up equations so that units cancel appropriately.

**SOLUTION**

(a) Select the conversion factor from Worked Example 1.9(b) so that the “lb” units cancel and “g” remains:

\[
0.75 \text{ lb} \times \frac{454 \text{ g}}{1 \text{ lb}} = 340 \text{ g}
\]

(b) In this, as in many problems, it is convenient to use more than one conversion factor. As long as the unwanted units cancel correctly, two or more conversion factors can be strung together in the same calculation. In this case, we can convert first between quarts and milliliters and then between milliliters and deciliters:

\[
0.50 \text{ qt} \times \frac{946.4 \text{ mL}}{1 \text{ qt}} \times \frac{1 \text{ dL}}{100 \text{ mL}} = 4.7 \text{ dL}
\]

**Worked Example 1.13 Factor Labels: Unit Conversions**

A child is 21.5 inches long at birth. How long is this in centimeters?

**ANALYSIS** This problem calls for converting from inches to centimeters, so we will need to know how many centimeters are in an inch and how to use this information as a conversion factor.

—continued on next page
BALLPARK ESTIMATE It takes about 2.5 cm to make 1 in., and so it should take two and a half times as many centimeters to make a distance equal to approximately 20 in., or about 20 in. × 2.5 = 50 cm.

**SOLUTION**

**STEP 1:** Identify given information.
**STEP 2:** Identify answer and units.
**STEP 3:** Identify conversion factor.
**STEP 4:** Solve. Multiply the known length (in inches) by the conversion factor so that units cancel, providing the answer (in centimeters)

Length = 21.5 in.
Length = ?? cm

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

\[ 21.5 \text{ in.} \times \frac{2.54 \text{ cm}}{1 \text{ in.}} = 54.6 \text{ cm} \]
(Rounded off from 54.61)

BALLPARK CHECK How does this value compare with the ballpark estimate we made at the beginning? Are the final units correct? 54.6 cm is close to our original estimate of 50 cm.

---

**Worked Example 1.14 Factor Labels: Concentration to Mass**

A patient requires an injection of 0.012 g of a pain killer available as a 15 mg/mL solution. How many milliliters of solution should be administered?

**ANALYSIS** Knowing the amount of pain killer in 1 mL allows us to use the concentration as a conversion factor to determine the volume of solution that would contain the desired amount.

**BALLPARK ESTIMATE** One milliliter contains 15 mg of the pain killer, or 0.015 g. Since only 0.012 g is needed, a little less than 1.0 mL should be administered.

**SOLUTION**

**STEP 1:** Identify known information.
**STEP 2:** Identify answer and units.
**STEP 3:** Identify conversion factors. Two conversion factors are needed. First, g must be converted to mg. Once we have the mass in mg, we can calculate mL using the conversion factor of mL/mg.
**STEP 4:** Solve. Starting from the desired dosage, we use the conversion factors to cancel units, obtaining the final answer in mL.

\[ \text{Dosage} = 0.012 \text{ g} \]
\[ \text{Concentration} = 15 \text{ mg/mL} \]
\[ \text{Volume to administer} = ?? \text{ mL} \]

\[ 1 \text{ mg} = \frac{0.001 \text{ g}}{1 \text{ mg}} \]
\[ 15 \text{ mg/mL} \Rightarrow \frac{1 \text{ mL}}{15 \text{ mg}} \]

\[ (0.012 \text{ g}) \left( \frac{1 \text{ mg}}{0.001 \text{ g}} \right) \left( \frac{1 \text{ mL}}{15 \text{ mg}} \right) = 0.80 \text{ mL} \]

BALLPARK CHECK Consistent with our initial estimate of a little less than 1 mL.

---

**Worked Example 1.15 Factor Labels: Multiple Conversion Calculations**

Administration of digitalis to control atrial fibrillation in heart patients must be carefully regulated because even a modest overdose can be fatal. To take differences between patients into account, dosages are sometimes prescribed in micrograms per kilogram of body weight (µg/kg). Thus, two people may differ greatly in weight, but both will receive the proper dosage. At a dosage of 20 µg/kg body weight, how many milligrams of digitalis should a 160 lb patient receive?

**ANALYSIS** Knowing the patient's body weight (in kg) and the recommended dosage (in µg/kg), we can calculate the appropriate amount of digitalis.

**BALLPARK ESTIMATE** Since a kilogram is roughly equal to 2 lb, a 160 lb patient has a mass of about 80 kg. At a dosage of 20 µg/kg, an 80 kg patient should receive 80 × 20 µg, or about 1600 µg of digitalis, or 1.6 mg.
SOLUTION

STEP 1: Identify known information.

STEP 2: Identify answer and units.

STEP 3: Identify conversion factors. Two conversions are needed. First, convert the patient’s weight in pounds to weight in kg. The correct dose can then be determined based on μg digitalis/kg of body weight. Finally, the dosage in μg is converted to mg.

STEP 4: Solve. Use the known information and the conversion factors so that units cancel, obtaining the answer in mg.

BALLPARK CHECK Close to our estimate of 1.6 mg.

PROBLEM 1.15
Write appropriate conversion factors and carry out the following conversions:
(a) 16.0 oz = ? g  
(b) 2500 mL = ? L  
(c) 99.0 L = ? qt

PROBLEM 1.16
Convert 0.840 qt to milliliters in a single calculation using more than one conversion factor.

PROBLEM 1.17
A patient is to receive 20 mg of methimazole, a drug used to treat hyperthyroid conditions. The drug is dissolved in solution containing 8 mg/mL. What volume of solution should be administered?

PROBLEM 1.18
Calculate the dosage in milligrams per kilogram body weight for a 135 lb adult who takes two aspirin tablets containing 0.324 g of aspirin each. Calculate the dosage for a 40 lb child who also takes two aspirin tablets.

1.11 Temperature, Heat, and Energy

Learning Objectives:
• Define the relationship between temperature and heat energy and convert temperatures between various temperature scales.
• Use temperature and specific heat to evaluate the flow of heat/energy in matter.

All chemical reactions are accompanied by a change in energy, which is defined in scientific terms as the capacity to do work or supply heat (Figure 1.8). Detailed discussion of the various kinds of energy will be included in Chapter 7, but for now we will look at the various units used to describe energy and heat, and how heat energy can be gained or lost by matter.

Temperature, the measure of the amount of heat energy in an object, is commonly reported either in Fahrenheit (°F) or Celsius (°C) units. The SI unit for reporting temperature, however, is the kelvin (K). (Note that we say only “kelvin,” not “degrees kelvin.”)
The kelvin and the Celsius degree are the same size—both are 1/100 of the interval between the freezing point of water and the boiling point of water at atmospheric pressure. Thus, a change in temperature of 1 °C is equal to a change of 1 K. The only difference between the Kelvin and Celsius temperature scales is that they have different zero points. The Celsius scale assigns a value of 0 °C to the freezing point of water, but the Kelvin scale assigns a value of 0 K to the coldest possible temperature, sometimes called absolute zero, which is equal to -273.15 °C. Thus, 0 K = -273.15 °C, and +273.15 K = 0 °C. For example, a warm spring day with a temperature of 25 °C has a Kelvin temperature of 298 K (for most purposes, rounding off to 273 is sufficient).  

\[
\begin{align*}
\text{Temperature in K} &= \text{Temperature in } °C + 273.15 \\
\text{Temperature in } °C &= \text{Temperature in K} - 273.15
\end{align*}
\]

For practical applications in medicine and clinical chemistry, the Fahrenheit and Celsius scales are used almost exclusively. The Fahrenheit scale defines the freezing point of water as 32 °F and the boiling point of water as 212 °F, whereas 0 °C and 100 °C are the freezing and boiling points of water on the Celsius scale. Thus, it takes 180 °F to cover the same range encompassed by only 100 °C, and a Celsius degree is therefore exactly \(\frac{180}{100} = \frac{9}{5} = 1.8\) times as large as a Fahrenheit degree. In other words, a change in temperature of 1.0 °C is equal to a change of 1.8 °F. Figure 1.9 gives a comparison of all three scales.

Converting between the Fahrenheit and Celsius scales is similar to converting between different units of length or volume, but is a bit more complex because two corrections need to be made—one to adjust for the difference in degree size and one to adjust for the different zero points. The degree-size correction is made by using the relationship 1 °C = 1.8 °F and 1 °F = \(\frac{1}{1.8}\) °C. The zero-point correction is made by

\(\text{Fahrenheit } (°F)\)

\(\text{Celsius } (°C)\)

\(\text{Kelvin } (K)\)

\(\text{Boiling water}\)

\(\text{Body temperature}\)

\(\text{Room temperature}\)

\(\text{Freezing water}\)

\(\text{A cold day}\)

\(\text{"Crossover point"}\)

\(\text{"Absolute zero"}\)

\(212\)

\(98.6\)

\(68\)

\(32\)

\(-4\)

\(-40\)

\(-273.15\)

\(100\)

\(37\)

\(20\)

\(0\)

\(-20\)

\(-40\)

\(0\)

\(373.15\)

\(310\)

\(293\)

\(273.15\)

\(253\)

\(233\)

\(\text{Figure 1.9}\)

A comparison of the Fahrenheit, Celsius, and Kelvin temperature scales.

One Celsius degree is 1.8 times the size of one Fahrenheit degree.
CHEMISTRY IN ACTION

Temperature-Sensitive Materials

The physical properties of many materials change with the ambient temperature. Substances known as thermochromic materials change color as their temperature increases, and they change from the liquid phase to a semicrystalline-ordered state. These "liquid crystals" can be incorporated into plastics or paints and can be used to monitor temperature. For example, some meat packaging now includes a temperature strip that darkens when the meat is stored above a certain temperature, which makes the meat unsafe to eat. Hospitals and other medical facilities now routinely use strips that, when placed under the tongue or applied to the forehead, change color to indicate the patient's body temperature.

Other temperature-sensitive materials, called shape-memory alloys (SMAs), can be bent out of shape and will recover their original shape when heated above a certain temperature. These materials have many practical and clinical applications, including orthodontic wires that do not need to be tightened. The SMA is bent to fit into the orthodontic form, but once in the mouth its temperature increases and it contracts back to its original shape, applying constant force to align the teeth. SMAs are also used in stents. A collapsed stent can be inserted into an artery or vein; at body temperature the stent expands to its original shape and provides support for the artery or vein, improving blood flow.

CIA Problem 1.4 A thermochromic plastic chip included in a shipping container for beef undergoes an irreversible color change if the storage temperature exceeds 28 °F. What is this temperature on the Celsius and Kelvin scales?

CIA Problem 1.5 A temperature-sensitive bath toy undergoes several color changes in the temperature range from 37 °C to 47 °C. What is the corresponding temperature range on the Fahrenheit scale?

remembering that the freezing point is higher by 32 on the Fahrenheit scale than on the Celsius scale. These corrections are incorporated into the following formulas, which show the conversion methods:

Celsius to Fahrenheit: °F = \left( \frac{1.8 \, ^\circ C}{^\circ F} \times ^\circ C \right) + 32 ^\circ F

Fahrenheit to Celsius: °C = \frac{^\circ C}{1.8 \, ^\circ F} \times (^\circ F - 32 ^\circ F)

Energy is represented in SI units by the unit joule (J; pronounced "jool"), but the metric unit calorie (cal) is still widely used in medicine. In this text we will present energy values in both units. One calorie is the amount of heat necessary to raise the temperature of 1 g of water by 1 °C. A kilocalorie (kcal), often called a large calorie (Cal) or food calorie by nutritionists, equals 1000 cal:

1000 cal = 1 kcal
1 kcal = 4.184 J
1000 J = 1 kJ
1 kcal = 4.184 kJ

Not all substances have their temperatures raised to the same extent when equal amounts of heat energy are added. One calorie raises the temperature of 1 g of water by 1 °C but raises the temperature of 1 g of iron by 10 °C. The amount of heat needed
Specific heat The amount of heat that will raise the temperature of 1 g of a substance by 1 °C.

Specific heat (calories) = \( \frac{\text{calories}}{\text{grams} \times \text{°C}} \)

Specific heats vary greatly from one substance to another, as shown in Table 1.10. The specific heat of water, 1.00 cal/(g • °C) or 4.184 J/g °C, is higher than that of most other substances, which means that a large transfer of heat is required to change the temperature of a given amount of water by a given number of degrees. One consequence is that the human body, which is about 60% water, is able to withstand changing outside conditions.

Knowing the mass and specific heat of a substance makes it possible to calculate how much heat must be added or removed to accomplish a given temperature change, as shown in Worked Example 1.17.

**Worked Example 1.17 Specific Heat: Mass, Temperature, and Energy**

Taking a bath might use about 95 kg of water. How much energy (in calories and Joules) is needed to heat the water from a cold 15 °C to a warm 40 °C?

**ANALYSIS** From the amount of water being heated (95 kg) and the amount of the temperature change (40 °C − 15 °C = 25 °C), the total amount of energy needed can be calculated by using specific heat [1.00 cal/(g • °C)] as a conversion factor.

**BALLPARK ESTIMATE** The water is being heated by 25 °C (from 15 °C to 40 °C), and it therefore takes 25 cal to heat each gram. The tub contains nearly 100,000 g (95 kg is 95,000 g), and so it takes about 25 × 100,000 cal, or 2,500,000 cal, to heat all the water in the tub.
SOLUTION

STEP 1: Identify known information.

STEP 2: Identify answer and units.

STEP 3: Identify conversion factors. The amount of energy (in cal) can be calculated using the specific heat of water (cal/g °C), and it will depend on both the mass of water (in g) to be heated and the total temperature change (in °C). In order for the units in specific heat to cancel correctly, the mass of water must first be converted from kg to g.

STEP 4: Solve. Starting with the known information, use the conversion factors to cancel unwanted units.

BALLPARK CHECK Close to our estimate of $2.5 \times 10^6$ cal.

PROBLEM 1.19

The highest land temperature ever recorded was 136 °F in Al Aziziyah, Libya, on September 13, 1922. What is this temperature on the Kelvin scale?

PROBLEM 1.20

A patient exhibits a temperature of 39 °C. What is the body temperature of the patient in °F?

PROBLEM 1.21

Assuming that Coca-Cola has the same specific heat as water, how much energy in calories is removed when 350 g of Coca-Cola (about the contents of one 12 oz can) is cooled from room temperature (25 °C) to refrigerator temperature (3 °C)?

PROBLEM 1.22

What is the specific heat of aluminum if it takes 161 cal (674 J) to raise the temperature of a 75 g aluminum bar by 10.0 °C?

### 1.12 Density and Specific Gravity

**Learning Objective:**

- Define density and specific gravity and use these quantities in mass/volume calculations.

One further physical quantity that we will take up in this chapter is **density**, which relates the mass of an object to its volume. Density is usually expressed in units of grams per cubic centimeter (g/cm$^3$) for solids and grams per milliliter (g/mL) for liquids. Thus, if we know the density of a substance, we know both the mass of a given volume and the volume of a given mass. The densities of some common materials are listed in Table 1.11.

\[
\text{Density} = \frac{\text{Mass (g)}}{\text{Volume (mL or cm}^3)}
\]

Although most substances contract when cooled and expand when heated, water behaves differently. Water contracts when cooled from 100 °C to 3.98 °C but below this temperature it begins to **expand** again. The density of liquid water is at its maximum of 1.0000 g/mL at 3.98 °C but decreases to 0.999 87 g/mL at 0 °C. When freezing occurs, the density drops still further to a value of 0.917 g/cm$^3$ for ice at 0 °C. Since a
A Galileo thermometer contains several weighted bulbs that rise or fall as the density of the liquid changes with temperature.

**Specific gravity** The density of a substance divided by the density of water at the same temperature.

A The Galileo thermometer contains several weighted bulbs that rise or fall as the density of the liquid changes with temperature.

Specific gravity The density of a substance divided by the density of water at the same temperature.

Table 1.11 Densities of Some Common Materials at 25°C

<table>
<thead>
<tr>
<th>Substance</th>
<th>Density* (g/mL)</th>
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<tr>
<td>Gases</td>
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<td>Helium</td>
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<td>Liquids</td>
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<td>Human fat</td>
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<td>1.0000</td>
<td>Cork</td>
<td>0.22–0.26</td>
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<tr>
<td>Urine</td>
<td>1.003–1.030</td>
<td>Table sugar</td>
<td>1.59</td>
</tr>
<tr>
<td>Blood plasma</td>
<td>1.027</td>
<td>Balsa wood</td>
<td>0.12</td>
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</table>

*Densities are in g/cm³ for solids and g/mL for liquids and gases. As noted in Section 1.7, 1 mL = 1 cm³.

