1.1 Chemical Fundamentals

Living things are composed of matter. Matter has mass, occupies space, and is found in a bewildering variety of forms. According to the Bohr–Rutherford model, the atom is composed of an extremely small nucleus containing positively charged protons \((p^+)\) and neutral neutrons \((n^0)\) surrounded by tiny negatively charged electrons \((e^-)\). The atomic number, \(Z\), is equal to the number of protons in the nucleus and also the number of electrons in a neutral atom. The mass number, \(A\), equals the sum of the protons and neutrons in the nucleus. The elements are arranged in order of increasing atomic number in the periodic table. Carbon is element six on the table. It has six protons and six neutrons in its nucleus and six electrons surrounding the nucleus. Thus, for carbon, \(Z = 6\) and \(A = 12\). It is useful to symbolize individual atoms with a shorthand notation that places the atomic number at the bottom left corner of the element’s chemical symbol and the mass number at the top left corner (Figure 1). An atom of sulfur may be represented as \(^{32}_{16}\text{S}\). This atom may also be symbolized as S-32 or sulfur-32.

Isotopes

The masses of the elements given in the periodic table are not whole number values because, in nature, most elements contain atoms with different numbers of neutrons. We call these atoms isotopes. Isotopes are atoms of an element with the same atomic number but a different mass number. Since the atomic number is the same, the number of protons and electrons are the same. A difference in the number of neutrons in the nucleus distinguishes isotopes from one another.

The element carbon, C, a major constituent of living organisms, consists of three isotopes: \(^{12}\text{C}\), \(^{13}\text{C}\), and \(^{14}\text{C}\). Carbon-12 accounts for about 99% of the carbon atoms in nature. Carbon-13 makes up most of the rest, with carbon-14 present in trace amounts. In a series of investigations led by Ernest Rutherford (1871–1937) and Canadian scientist Harriet Brooks (1876–1933), it was discovered that the nucleus of some isotopes spontaneously breaks apart or decays. Isotopes that can decompose in this way are called radioisotopes and are said to be radioactive. Radioactivity results in the formation of different elements, the release of a number of subatomic particles, and radiation. Carbon-12 and carbon-13 are stable isotopes of carbon, but carbon-14 is radioactive. It spontaneously decays into nitrogen-14. Hydrogen is composed of three isotopes: \(^1\text{H}\) (protium), \(^2\text{H}\) (deuterium), and radioactive \(^3\text{H}\) (tritium). Table 1 summarizes the basic characteristics of the three isotopes of carbon and hydrogen.

Table 1 Three Isotopes of Carbon and Hydrogen

<table>
<thead>
<tr>
<th>Name, symbol</th>
<th>Atomic number ((Z))</th>
<th>Mass number ((A))</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Relative abundance</th>
<th>Structural stability</th>
</tr>
</thead>
<tbody>
<tr>
<td>carbon-12, (^{12}\text{C})</td>
<td>6</td>
<td>12</td>
<td>6</td>
<td>6</td>
<td>98.9%</td>
<td>stable</td>
</tr>
<tr>
<td>carbon-13, (^{13}\text{C})</td>
<td>6</td>
<td>13</td>
<td>6</td>
<td>7</td>
<td>1.1%</td>
<td>stable</td>
</tr>
<tr>
<td>carbon-14, (^{14}\text{C})</td>
<td>6</td>
<td>14</td>
<td>6</td>
<td>8</td>
<td>trace</td>
<td>radioactive</td>
</tr>
<tr>
<td>hydrogen-1, (^1\text{H})</td>
<td>1</td>
<td>1</td>
<td>1</td>
<td>0</td>
<td>99.8%</td>
<td>stable</td>
</tr>
<tr>
<td>hydrogen-2, (^2\text{H})</td>
<td>1</td>
<td>2</td>
<td>1</td>
<td>1</td>
<td>0.2%</td>
<td>stable</td>
</tr>
<tr>
<td>hydrogen-3, (^3\text{H})</td>
<td>1</td>
<td>3</td>
<td>1</td>
<td>2</td>
<td>trace</td>
<td>radioactive</td>
</tr>
</tbody>
</table>
Every radioactive isotope has a characteristic property called its **half-life**. The half-life of a radioisotope is the time it takes for one half of the atoms in a sample to decay. The half-life of different radioisotopes varies considerably, but the rate of decay of a particular isotope is constant. Radioisotopes are both useful and dangerous. Two useful applications of radioisotopes are radiometric dating and radioactive tracers.

Carbon-14 makes its way into plants as they absorb a mixture of radioactive and nonradioactive carbon dioxide from the air or water for photosynthesis. It enters the bodies of other organisms through the food chain. When the organism is alive, the ratio of carbon-12 to carbon-14 is the same as it is in the atmosphere. When an organism dies, it stops absorbing carbon from the environment. The amount of carbon-12 remains constant, but the amount of carbon-14 decreases in a predictable way because it is radioactive. Measuring the ratio of carbon-12 to carbon-14 in a dead or fossilized organism allows scientists to calculate the time that has elapsed since the organism's death. This process is known as carbon-14 dating, one type of radiometric dating.

Cells generally use radioactive atoms in the same way that they use nonradioactive isotopes of the same element. However, since radioisotopes emit radiation as they decay, their location may be readily detected. **Radioactive tracers** are radioisotopes used to follow chemicals through chemical reactions and to trace their path as they move through the cells and bodies of organisms. Radioactive isotopes have found many applications in biological, chemical, and medical research.