Knowing the density of a liquid is useful because it is often easier to measure a liquid's volume rather than its mass. Suppose, for example, that you need 1.50 g of ethanol. Rather than use a dropper to weigh out exactly the right amount, it would be much easier to look up the density of ethanol (0.7893 g/mL at 20°C) and measure the correct volume (1.90 mL) with a syringe or graduated cylinder. Thus, density acts as a conversion factor between mass (g) and volume (mL).

\[
1.50 \text{ g ethanol} \times \frac{1 \text{ mL ethanol}}{0.7893 \text{ g ethanol}} = 1.90 \text{ mL ethanol}
\]

For many purposes, ranging from winemaking to medicine, it is more convenient to use specific gravity than density. The specific gravity (sp gr) of a substance (usually a liquid) is simply the density of the substance divided by the density of water at the same temperature. Because all units cancel, specific gravity is unitless:

\[
\text{Specific gravity} = \frac{\text{Density of substance (g/mL)}}{\text{Density of water at the same temperature (g/mL)}}
\]

At typical temperatures, the density of water is very close to 1 g/mL. Thus, the specific gravity of a substance is numerically equal to its density and is used in the same way.

The specific gravity of a liquid can be measured using an instrument called a hydrometer, which consists of a weighted bulb on the end of a calibrated glass tube, as shown in Figure 1.10. The depth to which the hydrometer sinks in a fluid indicates the fluid's specific gravity: the lower the bulb sinks, the lower the specific gravity of the fluid.

In medicine, a hydrometer called a urinometer is used to indicate the amount of solids dissolved in urine. Although the specific gravity of normal urine is about 1.003–1.030, conditions such as diabetes mellitus or a high fever cause an abnormally high urine specific gravity, indicating either excessive elimination of solids or decreased elimination of water. Abnormally low specific gravity is found in individuals using diuretics—drugs that increase water elimination.

**Worked Example 1.18 Density: Mass-to-Volume Conversion**

What volume of isopropyl alcohol (rubbing alcohol) would you use if you needed 25.0 g? The density of isopropyl alcohol is 0.7855 g/mL at 20°C.

**ANALYSIS** The known information is the mass of isopropyl alcohol needed (25.0 g). The density (0.7855 g/mL) acts as a conversion factor between mass and the unknown volume of isopropyl alcohol.
BALLPARK ESTIMATE Because 1 mL of isopropyl alcohol contains only 0.7885 g of the alcohol, obtaining 1 g of alcohol requires almost 20% more than 1 mL, or about 1.2 mL. Therefore, a volume of about $25 \times 1.2 \text{ mL} = 30 \text{ mL}$ is needed to obtain 25 g of alcohol.

SOLUTION
STEP 1: Identify known information.

Mass of rubbing alcohol = 25.0 g
Density of rubbing alcohol = 0.7855 g/mL
Volume of rubbing alcohol = ?? mL

Density = g/mL $\rightarrow \frac{1}{\text{density}} = \frac{\text{mL}}{g}$

STEP 2: Identify answer and units.

$25.0 \text{ g alcohol} \times \frac{1 \text{ mL alcohol}}{0.7855 \text{ g alcohol}} = 31.8 \text{ mL alcohol}$

STEP 3: Identify conversion factors. Starting with the mass of isopropyl alcohol (in g), the corresponding volume (in mL) can be calculated using density (g/mL) as the conversion factor.

STEP 4: Solve. Starting with the known information, set up the equation with conversion factors so that unwanted units cancel.

BALLPARK CHECK Our estimate was 30 mL.

PROBLEM 1.23
A sample of pumice, a porous volcanic rock, weighs 17.4 grams and has a volume of 27.3 cm$^3$. If this sample is placed in a container of water, will it sink or will it float? Explain.

PROBLEM 1.24
Chloroform, once used as an anesthetic agent, has a density of 1.474 g/mL. What volume would you use if you needed 12.37 g?

PROBLEM 1.25
The sulfuric acid solution in an automobile battery typically has a specific gravity of about 1.27. Is battery acid more dense or less dense than pure water?

CHEMISTRY IN ACTION

A Measurement Example: Obesity and Body Fat
At the beginning of the chapter, we mentioned that some fat is good, but how much is too much and what are the health risks of too much body fat? The impacts of obesity include significant adverse health effects—heart disease, stroke, type 2 diabetes, and certain types of cancer—as well as annual medical costs related to obesity in excess of US$147 billion. Of particular concern is childhood obesity; the percentage of children aged 6–11 who were obese increased from 7% in 1980 to nearly 18% in 2012. For teenagers (ages 12–19), the increase was even more dramatic, from 5% to nearly 21% over the same period. In 2012, more than one-third of children and adolescents were overweight or obese.

At the beginning of the chapter, we learned that obesity is ancessive amount of body fat and one way to measure body fat is

A person's percentage body fat can be estimated by measuring the thickness of the fat layer under the skin.
through buoyancy testing. But obesity is also defined by reference to body mass index (BMI), which is equal to a person’s mass in kilograms divided by the square of his or her height in meters. BMI can also be calculated by dividing a person’s weight in pounds by the square of her or his height in inches multiplied by 703. For instance, someone 5 ft 7 in. (67 inches; 1.70 m) tall weighing 147 lb (66.7 kg) has a BMI of 23:

\[
\text{BMI} = \frac{\text{weight (kg)}}{\text{height (m)}^2} \quad \text{or} \quad \frac{\text{weight (lb)}}{\text{height (in.)}^2} \times 703
\]

A BMI of 25 or above is considered overweight, and a BMI of 30 or above is obese. By these standards, approximately 61% of the U.S. population is overweight. Health professionals are concerned by the rapid rise in obesity in the United States because of the link between BMI and health problems. Many reports have documented the correlation between health and BMI, including a recent study on more than 1 million adults. The lowest death risk from any cause, including cancer and heart disease, is associated with a BMI between 22 and 24. Risk increases steadily as BMI increases, more than doubling for a BMI above 29.

An individual’s percentage of body fat is most easily measured by the skinfold-thickness method. The skin at several locations on the arm, shoulder, and waist is pinched, and the thickness of the fat layer beneath the skin is measured with calipers. Comparing the measured results to those in a standard table gives an estimation of percentage body fat. As an alternative to skinfold measurement, a more accurate assessment of body fat can be made by underwater immersion, or buoyancy testing as we learned at the beginning of the chapter.

There is good news—campaigns to increase awareness of the negative effects of obesity and a renewed emphasis on healthy eating and exercise are producing results. The prevalence of obesity in children aged 2–5 years old decreased from 13.9% in 2004 to 8.4% in 2012.

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Body Mass Index (numbers in boxes)

CIA Problem 1.6 Calculate the BMI for an individual who is

(a) 5 ft 1 in. tall and weighs 155 lb
(b) 5 ft 11 in. tall and weighs 170 lb
(c) 6 ft 3 in. tall and weighs 195 lb

Which of these individuals is likely to have increased health risks?

CIA Problem 1.7 Liposuction is a technique for removing fat deposits from various areas of the body. How many liters of fat would have to be removed to result in a 5.0 lb weight loss? The density of human fat is 0.94 g/mL

SUMMARY REVISITING THE LEARNING OBJECTIVES

- Identify properties of matter and differentiate between chemical and physical changes. Matter is anything that has mass and occupies volume—that is, anything physically real. A property is any characteristic that can be used to describe or identify something; physical properties can be seen or measured without changing the chemical identity of the substance (i.e., color, melting point), while chemical properties can only be seen or measured when the substance undergoes a chemical change, such as a chemical reaction (see Problems 33–35).

- Identify the three states of matter and describe their properties. Matter can be classified by its physical state as solid, liquid, or gas. A solid has a definite volume and shape, a liquid has a definite volume but indefinite shape, and a gas has neither a definite volume nor a definite shape (see Problems 26, 36–39, and 41).

- Distinguish between mixtures and pure substances and classify pure substances as elements or compounds. Matter can also be classified by composition as being either pure or a mixture. Every pure substance is either an element or a chemical compound. Elements are fundamental substances that cannot be chemically changed into anything simpler. A chemical compound, by contrast, can be broken down by chemical change into simpler substances. Mixtures are composed of two or more pure substances and can be separated into component parts by physical means (see Problems 40–43, 53, and 92).

- Identify the symbols and names of the common elements. Elements are represented by one- or two-letter symbols, such as H for hydrogen, Ca for calcium, Al for aluminum, and so on. Most symbols are the first one or two letters of the element name, but some symbols are
derived from Latin names—Na [sodium], for example (see Problems 26, 27, 44–52, and 114).

- **Identify a chemical change as a chemical reaction.** A chemical reaction is a symbolic representation of a chemical change. The starting materials (reactants) are on the left, the final materials (products) are on the right. An arrow is used to indicate a chemical change as reactants are converted to products; reaction conditions written above/below the arrow. The reactants and products are identified using chemical symbols to represent the elements or compounds, and their physical states are indicated using appropriate abbreviations (s, l, g, aq) (see Problems 34, 35, 42, 43, and 92).

- **Write very large or very small numbers using scientific notation or units with appropriate numerical prefixes.** Measurements of small and large quantities are usually written in scientific notation as the product of a number between 1 and 10, times a power of 10. Numbers greater than 10 have a positive exponent, and numbers less than 1 have a negative exponent. For example, 3562 = 3.562 × 10^3, and 0.00391 = 3.91 × 10^{-3} (see Problems 55–58 and 67).

- **Name and correctly use the metric and SI units of measurement for mass, length, volume, and temperature, and convert units appropriately.** A property that can be measured is called a physical quantity and is described by both a number and a label, or unit. The preferred units are either those of the International System of Units (SI units) or the metric system. Mass, the amount of matter an object contains, is measured in kilograms (kg) or grams (g). Length is measured in meters (m). Volume is measured in cubic meters (m^3) in the SI system and in liters (L) or milliliters (mL) in the metric system. Temperature is measured in kelvins (K) in the SI system and in degrees Celsius (°C) in the metric system. A measurement in one unit can be converted to another unit by multiplying by a conversion factor that expresses the exact relationship between the units. The conversion factor should be arranged so that the starting unit is canceled and the desired unit is carried over to the answer (see Problems 54–56, 58, 67–77, 93, 94, 96–98, 103, 106, 110, and 115).

- **Use significant figures and scientific notation to represent the precision of a measurement.** When measuring physical quantities or using them in calculations, it is important to indicate the exactness of the measurement by using significant figures or numbers to represent those decimal places that are known with certainty, plus one additional decimal place indicating the point at which the measured value is uncertain. For example, a mass that was recorded as 15.34 g has an uncertainty in the last decimal place of ± 0.01 g (see Problems 29–31, 59, 60, and 97).

- **Determine the appropriate number of significant figures in a calculated result and round off numbers in calculations involving measurements.** For multiplication and division, the number of significant figures in the calculated result is the same as the number with the fewest significant figures involved in the calculation. For addition and subtraction, the number of significant figures in the calculated result is determined by the least precise decimal place for the numbers involved in the calculation. If necessary, the calculated result is rounded off to obtain the final answer to the correct number of significant figures (see Problems 42, 43, 59–66, 101, and 105).

- **Use the factor-label method [conversion factors] to solve a problem and check the result to ensure that it makes sense chemically and physically.** Problems are best solved by applying the factor-label method in which units can be multiplied and divided just as numbers can. The idea is to set up an equation so that all unwanted units cancel, leaving only the desired units. Usually it is best to start by identifying the known and needed information, then decide how to convert the known information to the answer, and finally check to make sure the answer is reasonable both chemically and physically (see Problems 67–77, 96–101, 103–107, 112, and 113).

- **Define the relationship between temperature and heat energy and be able to convert temperatures between various temperature scales.** Temperature is a measure of the amount of heat energy in an object. Temperature is reported using the Fahrenheit, Celsius, and Kelvin scales, with conversions between scales as shown on pages 28–29 (see Problems 78, 91, 102, and 111).

- **Use temperature and specific heat to evaluate the flow of heat/energy in matter.** Heat flows from a hot object to a cold object. The specific heat of a substance is the amount of heat necessary to raise the temperature of 1 g of the substance by 1 °C (1 cal/g°C or 4.184 J/g°C). Water has an unusually high specific heat, which helps our bodies to maintain an even temperature (see Problems 79–84, 95, 104, 107, and 108).

- **Define density and specific gravity and to use these quantities in mass/volume calculations.** Density, the physical property that relates mass to volume, is expressed in units of grams per milliliter (g/mL) for a liquid or grams per cubic centimeter (g/cm^3) for a solid. The specific gravity of a liquid is the density of the liquid divided by the density of water at the same temperature. Because the density of water is approximately 1 g/mL, specific gravity and density have the same numerical value (see Problems 28, 32, 85–90, 102, 107, 109, 110, 115, and 118).

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**KEY WORDS**

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### UNDERSTANDING KEY CONCEPTS

The problems in this section are intended as a bridge between the Chapter Summary and the Additional Problems that follow. Primarily visual in nature, they are designed to help you test your grasp of the chapter's most important principles before attempting to solve quantitative problems. Answers to all Key Concept Problems are at the end of the book following the appendixes.

**1.26** The six elements in blue at the far right of the periodic table are gases at room temperature. The red elements in the middle of the table are the so-called coinage metals. Identify each of these elements using the periodic table inside the front cover of this book.

**1.27** Identify the three elements indicated on the following periodic table. Do an Internet search to identify the common sources of these elements and some of their common uses or applications.

**1.28** The radioactive element indicated on the following periodic table is used in smoke detectors. Identify it.
1.29  (a) What is the specific gravity of the following solution?  
(b) How many significant figures does your answer have?  
(c) Is the solution more dense or less dense than water?

1.30  Assume that you have two graduated cylinders, one with a capacity of 5 mL (a) and the other with a capacity of 50 mL (b). Draw a line in each showing how much liquid you would add if you needed to measure 2.64 mL of water. Which cylinder do you think is more precise? Explain.

1.31  State the length of the pencil depicted in the accompanying figure in both inches and centimeters using appropriate numbers of significant figures.

1.32  Assume that you are delivering a solution sample from a pipette. Figures (a) and (b) show the volume level before and after dispensing the sample, respectively. State the liquid level (in mL) before and after dispensing the sample, and calculate the volume of the sample.

1.33  Assume that identical hydrometers are placed in ethanol (sp gr 0.7893) and in chloroform (sp gr 1.4832). In which liquid will the hydrometer float higher? Explain.

ADDITIONAL SECTION PROBLEMS

These exercises are divided into sections by topic. Each section begins with review and conceptual questions, followed by numerical problems of varying levels of difficulty. Many of the problems dealing with more difficult concepts or skills are presented in pairs, with each even-numbered problem followed by an odd-numbered one requiring similar skills. The final section consists of unpaired Conceptual Problems that draw on various parts of the chapter and, in future chapters, may even require the use of concepts from previous chapters. An additional feature in this edition is the incorporation of Group Questions that may sometimes require using resources other than the textbook and are suitable as small group activities. Answers to all even-numbered problems are given at the end of the book following the appendices.

CHEMISTRY AND THE PROPERTIES OF MATTER (SECTION 1.1)

1.34  What is the difference between a physical change and a chemical change?

1.35  Which of the following is a physical change and which is a chemical change?  
(a) Boiling water  
(b) Decomposing water by passing an electric current through it  
(c) Exploding of potassium metal when placed in water  
(d) Breaking of glass

1.36  Which of the following is a physical change and which is a chemical change?  
(a) Making lemonade (lemons + water + sugar)  
(b) Frying eggs  
(c) Burning a candle  
(d) Whipping cream  
(e) Leaves changing color
1.37 Name and describe the three states of matter.

1.38 Name two changes of state and describe what causes each to occur.

1.39 Sulfur dioxide is a compound produced when sulfur burns in air. It has a melting point of -72.7 °C and a boiling point of -10 °C. In what state does it exist at room temperature (298 K)? (Refer to Figure 1.9.)

1.40 Butane (C\_4H\_8) is an easily compressible gas used in cigarette lighters. It has a melting point of -138.4 °C and a boiling point of -0.5 °C. Would you expect a butane lighter to work in winter when the temperature outdoors is 25 °F? Why or why not? (Refer to Figure 1.9.)

1.41 Classify each of the following as a mixture or a pure substance:
   (a) Pea soup
   (b) Seawater
   (c) The contents of a propane tank
   (d) Urine
   (e) Lead
   (f) A multivitamin tablet

1.42 Which of these terms, (i) mixture, (ii) solid, (iii) liquid, (iv) gas, (v) chemical element, (vi) chemical compound, applies to the following substances at room temperature?
   (a) Gasoline
   (b) Iodine
   (c) Water
   (d) Air
   (e) Blood
   (f) Sodium bicarbonate
   (g) Gaseous ammonia
   (h) Silicon

1.43 Hydrogen peroxide, often used in solutions to cleanse cuts and scrapes, breaks down to yield water and oxygen:
   \[ \text{Hydrogen peroxide, } H_2O_2(\text{aq}) \rightarrow \text{Hydrogen, } H_2(g) + \text{Oxygen, } O_2(g) \]
   (a) Identify the reactants and products.
   (b) Which of the substances are chemical compounds, and which are elements?

1.44 When sodium metal is placed in water, the following change occurs:
   \[ \text{Sodium, } Na(s) + \text{Water, } H_2O(l) \rightarrow \text{Hydrogen, } H_2(g) + \text{Sodium hydroxide, } NaOH(\text{aq}) \]
   (a) Identify the reactants and products and their physical states.
   (b) Which of the substances are elements, and which are chemical compounds?

1.45 What is the most abundant element in the earth's crust? In the human body? List the name and symbol for each.

1.46 What are the symbols for the following elements? Perform a web search to identify some of the common uses of the elements listed.
   (a) Iodine
   (b) Chromium
   (c) Technetium
   (d) Arsenic
   (e) Barium

1.47 Supply the missing names or symbols for the elements in the spaces provided:
   (a) N _____
   (b) K _____
   (c) Cl _____
   (d) _____ Calcium
   (e) _____ Phosphorus
   (f) _____ Manganese

1.48 Correct the following statements.
   (a) The symbol for bromine is BR.
   (b) The symbol for manganese is Mg.
   (c) The symbol for carbon is Ca.
   (d) The symbol for potassium is Po.

1.49 Correct the following statements.
   (a) Carbon dioxide has the formula CO2.
   (b) Carbon dioxide has the formula Co\_2.
   (c) Table salt, NaCl, is composed of nitrogen and chlorine.

1.50 The amino acid, glycine, has the formula C\_2H\_5NO\_2. Which elements are present in glycine? What is the total number of atoms represented by the formula?

1.51 Glucose, a form of sugar, has the formula C\_6H\_12O\_6. Which elements are included in this compound, and how many atoms of each are present?

1.52 Write the formula for ibuprofen: 13 carbons, 18 hydrogens, and 2 oxygens. What are the common uses of ibuprofen?

1.53 The atmosphere consists of a number of permanent gases: oxygen (O\_2), nitrogen (N\_2), carbon dioxide (CO\_2), water vapor (H\_2O), and argon (Ar). Identify each substance as an element or a compound. Would you consider the atmosphere to be a heterogeneous or a homogeneous mixture?

1.54 What is the difference between a physical quantity and a number?

1.55 What are the units used in the SI system to measure mass, volume, length, and temperature? In the metric system?

1.56 Give the full name of the following units:
   (a) cc
   (b) dm
   (c) mm
   (d) nL
   (e) mg
   (f) m^3

1.57 Write the symbol for the following units:
   (a) nanogram
   (b) centimeter
   (c) microliter
   (d) micrometer
   (e) milligram

1.58 How many picograms are in 1 mg? In 35 ng?

1.59 How many microliters are in 1 L? In 20 mL?

1.60 Express the following numbers in scientific notation with the correct number of significant figures:
   (a) 9457
   (b) 0.000 07
   (c) 20,000,000,000 (four significant figures)
   (d) 0.012 345
   (e) 652.38
1.61 Convert the following numbers from scientific notation to standard notation:

(a) $5.28 \times 10^3$
(b) $8.205 \times 10^{-2}$
(c) $1.84 \times 10^{-5}$
(d) $6.37 \times 10^4$

1.62 How many significant figures does each of the following numbers have?

(a) 237,401
(b) 0.300
(c) 3.01
(d) 244.4
(e) 50,000
(f) 660

1.63 How many significant figures are there in each of the following quantities?

(a) Distance from New York City to Wellington, New Zealand, 14,397 km
(b) Average body temperature of a crocodile, 25.6 °C
(c) Melting point of gold, 1064 °C
(d) Diameter of an influenza virus, 0.000 01 mm
(e) Radius of a phosphorus atom, 0.110 nm

1.64 The diameter of the earth at the equator is 7926.381 mi.

(a) Round off the earth's diameter to four significant figures, to two significant figures, and to six significant figures.
(b) Express the earth's diameter in scientific notation.

1.65 Carry out the following calculations, express each answer to the correct number of significant figures, and express them in scientific notation.

(a) 9.02 g + 3.1 g
(b) 88.80 cm + 7.391 cm
(c) 362 mL - 99.5 mL
(d) 12.4 mg + 6.378 mg + 2.089 mg

1.66 Carry out the following calculations, express the answers to the correct number of significant figures, and include units in the answers.

(a) $9.02 \times 3.1$
(b) $88.80 + 7.391$
(c) $362 - 99.5$
(d) $12.4 + 6.378 + 2.089$

1.67 Carry out the following calculations, express the answers to the correct numbers of significant figures, and include units in the answers.

(a) $5280 \times 6.2$ mi
(b) $4.5 \times 3.25$ m
(c) $2.50 + 8.3 \times 10^{-3}$ g
(d) $4.70 \times 6.8 \times 2.54$ cm

UNIT CONVERSIONS AND PROBLEM SOLVING [SECTION 1.10]

1.68 Carry out the following conversions:

(a) 3.614 mg to centigrams
(b) 12.0 kL to megaliters
(c) 14.4 μm to millimeters
(d) $6.03 \times 10^{-6}$ cg to nanograms
(e) 174.5 mL to deciliters
(f) $1.5 \times 10^{-2}$ km to centimeters

1.69 Carry out the following conversions. Consult Tables 1.7–1.9 as needed.

(a) 56.4 mi to kilometers and to megameters
(b) 2.0 L to quarts and to fluid ounces
(c) 7 ft 2.0 in. to centimeters and to meters
(d) 1.35 lb to kilograms and to decigrams

1.70 Express the following quantities in more convenient units by using SI unit prefixes:

(a) $9.78 \times 10^4$ g
(b) $1.33 \times 10^{-4}$ L
(c) 0.000 000 000 46 g
(d) $2.99 \times 10^8$ cm

1.71 Fill in the blanks to complete the equivalencies either with appropriate unit prefixes or with the appropriate scientific notation. The first blank is filled in as an example.

(a) $125 \text{km} = 1.25 \times 10^5 \text{m}$
(b) $6.285 \times 10^1 \text{mg} = \text{？} \text{kg}$
(c) $47.35 \text{dL} = 4.735 \times \text{？} \text{mL}$
(d) $67.4 \text{cm} = 6.7 \times 10^{-4} \text{？}$

1.72 The speed limit in Canada is 100 km/h.

(a) How many miles per hour is this?
(b) How many feet per second?

1.73 The muzzle velocity of a projectile fired from a 9 mm handgun is 1200 ft/s.

(a) How many miles per hour is this?
(b) How many meters per second?

1.74 The diameter of a red blood cell is $6 \times 10^{-6}$ m.

(a) How many centimeters is this?
(b) How many red blood cells are needed to make a line 1 cm long? 1 in. long?

1.75 The Willis Tower in Chicago has an approximate floor area of 418,000 m$^2$. How many square feet of floor space is this?

1.76 A normal value for blood cholesterol is 200 mg/dL of blood. If a normal adult has a total blood volume of 5 L, how much total cholesterol is present?

1.77 The recommended daily dose of calcium for an 18-year-old male is 1200 mg. If 1.0 cup of whole milk contains 290 mg of calcium and milk is his only calcium source, how much milk should an 18-year-old male drink each day?

1.78 The white blood cell concentration in normal blood is approximately 12,000 cells/mm$^3$ of blood. How many white blood cells does a normal adult with 5 L of blood have? Express the answer in scientific notation.

ENERGY, HEAT, AND TEMPERATURE [SECTION 1.11]

1.79 The boiling point of liquid nitrogen, used in the removal of warts and in other surgical applications, is $-195.8 \degree \text{C}$. What is this temperature in kelvins and in degrees Fahrenheit?

1.80 Diethyl ether, a substance once used as a general anesthetic, has a specific heat of 0.895 cal/(g °C). How many calories and how many kilocalories of heat are needed to raise the temperature of 30.0 g of diethyl ether from 10.0 °C to 30.0 °C? How many joules and kilojoules?
1.81 Aluminum has a specific heat of 0.215 cal/(g °C). When 25.7 cal (108.5 J) of heat is added to 18.4 g of aluminum at 20.0 °C, what is the final temperature of the aluminum?

1.82 Calculate the specific heat of copper if it takes 23 cal (96 J) to heat a 5.0 g sample from 25 °C to 75 °C.

1.83 The specific heat of fat is 0.45 cal/(g • °C) (1.9 J/g °C) and the density of fat is 0.94 g/cm³. How much energy (in calories and joules) is needed to heat 10 cm³ of fat from room temperature (25 °C) to its melting point (35 °C)?

1.84 A 150 g sample of mercury and a 150 g sample of iron are at an initial temperature of 25.0 °C. If 250 cal (1050 J) of heat is applied to each sample, what is the final temperature of each? (See Table 1.10.)

1.85 When 100 cal (418 J) of heat is applied to a 125 g sample, the temperature increases by 28 °C. Calculate the specific heat of the sample and compare your answer to the values in Table 1.10. What is the identity of the sample?

1.86 Aspirin has a density of 1.40 g/cm³. What is the volume in cubic centimeters of a tablet weighing 250 mg?

1.87 Gaseous hydrogen has a density of 0.0899 g/L at 0 °C. How many liters would you need if you wanted 1.0078 g of hydrogen?

1.88 What is the density of lead (in g/cm³) if a rectangular bar measuring 0.500 cm in height, 1.55 cm in width, and 25.00 cm in length has a mass of 220.9 g?

1.89 What is the density of lithium metal (in g/cm³) if a cube measuring 0.82 cm × 1.45 cm × 1.25 cm has a mass of 0.794 g?

1.90 Ethanol produced by fermentation has a specific gravity of 0.787 at 25 °C. What is the volume of 125 g of ethanol at this temperature? (The density of water at 25 °C is 0.997 g/mL.)

1.91 Ethylene glycol, commonly used as automobile antifreeze, has a specific gravity of 1.1088 at room temperature (25 °C). What is the mass of 1.00 L of ethylene glycol at this temperature?

1.92 Another temperature scale is the Rankine scale. It represents an absolute temperature scale similar to the Kelvin scale, with a common absolute zero (i.e., 0.0 K = 0.0 °R). However, whereas a change of 1.0 K is the same as a change of 1.0 °C, a change of 1.0 °R is the same as 1.0 °F. Absolute zero on the Rankine scale equals −459.67 °F. Water freezes at 32 °F (or 0.0 °C) and boils at 212 °F (100.0 °C). Convert these temperatures to their equivalent temperatures on the Rankine scale.

1.93 A white solid with a melting point of 730 °C is melted. When electricity is passed through the resultant liquid, a brown gas and a molten metal are produced. Neither the metal nor the gas can be broken down into anything simpler by chemical means. Classify each—the white solid, the molten metal, and the brown gas—as a mixture, a compound, or an element.

1.94 Refer to the pencil in Problem 1.31. Using the equivalent values in Table 1.8 as conversion factors, convert the length measured in inches to centimeters. Compare the calculated length in centimeters to the length in centimeters measured using the metric ruler. How do the two values compare? Explain any differences.

1.95 Gemstones are weighed in carats, where 1 carat = 200 mg exactly. What is the mass in grams of the Hope diamond, the world's largest blue diamond, at 44.4 carats?

1.96 The relationship between the nutritional unit for energy and the metric unit is 1 Calorie = 1 kcal.

(a) One donut contains 350 Calories. Convert this to calories and joules.

(b) If the energy in one donut was used to heat 35.5 kg of water, calculate the increase in temperature of the water (in °C).

1.97 Drug dosages are typically prescribed in units of milligrams per kilogram of body weight. A new drug has a recommended dosage of 9 mg/kg.

(a) How many milligrams would a 130 lb woman have to take to obtain this dosage?

(b) How many 125 mg tablets should a 40 lb child take to receive the recommended dosage?

1.98 A clinical report gave the following data from a blood analysis: iron, 39 mg/dL; calcium, 8.3 mg/dL; cholesterol, 224 mg/dL. Express each of these quantities in grams per deciliter, writing the answers in scientific notation.

1.99 The Spirit of America Goodyear blimp has a volume of 2.027 × 10⁵ ft³.

(a) Convert this volume to L.

(b) When in operation it is filled with helium gas. If the density of helium at room temperature is 0.179 g/L, calculate the mass of helium in the blimp.

(c) What is the mass of air occupying the same volume? The density of air at room temperature is 1.20 g/L.

1.100 Approximately 75 mL of blood is pumped by a normal human heart at each beat. Assuming an average pulse of 72 beats per minute, how many milliliters of blood are pumped in one day?

1.101 A doctor has ordered that a patient be given 15 g of glucose, which is available in a concentration of 50.00 g glucose/1000.0 mL of solution. What volume of solution should be given to the patient?

1.102 Reconsider the volume of the sample dispensed by pipette in Problem 1.32. Assuming that the solution in the pipette has a density of 0.963 g/mL, calculate the mass of solution dispensed in the problem to the correct number of significant figures.

1.103 Today, thermometers containing mercury are used less frequently than in the past because of concerns regarding the toxicity of mercury and because of its relatively high melting point (−39 °C). This means that mercury thermometers cannot be used in very cold environments because the mercury is a solid under such conditions. Alcohol
thermometers, however, can be used over a temperature range from \(-115 \, ^\circ C\) (the melting point of alcohol) to \(78.5 \, ^\circ C\) (the boiling point of alcohol).

(a) What is the effective temperature range of the alcohol thermometer in °F?

(b) The densities of alcohol and mercury are 0.79 g/mL and 13.6 g/mL, respectively. If the volume of liquid in a typical laboratory thermometer is 1.0 mL, what mass of alcohol is contained in the thermometer? What mass of mercury?

1.104 In a typical person, the level of blood glucose (also known as blood sugar) is about 85 mg/100 mL of blood. If an average body contains about 11 pints of blood, how many grams and how many pounds of glucose are present in the blood?

1.105 A patient is receiving 3000 mL/day of a solution that contains 5 g of dextrose (glucose) per 100 mL of solution. If glucose provides 4 kcal/g of energy, how many kilocalories per day is the patient receiving from the glucose?

1.106 A rough guide to fluid requirements based on body weight is 100 mL/kg for the first 10 kg of body weight, 50 mL/kg for the next 10 kg, and 20 mL/kg for weight over 20 kg. What volume of fluid per day is needed by a 55 kg woman? Give the answer with two significant figures.

1.107 Chloral hydrate, a sedative and sleep-inducing drug, is available as a solution labeled 10.0 gr/fluidram. What volume in milliliters should be administered to a patient who is meant to receive 7.5 gr per dose? (1 gr = 64.8 mg; 1 fluidram = 3.72 mL)

1.108 When 1.0 tablespoon of butter is burned or used by our body, it releases 100 kcal (100 food Calories or 418.4 kJ) of energy. If we could use all the energy provided, how many tablespoons of butter would have to be burned to raise the temperature of 3.00 L of water from 18.0 °C to 90.0 °C?

1.109 An archeologist finds a 1.62 kg goblet that she believes to be made of pure gold. When 1350 cal (5650 J) of heat is added to the goblet, its temperature increases by 7.8 °C. Calculate the specific heat of the goblet. Is it made of gold? Explain.

1.110 In another test, the archeologist in Problem 1.109 determines that the volume of the goblet is 205 mL. Calculate the density of the goblet and compare it with the density of gold (19.3 g/mL), lead (11.4 g/mL), and iron (7.86 g/mL). What is the goblet probably made of?

1.111 Imagine that you place a piece of cork measuring 1.30 cm × 5.50 cm × 3.00 cm in a pan of water and that on top of the cork you place a small cube of lead measuring 1.15 cm on each edge. The density of cork is 0.235 g/cm³ and the density of lead is 11.35 g/cm³. Will the combination of cork plus lead float or sink?

1.112 At a certain point, the Celsius and Fahrenheit scales “cross” and the numerical value of the Celsius temperature is the same as the numerical value of the Fahrenheit temperature. At what temperature does this crossover occur?

GROUP PROBLEMS

1.113 In the chapter, the conversion of currency was used as an example for unit conversion. Find out what the current monetary conversion rates are and convert US$500 into (a) euros, (b) British pounds, (c) rupees, and (d) Canadian dollars.

1.114 Look up the chemical formula for chloral hydrate mentioned in Problem 1.107. How many different elements are included in the compound, and how many atoms of each element?

1.115 The specific gravity of ethanol is 0.787, while the specific gravity of water is 1.0. Alcoholic beverages are a mixture of water and alcohol and have a specific gravity somewhere between 0.787 and 1.0 density of ethanol. Look up the average alcohol content, typically reported as % by volume, and the specific gravity of each of the following: 80 proof whiskey; red table wine; domestic beer.