Radiolabelled molecules (molecules containing specific radioisotopes) are synthesized to investigate a variety of reaction mechanisms and biochemical processes. Melvin Calvin, a pioneer in photosynthesis research, used carbon-14-labelled molecules to determine the sequence of reactions in photosynthesis (Chapter 3). Radioisotopes are used in the study of many biochemical reactions (covered in Units 1 and 3) and in DNA sequencing procedures (covered in Unit 2). Since most compounds of biological importance contain carbon and hydrogen, carbon-14 and hydrogen-3 (tritium) are commonly used as tracers in biological research.

Besides being research tools, radioisotopes are used in the relatively new field of nuclear medicine for diagnosis and treatment (**Table 2**). For example, the thyroid gland produces hormones that have great influence over growth and metabolism. This gland is located in front of the windpipe and is the only organ of the body that actively absorbs iodine. If a patient's symptoms point to abnormal levels of thyroid hormone output, the physician may administer a small amount of radioactive iodine-131 and then use a photographic device to scan the thyroid gland. The resulting images help to identify the possible causes of the condition. **Figure 2** shows three images obtained using this technique.

---

**Figure 2**
Location of thyroid gland and scans of the thyroid gland from three patients
Although very useful in biological research and medicine, radiation from decaying radioisotopes poses a hazard to living organisms. High-energy radiation penetrates tissues and damages cellular molecules, and regular exposure can lead to radiation sickness, genetic mutations, and, at high enough doses, cell death. Because of this, living organisms must be shielded from sources of high-energy radiation. Researchers and technicians who regularly work with radioactive materials wear radiation-sensitive badges called dosimeters to monitor their daily exposure (Figure 3).

### Table 2  Radioisotopes Used in Nuclear Medicine

<table>
<thead>
<tr>
<th>Radioisotope</th>
<th>Medical uses</th>
<th>Half-life</th>
</tr>
</thead>
<tbody>
<tr>
<td>technetium-99</td>
<td>to view the skeleton and heart muscle but also the brain, thyroid, lungs, liver, spleen, kidney, gall bladder, bone marrow, and salivary glands</td>
<td>6.02 h</td>
</tr>
<tr>
<td>iodine-125</td>
<td>to evaluate the filtration rate of the kidneys and to determine bone density measurements</td>
<td>42 d</td>
</tr>
<tr>
<td>iodine-131</td>
<td>to view and treat thyroid, liver, and kidney diseases and various cancers</td>
<td>8.0 d</td>
</tr>
<tr>
<td>phosphorus-32</td>
<td>to treat polycythemia vera (excess red blood cells)</td>
<td>14.3 d</td>
</tr>
<tr>
<td>strontium-89</td>
<td>to relieve the pain of secondary cancers lodged in the bone</td>
<td>46.7 h</td>
</tr>
<tr>
<td>indium-111</td>
<td>to study the brain, the colon, and sites of infection</td>
<td>2.8 d</td>
</tr>
<tr>
<td>fluorine-18</td>
<td>to image tumours and localized infections</td>
<td>110 min</td>
</tr>
</tbody>
</table>

Although very useful in biological research and medicine, radiation from decaying radioisotopes poses a hazard to living organisms. High-energy radiation penetrates tissues and damages cellular molecules, and regular exposure can lead to radiation sickness, genetic mutations, and, at high enough doses, cell death. Because of this, living organisms must be shielded from sources of high-energy radiation. Researchers and technicians who regularly work with radioactive materials wear radiation-sensitive badges called dosimeters to monitor their daily exposure (Figure 3).

### Practice

#### Understanding Concepts

1. Determine the name of the element, the atomic number, and the number of neutrons in the nucleus of an atom that has 19 protons and a mass number of 39.

2. Write the shorthand notation for the element containing
   (a) 24 protons
   (b) 15 electrons, and 17 neutrons

3. Which subatomic particle of the nucleus is present in the same amount in all isotopes of a particular element?

4. Iodine-131 has a half-life of eight days. What mass of iodine-131 would remain from a 20-g sample after 32 days?

5. Why is carbon-14 useful in radioactive dating, but not as useful in nuclear medicine?

#### Applying Inquiry Skills

6. In humans, dietary calcium helps maintain strong bones and teeth. Calcium is absorbed into the bloodstream by the cells that line the small intestine. Bone cells called osteoblasts absorb calcium from the blood and use it to form the mineralized part of bone called bone matrix. Osteoporosis is a condition in which bones become thin and brittle because of a decrease in the density of bone matrix. Explain how radioisotopes could be used to determine whether intestinal cells or osteoblasts lack the ability to process calcium in a person with osteoporosis.

#### Making Connections

7. Comment on this statement: “The health benefits gained from the use of radioactive pharmaceuticals in this generation threaten the health of generations to come.”
Chemical Bonding

Electrons are not arranged haphazardly about the nucleus of an atom. They occupy positions at various distances from the nucleus, depending on their energy content. These positions are called energy levels because as the electron moves to new positions farther from the nucleus, its potential energy increases. The energy levels are called \( n = 1, n = 2, n = 3 \), and so forth (\( n = 1 \) is closest to the nucleus). Early experiments in atomic structure conducted by Ernest Rutherford (1911), Niels Bohr (1913), and others seemed to indicate that electrons move in the space about the nucleus in circular orbits like planets revolve around the sun.