1.116 Sulfuric acid (H₂SO₄, density 1.83 g/mL) is produced in larger amounts than any other chemical: Global production exceeded 230 million metric tonnes in 2012 and is projected to exceed 267 million tonnes by 2016. What was the volume of this amount (in liters) produced in 2012? What are the most common applications of sulfuric acid?
A patient visits his or her local clinic complaining of headaches and lethargy. A blood sample is taken and analyzed to determine the relative amounts of certain elements, including many metals identified as micronutrients or trace nutrients. Not enough iron, for example, could indicate anemia, while elevated levels of heavy metals, such as lead or cadmium, could be indicators of toxicity effects, such as headache or a feeling of fatigue. Knowing atomic structure and how elemental properties are related to the arrangement of electrons in a given atom allow us to identify and detect substances in the blood, including oxygen and essential nutrients, even at very low levels. Spectrometers, such as the portable blood oximeter featured above, measure the interaction of atoms or molecules with energy (such as a flame or light source) to determine the identity and concentrations of these substances, which should be in a certain range to ensure good health. As we will see in more detail in the Chemistry in Action feature on page 66, atoms will absorb or emit light of a specific wavelength based on the electron configuration and excitation in the atom. The color of the light can be used to determine the identity of certain elements, which can be used to determine the cause of the patient’s symptoms.

Chemistry is studied on two levels. In the previous chapter, we learned about chemistry on the large-scale, or *macroscopic*, level, looking at the properties and transformations of matter that we can see and measure. We also introduced the elements that make up all matter and how we can use symbols to represent the many different elements and compounds of which matter is made. But what makes one element different from another? To answer that question, we need to look at the submicroscopic or atomic level, studying the behavior and properties of individual atoms. Although scientists have long been convinced of their existence, only within the past 20 years have powerful new instruments made it possible to see individual atoms. In this chapter, we will learn about modern atomic theory and how the structure of atoms influences macroscopic properties.
2.1 Atomic Theory and the Structure of Atoms

Learning Objective:
- Explain the major assumptions of atomic theory, and name and identify the properties of the subatomic particles that make up an atom.

Take a piece of aluminum foil, and cut it in two. Then, take one of the pieces and cut it in two, and so on. Assuming that you have extremely small scissors and extraordinary dexterity, how long can you keep dividing the foil? Is there a limit, or is matter infinitely divisible into ever smaller and smaller pieces? Historically, this argument dates as far back as the ancient Greek philosophers. Aristotle believed that matter could be divided infinitely, while Democritus argued (correctly) that there is a limit. The smallest and simplest bit that aluminum (or any other element) can be divided and still be identifiable as aluminum is called an atom, a word derived from the Greek atomos, meaning "indivisible."

Chemistry is built on four fundamental assumptions about atoms and matter, proposed by English scientist John Dalton in 1808, which together make up modern atomic theory:
- All matter is composed of atoms.
- Atoms of any given element share the same chemical properties while atoms of different elements have different properties.
- Chemical compounds consist of atoms combined in specific ratios. That is, only whole atoms can combine—one A atom with one B atom, or one A atom with two B atoms, and so on. The vast number of ways that atoms can combine with one another results in the enormous diversity in the substances around us.
- Chemical reactions change only the way that atoms are combined in compounds. The atoms themselves are unchanged and do not disappear.

Atoms are extremely small, ranging from about $7.4 \times 10^{-11}$ m in diameter for a hydrogen atom to $5.24 \times 10^{-10}$ m for a cesium atom. In mass, atoms vary from $1.67 \times 10^{-24}$ g for hydrogen to $3.95 \times 10^{-22}$ g for uranium, one of the heaviest naturally occurring atoms. It is difficult to appreciate just how small atoms are, although it might help if you realize that a fine pencil line is about 3 million atoms across and that even the smallest speck of dust contains about $10^{16}$ atoms. Our current understanding of atomic structure is the result of many experiments performed in the late 1800s and early 1900s (see Chemistry in Action on p. 44).

Atoms are composed of tiny subatomic particles called protons, neutrons, and electrons. A proton has a mass of $1.672 \times 10^{-24}$ g and carries a positive (+) electrical charge, a neutron has a mass similar to that of a proton ($1.674 \times 10^{-24}$ g) but is electrically neutral, and an electron has a mass that is only $1/1836$ that of a proton ($9.109 \times 10^{-28}$ g) and carries a negative (−) electrical charge. In fact, electrons are so much lighter than protons and neutrons that their mass is usually ignored.

Table 2.1 A Comparison of Subatomic Particles

<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>p</td>
<td>$1.672622 \times 10^{-24}$</td>
<td>1.007276</td>
<td>+1</td>
</tr>
<tr>
<td>Neutron</td>
<td>n</td>
<td>$1.674927 \times 10^{-24}$</td>
<td>1.008665</td>
<td>0</td>
</tr>
<tr>
<td>Electron</td>
<td>e⁻</td>
<td>$9.109328 \times 10^{-28}$</td>
<td>$5.485799 \times 10^{-4}$</td>
<td>−1</td>
</tr>
</tbody>
</table>

The masses of atoms and their constituent subatomic particles are so small when measured in grams that it is more convenient to express them on a relative mass scale. The basis for the relative atomic mass scale is an atom of carbon that contains six protons and six neutrons. Such an atom is assigned a mass of exactly 12 atomic mass units (amu; also called a dalton in honor of John Dalton), where 1 amu = 1.660 539 × 10⁻²⁴ g.

Atomic mass unit (amu) The unit for describing the mass of an atom; 1 amu = $\frac{1}{12}$ the mass of a carbon-12 atom.
Thus, for all practical purposes, both a proton and a neutron have a mass of 1 amu (Table 2.1). Hydrogen atoms are only about one-twelfth as heavy as carbon atoms and have a mass close to 1 amu, magnesium atoms are about twice as heavy as carbon atoms and have a mass close to 24 amu, and so forth.

Subatomic particles are not distributed at random throughout an atom. Rather, the protons and neutrons are packed closely together in a dense core called the nucleus. Surrounding the nucleus, the electrons move about rapidly through a large, mostly empty volume of space (Figure 2.1). Measurements show that the diameter of a nucleus is only about $10^{-15}$ m, whereas that of the atom itself is about $10^{-10}$ m. For comparison, if an atom were the size of a large domed stadium, the nucleus would be approximately the size of a small pea in the center of the playing field.

The positively charged protons in the nucleus also repel one another but are nevertheless held together by a unique attraction called the nuclear strong force, which we will discuss further in Chapter 11.

The structure of the atom is determined by an interplay of different attractive and repulsive forces. Because unlike charges attract one another, the negatively charged electrons are held near the positively charged nucleus. But because like charges repel one another, the electrons also try to get as far away from one another as possible, accounting for the relatively large volume they occupy.

**CHEMISTRY IN ACTION**

**Are Atoms Real?**

Chemistry rests on the premise that matter is composed of the tiny particles we call atoms. Every chemical reaction and every physical law that governs the behavior of matter is explained by chemists in terms of atomic theory. But how do we know that atoms are real and not just an imaginary concept? And how do we know the structure of the atom?

The development of our understanding of atomic structure is another example of the scientific method at work, with several scientists contributing to our understanding of atomic structure. J. J. Thomson demonstrated that matter contained negatively charged particles that were 1000 times lighter than H\(^+\), the lightest positively charged particles found in aqueous solution, and that the mass-to-charge ratio of these particles was the same regardless of the material used to produce the particles [Section 5.5 and Chapter 10]. Ernest Rutherford deduced that an atom consists mostly of empty space [occupied by the negatively charged electrons] and that most of the mass and all of the positive charges are contained in a relatively small, dense region that he called the "nucleus."
**SECTION 2.2 Elements and Atomic Number**

**Learning Objective:**

- Identify atoms of an element based on the number of protons in the nucleus.

All atoms contain proton, neutrons, and electrons, but how do we distinguish an atom of carbon from an atom of oxygen, or sodium? Each atom has a specific number of protons, neutrons, and electrons, and the identity of the element is determined by the number of protons within the nucleus, also called the element's **atomic number** \(Z\). Every element has a different number of protons within its nucleus, thus every element has a different atomic number. If we know the number of protons in an atom, we can identify the element. Any atom with six protons, for example, is a carbon atom because the atomic number for carbon is 6 \((Z = 6)\).

Atoms are neutral and have no net charge because the number of positively charged protons in an atom is the same as the number of negatively charged electrons. Thus, the atomic number also equals the number of electrons in every atom of a given element. Hydrogen, \(Z = 1\), has only 1 proton and 1 electron; carbon, \(Z = 6\), has 6 protons and 6 electrons; sodium, \(Z = 11\), has 11 protons and 11 electrons; and so on, up

**Atomic number \((Z)\)** The number of protons in the nucleus of an atom of a given element.

**LOOKING AHEAD** In a neutral atom, the number of electrons is equal to the number of protons. However, most elements can gain or lose electrons to form charged particles, called **ions**, which will be discussed in Chapter 3.
Mass number \((A)\) The total number of protons and neutrons in an atom.

to the element with the largest known atomic number \((Z = 118)\). In a periodic table, elements are listed in order of increasing atomic number, beginning at the upper left and ending at the lower right.

The sum of the protons and neutrons in an atom is called the atom's **mass number** \((A)\). For example, hydrogen atoms with 1 proton and no neutrons have mass number 1, carbon atoms with 6 protons and 6 neutrons have mass number 12, sodium atoms with 11 protons and 12 neutrons have mass number 23. Atomic number and mass number can be written using chemical symbols by showing the element's mass number \((A)\) as a superscript and its atomic number \((Z)\) as a subscript in front of the atomic symbol. For example, \(\frac{A}{Z}X\), where \(X\) represents the symbol for the element, \(A\) represents the mass number, and \(Z\) represents the atomic number.

**Worked Example 2.1** Atomic Structure: Protons, Neutrons, and Electrons

Phosphorus has the atomic number \(Z = 15\). How many protons, electrons, and neutrons are there in phosphorus atoms, which have mass number \(A = 31\)?

**ANALYSIS** The atomic number gives the number of protons, which is the same as the number of electrons, and the mass number gives the total number of protons plus neutrons.

**SOLUTION**

Phosphorus atoms, with \(Z = 15\), have 15 protons and 15 electrons. To find the number of neutrons, subtract the atomic number from the mass number.

\[
\text{Mass number (sum of protons and neutrons)} - \text{Atomic number (number of protons)} = \text{16 neutrons}
\]

**Worked Example 2.2** Atomic Structure: Atomic Number and Atomic Mass

An atom contains 28 protons and has \(A = 60\). Give the number of electrons and neutrons in the atom, and identify the element.

**ANALYSIS** The number of protons and the number of electrons are the same and are equal to the atomic number \(Z\), 28 in this case. Subtracting the number of protons (28) from the total number of protons plus neutrons (60) gives the number of neutrons.

**SOLUTION**

The atom has 28 electrons and \(60 - 28 = 32\) neutrons. The list of elements inside the front cover shows that the element with atomic number 28 is nickel (Ni).

**PROBLEM 2.1**

Use the list inside the front cover to identify the following elements:

(a) \(A = 186\), with 111 neutrons
(b) \(A = 59\), with 21 neutrons
(c) \(A = 127\), with 75 neutrons

### 2.3 Isotopes and Atomic Weight

**Learning Objective:**

- Write the symbols for different isotopes of an element, and use relative abundances and atomic masses of isotopes to calculate the average atomic weight of an element.

All atoms of a given element have the same number of protons, equal to the atomic number \((Z)\) of that element; however, different atoms of an element can have different numbers of neutrons and, therefore, different mass numbers. Atoms with identical atomic numbers but different mass numbers are called **isotopes**. Hydrogen, for example, has three isotopes. The most abundant hydrogen isotope, called **protium**, has...
one proton but no neutrons and thus has a mass number of 1. A second hydrogen isotope, called deuterium, also has one proton, but has one neutron and a mass number of 2; and a third isotope, called tritium, has two neutrons and a mass number of 3.

### Protium
- One proton
- No neutrons
- Mass number = 1

### Deuterium
- One proton
- One neutron
- Mass number = 2

### Tritium
- One proton
- Two neutrons
- Mass number = 3

A specific isotope is represented by showing its mass number \( A \) as a superscript and its atomic number \( Z \) as a subscript in front of the atomic symbol, for example, \( ^1\text{H} \). Thus, protium is \( ^1\text{H} \), deuterium is \( ^2\text{H} \), and tritium is \( ^3\text{H} \).

Unlike the three isotopes of hydrogen, the isotopes of most elements do not have distinctive names. Instead, the mass number of the isotope is given after the name of the element. The \( ^{235}\text{U} \) isotope used in nuclear reactors, for example, is usually referred to as uranium-235, or U-235.

Most naturally occurring elements are mixtures of isotopes. In a large sample of naturally occurring hydrogen atoms, for example, 99.985% have mass number \( A = 1 \) (protium) and 0.015% have mass number \( A = 2 \) (deuterium). Therefore, it is useful to know the average mass of the atoms in a large sample, a value called the element’s atomic weight. For hydrogen, the atomic weight is 1.008 amu. Atomic weights for all elements are given on the inside of the front cover of this book.

To calculate the atomic weight of an element, the individual masses of the naturally occurring isotopes and the percent abundance of each must be known. The atomic weight can then be calculated as the sum of the masses of the individual isotopes for that element, or

\[
\text{Atomic weight} = \sum [\text{(isotopic abundance)} \times \text{(isotopic mass)}]
\]

where the Greek symbol \( \Sigma \) indicates the mathematical summing of terms.

Chlorine, for example, occurs on earth as a mixture of 75.77% \( ^{35}\text{Cl} \) atoms (mass = 34.97 amu) and 24.23% \( ^{37}\text{Cl} \) atoms (mass = 36.97 amu). This can also be expressed in terms of fractional composition (i.e., 75.77% of all chlorine atoms is the same as a fraction of 0.7577). The atomic weight is found by calculating the percentage of the mass contributed by each isotope. For chlorine, the calculation is done in the following way (to four significant figures), giving an atomic weight of 35.45 amu:

\[
\begin{align*}
\text{Contribution from } ^{35}\text{Cl: } (0.7577 \times 34.97 \text{ amu}) &= 26.4968 \text{ amu} \\
\text{Contribution from } ^{37}\text{Cl: } (0.2423 \times 36.97 \text{ amu}) &= 8.9578 \text{ amu} \\
\text{Atomic weight} = 35.4546 &= 35.45 \text{ amu} \\
&\text{(Rounded to four significant figures)}
\end{align*}
\]

The final number of significant figures in this case (four) was determined by the rounding rules presented in Chapter 1. Note that the final rounding to four significant figures was not done until after the final answer was obtained.
**Worked Example 2.3** Average Atomic Mass: Weighted-Average Calculation

Gallium is a metal with a very low melting point—it will melt in the palm of your hand. It has two naturally occurring isotopes: 60.4% is Ga-69 (mass = 68.9257 amu) and 39.6% is Ga-71 (mass = 70.9248 amu). Calculate the atomic weight for gallium.

**ANALYSIS** We can calculate the average atomic mass for the element by summing up the contributions from each of the naturally occurring isotopes.

**BALLPARK ESTIMATE** The masses of the two naturally occurring isotopes of gallium differ by 2 amu (68.9 and 70.9 amu). Since slightly more than half of the Ga atoms are the lighter isotope (Ga-69), the average mass will be slightly less than halfway between the two isotopic masses; estimate = 69.8 amu.

**SOLUTION**

**STEP 1: Identify known information.**

Ga-69 (60.4% at 68.9257 amu)
Ga-71 (39.6% at 70.9248 amu)

**STEP 2: Identify the unknown answer and units.**

Atomic weight for Ga (in amu) = ?

**STEP 3: Identify conversion factors or equations.** This equation calculates the average atomic weight as a weighted average of all naturally occurring isotopes.

\[
\text{Atomic weight} = \sum [(\text{isotopic abundance}) \times (\text{isotopic mass})]
\]

**STEP 4: Solve.** Substitute known information and solve.

\[
\text{Atomic weight} = (0.604) \times (68.9257 \text{ amu}) = 41.6311 \text{ amu} \\
+ (0.396) \times (70.9248 \text{ amu}) = 28.0862 \text{ amu} \\
\text{Atomic weight} = 69.7 \text{ amu (3 significant figures)}
\]

**BALLPARK CHECK** Our estimate (69.8 amu) is close!

**Worked Example 2.4** Identifying Isotopes from Atomic Mass and Atomic Number

Identify element X in the symbol $^{194}_{78}X$ and give its atomic number, mass number, number of protons, number of electrons, and number of neutrons.

**ANALYSIS** The identity of the atom corresponds to the atomic number—78.

**SOLUTION**

Element X has $Z = 78$, which shows that it is platinum. (Look inside the front cover for the list of elements.) The isotope $^{194}_{78}$Pt has a mass number of 194, and we can subtract the atomic number from the mass number to get the number of neutrons. This platinum isotope therefore has 78 protons, 78 electrons, and 194 − 78 = 116 neutrons.

**PROBLEM 2.2**

Potassium (K) has two naturally occurring isotopes: K-39 (93.12% mass = 38.9637 amu) and K-41 (6.88%; 40.9618 amu). Calculate the atomic weight for potassium. How does your answer compare with the atomic weight given in the list inside the front cover of this book?

**PROBLEM 2.3**

Bromine, an element present in compounds used as sanitizers and fumigants (for example, ethylene bromide), has two naturally occurring isotopes. Look up the mass numbers of the two naturally occurring isotopes of bromine, along with their percent abundance.