However, more recently we have learned that electrons are so small that it is impossible to know exactly where they are and how they are moving at any given time. Today, using statistical analyses, scientists can determine locations around the nucleus where electrons are most likely to be found. These volumes of space are called orbitals. An orbital can accommodate no more than two electrons. The energy levels around the nucleus of an atom have different orbitals; however, similarly shaped orbitals may be found at each level. An electron in the first energy level, \( n = 1 \), most likely occupies a spherical orbital called the \( 1s \) orbital, as shown in Figure 4(a) (electron shell diagrams). When two electrons occupy an orbital, they pair up, forming a more stable arrangement than does an electron by itself. Although electrons occupy positions within three-dimensional spaces about the nucleus, it is still convenient to illustrate them as dots on shells surrounding the nucleus (electron-shell diagrams), as illustrated in Figure 4. In the second energy level, \( n = 2 \), electrons may occupy a spherical orbital called the \( 2s \) orbital (larger than the \( 1s \) orbital), or any one of three dumbbell-shaped \( 2p \) orbitals, as depicted in Figure 4(b). Since only two electrons can occupy the same orbital, energy level \( n = 1 \) may contain up to two electrons (one pair), and energy level \( n = 2 \) may contain up to eight electrons (four pairs). Higher energy levels contain more electrons with higher energy values, and the orbitals are called \( 3s, 3p, 3d \), and so on. The arrangement of electrons in the orbitals is called the atom’s electron configuration. Table 3 (page 12) lists the number of electrons that may occupy orbitals in energy levels 1 to 3.

### Table 3

<table>
<thead>
<tr>
<th>electron orbitals</th>
<th>(a) 1s orbital</th>
<th>(b) 2s and 2p orbitals</th>
<th>(c) Neon ((^{10})Ne): 1s, 2s, and 2p</th>
</tr>
</thead>
<tbody>
<tr>
<td>electron orbitals</td>
<td>1s orbital ((2e^-))</td>
<td>2s orbital ((2e^-))</td>
<td>2p orbitals (2p_x, 2p_y, 2p_z)</td>
</tr>
<tr>
<td>electron-shell diagrams</td>
<td><img src="image.png" alt="1s orbital" /></td>
<td><img src="image.png" alt="2s orbital" /></td>
<td><img src="image.png" alt="2p orbitals" /></td>
</tr>
</tbody>
</table>

**Figure 4**

Each orbital can hold up to two electrons.

(a) The 1s orbital is in the energy level closest to the nucleus, \( n = 1 \).

(b) The 2s orbital is in energy level \( n = 2 \) and is spherical like the 1s orbital. Energy level \( n = 2 \) also contains three dumbbell-shaped \( 2p \) orbitals oriented at right angles to one another along the \( x \)-axis \((2p_x)\), along the \( y \)-axis \((2p_y)\), and along the \( z \)-axis \((2p_z)\).

(c) The noble gas neon has 10 electrons, two in a 1s orbital, two in a 2s orbital, and one pair in each of the three 2p orbitals. It is an unreactive element because it has full outer \( s \) and \( p \) orbitals.
The outermost \( s \) and \( p \) orbitals are called valence orbitals, and the electrons in these orbitals are called \textbf{valence electrons}. The chemical behaviour of an atom is determined by the number and arrangement of its valence electrons. A special stability is associated with having full outermost \( s \) and \( p \) orbitals. Helium (Z = 2) contains two valence electrons in a 1\( s \) orbital. Since there are no \( p \) orbitals in energy level \( n = 1 \), the 1\( s \) orbital is the only valence orbital. Helium is inert (nonreactive) because the 1\( s \) orbital is full with a pair of electrons. Neon (Z = 10), shown in Figure 4(c) on page 11, is inert because its outermost \( s \) and \( p \) orbitals, the 2\( s \) and 2\( p \) orbitals, are full with four pairs of electrons (one pair in the 2\( s \) orbital and one pair in each of the three 2\( p \) orbitals). These elements are called noble gases because they do not attempt to gain, lose, or share electrons with other atoms; they are stable and do not normally participate in chemical reactions. All other elements attempt to gain, lose, or share valence electrons to achieve the same stable electron configuration as a noble gas. These interactions are the root cause of chemical reactions and are responsible for the formation of chemical bonds between atoms.

It is convenient to draw diagrams of the elements showing only the valence electrons. Such diagrams are called Lewis dot diagrams (Figures 5 and 6).

The periodic table is an arrangement of the elements in order of increasing atomic number. Vertical columns on the table are called groups or families; horizontal rows are called periods. Elements in the same group contain the same number of valence electrons. Thus, group 1 elements all contain one valence electron, group 2 elements contain two valence electrons, and group 17 elements contain seven valence electrons. The noble gases constitute group 18 and contain eight valence electrons (a stable octet).

---

**Table 3** Number of Electrons in Various Energy Levels

<table>
<thead>
<tr>
<th>Energy level (( n ))</th>
<th>Orbital types</th>
<th>Maximum number of electrons allowed</th>
</tr>
</thead>
<tbody>
<tr>
<td>( 1 )</td>
<td>1( s )</td>
<td>2</td>
</tr>
<tr>
<td>( 2 )</td>
<td>2( s ), 2( p )</td>
<td>8</td>
</tr>
<tr>
<td>( 3 )</td>
<td>3( s ), 3( p ), 3( d )</td>
<td>18</td>
</tr>
</tbody>
</table>

---

**Learning Tip**

**Lewis Dot Diagrams**

When drawing Lewis dot diagrams, first place electrons one at a time at the 12, 3, 6, and 9 o’clock positions, as required. Then, add more electrons by pairing them one at a time in the same order as you placed the first four (Figure 6).

---

**Figure 5**

Lewis dot diagrams of the first 20 elements

**Figure 6**

The progressive placement of electrons in Lewis dot diagrams
Compounds are stable combinations of atoms of different elements held together by chemical bonds. Two of the most common bonds are the ionic bond and the covalent bond. When atoms lose electrons, they become positively charged ions called cations. When atoms gain electrons, they become negatively charged ions called anions. An ionic bond is a force of attraction between cations and anions. An ionic compound, such as sodium chloride, NaCl(s), is also called a salt or an ionic solid. A covalent bond is formed when two atoms share one or more pairs of valence electrons. In a hydrogen molecule, two hydrogen atoms share a pair of electrons (one from each atom) and, therefore, possess a single covalent bond. Some molecules contain double and triple covalent bonds. For example, oxygen molecules, O2(g), are composed of two oxygen atoms held together with a double covalent bond (two shared pairs of electrons). Nitrogen molecules, N2(g), are composed of two nitrogen atoms with a triple covalent bond between them (three shared pairs of electrons).