(a) Write the symbols for both isotopes.

(b) Using the masses and natural percent abundances, calculate the average molecular weight for bromine and compare your value to the value found in the periodic table on page 50.
Isotopes are used in many applications, including diagnosis and treatment of cancer and other diseases. In this activity, we will explore the structure of some isotopes of a specific element. Take two pieces of construction paper of different colors and cut each into about 25 pieces. Label each piece of one color with an "n" for neutron and each piece of the other color with a "p" for proton.

**a.** Distribute the pieces of construction paper into three piles as follows. Into pile 1, place six "p" and six "n" pieces. Into pile 2, place six "p" and seven "n" pieces. Into pile 3, place six "p" and eight "n" pieces. How is each pile similar, and how are they different?

**b.** The three piles represent isotopes of a particular element. Which element? Write the atomic symbols for each isotope.

**c.** Look up the natural abundance of each isotope and calculate the average atomic mass for this element. How does your answer compare with the atomic weight given in the periodic table?

### PROBLEM 2.4

An element used to sanitize water supplies has two naturally occurring isotopes with mass numbers of 35 and 37, and 17 electrons. Write the symbols for both isotopes, including their atomic numbers and mass numbers.

### 2.4 The Periodic Table

**Learning Objective:**
- Locate elements on the periodic table and classify them as metals, nonmetals, or metalloids based on their location.

Ten elements have been known since the beginning of recorded history: antimony (Sb), carbon (C), copper (Cu), gold (Au), iron (Fe), lead (Pb), mercury (Hg), silver (Ag), sulfur (S), and tin (Sn). It is worth noting that the symbols for many of these elements are derived from their Latin names, a reminder that they have been known since the time when Latin was the language used for all scholarly work. The first "new" element to be found in several thousand years was arsenic (As), discovered in about 1250. In fact, only 24 elements were known up to the time of the American Revolution in 1776.

As the pace of discovery quickened in the late 1700s and early 1800s, chemists began to look for similarities among elements that might make it possible to draw general conclusions. Numerous attempts were made in the mid-1800s to account for the similarities among groups of elements, but the great breakthrough came in 1869 when the Russian chemist Dmitri Mendeleev organized the elements in order of increasing mass and then organized elements into groups based on similarities in chemical behavior. His table is a forerunner of the modern periodic table. The table has boxes for each element that give the symbol, atomic number, and atomic mass of the element:

Periodic table A tabular format listing all known elements where the atomic symbol (top), name of the element (middle), and atomic mass (bottom) are given in each box that represents the element.

The atomic masses for each element in the table are the average masses calculated based on the mass and percent abundance of the naturally occurring stable isotopes. The boxes are arranged in order of increasing atomic number, with the elements arranged in rows and columns as shown in Figure 2.2. An enormous amount of information is embedded in the periodic table, information that gives chemists the ability to explain known chemical behavior of elements and to predict new behavior.
Metal A malleable element, with a lustrous appearance, that is a good conductor of heat and electricity.

Nonmetal An element that is a poor conductor of heat and electricity.

Metalloid An element whose properties are intermediate between those of a metal and a nonmetal.

One way of classifying the elements is by similarities in physical properties. Of the 118 currently known elements, 94 are classified as metals—aluminum, gold, copper, and zinc, for example. Metals are solid at room temperature (except for mercury), usually have a lustrous appearance when freshly cut, are good conductors of heat and electricity, and are malleable rather than brittle. That is, metals can be pounded into different shapes rather than shattering when struck. Note that metals occur on the left side of the periodic table.

Eighteen elements are nonmetals. All are poor conductors of heat and electricity. Eleven are gases at room temperature, six are brittle solids, and one is a liquid. Oxygen and nitrogen, for example, are gases present in air; sulfur is a solid found in large underground deposits. Bromine is the only liquid nonmetal. Note that nonmetals occur on the upper right side of the periodic table.

The metalloids are located in a zigzag band between the metals on the left and nonmetals on the right side of the periodic table. Although there is some debate as to which elements to include in this list, we include only six in this text: boron, silicon, arsenic, germanium, antimony, and tellurium. The metalloids are so named because their properties are intermediate between those of metals and nonmetals. Pure silicon, for example, has a lustrous or shiny surface, like a metal, but it is brittle, like a nonmetal, and its electrical conductivity lies between that of metals and nonmetals. Some chemistry texts identify polonium as a metalloid, but its chemical behavior and conductivity more closely resemble that of other metals. Others include astatine in the list, but this is purely academic: as a very rare and unstable element, it would be difficult to collect a sample of astatine large enough to obtain reliable data regarding its chemical and physical behavior.

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### Figure 2.2

The periodic table of the elements.

Elements are organized into groups, indicated with numbers and letters. Main group elements are in columns labeled 1A–8A, while the transition metal groups are in columns labeled 1B–8B. Elements to the left and bottom of the periodic table are classified as metals, while elements in the upper right portion are classified as nonmetals.
Metals: Gold, zinc, and copper.
(a) Known for its beauty, gold is very unreactive and is used primarily in jewelry and in electronic components. (b) Zinc, an essential trace element in our diets, has industrial uses ranging from the manufacture of brass, to roofing materials, to batteries. (c) Copper is widely used in electrical wiring, in water pipes, and in coins.

Nonmetals: Nitrogen, sulfur, and iodine.
(a) Nitrogen, (b) sulfur, and (c) iodine are essential to all living things. Pure nitrogen, which constitutes almost 80% of air, is a gas at room temperature and does not condense to a liquid until it is cooled to −328 °C. Sulfur, a yellow solid, is found in large underground deposits in Texas and Louisiana. Iodine is a dark violet crystalline solid that was first isolated from seaweed.

Metalloids: Boron and silicon.
(a) Boron is a strong, hard metalloid used in making the composite materials found in military aircraft. (b) Silicon is well known for its use in making computer chips.
Another way of classifying the elements in the periodic table is based on similarities in chemical behavior. Beginning at the upper left corner of the periodic table, elements are arranged by increasing atomic number into seven horizontal rows, called **periods**, and 18 vertical columns, called **groups**. When organized in this way, *the elements in a given group have similar chemical properties*. Lithium, sodium, potassium, and the other elements in group 1A behave similarly. Chlorine, bromine, iodine, and the other elements in group 7A behave similarly and so on throughout the table.

Note that different periods (rows) contain different numbers of elements. The first period contains only two elements, hydrogen and helium; the second and third periods each contain eight elements; the fourth and fifth periods each contain 18; the sixth and seventh periods contain 32. Note also that the 14 elements following lanthanum (the **lanthanides**) and the 14 following actinium (the **actinides**) are pulled out and shown below the others.

Groups are numbered in two ways, both shown in Figure 2.2. The two large groups on the far left and the six on the far right are called the **main group elements** and are numbered 1A through 8A. The 10 smaller groups in the middle of the table are called the **transition metal elements** and are numbered 1B through 8B. Alternatively, all 18 groups are numbered sequentially from 1 to 18. The 14 groups shown separately at the bottom of the table are called the **inner transition metal elements** and are not numbered.

**Problem 2.5**
Locate aluminum in the periodic table and give its group number and period number.

**Problem 2.6**
Identify the group 1B element in period five and the group 2A element in period four.

**Problem 2.7**
There are five elements in group 5A of the periodic table. Identify them and give the period of each.

**Problem 2.8**
The six metalloids are boron (B), silicon (Si), germanium (Ge), arsenic (As), antimony (Sb), and tellurium (Te). Locate them in the periodic table and tell where they appear with respect to metals and nonmetals.

**Problem 2.9**
Locate the following elements in the periodic table, give the corresponding name for each, and classify them according to group (i.e., halogen, noble gas, alkali metal, etc.).

(a) Ti  (b) Te  (c) Se  (d) Sc  (e) At  (f) Ar

### 2.5 Some Characteristics of Different Groups

**Learning Objective:**
- Classify elements and describe chemical behavior based on group membership.

To see why the periodic table has the name it does, look at the graph of atomic radius versus atomic number in Figure 2.3. The graph shows an obvious **periodicity**—a repeating rise-and-fall pattern. Beginning on the left with atomic number 1 (hydrogen), the sizes of the atoms increase to a maximum at atomic number 3 (lithium), then decrease to a minimum, then increase again to a maximum at atomic number 11 (sodium), then decrease, and
so on. It turns out that the local maximum values occur for atoms of group 1A elements—Li, Na, K, Rb, Cs, and Fr—and the local minimum values occur for atoms of the group 7A elements.

There is nothing unique about the periodicity of atomic radii shown in Figure 2.3. The melting points of the first 100 elements, for example, exhibit similar periodic behavior, as shown in Figure 2.4, with a systematic trend of peaks and valleys as you progress through the elements in the periodic table. Many other physical and chemical properties can be plotted in a similar way with similar results. In fact, the various elements in a given group of the periodic table usually show remarkable similarities in many of their chemical and physical properties. Look at the following four groups, for example:

- **Group 1A—Alkali metals:** Lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr) are shiny, soft metals with low melting points. All react rapidly (often violently) with water to form products that are highly alkaline, or basic—hence the name alkali metals. Because of their high reactivity, the alkali metals are never found in nature in the pure state but only in combination with other elements.

- **Group 2A—Alkaline earth metals:** Beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra) are also lustrous, silvery metals but are less reactive than their neighbors in group 1A. Like the alkali metals, the alkaline earths are never found in nature in the pure state.
Halogen An element in group 7A of the periodic table.

Noble gas An element in group 8A of the periodic table.

- **Group 7A—Halogens:** Fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At) are colorful and corrosive nonmetals. All are found in nature only in combination with other elements, such as with sodium in table salt (sodium chloride, \( \text{NaCl} \)). In fact, the group name halogen is taken from the Greek word \( \text{hals} \), meaning salt.

- **Group 8A—Noble gases:** Helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn) are colorless gases. The elements in this group were labeled the “noble” gases because of their lack of chemical reactivity—helium, neon, and argon do not combine with any other elements, whereas krypton and xenon combine with a very few.

**PROBLEM 2.10**
Locate (a) krypton, (b) strontium, (c) nitrogen, and (d) cobalt in the periodic table. Indicate which categories apply to each: (i) metal, (ii) nonmetal, (iii) transition element, (iv) main group element, and (v) noble gas.

**PROBLEM 2.11**
For each of the following sets of elements, arrange in order of increasing atomic radius:

a) Na, Li, Rb, K  
b) Li, O, C, F  
c) Cl, Br, I, F

**PROBLEM 2.12**
For each set of elements presented in the previous problem, arrange in order of increasing melting point.

**KEY CONCEPT PROBLEM 2.13**
Identify the elements whose nuclei are shown next. For each, tell its group number, its period number, and whether it is a metal, nonmetal, or metalloid.

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**CHEMISTRY IN ACTION**

**Essential Elements and Group Chemistry**

In Chapter 1, we introduced the essential elements—elements that are vital to good health and fitness. In this chapter, we demonstrated how elements in a group exhibit similar chemical properties. As you might expect, the properties of a group influence their role in the metabolism of an organism. As we noted in Chapter 1, these elements are typically not present in the body as free atoms but as ions or combined with other elements in compounds (discussed in Chapter 3). Let's take a look at some of the major players in the different groups:

- **1A. Alkali metals:** Lithium (Li) plays no known physiological role in the body, but the apparent neurological effects of the Li ion in many lithium compounds explains why they are often used as mood stabilizing drugs. Sodium (Na) is considered a macronutrient because of the vital role it plays in regulation of blood volume and blood pressure and transmission of nerve impulses. Potassium (K), like sodium, is also involved in neurological functions and nerve impulse transmission, and a deficiency in K ions can lead to cardiac dysfunctions.
2A. Alkaline earth metals: Magnesium (Mg) ions are important in many enzymatic processes, including energy production and DNA synthesis. Magnesium compounds also are used therapeutically as laxatives and to relieve the symptoms of fibromyalgia, migraines, and premenstrual syndrome. Calcium (Ca) is a major component in teeth and bones and also plays a role in neurotransmission and muscle contraction. An excess of calcium ions in the blood can lead to impaired kidney function and decreased absorption of other important minerals.

7A. Halogens: Fluorine (F) compounds have been added to toothpaste and municipal drinking water supplies to strengthen tooth enamel and increase dental health. Fluorine-containing drugs are used to lower cholesterol, as antidepressants, antibiotics, and anesthetics. Chlorine (Cl) ions play an important role in maintaining salt balance in bodily fluids. Bromine (Br) was initially thought to have no biological function; recent research indicates that it is necessary for tissue development. Iodine (I) is an essential trace element, primarily because of its role as a constituent in thyroxine, a thyroid hormone responsible for regulation of basal metabolism.

1–8B. Transition metals: Probably the most familiar essential transition metal is iron (Fe), a major component in hemoglobin that is responsible for oxygen transport in the blood. But many other transition metals are constituents of enzymes [biological catalysts], including chromium (Cr), cobalt (Co), copper (Cu), molybdenum (Mo), manganese (Mn), and zinc (Zn). Zinc, in particular, is recognized as an essential mineral important to public health. Zinc deficiencies in children are linked to delayed growth and sexual maturation, severe dermatitis, and diarrhea.

The infant is suffering from acrodermatitis enteropathica, a skin condition resulting from an inability to metabolize zinc.

CIA Problem 2.3 Elements from other groups also play important biological roles. Identify each of the following:

a) A group 5A element that is a major component of cell membranes and bones.

b) A group 6A element that is involved in thyroid function and a constituent of enzymes involved in fat metabolism but is toxic in large doses.

CIA Problem 2.4 Locate and identify the group number for each of the transition metals mentioned earlier.

2.6 Electronic Structure of Atoms

Learning Objective:

- Describe the distribution of electrons into shells, subshells, and orbitals around the nucleus of an atom.

Why does the periodic table have the shape it does, with periods of different length? Why are periodic variations observed in atomic radii and in so many other characteristics of the elements? And why do elements in a given group of the periodic table show similar chemical behavior? These questions occupied the thoughts of chemists for more than 50 years after Mendeleev, and it was not until well into the 1920s that the answers were established. Today, we know that the properties of the elements are determined by the arrangement of electrons in their atoms.

Our current understanding of the electronic structure of atoms is based on the quantum mechanical model, developed by Austrian physicist Erwin Schrödinger in 1926. One of the fundamental assumptions of the model is that electrons have both particle-like and wave-like properties, and that the behavior of electrons can be described using a mathematical equation called a wave function. One consequence of this assumption is that electrons are not perfectly free to move about in an atom.
Stairs are quantized because they change height in discrete amounts. A ramp, by contrast, is not quantized because it changes height continuously.

**Shell (electron)** A grouping of electrons in an atom according to energy.

Instead, each electron is restricted to a certain region of space within the atom, depending on the energy level of the electron. Different electrons have different amounts of energy and thus occupy different regions within the atom. Furthermore, the energies of electrons are quantized or restricted to having only certain values.

To understand the idea of quantization, think about the difference between stairs and a ramp. A ramp is not quantized because it changes height continuously. Stairs, by contrast, are quantized because they change height only by a fixed amount. When you walk up a flight of stairs, you can put your foot on each step, but you cannot stand any place between the two steps. Conversely, on a ramp, you can step anywhere on the ramp you like. In the same way, the energy values available to electrons in an atom change only in steps rather than continuously.

The wave functions derived from the quantum mechanical model also provide important information about the location of electrons in an atom. Just as a person can be found by giving his or her address within a state, an electron can be found by giving its "address" within an atom. Furthermore, just as a person's address is composed of several successively narrower categories—city, street, and house number—an electron's address is also composed of successively narrower categories—shell, subshell, and orbital, which are defined by the quantum mechanical model.

The electrons in an atom are grouped around the nucleus into shells, like the layers in an onion, according to the energy of the electrons. The shell is designated using the letter \( n \); \( n = 1 \) for the first shell (period 1), \( n = 2 \) for the second shell (period 2), and so on. The farther a shell is from the nucleus, the larger it is, the more electrons it can hold, the higher the energies of those electrons, and thus the easier they are to remove because they are the farthest away from the positively charged nucleus. The first shell (the one nearest the nucleus) can hold only 2 electrons, the second shell can hold 8, the third shell can hold 18, and the fourth shell can hold 32 electrons.

<table>
<thead>
<tr>
<th>Shell number:</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electron capacity:</td>
<td>2</td>
<td>8</td>
<td>18</td>
<td>32</td>
</tr>
</tbody>
</table>

Within shells, electrons are further grouped into subshells of four different types, identified in order of increasing energy by the letters \( s \), \( p \), \( d \), and \( f \). The first shell has only one subshell, \( s \). The second shell has two subshells: an \( s \) subshell and a \( p \) subshell. The third shell has an \( s \), \( p \), and \( d \) subshell. The fourth shell has an \( s \), \( p \), \( d \), and \( f \) subshell. Of the four types, we will be concerned mainly with \( s \) and \( p \) subshells because most of the elements found in living organisms use only these. A specific subshell is symbolized by writing the number of the shell followed by the letter for the subshell. For example, the designation \( 3p \) refers to the \( p \) subshell in the third shell \( (n = 3) \). Note that the number of subshells in a given shell is equal to the shell number. For example, shell number 3 has three subshells \( (s, p, \text{and} \ d) \).