Covalent bonds are stronger than ionic bonds. Ionic and covalent bonds are generally called **intramolecular forces of attraction** (Because they hold the atoms of a molecule or ions of an ionic solid together). Compounds may be described by using chemical formulas and Lewis diagrams (Table 4).

**Table 4**  Ionic and Covalent Substances

<table>
<thead>
<tr>
<th>Bond type</th>
<th>Constituent entities</th>
<th>Force of attraction</th>
<th>Chemical formula</th>
<th>Lewis diagram(^1)</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>ionic</td>
<td>metal cations(^2) and nonmetal anions(^3)</td>
<td>electrostatic attraction between oppositely charged ions</td>
<td>NaCl(_{(s)})</td>
<td>Na(^+[Cl^-])</td>
<td>sodium chloride</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>MgF(_2{(s)})</td>
<td>[F(^-)][Mg(^2+)][F(^-)]</td>
<td>magnesium fluoride</td>
</tr>
<tr>
<td>covalent</td>
<td>neutral atoms</td>
<td>electrostatic attraction between nuclei and valence electrons of neutral atoms</td>
<td>H(_2\text{O}(l))</td>
<td>(\overset{-}{\text{O}}) (-\text{H} ) (\text{H}) (\text{H})</td>
<td>water</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>CO(_2(g))</td>
<td>(\text{O} \equiv \text{C} \equiv \text{O})</td>
<td>carbon dioxide</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>C(<em>6\text{H}</em>{12}\text{O}_6(s))</td>
<td>(\overset{-}{\text{O}}) (-\text{C} \equiv \text{O} ) (\overset{-}{\text{C}}) (-\text{O} \equiv \text{C} \equiv \text{O}) (\overset{-}{\text{O}}) (-\text{H}) (\overset{-}{\text{H}}) (\overset{-}{\text{O}})</td>
<td>glucose</td>
</tr>
</tbody>
</table>

\(^1\) For covalent molecules, each horizontal or vertical line represents a pair of electrons.

\(^2\) From groups 1, 2, or 3 of the periodic table

\(^3\) From groups 16 or 17 of the periodic table
Electronegativity and the Polarity of Covalent Bonds

All the atoms in the periodic table have been assigned an electronegativity number (E_n) on the basis of experimental results. The larger the electronegativity number, the stronger the atom attracts the electrons of a covalent bond. Figure 7 shows a periodic table that indicates electronegativity numbers. Note that the noble gases are not assigned an E_n because they do not participate in chemical bonding (E_n = 0 for noble gases). When an electron pair is unequally shared, the atom that attracts the pair more strongly takes on a partial negative charge (δ−) and the atom that attracts more weakly takes on a partial positive charge (δ+). The lowercase Greek letter delta (δ) denotes a partial charge. The charge is partial because the electron pair spends time around both atoms, but spends more time around one atom than it does around the other. This difference in electron attraction forms a polar covalent bond.

The electronegativity difference (ΔE_n) is the difference in electronegativity number between two atoms participating in a covalent bond. This difference will be zero only if the electronegativity number of two atoms is the same. In this case, the electron pair will be shared equally, and a nonpolar covalent bond is formed. If the electronegativity difference is greater than zero but less than 1.7, the bond is polar covalent. The polarity of the bond increases as the ΔE_n approaches 1.7. When the ΔE_n is equal to or greater than 1.7, the bond is considered ionic (Figure 8). Thus, the ionic bond may be considered an extreme form of polar covalency. You will notice that the electronegativity differences between atoms in groups 1 or 2, and atoms in groups 16 or 17, are generally greater than or equal to 1.7. Atoms from these two sets of groups generally form ionic compounds.
**Molecular Shape**

A molecule’s biological function is determined by the types of bonds between its atoms, and by its overall shape and polarity. The types of atoms determine the types of bonds, and the orientation of bonding electron pairs determines molecular shape. When atoms react to form covalent bonds, their valence electron orbitals undergo a process called **hybridization** that changes the orientation of the valence electrons (Figure 9).

Hybridization is a complex process. The Valence Shell Electron Pair Repulsion (VSEPR) theory, developed by Canadian chemist Ronald J. Gillespie (1939– ), is a useful and relatively simple way of predicting molecular shape. This theory states that, since electrons are all negatively charged, valence electron pairs repel one another and will move as far apart as possible. The majority of molecules of biological importance possess up to four valence electron pairs around a central atom (an exception is phosphorous, which has five pairs). Table 5 illustrates the resulting shape of molecules containing one to four valence electron pairs. Note that nonbonding pairs occupy more space than do bonding pairs and will repel and compress the bond angles of the bonding pairs. Also note that the orbital model and the VSEPR model predict similar molecular shapes.

**Table 5** Molecular Shapes

<table>
<thead>
<tr>
<th>Name and formula (shape)</th>
<th>Orbital model</th>
<th>VSEPR model</th>
<th>Bonding valence electron pairs</th>
<th>Nonbonding valence electron pairs</th>
<th>Ball-and-stick diagram</th>
<th>Space-filling diagram</th>
</tr>
</thead>
<tbody>
<tr>
<td>methane, CH₄ (tetrahedral)</td>
<td><img src="image1.png" alt="Orbital model" /></td>
<td><img src="image2.png" alt="VSEPR model" /></td>
<td>4</td>
<td>0</td>
<td><img src="image3.png" alt="Ball-and-stick diagram" /></td>
<td><img src="image4.png" alt="Space-filling diagram" /></td>
</tr>
<tr>
<td>ammonia, NH₃ (pyramidal)</td>
<td><img src="image5.png" alt="Orbital model" /></td>
<td><img src="image6.png" alt="VSEPR model" /></td>
<td>3</td>
<td>1</td>
<td><img src="image7.png" alt="Ball-and-stick diagram" /></td>
<td><img src="image8.png" alt="Space-filling diagram" /></td>
</tr>
<tr>
<td>water, H₂O (angular)</td>
<td><img src="image9.png" alt="Orbital model" /></td>
<td><img src="image10.png" alt="VSEPR model" /></td>
<td>2</td>
<td>2</td>
<td><img src="image11.png" alt="Ball-and-stick diagram" /></td>
<td><img src="image12.png" alt="Space-filling diagram" /></td>
</tr>
<tr>
<td>hydrogen chloride, HCl (linear)</td>
<td><img src="image13.png" alt="Orbital model" /></td>
<td><img src="image14.png" alt="VSEPR model" /></td>
<td>1</td>
<td>3</td>
<td><img src="image15.png" alt="Ball-and-stick diagram" /></td>
<td><img src="image16.png" alt="Space-filling diagram" /></td>
</tr>
</tbody>
</table>

**Figure 9**

Hybridization of one s orbital and three p orbitals forms a new orbital system (sp³ hybrid) with four orbitals that point toward the vertices of a regular tetrahedron whose internal angles are all 109.5°. As usual, only two valence electrons can occupy each orbital.
Molecular Polarity

Covalent bonds may be polar or nonpolar. However, the polarity of a molecule as a whole is dependent on bond polarity and molecular shape. **Figure 10(a)** shows that symmetrical molecular structures produce nonpolar molecules (whether the bonds are polar or not). Asymmetrical molecular shapes produce nonpolar molecules if the bonds are all nonpolar, and they produce polar molecules if at least one bond is polar, as **Figure 10(b)** illustrates.

(a) Tetrahedral shape  
- symmetrical arrangement of polar covalent bonds  
- nonpolar molecule

(b) Asymmetrical arrangement of nonpolar bonds  
- nonpolar molecule

**carbon tetrachloride, CCl₄(g)**

**ozone, O₃(g)**

- asymmetrical arrangement of nonpolar bonds  
- nonpolar molecule

**ammonia, NH₃(g)**

- asymmetrical arrangement of polar bonds  
- polar molecules

**methanol, CH₃OH(l)**

### Practice

**Understanding Concepts**

8. Hydrogen sulfide, H₂S(g), also known as “sewer gas” is released when bacteria break down sulfur-containing waste.  
(a) Draw a Lewis diagram for hydrogen sulfide.  
(b) Determine whether the S–H bonds in hydrogen sulfide are nonpolar covalent, polar covalent, or ionic.  
(c) Use the VSEPR theory to determine the shape of the hydrogen sulfide molecule.  
(d) Determine whether the hydrogen sulfide molecule is polar or nonpolar. Explain.

9. How can a molecule with polar covalent bonds be nonpolar?

### Water

Life as we know it could not exist without water. Approximately two thirds of Earth’s surface is covered in water and about two thirds of any organism is composed of water. Bacterial cells are about 70% water by mass. Herbaceous plants, such as celery, are about 90% water and woody plants, such as trees, are about 50% water. Water keeps plant cells rigid and is required for photosynthesis. In humans, lungs are about 90% water, and water makes up about 70% of the brain. Fat tissue is about 25% water, and even bones contain more than 20% water by mass. Water helps control body temperature, keeps eyeballs moist, and lubricates joints. It also acts as a shock absorber, protecting your brain and spinal cord from bruising when your head and back are jarred. Water is so common that its extraordinary properties are easily overlooked.

Water is a colourless, tasteless, odourless substance that can exist as a solid, liquid, or gas in the temperature ranges normally found near Earth’s surface. Its physical and chemical properties are a direct result of its simple composition and structure. Water, H₂O, comprises two hydrogen atoms attached to one oxygen atom. Water’s polar covalent bonds and asymmetrical structure create a highly polar molecule (**Figure 11**).
The polarity of water allows it to form chemical bonds with other molecules and ions, including other water molecules. Bonds between molecules are called intermolecular bonds. These are the bonds that determine the physical state of molecular substances at a given temperature and pressure. They are broken when solids melt into liquids and liquids evaporate into gases. Intermolecular bonds are weaker than the intramolecular covalent and ionic bonds that hold atoms and ions together in molecules and ionic solids. There are three types of intermolecular bonds: London forces (London dispersion forces), dipole–dipole forces, and hydrogen bonds (H-bonds). London dispersion forces are the weakest, and exist between all atoms and molecules. London forces constitute the only intermolecular forces of attraction between noble gas atoms and between nonpolar molecules. These bonds are formed by the temporary unequal distribution of electrons as they randomly move about the nuclei of atoms. The unequal electron clouds allow the electrons of one neutral atom to attract the nucleus of neighbouring atoms. These bonds are very weak between single atoms, such as helium, and small nonpolar molecules, such as methane, CH₄(g) (in natural gas) Figure 12(a). That is why these materials are gases at room temperature. London forces become more significant between large nonpolar molecules, such as octane C₈H₁₈(l) (in gasoline), a liquid at room temperature, because of the cumulative effect of many London forces between their atoms.

Dipole–dipole forces, illustrated in Figure 12(b), hold polar molecules to one another. In this case, the partially positive side of one polar molecule attracts the partially negative side of adjacent polar molecules. These intermolecular forces of attraction are stronger than London forces.