Finally, within each subshell, electrons are grouped into orbitals, regions of space within an atom where the specific electrons are most likely to be found. There are different numbers of orbitals within the different kinds of subshells. A given \( s \) subshell has only one orbital, a \( p \) subshell has three orbitals, a \( d \) subshell has five orbitals, and an \( f \) subshell has seven orbitals. Each orbital can hold only two electrons, which differ in a property known as spin. If one electron in an orbital has a clockwise spin, the other electron in the same orbital must have a counterclockwise spin. Since the number of orbitals in a shell increases as \( n \) increases, the number of electrons that can be placed in a shell also increases with \( n \), as seen in Table 2.2. The following figure summarizes the configuration of shells, subshells, and orbitals.

<table>
<thead>
<tr>
<th>Shell number:</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
</tr>
</thead>
<tbody>
<tr>
<td>Subshell designation:</td>
<td>( s )</td>
<td>( s, p )</td>
<td>( s, p, d )</td>
<td>( s, p, d, f )</td>
</tr>
<tr>
<td>Number of orbitals:</td>
<td>1</td>
<td>1, 3</td>
<td>1, 3, 5</td>
<td>1, 3, 5, 7</td>
</tr>
</tbody>
</table>
### Table 2.2 Electron Distribution in Atoms

<table>
<thead>
<tr>
<th>Shell Number</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
</tr>
</thead>
<tbody>
<tr>
<td>Subshell designation</td>
<td>s</td>
<td>s, p</td>
<td>s, p, d</td>
<td>s, p, d, f</td>
</tr>
<tr>
<td>Number of orbitals</td>
<td>1</td>
<td>1, 3</td>
<td>1, 3, 5</td>
<td>1, 3, 5, 7</td>
</tr>
<tr>
<td>Number of electrons</td>
<td>2</td>
<td>2, 6</td>
<td>2, 6, 10</td>
<td>2, 6, 10, 14</td>
</tr>
<tr>
<td>Total electron capacity</td>
<td>2</td>
<td>6</td>
<td>18</td>
<td>32</td>
</tr>
</tbody>
</table>

In the quantum mechanical model, different orbitals have different shapes and orientations. Orbitals in s subshells are spherical regions centered about the nucleus, whereas orbitals in p subshells are roughly dumbbell-shaped regions where the nucleus is at the midpoint of the dumbbells (Figure 2.5). As shown in Figure 2.5b, the three p orbitals in a given subshell are oriented at right angles to one another.

The overall electron distribution within an atom is summarized in Table 2.2 and in the following list:

- The first shell has a maximum capacity of only two electrons. The two electrons have different spins and are in a single 1s orbital.
- The second shell has a maximum capacity of eight electrons. Two are in a 2s orbital, and 6 are in the three different 2p orbitals (two per 2p orbital).
- The third shell has a maximum capacity of 18 electrons. Two are in a 3s orbital, 6 are in three 3p orbitals, and 10 are in five 3d orbitals.
- The fourth shell has a maximum capacity of 32 electrons. Two are in a 4s orbital, 6 are in three 4p orbitals, 10 are in five 4d orbitals, and 14 are in seven 4f orbitals.

### Worked Example 2.5 Atomic Structure: Electron Shells

How many electrons are present in an atom that has its first and second shells filled and has four electrons in its third shell? Name the element.

**Analysis** The number of electrons in the atom is calculated by adding the total electrons in each shell. We can identify the element from the number of protons in the nucleus, which is equal to the number of electrons in the atom.

**Solution**

The first shell of an atom holds two electrons in its 1s orbital, and the second shell holds eight electrons (two in a 2s orbital and six in three 2p orbitals). Thus, the atom has a total of 2 + 8 + 4 = 14 electrons. Since the number of electrons is equal to the number of protons, the element's atomic number \( Z = 14 \) and must be silicon (Si).

**Problem 2.14**

How many electrons are present in an atom in which the first and second shells and the 3s subshell are filled? Name the element.
This exercise is designed to help visualize the structure of the atom more closely. The manipulation of an onion will simulate the phenomenal behavior and properties of individual atoms.

a. Cut a medium-sized whole onion in half and remove the outer dry peeling/skin. If we consider the central kernel of the onion as the nucleus, each “layer” would then correspond to a shell containing varying numbers of electrons.

b. How many shells/layers are there in your onion? To which period in the periodic table does this correspond?

c. Note how far each layer is from the “nucleus.” What does this imply about the relative attractive forces between each layer and the nucleus?

d. Now peel the successive layers of the onion. How big is the outermost layer compared with the inner layers? What does this imply about the number of electrons that can fit in each layer?

2.7 Electron Configurations

Learning Objective:
- Write the electronic configuration for an atom to describe how electrons are distributed into specific orbitals.

The exact arrangement of electrons in an atom’s shells and subshells is called the atom’s electron configuration and can be predicted by applying three rules:

RULE 1: Electrons occupy the lowest-energy orbitals available, beginning with 1s. Within each shell, the orbital energies increase in the order s, p, d, and f. For the first three periods, the order of energy is as follows: 1s, 2s, 2p, 3s, and 3p. Across shells, the orbital closer to the nucleus is lower in energy. For example, a 2s orbital has lower energy than a 3s orbital. As a result, above the 3p level the order of how shells are filled is not as straightforward. For example, the 4s orbital is lower in energy than the 3d orbitals and is therefore filled first. The energy level diagram and simple scheme shown in Figure 2.6 can be used to predict the order in which orbitals are filled. Neither of these diagrams need to be memorized, however, as you can also use the periodic table to determine the order in which orbitals are filled in relation to their placement, shown later in Section 2.8.

RULE 2: Each orbital can hold only two electrons, which must be of opposite spin.

RULE 3: Two or more orbitals with the same energy are each half-filled by one electron before any one orbital is completely filled by the addition of the second electron. For example, one electron is added to each of the three p orbitals before a second electron is added to fill an orbital.

Electron configurations of the first 20 elements are shown in Table 2.3. Notice that the number of electrons in each subshell is indicated by a superscript. For example, the notation 1s\(^2\) 2s\(^2\) 2p\(^6\) 3s\(^2\) for magnesium means that magnesium atoms have two electrons in the first shell, eight electrons in the second shell, and two electrons in the third shell.

In addition to writing the configurations as shown above, we can also use orbital diagrams. In the written representation, the superscript in the notation 1s\(^2\) means that the 1s orbital is occupied by only one electron. In an orbital diagram, the 1s orbital is indicated by a line or a box and the single electron in this orbital is shown by a single arrow pointing up (↑). A single electron in an orbital is often referred to as being unpaired. Two electrons in an orbital are paired, with spins in opposite directions, so they are represented by two arrows pointing in opposite directions (one up, one down).
### Table 2.3 Electron Configurations of the First 20 Elements

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Number</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1</td>
<td>1s¹</td>
</tr>
<tr>
<td>He</td>
<td>2</td>
<td>1s²</td>
</tr>
<tr>
<td>Li</td>
<td>3</td>
<td>1s² 2s¹</td>
</tr>
<tr>
<td>Be</td>
<td>4</td>
<td>1s² 2s² 2p¹</td>
</tr>
<tr>
<td>B</td>
<td>5</td>
<td>1s² 2s² 2p¹</td>
</tr>
<tr>
<td>C</td>
<td>6</td>
<td>1s² 2s² 2p²</td>
</tr>
<tr>
<td>N</td>
<td>7</td>
<td>1s² 2s² 2p³</td>
</tr>
<tr>
<td>O</td>
<td>8</td>
<td>1s² 2s² 2p⁴</td>
</tr>
<tr>
<td>F</td>
<td>9</td>
<td>1s² 2s² 2p⁵</td>
</tr>
<tr>
<td>Ne</td>
<td>10</td>
<td>1s² 2s² 2p⁶</td>
</tr>
<tr>
<td>Na</td>
<td>11</td>
<td>1s² 2s² 2p⁶ 3s¹</td>
</tr>
<tr>
<td>Mg</td>
<td>12</td>
<td>1s² 2s² 2p⁶ 3s² 3p¹</td>
</tr>
<tr>
<td>Al</td>
<td>13</td>
<td>1s² 2s² 2p⁶ 3s² 3p³</td>
</tr>
<tr>
<td>Si</td>
<td>14</td>
<td>1s² 2s² 2p⁶ 3s² 3p⁴</td>
</tr>
<tr>
<td>P</td>
<td>15</td>
<td>1s² 2s² 2p⁶ 3s² 3p⁵</td>
</tr>
<tr>
<td>S</td>
<td>16</td>
<td>1s² 2s² 2p⁶ 3s² 3p⁶</td>
</tr>
<tr>
<td>Cl</td>
<td>17</td>
<td>1s² 2s² 2p⁶ 3s² 3p⁶ 4s¹</td>
</tr>
<tr>
<td>Ar</td>
<td>18</td>
<td>1s² 2s² 2p⁶ 3s² 3p⁶ 4s²</td>
</tr>
<tr>
<td>K</td>
<td>19</td>
<td>1s² 2s² 2p⁶ 3s² 3p⁶ 4s²</td>
</tr>
<tr>
<td>Ca</td>
<td>20</td>
<td>1s² 2s² 2p⁶ 3s² 3p⁶ 4s²</td>
</tr>
</tbody>
</table>

As you read through the following electron configurations, check the atomic number and the location of each element in the periodic table (Figure 2.2). See if you can detect the relationship between electron configuration and position in the table.

- **Hydrogen (Z = 1):** The single electron in a hydrogen atom is in the lowest-energy, 1s, level. The configuration can be represented in either of two ways:

  \[
  H \quad 1s^1 \quad \text{or} \quad \uparrow \downarrow 1s^1
  \]

- **Helium (Z = 2):** The two electrons in helium are both in the lowest-energy, 1s, orbital, and their spins are paired, as represented by up and down arrows (\(\uparrow \downarrow\)). Helium has a completely filled first shell \((n = 1)\) of electrons.

  \[
  \text{He} \quad 1s^2 \quad \text{or} \quad \uparrow \downarrow 1s^2
  \]

- **Lithium (Z = 3):** Lithium has three electrons, so we must now use the orbitals in the second shell, starting with 2s. Since electrons are always added to the lowest energy level first, the first two electrons fill the first shell. Next, the second shell begins to fill. The third electron goes into the 2s orbital and is unpaired:

  \[
  \text{Li} \quad 1s^2 2s^1 \quad \text{or} \quad \uparrow \uparrow \uparrow 1s^2 2s^1
  \]

Because \([\text{He}]\) has the configuration of a filled 1s² orbital, it is sometimes substituted for the 1s² orbital in depictions of electron pairing. Using this alternative shorthand notation, the electron configuration for Li is written \([\text{He}] 2s^1\).
- **Beryllium (Z = 4):** For beryllium's four electrons, we continue to use the second shell. The three electrons are configured as they were for lithium, and the fourth electron pairs up to fill the 2s orbital:

\[
\text{Be} \quad 1s^2 \ 2s^2 \text{ or } \begin{array}{c}
1s^2 \\
2s^2 \\
\end{array} \text{ or } \begin{array}{c}
\text{He} \\
2s^2 \\
\end{array}
\]

- **Boron (Z = 5), Carbon (Z = 6), Nitrogen (Z = 7):** The next three elements use the three 2p orbitals, one at a time. For boron, the fifth electron starts to fill the first 2p orbital. Carbon and nitrogen's sixth and seventh electron are placed in the next two orbitals, respectively (instead of filling the first p orbital). Note that representing the configurations with lines and arrows gives more information than the alternative written notations because the filling and pairing of electrons in individual orbitals within the p subshell is shown.

\[
\begin{array}{c}
\text{B} \\
1s^2 \ 2s^2 \ 2p^1 \\
\end{array} \text{ or } \begin{array}{c}
1s^2 \\
2s^2 \\
\end{array} \begin{array}{c}
\text{He} \\
2s^2 \\
\end{array} \\
\]

\[
\begin{array}{c}
\text{C} \\
1s^2 \ 2s^2 \ 2p^2 \\
\end{array} \text{ or } \begin{array}{c}
1s^2 \\
2s^2 \\
\end{array} \begin{array}{c}
\text{He} \\
2s^2 \\
\end{array} \\
\]

\[
\begin{array}{c}
\text{N} \\
1s^2 \ 2s^2 \ 2p^3 \\
\end{array} \text{ or } \begin{array}{c}
1s^2 \\
2s^2 \\
\end{array} \begin{array}{c}
\text{He} \\
2s^2 \\
\end{array} \\
\]

- **Oxygen (Z = 8), Fluorine (Z = 9), Neon (Z = 10):** Electrons now pair up one by one to fill the three 2p orbitals and fully occupy the second shell.

\[
\begin{array}{c}
\text{O} \\
1s^2 \ 2s^2 \ 2p^4 \\
\end{array} \text{ or } \begin{array}{c}
1s^2 \\
2s^2 \\
\end{array} \begin{array}{c}
\text{He} \\
2s^2 \\
\end{array} \\
\]

\[
\begin{array}{c}
\text{F} \\
1s^2 \ 2s^2 \ 2p^5 \\
\end{array} \text{ or } \begin{array}{c}
1s^2 \\
2s^2 \\
\end{array} \begin{array}{c}
\text{He} \\
2s^2 \\
\end{array} \\
\]

\[
\begin{array}{c}
\text{Ne} \\
1s^2 \ 2s^2 \ 2p^6 \\
\end{array} \text{ or } \begin{array}{c}
1s^2 \\
2s^2 \\
\end{array} \begin{array}{c}
\text{He} \\
2s^2 \\
\end{array} \\
\]

Just as [He] was used as a shorthand notation to indicate the closed-shell configuration 1s², we may also use [Ne] to represent the electron configuration for a completely filled set of orbitals in the second shell, or 1s²2s²2p⁶. Both helium and neon are noble gases and are located in Group 8A on the periodic table. All of electron configurations of the elements in one period can be written in shorthand using the noble gas that immediately precedes it in the periodic table.

- **Sodium to Calcium (Z = 11 – 20):** The pattern seen for lithium through neon is seen again for sodium (Z = 11) through argon (Z = 18) as the 3s and 3p subshells fill up. For elements having a third filled shell, we may use [Ar] to represent a completely filled third shell. After argon, however, the first crossover in subshell energies occurs. As indicated in Figure 2.6, the 4s subshell is lower in energy than the 3d subshell and is filled first. Potassium (Z = 19) and calcium (Z = 20), therefore, have the following electron configurations:

\[
\begin{array}{c}
\text{K} \\
1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^1 \\
\end{array} \text{ or } \begin{array}{c}
\text{Ar} \ 4s^1 \\
\end{array} \\
\]

\[
\begin{array}{c}
\text{Ca} \\
1s^2 \ 2s^2 \ 2p^6 \ 3s^2 \ 3p^6 \ 4s^2 \\
\end{array} \text{ or } \begin{array}{c}
\text{Ar} \ 4s^2 \\
\end{array} \\
\]

After calcium we enter the transition metals, and the subsequent electrons for these elements would be placed into the next lowest energy orbitals, or the 3d.
**Worked Example 2.6** Atomic Structure: Electron Configurations

Show how the electron configuration of magnesium can be assigned.

**ANALYSIS** Magnesium, Z = 12, has 12 electrons to be placed in specific orbitals. Assignments are made by putting two electrons in each orbital, according to the order shown in Figure 2.6.

- The first two electrons are placed in the 1s orbital (1s²).
- The next two electrons are placed in the 2s orbital (2s²).
- The next six electrons are placed in the three available 2p orbitals (2p⁶).
- The remaining two electrons are both put in the 3s orbital (3s²).

**SOLUTION**
Magnesium has the configuration 1s²2s²2p⁶3s² or [Ne]3s².

**Worked Example 2.7** Electron Configurations: Orbital-Filling Diagrams

Write the electron diagram of phosphorus, Z = 15, using up and down arrows to show how the electrons in each orbital are paired.

**ANALYSIS** Phosphorus has 15 electrons, which occupy orbitals according to the order shown in Figure 2.6.

- The first two are paired and fill the first shell (1s²).
- The next eight fill the second shell (2s²2p⁶). All electrons are paired.
- The remaining five electrons enter the third shell, where two fill the 3s orbital (3s²) and three occupy the 3p subshell, one in each of the three p orbitals.

**SOLUTION**

PROBLEM 2.15
An element has completely filled n = 1 and n = 2 shells and has six electrons in the n = 3 shell. Identify the element and its major group (i.e., main group, transition, etc.). Is it a metal or a nonmetal? Identify the orbital in which the last electron is found.