![Figure 12](image)

**Figure 12**

(a) London forces are weak forces of attraction between all atoms and molecules. They are the only intermolecular forces that hold nonpolar molecules to one another. Small nonpolar molecules, such as methane, are gases at room temperature because of the relative weakness of the London forces between the molecules. Larger nonpolar molecules, such as octane, are liquids at room temperature because of the cumulative effect of many London forces between their atoms.

(b) Dipole–dipole forces hold polar molecules together. These intermolecular forces of attraction are stronger than London forces.

Hydrogen bonds are the strongest intermolecular forces of attraction. They are, in fact, especially strong dipole–dipole forces that only form between an electropositive H of one polar molecule and an electronegative N, O, or F of a neighbouring polar molecule (Figure 13, page 18). Water molecules hold onto each other by H-bonds. Hydrogen bonds and an angular shape give water its many unique properties (Figure 14, page 18).

London dispersion forces, dipole–dipole attractions, and hydrogen bonds are collectively called van der Waals forces.
Water: The Versatile Solvent

More substances dissolve in water than in any other liquid. Practically all chemicals of biological interest occur in a water medium. The reason for water’s excellent dissolving ability relates to its polarity—water molecules provide partial positive and negative charges to which other polar molecules or ions can attach. Figure 15 illustrates the dissolving of sodium chloride in water. When ionic solids dissolve, the anions and cations dissociate from one another. Dissociation involves the breaking of ionic bonds.

Materials do not have to be ionic to dissolve in water. Polar molecules generally dissolve readily in water. Polar covalent substances, such as sugars and alcohols, dissolve easily in water. Any substance that dissolves in another substance is referred to as soluble. Water has been called the universal solvent because so many different substances dissolve in it. However, some things are easier to dissolve in water than others, and some materials do not dissolve in water to any appreciable degree. Substances that dissolve very little, such as iron and chalk, are called insoluble. The terms **miscible** and **immiscible** are used to describe whether liquids dissolve in each other or not. Ethanol (the alcohol in alcoholic beverages) and ethylene glycol (the alcohol in antifreeze) are miscible with water. Gasoline and oil are immiscible with water but are miscible with each other. The subscript \(\text{(aq)}\), meaning “aqueous,” is used to identify a molecule or ion that is dissolved in water. The formula \(\text{C}_6\text{H}_{12}\text{O}_6\text{(aq)}\) symbolizes an aqueous solution of glucose.

Small nonpolar molecules, such as oxygen \(\text{(O}_2\) and carbon dioxide \(\text{(CO}_2\), cannot form hydrogen bonds with water and, thus, are only slightly soluble. That is why a soluble protein carrier molecule, such as hemoglobin, is needed to transport oxygen, and to a lesser degree, carbon dioxide, within the circulatory systems of animals. Large nonpolar molecules, such as fats and oils, also do not form hydrogen bonds with water. When placed in water, they are excluded from associating with water molecules because the
water molecules form hydrogen bonds with one another. For this reason, nonpolar molecules are said to be **hydrophobic** (meaning “water fearing”). Conversely, polar molecules that can form hydrogen bonds with water are described as **hydrophilic** (“water loving”). Oil floats on water because it is less dense than water, but it fails to dissolve in water because it is nonpolar and hydrophobic. Although unable to form solutions with water, nonpolar substances do dissolve in other nonpolar materials. Oil and gasoline are nonpolar substances that do not dissolve in water but dissolve well into each other. “Like dissolves like” is a generalization that describes the fact that polar substances dissolve in other polar substances and nonpolar materials dissolve in other nonpolar materials.

**The Unique Properties of Water**

Its angular shape and hydrogen-bonding characteristics give water a number of extraordinary properties. **Table 6** summarizes some of water’s unique characteristics.

**Table 6** The Unique Properties of Water

<table>
<thead>
<tr>
<th>Descriptive characteristic</th>
<th>Property</th>
<th>Explanation</th>
<th>Effect</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water clings.</td>
<td>cohesion</td>
<td>Water molecules form hydrogen bonds with one another.</td>
<td>high surface tension</td>
<td>A water strider (Gerris species) walks on the surface of a pond.</td>
</tr>
<tr>
<td></td>
<td>adhesion</td>
<td>Water molecules form hydrogen bonds with other polar materials.</td>
<td>capillary action</td>
<td>Capillary action causes water to creep up a narrow glass tube and paper.</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Water absorbs lots of heat.</td>
<td>high specific heat capacity¹</td>
<td>Hydrogen bonding causes water to absorb a large amount of heat before its temperature increases appreciably and also causes it to lose large amounts of heat before its temperature decreases significantly.</td>
<td>temperature moderation</td>
<td>High heat capacity helps organisms maintain a constant body temperature.</td>
</tr>
<tr>
<td></td>
<td>high specific heat of vaporization²</td>
<td>Hydrogen bonding causes liquid water to absorb a large amount of heat to become a vapour (gas).</td>
<td>evaporative cooling</td>
<td>Many organisms, including humans, dissipate body heat by evaporation of water from surfaces, such as skin (by sweating) and tongue (by panting).</td>
</tr>
<tr>
<td>Solid water is less dense than liquid water.</td>
<td>highest density at 4°C</td>
<td>As water molecules cool below 0°C, they form a crystalline lattice (freezing). The hydrogen bonds between the V-shaped molecules spread the molecules apart, reducing the density below that of liquid water.</td>
<td>Ice floats on liquid water.</td>
<td>Fish and other aquatic organisms are able to survive in winter.</td>
</tr>
</tbody>
</table>

¹ Heat energy, like all other forms of energy, is measured in joules (J). Specific heat capacity is a measure of the extent to which a substance resists changes in temperature when it absorbs or releases heat. The specific heat capacity of water is 4.18 J/(g°C), approximately twice that of most organic compounds.

² The amount of heat energy required to convert 1 g of liquid water into a vapour is 2.4 kJ, nearly twice as much heat energy as is needed to vaporize 1 g of ethanol.
Acids, Bases, and Buffers

Pure water never contains only H₂O molecules. At 25°C, two H₂O molecules in every 550 million react with each other. One H₂O molecule transfers an H⁺ ion to the other H₂O molecule. This produces an OH⁻ (hydroxide) ion and an H₃O⁺ (hydronium) ion. The process is called the autoionization of water and is summarized in Figure 16.

The hydronium ion has been characterized as a substance that gives rise to the properties of acidic solutions—sour taste, the ability to conduct electricity, and the ability to turn blue litmus red. Acids are defined as substances that increase the concentration of H₃O⁺(aq) when dissolved in water and that contain at least one ionizable hydrogen atom in their chemical structure. The following equation describes the reaction of hydrogen chloride with water to produce hydrochloric acid:

\[ \text{HCl}(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq) \]

The hydroxide ion has the properties of a base—bitter taste, slippery feel, and the ability to conduct electricity and to change the colour of red litmus to blue. Bases are substances that increase the concentration of OH⁻(aq) ions in solution. This is accomplished in one of two ways. Ionic bases containing OH⁻ ions, such as sodium hydroxide (NaOH), dissociate in water to produce OH⁻ ions directly.

\[ \text{NaOH}(s) \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq) \]

Some bases combine with H⁺ ions directly. Ammonia (NH₃), a product of decomposed plant and animal matter, is such a base. Instead of releasing OH⁻ ions into solution directly, ammonia combines with an H⁺(aq) ion, from H₂O(l), to provide an ammonium ion, NH₄⁺(aq), and a hydroxide ion, OH⁻(aq).

\[ \text{NH}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{NH}_4^+(aq) + \text{OH}^-(aq) \]

Since pure water contains equal numbers of hydronium and hydroxide ions, it is neutral. When equal amounts of hydronium and hydroxide ions react, water is formed.

\[ \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \rightarrow 2\text{H}_2\text{O}(l) \text{ or } \text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l) \]

Since water is neutral, this reaction is called a neutralization reaction, and occurs whenever an acid is mixed with a base. In these reactions, water and a salt are produced.

\[ \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq) \]

The acidity of an aqueous solution may be expressed in terms of hydronium ion concentration, symbolized as [H₃O⁺(aq)]. The concentration of a solute in aqueous solution is measured in moles of the solute per litre of solution and is symbolized as mol/L, the mole being the amount of any substance that contains 6.02 × 10²³ entities of that substance. A hydronium ion concentration of 1.0 mol/L means that the solution contains 1.0 mol of H₃O⁺(aq) (6.02 × 10²³ H₃O⁺ ions) per litre of solution. In pure water at 25°C, the [H₃O⁺(aq)] is 1.0 × 10⁻⁷ mol/L. It is common in acid–base chemistry for concentrations to be very small values. The pH scale was devised as a more convenient way to express the concentration of H₃O⁺(aq). The pH of an aqueous solution is equal to the negative logarithm of the hydronium ion concentration, \(-\log_{10}[\text{H}_3\text{O}^+(aq)]\). Thus, the pH of pure water is 7 since pH = \(-\log_{10}(10^{-7} \text{ mol/L}) = 7\). Any aqueous solution with pH = 7 will be neutral, indicating a balance of hydronium and hydroxide ions. Solutions whose pH is less than 7 are acidic and solutions with pH greater than 7 are basic (Figure 17).
Strong and Weak Acids and Bases
Acids and bases may be classified as strong or weak, according to the degree to which they ionize when dissolved in water. **Strong acids**, such as $\text{HCl(g)}$, and **strong bases**, such as $\text{NaOH(s)}$, ionize completely when dissolved in water.

\[
\text{HCl(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{Cl}^-_{(aq)}
\]

\[
\text{NaOH(aq)} \rightleftharpoons \text{Na}^+_{(aq)} + \text{OH}^-_{(aq)}
\]

**Weak acids**, such as acetic acid, $\text{CH}_3\text{COOH(aq)}$, and **weak bases**, such as $\text{NH}_3(aq)$, ionize partially in water.

\[
\text{CH}_3\text{COOH(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{CH}_3\text{COO}^-_{(aq)}
\]

\[
\text{NH}_3(aq) + \text{H}_2\text{O(l)} \rightleftharpoons \text{NH}_4^+_{(aq)} + \text{OH}^-_{(aq)}
\]

The equations representing the reactions of weak acids and bases contain double arrows to indicate that these reactions may proceed in both directions—they are reversible reactions. When first placed in water, acetic acid molecules react with water molecules to form hydronium and acetate ions; the forward reaction is favoured. As the concentration of the ions increases, the reverse reaction occurs more frequently. When approximately 1.3% of the acetic acid molecules have ionized, the rates of the forward and reverse reactions become equal and the solution is said to be in a state of equilibrium. Once equilibrium is reached, the forward and backward reactions continue to occur at equal rates, and the concentrations of all entities in the solution remain constant. Most organic acids and bases are weak and reach equilibrium.

Conjugate Acids and Bases
According to the Brønsted–Lowry concept, reversible acid–base reactions involve the transfer of a proton. An acid is a proton donor; a base is a proton acceptor. In an acetic acid solution, the forward reaction involves a proton transfer from acetic acid to water. Acetic acid, $\text{CH}_3\text{COOH}$, acts as the acid; water, $\text{H}_2\text{O}$, acts as the base. In the reverse reaction, the hydronium ion, $\text{H}_3\text{O}^+$, acts as the acid, and the acetate ion, $\text{CH}_3\text{COO}^-$, acts as the base.