PROBLEM 2.16
Write electron configurations for the following elements. (You can check your answers in Table 2.3.)

(a) C  (b) P  (c) Cl   (d) K

PROBLEM 2.17
For an atom containing 33 electrons, identify the incompletely filled subshell and show the paired and/or unpaired electrons in this subshell using up and down arrows.

**KEY CONCEPT PROBLEM 2.18**
Identify the atom with the following orbital-filling diagram.

```
1s² 2s² 2p⁶ 3s² 3p⁶ 4s 3d 4p
```
2.8 Electron Configurations and the Periodic Table

Learning Objective:
- Identify the valence shell electrons for an atom, and which subshell of electrons \( [s, p, d, f] \) correlate with which groups in the periodic table.

How is an atom's electron configuration related to its chemical behavior, and why do elements with similar behavior occur in the same group of the periodic table? As shown in Figure 2.7, the periodic table can be divided into four regions, or blocks, of elements according to the electron shells and subshells occupied by the subshell filled last.

- The main group 1A and 2A elements on the left side of the table (plus He) are called the **s-block elements** because an \( s \) subshell is filled last in these elements.
- The main group 3A–8A elements on the right side of the table (except He) are the **p-block elements** because a \( p \) subshell is filled last in these elements.
- The transition metals in the middle of the table are the **d-block elements** because a \( d \) subshell is filled last in these elements.
- The inner transition metals detached at the bottom of the table are the **f-block elements** because an \( f \) subshell is filled last in these elements.

![Figure 2.7](image)

**The blocks of elements in the periodic table correspond to filling the different types of subshells.**

Beginning at the top left and going across successive rows of the periodic table provides a method for remembering the order of orbital filling:

- \( 1s \to 2s \to 2p \to 3s \to 3p \to 4s \to 3d \to 4p \), and so on.

Thinking of the periodic table as outlined in Figure 2.7 provides a simple way to remember the order of orbital filling shown previously in Figure 2.6. Beginning at the top left corner of the periodic table, the first row contains only two elements (H and He) because only two electrons are required to fill the \( 1s \) orbital in the first shell, \( 1s^2 \). The second row begins with two \( s \)-block elements (Li and Be) and continues with six \( p \)-block elements (B through Ne), so electrons fill the next available \( s \) orbital (\( 2s \)) and then the first available \( p \) orbitals (\( 2p \)). The third row is similar to the second row, so the \( 3s \) and \( 3p \) orbitals are filled next. The fourth row again starts with two \( s \)-block elements (K and Ca) but is then followed by 10 \( d \)-block elements (Sc through Zn) and 6 \( p \)-block elements (Ga through Kr). Thus, the order of orbital filling is \( 4s \) followed by the first
available d orbitals (3d) followed by 4p. Continuing through successive rows of the periodic table gives the entire filling order, identical to that shown in Figure 2.6.

\[ \begin{align*}
1s &\rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p \rightarrow 5s \\
4d &\rightarrow 5p \rightarrow 6s \rightarrow 4f \rightarrow 5d \rightarrow 6p \rightarrow 7s \rightarrow 5f \rightarrow 6d \rightarrow 7p
\end{align*} \]

But why do the elements in a given group of the periodic table have similar properties? The answer emerges when you look at Table 2.4, which gives electron configurations for elements in the main groups 1A, 2A, 7A, and 8A. Focusing only on the electrons in the outermost shell, or valence shell, elements in the same group of the periodic table have similar electron configurations in their valence shells. The group 1A elements, for example, all have one valence electron, \( ns^1 \) (where \( n \) represents the number of the valence shell: \( n = 2 \) for Li, \( n = 3 \) for Na, \( n = 4 \) for K, and so on). The group 2A elements have two valence electrons \( (ns)^2 \). The group 7A elements have seven valence electrons \( (ns^2 np^5) \). For example, fluorine (F) has the electron configuration of \( 1s^22s^22p^5 \) (valence electrons in bold). The group 8A elements (except He) have eight valence electrons \( (ns^2 np^6) \). You might also notice that the group numbers from 1A through 8A give the numbers of valence electrons for the elements in each main group. It is worth noting that the valence electrons are those in the outermost shell \( (n) \)—not necessarily in the orbitals that were filled last!

Table 2.4 Valence-Shell Electron Configurations for Groups 1A, 2A, 7A, and 8A Elements

<table>
<thead>
<tr>
<th>Group</th>
<th>Element</th>
<th>Atomic Number</th>
<th>Valence-Shell Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>1A</td>
<td>Li [lithium]</td>
<td>3</td>
<td>2s(^1)</td>
</tr>
<tr>
<td></td>
<td>Na [sodium]</td>
<td>11</td>
<td>3s(^1)</td>
</tr>
<tr>
<td></td>
<td>K [potassium]</td>
<td>19</td>
<td>4s(^1)</td>
</tr>
<tr>
<td></td>
<td>Rb [rubidium]</td>
<td>37</td>
<td>5s(^1)</td>
</tr>
<tr>
<td></td>
<td>Cs [cesium]</td>
<td>55</td>
<td>6s(^1)</td>
</tr>
<tr>
<td>2A</td>
<td>Be [beryllium]</td>
<td>4</td>
<td>2s(^2)</td>
</tr>
<tr>
<td></td>
<td>Mg [magnesium]</td>
<td>12</td>
<td>3s(^2)</td>
</tr>
<tr>
<td></td>
<td>Ca [calcium]</td>
<td>20</td>
<td>4s(^2)</td>
</tr>
<tr>
<td></td>
<td>Sr [strontium]</td>
<td>38</td>
<td>5s(^2)</td>
</tr>
<tr>
<td></td>
<td>Ba [barium]</td>
<td>56</td>
<td>6s(^2)</td>
</tr>
<tr>
<td>7A</td>
<td>F [fluorine]</td>
<td>9</td>
<td>2s(^1)2p(^5)</td>
</tr>
<tr>
<td></td>
<td>Cl [chlorine]</td>
<td>17</td>
<td>3s(^1)3p(^5)</td>
</tr>
<tr>
<td></td>
<td>Br [bromine]</td>
<td>35</td>
<td>4s(^2)4p(^5)</td>
</tr>
<tr>
<td></td>
<td>I [iodine]</td>
<td>53</td>
<td>5s(^2)5p(^5)</td>
</tr>
<tr>
<td>8A</td>
<td>He [helium]</td>
<td>2</td>
<td>1s(^2)</td>
</tr>
<tr>
<td></td>
<td>Ne [neon]</td>
<td>10</td>
<td>2s(^1)2p(^6)</td>
</tr>
<tr>
<td></td>
<td>Ar [argon]</td>
<td>18</td>
<td>3s(^2)3p(^6)</td>
</tr>
<tr>
<td></td>
<td>Kr [krypton]</td>
<td>36</td>
<td>4s(^2)4p(^6)</td>
</tr>
<tr>
<td></td>
<td>Xe [xenon]</td>
<td>54</td>
<td>5s(^2)5p(^6)</td>
</tr>
</tbody>
</table>

What is true for the main group elements is also true for the other groups in the periodic table: atoms within a given group have the same number of valence electrons and have similar electron configurations. Because the valence electrons are the most loosely held, they are the most important in determining an element's properties. Similar electron configurations thus explain why the elements in a given group of the periodic table have similar chemical behavior.

**Valence shell** The outermost electron shell of an atom.

**Valence electron** An electron in the valence shell of an atom.
Worked Example 2.8  Electron Configurations: Valence Electrons

Write the electron configuration for the following elements, using both the complete and the shorthand notations. Indicate which electrons are the valence electrons.

(a) Na

(b) Cl

(c) Zr

ANALYSIS Locate the row and the block in which each of the elements is found in Figure 2.7. The location can be used to determine the complete electron configuration and to identify the valence electrons.

SOLUTION

(a) Na (sodium) is located in the third row and in the first column of the s-block. Therefore, all orbitals up to the 3s are completely filled, and there is one electron in the 3s orbital.

\[
Na: 1s^2 2s^2 2p^6 3s^1 \text{ or } [\text{Ne}] 3s^1 \text{ (valence electrons are underlined)}
\]

(b) Cl (chlorine) is located in the third row and in the fifth column of the p-block. Therefore, there are five electrons in the 3p orbital.

\[
Cl: 1s^2 2s^2 2p^6 3s^2 3p^5 \text{ or } [\text{Ne}] 3s^2 3p^5
\]

(c) Zr (zirconium) is located in the fifth row and in the second column of the d-block. All orbitals up to the 4d are completely filled, and there are two electrons in the 4d orbitals. Note that the 4d orbitals are filled after the 5s orbitals in both Figures 2.6 and 2.7.

\[
Zr: 1s^2 2s^2 2p^6 3s^1 3p^6 4s^2 3d^{10} 4p^6 5s^2 5p^2 \text{ or } [\text{Kr}] 5s^2 4d^2
\]

Worked Example 2.9  Electron Configurations: Valence-Shell Configurations

Using \( n \) to represent the number of the valence shell, write a general valence-shell configuration for the elements in group 6A.

ANALYSIS The elements in group 6A have six valence electrons. In each element, the first two of these electrons are in the valence \( s \) subshell, giving \( ns^2 \), and the next four electrons are in the valence \( p \) subshell, giving \( np^4 \).

SOLUTION For group 6A, the general valence-shell configuration is \( ns^2 np^4 \).

Worked Example 2.10  Electron Configurations: Inner Shells versus Valence Shell

How many electrons are in a tin atom? Give the number of electrons in each shell. How many valence electrons are there in a tin atom? Write the valence-shell configuration for tin.

ANALYSIS The total number of electrons will be the same as the atomic number for tin \((Z = 50)\). The number of valence electrons will equal the number of electrons in the valence shell.

SOLUTION Checking the periodic table shows that tin (Sn) has atomic number 50 and is in group 4A. The number of electrons in each shell is

<table>
<thead>
<tr>
<th>Shell number</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of electrons</td>
<td>2</td>
<td>8</td>
<td>18</td>
<td>18</td>
<td>4</td>
</tr>
</tbody>
</table>

As expected from the group number, tin has four valence electrons. They are in the 5s and 5p subshells and have the configuration \( 5s^2 5p^2 \). Although there are \( f \) orbitals available in the \( n = 4 \) shell, the 5s orbital is of lower energy than the 4f orbitals, and so will fill first. Hence, there are only 18 electrons in the \( n = 4 \) shell.
### PROBLEM 2.19
Write the electron configuration for the following elements, using both the complete and the shorthand notations. Indicate which electrons are the valence electrons.

(a) F  
(b) Al  
(c) As

### PROBLEM 2.20
Identify the group in which all the elements have the valence-shell configuration $ns^2$.

### PROBLEM 2.21
For chlorine, identify the group number, give the number of electrons in each occupied shell, and write its valence-shell configuration.

#### KEY CONCEPT PROBLEM 2.22
Identify the group number and write the general valence-shell configuration (e.g., $ns^1$ for group 1A elements) for the elements indicated in red in the following periodic table.

![](image)

### 2.9 Electron-Dot Symbols

**Learning Objective:**
- Write Lewis dot symbols to represent the valence electrons for a given atom.

Valence electrons play such an important role in the behavior of atoms that it is useful to have a method for including them with atomic symbols. In an electron-dot symbol (also called Lewis symbols), dots are placed around the atomic symbol to indicate the number of valence electrons present. A group 1A atom, such as sodium, has a single dot; a group 2A atom, such as magnesium, has two dots; a group 3A atom, such as boron, has three dots; and so on.

Table 2.5 gives electron-dot symbols for atoms of the first few elements in each main group. As shown, the dots are distributed around the four sides of the element symbol, singly at first until each of the four sides has one dot. As more electron dots are added they will form pairs, with no more than two dots on a side. Note that helium differs from other noble gases in having only two valence electrons rather than eight. Nevertheless, helium is considered a member of group 8A because its properties resemble those of the other noble gases and because its highest occupied subshell is filled ($1s^2$).

**Table 2.5 Electron-Dot Symbols for Some Main Group Elements**

<table>
<thead>
<tr>
<th></th>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>Noble Gases</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>H⁺</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>He⁺</td>
</tr>
<tr>
<td>Li⁺</td>
<td>Be⁺</td>
<td>B⁺</td>
<td>C⁺</td>
<td>N⁺</td>
<td>O⁺</td>
<td>F⁺</td>
<td>Ne⁺</td>
<td></td>
</tr>
<tr>
<td>Na⁺</td>
<td>Mg⁺</td>
<td>Al⁺</td>
<td>Si⁺</td>
<td>P⁺</td>
<td>S⁺</td>
<td>Cl⁺</td>
<td>Ar⁺</td>
<td></td>
</tr>
<tr>
<td>K⁺</td>
<td>Ca⁺</td>
<td>Ga⁺</td>
<td>Ge⁺</td>
<td>As⁺</td>
<td>Se⁺</td>
<td>Br⁺</td>
<td>Kr⁺</td>
<td></td>
</tr>
</tbody>
</table>

**Electron-dot (Lewis) symbol** An atomic symbol with dots placed around it to indicate the number of valence electrons.
Worked Example 2.11  Electron Configurations: Electron-Dot Symbols

Write the electron-dot symbol for any element X in group 5A.

**ANALYSIS**  The group number, 5A, indicates five valence electrons. The first four are distributed singly around the four sides of the element symbol, and any additional are placed to form electron pairs.

**SOLUTION**  
\[ \text{X: (5 electrons)} \]

**PROBLEM 2.23**
Write the electron-dot symbol for any element X in group 3A.

**PROBLEM 2.24**
Write electron-dot symbols for radon, lead, xenon, and radium.

**PROBLEM 2.25**
When an electron in a strontium atom drops from the excited state to the ground state, it emits red light, as explained in the following Chemistry in Action feature. When an electron in a copper atom drops from the excited state to the ground state, it emits blue light. What are the approximate wavelengths of the red light and the blue light? Which color is associated with higher energy?

**CHEMISTRY IN ACTION**

**Atoms and Light**
What we see as light is really a wave of energy moving through space. The shorter the length of the wave (the wavelength), the higher the energy; the longer the wavelength, the lower the energy.

<table>
<thead>
<tr>
<th>Shorter wavelength (higher energy)</th>
<th>Longer wavelength (lower energy)</th>
</tr>
</thead>
</table>

What happens when a beam of electromagnetic energy collides with an atom? Remember that electrons are located in orbitals based on their energy levels. An atom with its electrons in their usual, lowest-energy locations is said to be in its ground state. If the amount of electromagnetic energy is just right, an electron can be kicked up from its usual energy level to a higher one. Energy from an electrical discharge or in the form of heat can also boost electrons to higher energy levels. With one of its electrons promoted to a higher energy, an atom is said to be excited. The excited state does not last long, though, because the electron quickly drops back to its more stable, ground-state energy level, releasing its extra energy in the process. If the released energy falls in the range of visible light (400–800 nm), we can see the result. Many practical applications, from neon lights to fireworks, are the result of this phenomenon.

[This chest X ray equipment uses high energy, short wavelength electromagnetic radiation to generate diagnostic images.]
Summary

The interaction of light with matter has many significant impacts. The UV radiation (200–350 nm) from the sun has enough energy to cause sunburn and, with chronic long-term exposure, can lead to skin cancers. Higher energy radiation (X rays) is used in many diagnostic applications, while even higher energies (gamma rays) can be used to kill cancer cells. In clinical applications, the concentration of certain biologically important metals in body fluids, such as blood or urine, is measured by sensitive instruments (such as the spectrophotometer mentioned at the beginning of this chapter), relying on the principle of electron excitation, where metal atoms will emit light of a specific wavelength corresponding to electronic transitions in the atom. These instruments measure the intensity of color produced in a flame by lithium (red), sodium (yellow), and potassium (violet), to determine the concentrations of these metals, which should be in a certain range to ensure good health. If the levels of these and other essential metals are outside the optimal range, it may be an indication of poor nutrition or certain diseases.

CIA Problem 2.5 Which type of electromagnetic energy in the following pairs is of higher energy?
(a) Infrared, ultraviolet
(b) Gamma waves, microwaves
(c) Visible light, X rays

CIA Problem 2.6 Why do you suppose ultraviolet rays from the sun are more damaging to the skin than visible light?

SUMMARY REVISING THE CHAPTER LEARNING OBJECTIVES

• Explain the major assumptions of atomic theory, and name and identify the properties of the subatomic particles that make up an atom. All matter is composed of atoms. An atom is the smallest and simplest unit into which a sample of an element can be divided while maintaining the properties of the element. Atoms are made up of subatomic particles called protons, neutrons, and electrons. Protons have a positive electrical charge, neutrons are electrically neutral, and electrons have a negative electrical charge. The protons and neutrons in an atom are present in a dense, positively charged central region called the nucleus. Electrons are situated a relatively large distance away from the nucleus, leaving most of the atom as empty space (see Problems 31–40, 83, 85, 89, and 95).

• Identify atoms of an element based on the number of protons in the nucleus. Elements differ according to the number of protons their atoms contain, a value called the element’s atomic number (Z). All atoms of a given element have the same number of protons and an equal number of electrons. The number of neutrons in an atom is not predictable but is generally equal to or greater than the number of protons. The total number of protons plus neutrons in an atom is called the atom’s mass number (A) (see Problems 43, 46, and 82).

• Write the symbols for different isotopes of an element and use relative abundances and atomic masses of isotopes to calculate the average atomic weight of an element. The symbol for an atom is written using the symbol for the element (e.g., C for carbon), including
the atomic number \(Z\) as a subscript on the left, and the atomic mass \(A\) as a superscript. An atom of carbon-12 [6 protons + 6 neutrons] would be represented as \(^{12}\text{C}\). Atoms with identical numbers of protons and electrons but different numbers of neutrons are called isotopes. The atomic weight of an element is the weighted average mass of atoms of the element’s naturally occurring isotopes (see Problems 27, 40–49, 52, and 96).

- **Locate elements on the periodic table and classify them as metals, nonmetals, or metalloids based on their location.** The majority of elements are identified as metals and are located to the left/bottom of the periodic table. Only 18 elements are identified as nonmetals, and they are located to the upper right of the periodic table. Metalloids are located on a diagonal between the metals and nonmetals (see Problems 26, 52–57, 86, and 88).

- **Classify elements and describe chemical behavior based on group membership.** Elements are organized into the periodic table, consisting of 7 rows, or periods, and 18 columns, or groups. The two columns on the left side of the table and the six columns on the right are called the main groups. The 10 columns in the middle are the transition metal groups, and the 14 columns pulled out and displayed below the main part of the table are called the inner transition metal groups (see Problems 27, 56–61, 80–82, 97, and 98).

- **Describe the distribution of electrons into shells, subshells, and orbitals around the nucleus of an atom.** The electrons surrounding an atom are grouped into layers, or shells: Within each shell, electrons are grouped into subshells, and within each subshell into orbitals—regions of space in which electrons are most likely to be found. The s orbitals are spherical, and the p orbitals are dumbbell-shaped. Each shell can hold a specific number of electrons. The first shell can hold 2 electrons, the second shell can hold 8 electrons, the third shell can hold 18 electrons, and so on (see Problems 50, 51, and 62–69).

- **Write the electronic configuration for an atom to describe how electrons are distributed into specific orbitals.** The electron configuration of an element is predicted by assigning the element’s electrons into shells and orbitals, beginning with the lowest-energy orbital. For example, the first shell can hold 2 electrons in an s orbital \((1s^2)\); the second shell can hold 8 electrons in one s and three p orbitals \((2s^22p^6)\); the third shell can hold 18 electrons in one s, three p, and five d orbitals \((3s^23p^63d^{10})\); and so on (see Problems 29, 30, 68–74, 83, 84, 86, 87, and 90–94).

- **Identify the valence shell electrons for an atom, and which subshell of electrons \([s, p, d, f]\) correlate with which groups in the periodic table.** The valence electrons for an atom are found in the outermost shell and correspond to the location of the element in the periodic table. The number of valence electrons for the main group elements corresponds to the group number. The valence electrons for the 1A and 2A elements are located in s orbitals, while the valence electrons for groups 3A–8A are in p orbitals. Electrons in the d orbitals are associated with the transition metals, while f orbitals are associated with the inner transition metals [lanthanide and actinide series]. Within a given group in the table, elements have the same number of valence electrons in their valence shell and similar electron configurations (see Problems 28, 53, 74–79, 84, 86–88, and 94).

- **Write Lewis dot symbols to represent the valence electrons for a given atom.** The number of valence electrons is determined by the location of the element in the periodic table. The Lewis dot symbol for an atom is written as the chemical symbol for the element \((C\) for carbon) with the valence electrons represented as dots around the symbol. If there are only four [or fewer] valence electrons, they are written as single dots above, below, and to the left and right sides of the symbol. If there are more than four valence electrons, then the extra electrons are added to form electron pairs (see Problems 78, 81, and 86).

### KEY WORDS

- Alkali metal, p. 53
- Alkaline earth metal, p. 53
- Atom, p. 43
- Atomic mass unit (amu), p. 43
- Atomic number \((Z)\), p. 45
- Atomic theory, p. 43
- Atomic weight, p. 47
- d-Block element, p. 62
- Electron, p. 43
- Electron configuration, p. 58
- Electron-dot (Lewis) symbol, p. 65
- f-Block element, p. 62
- Group, p. 52
- Halogen, p. 54
- Inner transition metal element, p. 52
- Isotopes, p. 46
- Main group element, p. 52
- Mass number \((A)\), p. 46
- Metal, p. 50
- Metalloid, p. 50
- Neutron, p. 43
- Noble gas, p. 54
- Nonmetal, p. 50
- Nucleus, p. 44
- Orbital, p. 56
- Orbital diagram, p. 58
- Periodic table, p. 49
- p-Block element, p. 62
- Period, p. 52
- Proton, p. 43
- s-Block element, p. 62
- Shell (electron), p. 56
- Subatomic particles, p. 43
- Subshell (electron), p. 56
- Transition metal element, p. 52
- Valence electron, p. 63
- Valence shell, p. 63

### UNDERSTANDING KEY CONCEPTS

2.26 Where on the following outline of a periodic table do the indicated elements or groups of elements appear?

(a) Alkali metals  (b) Halogens
(c) Alkaline earth metals  (d) Transition metals
(e) Hydrogen  (f) Helium
(g) Metalloids
2.27 Is the element marked in red on the following periodic table likely to be a gas, a liquid, or a solid? What is the atomic number of the element in blue? Name at least one other element that is likely to be similar to the element in green.

Additional Problems

2.29 What atom has the following orbital-filling diagram?

\[ 1s^2 \, 2s^2 \, 2p^6 \, 3s^2 \, 3p^6 \]

2.30 Use the following orbital-filling diagram to show the electron configuration for As:

\[ 1s^2 \, 2s^2 \, 2p^6 \, 3s^2 \, 3p^6 \]

2.31 What four fundamental assumptions about atoms and matter make up modern atomic theory?

2.32 How do atoms of different elements differ?

2.33 Find the mass in grams of one atom of the following elements:
   (a) Bi, atomic weight 208.9804 amu
   (b) Xe, atomic weight 131.29 amu
   (c) He, atomic weight 4.0026 amu

2.34 Find the mass in atomic mass units of the following:
   (a) O atom, with a mass of \( 2.66 \times 10^{-23} \) g
   (b) Br atom, with a mass of \( 1.31 \times 10^{-22} \) g

2.35 What is the mass in grams of 6.022 \( \times 10^{23} \) N atoms of mass 14.01 amu?

2.36 What is the mass in grams of 6.022 \( \times 10^{23} \) O atoms of mass 16.00 amu?

2.37 How many O atoms of mass 15.99 amu are in 15.99 g of oxygen?

2.38 How many C atoms of mass 12.00 amu are in 12.00 g of carbon?

2.39 What are the names of the three subatomic particles? What are their approximate masses in atomic mass units, and what electrical charge does each have?

2.40 Where within an atom are the three types of subatomic particles located?

2.41 Give the number of neutrons in each naturally occurring isotope of argon: argon-36, argon-38, argon-40.

2.42 Give the number of protons, neutrons, and electrons in the following isotopes:
   (a) Al-27
   (b) Si-28
   (c) B-11
   (d) Ag-107

2.43 Which of the following symbols represent isotopes of the same element? Explain.
   (a) \( ^{18}_8\text{X} \)
   (b) \( ^{16}_8\text{X} \)
   (c) \( ^{20}_8\text{X} \)
   (d) \( ^{19}_8\text{X} \)

2.44 Give the name and the number of neutrons in each isotope listed in Problem 2.43.

2.45 Write the symbols for the following isotopes:
   (a) Its atoms contain 6 protons and 8 neutrons.
   (b) Its atoms have mass number 39 and contain 19 protons.
   (c) Its atoms have mass number 20 and contain 10 electrons.

2.46 Write the symbols for the following isotopes:
   (a) Its atoms contain 50 electrons and 70 neutrons.
   (b) Its atoms have \( A = 56 \) and \( Z = 26 \).
   (c) Its atoms have \( A = 226 \) and contain 88 electrons.

2.47 One of the most widely used isotopes in medical diagnostics is technetium-99m (the \( m \) indicates that it is a metastable isotope). Write the symbol for this isotope, indicating both mass number and atomic number.

2.48 Naturally occurring copper is a mixture of 69.17% Cu-63 with a mass of 62.93 amu and 30.83% Cu-65 with a mass of 64.93 amu. What is the atomic weight of copper?

2.49 Naturally occurring lithium is a mixture of 92.58% Li-7 with a mass of 7.016 amu and 7.42% Li-6 with a mass of 6.015 amu. What is the atomic weight of lithium?

The Periodic Table (Sections 2.4–2.6)

2.50 Why does the third period in the periodic table contain eight elements?

2.51 Why does the fourth period in the periodic table contain 18 elements?
2.52 Americium, atomic number 95, is used in household smoke detectors. What is the symbol for americium? Is americium a metal, a nonmetal, or a metalloid?

2.53 What subshell is being filled for the metalloid elements?

2.54 Answer the following questions for the elements from scandium through zinc:
(a) Are they metals or nonmetals?
(b) To what general class of elements do they belong?
(c) What subshell is being filled by electrons in these elements?

2.55 Answer the following questions for the elements from cerium through lutetium:
(a) Are they metals or nonmetals?
(b) To what general class of elements do they belong?
(c) What subshell is being filled by electrons in these elements?

2.56 For (a) rubidium (b) tungsten, (c) germanium, and (d) krypton, which of the following terms apply? (i) metal, (ii) nonmetal, (iii) metalloid (iv) transition element, (v) main group element, (vi) noble gas, (vii) alkali metal, (viii) alkaline earth metal.

2.57 For (a) calcium, (b) palladium, (c) carbon, and (d) radon, which of the following terms apply? (i) metal, (ii) nonmetal, (iii) metalloid, (iv) transition element, (v) main group element, (vi) noble gas, (vii) alkali metal, (viii) alkaline earth metal.

2.58 Name an element in the periodic table that you would expect to be chemically similar to sulfur.

2.59 Name an element in the periodic table that you would expect to be chemically similar to potassium.

2.60 What elements in addition to lithium make up the alkali metal family?

2.61 What elements in addition to fluorine make up the halogen family?

**ELECTRON CONFIGURATIONS (SECTIONS 2.6–2.9)**

2.62 What is the maximum number of electrons that can go into an orbital?

2.63 What are the shapes and locations within an atom of $s$ and $p$ orbitals?

2.64 What is the maximum number of electrons that can go into the first shell? The second shell? The third shell?

2.65 What is the total number of orbitals in the third shell? The fourth shell?

2.66 How many subshells are there in the third shell? The fourth shell? The fifth shell?

2.67 How many orbitals would you expect to find in the last subshell of the fifth shell? How many electrons would you need to fill this subshell?

2.68 How many electrons are present in an atom with its $1s$, $2s$, and $2p$ subshells filled? What is this element?

2.69 How many electrons are present in an atom with its $1s$, $2s$, $2p$, $3s$, $3p$, and $4s$ subshells filled and with two electrons in the $3d$ subshell? What is this element?

2.70 Use arrows to show electron pairing in the valence $p$ subshell of:
(a) Sulfur (b) Bromine (c) Silicon

2.71 Use arrows to show electron pairing in the $5s$ and $4d$ orbitals of:
(a) Rubidium (b) Niobium (c) Rhodium

2.72 Determine the number of unpaired electrons for each of the atoms in Problems 2.70 and 2.71.

2.73 Without looking back in the text, write the electron configurations for the following:
(a) Titanium $Z = 22$ (b) Phosphorus, $Z = 15$
(c) Argon, $Z = 18$ (d) Lanthanum, $Z = 57$

2.74 How many electrons does the element with $Z = 12$ have in its valence shell? Write the electron-dot symbol for this element.

2.75 How many valence electrons does the element with $Z = 12$ have? Explain. Write a generic electron-dot symbol for elements in this group.

2.76 Identify the valence subshell occupied by electrons in beryllium and arsenic atoms.

2.77 What group in the periodic table has the valence-shell configuration $ns^2 np^6$?

2.78 Give the number of valence electrons and draw electron-dot symbols for atoms of the following elements:
(a) Kr (b) C (c) Ca (d) K (e) B (f) Cl

2.79 Using $n$ for the number of the valence shell and write a general valence-shell configuration for the elements in group 6A and in group 2A.

**CONCEPTUAL PROBLEMS**

2.80 What elements in addition to helium make up the noble gas family?

2.81 Hydrogen is placed in group 1A on many periodic charts, even though it is not an alkali metal. On other periodic charts, however, hydrogen is included with group 7A even though it is not a halogen. Explain. (Hint: Draw electron-dot symbols for H and for the 1A and 7A elements.)

2.82 What is the atomic number of the yet-undiscovered element directly below francium (Fr) in the periodic table?

2.83 Give the number of electrons in each shell for lead.

2.84 Identify the highest-energy occupied subshell in atoms of the following elements:
(a) Iodine (b) Scandium (c) Arsenic (d) Aluminum

2.85 (a) What is the mass (in amu and in grams) of a single atom of Carbon-12?
(b) What is the mass (in grams) of 6.02 x $10^{23}$ atoms of Carbon-12?

2.86 An unidentified element is found to have an electron configuration by shell of 2 8 18 8 2. To what group and period does this element belong? Is the element a metal or a
nonmetal? How many protons does an atom of the element have? What is the name of the element? Write its electron-dot symbol.

2.87 Germanium, atomic number 32, is used in building semiconductors for microelectronic devices, and has an electron configuration by shell of 2 8 18 4.
(a) Write the electronic configuration for germanium.
(b) In what shell and orbitals are the valence electrons?

2.88 Tin, atomic number 50, is directly beneath germanium (Problem 2.87) in the periodic table. What electron configuration by shell would you expect tin to have? Is tin a metal or a nonmetal?

2.89 A blood sample is found to contain 8.6 mg/dL of Ca. How many atoms of Ca are present in 8.6 mg? The atomic weight of Ca is 40.08 amu.

2.90 What is wrong with the following electron configurations?
(a) Ni 1s² 2s² 2p⁶ 3s² 3p⁶ 3d¹⁰
(b) N 1s² 2p⁵
(c) Si 1s² 2s² 2p⁶
(d) Mg 1s² 2s² 2p⁶ 3s²

2.91 Not all elements follow exactly the electron-filling order described in Figure 2.6. Atoms of which elements are represented by the following electron configurations?
(a) 1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁴ 4s¹
(b) 1s² 2s² 2p⁶ 3s² 3p⁶ 3d¹⁰ 4s¹
(c) 1s² 2s² 2p⁶ 3s² 3p⁶ 3d¹⁰ 4s² 4p⁶ 4d⁵ 5s¹
(d) 1s² 2s² 2p⁶ 3s² 3p⁶ 3d¹⁰ 4s² 4p⁶ 4d¹⁰ 5s¹

2.92 What similarities do you see in the electron configurations for the atoms in Problem 2.91? How might these similarities explain their anomalous electron configurations?

2.93 Based on the identity of the elements whose electron configurations are given in Problem 2.91, write the electron configurations for the element with atomic number Z = 79.

2.94 What orbital is filled last in the most recently discovered element 117?

GROUP PROBLEMS

2.95 Look up one of the experiments by the scientists discussed in the Chemistry in Action on page 44, and explain how it contributed to our understanding of atomic structure.

2.96 Do a web search to identify each of the following elements/isotopes and indicate the number of neutrons, protons, and electrons in an atom of the element/isotope:
(a) A radioactive isotope used in cancer treatments. (There may be more than one answer!)
(b) The element having the greatest density.
(c) An element with Z < 90 that is not found in nature.

2.97 Tellurium (Z = 52) has a lower atomic number than iodine (Z = 53), yet it has a higher atomic weight (127.60 amu for Te vs. 126.90 amu for I). How is this possible? Can you find any other instances in the periodic table where two adjacent elements exhibit a similar behavior, that is, the element with the lower atomic number has a higher atomic mass?

2.98 Look again at the trends illustrated in Figures 2.3 and 2.4.
(a) How do the peaks/valleys correlate with locations in the periodic table?
(b) Are there other chemical properties that also exhibit periodic trends? What are they?
A. The Periodic Table
[Sections 2.4 and 2.5]
B. Electron Configurations
[Sections 2.7 and 2.8]

In previous chapters, we mentioned the importance of various elements for good health, identifying individual elements as macronutrients (needed in large amounts) or micronutrients (needed in lesser amounts). Of equal significance is the chemical form of the element; what is the chemical nature of the compounds in which an element is found? Many of these macro- and micronutrients, for example, exist as ions, or charged particles, and play critical roles in different cells within the body. Calcium ions, for example, are necessary for strong teeth and bones; sodium and potassium ions are necessary for signal transmission in nerve cells, such as those depicted in the artistic...