Acetic acid, acetate and hydronium ions, and water, are conjugate acid–base pairs.
Acid–Base Buffers

The components of living cells and the internal environments of multicellular organisms are sensitive to pH levels. Most cellular processes operate best at pH 7. Chemical reactions within cells normally produce acids and bases that have the potential to seriously disrupt function. Many of the foods we eat are acidic. Absorbing these acids could affect the pH balance of blood. Living cells use buffers to resist significant changes in pH. In living organisms, buffers usually consist of conjugate acid–base pairs in equilibrium. The most important buffer in human extracellular fluid and blood is the carbonic acid–bicarbonate (base) buffer (Figure 18). The components of this buffer are produced when carbon dioxide reacts with water molecules in body fluids. In the formation of the carbonic acid–bicarbonate buffer, carbon dioxide and water react to form carbonic acid, \( H_2CO_3(aq) \). Carbonic acid then ionizes to form bicarbonate ions \( HCO_3^- (aq) \) and \( H^+ (aq) \) ions.

Buffers in Living Systems (p. 78)
The vast majority of foods we eat are acidic. Fruit, vegetables, salad dressings, wine, and pop all add \( H_3O^+ (aq) \) ions to your blood. People also enjoy alkaline foods and beverages like shrimp, corn, and tonic water. Most cellular proteins, especially enzymes, work best at near-neutral pH, their activity dropping significantly when pH levels deviate from the ideal. How well does a buffer work to maintain a constant pH in cells? In Investigation 1.1.1, you will investigate the properties of acids, bases, and buffers.

INVESTIGATION 1.1.1

DID YOU KNOW??

Hyperventilation Improves Running Performance
Sprinters take advantage of the carbonic acid–bicarbonate buffer system in blood by hyperventilating shortly before a race. This increases the pH of the blood slightly, which allows better control of the short-term buildup of lactic acid during the sprint.

When \( H^+ (aq) \) ions enter the bloodstream, after someone eats a vinaigrette salad, for example, \( HCO_3^- (aq) \) ions react with the \( H^+ (aq) \) ions to produce \( H_2CO_3(aq) \). Similarly, if a base enters the blood and removes \( H^+ (aq) \) ions, \( H_2CO_3(aq) \) ionizes to replace missing \( H^+ (aq) \) ions in the blood. Together, the carbonic acid molecules (the conjugate acid) and bicarbonate ions (the conjugate base) help maintain the pH of blood at about 7.4. An increase or decrease in blood pH of 0.2 to 0.4 pH units is fatal if not treated immediately. A pH below 7.35 is described as acidosis and a pH above 7.45 is described as alkalosis.

Proteins can also act as buffers. Hemoglobin helps maintain a constant pH within red blood cells. Proteins are able to buffer solutions because some of the amino acids in their structure may be acidic and others may be basic (discussed further in Section 1.2). Together, the amino acids can remove excess \( H_3O^+ (aq) \) or \( OH^- (aq) \) from a solution.

INVESTIGATION 1.1.1

Buffers in Living Systems (p. 78)
The vast majority of foods we eat are acidic. Fruit, vegetables, salad dressings, wine, and pop all add \( H_3O^+ (aq) \) ions to your blood. People also enjoy alkaline foods and beverages like shrimp, corn, and tonic water. Most cellular proteins, especially enzymes, work best at near-neutral pH, their activity dropping significantly when pH levels deviate from the ideal. How well does a buffer work to maintain a constant pH in cells? In Investigation 1.1.1, you will investigate the properties of acids, bases, and buffers.

SUMMARY

Chemical Fundamentals

- The atomic number of an element is equal to the number of protons in the nucleus. The mass number of an atom is determined by the sum of the protons and neutrons in the nucleus.
- Isotopes are atoms of an element with the same atomic number but a different mass number. They differ from one another by the number of neutrons in their nuclei. Radioisotopes spontaneously break apart.
- Electronegativity is a measure of the ability of an atom to attract the pair of electrons in a covalent bond.
- VSEPR theory is a method for predicting molecular shape based on the mutual repulsion of electron pairs in a molecule (Figure 19).
- When \( \Delta E_n \geq 1.7 \), the bond is ionic. When \( 0 < \Delta E_n < 1.7 \), the bond is covalent.
Figure 19

- Hydrogen bonding accounts for the following properties of water: high solubility of polar and ionic substances, cohesion, adhesion, high surface tension, capillarity, high specific heat capacity, and high specific heat of vaporization.
- Acids are substances that increase the concentration of $\text{H}_3\text{O}^+\text{(aq)}$ when dissolved in water. Bases are substances that increase the concentration of $\text{OH}^-\text{ions}$ in solution.
- A buffer is a chemical system containing a substance that can donate $\text{H}^+\text{ions}$ when they are required and containing a substance that can remove $\text{H}^+\text{ions}$ when there are too many in a solution.

Section 1.1 Questions

Understanding Concepts

1. List the number of protons, neutrons, electrons, and valence electrons in the atoms of each of the following elements:
   (a) sulfur  (b) calcium  (c) nitrogen
2. Why do the elements with atomic numbers 8, 16, and 34 belong to the same family on the periodic table?
3. How many neutrons are found in an atom of cobalt-60?
4. What type of intermolecular forces of attraction must be overcome to melt each of the following solids?
   (a) ice, $\text{H}_2\text{O(s)}$  (b) iodine, $\text{I}_2(\text{s})$
5. State the principle on which the VSEPR theory is based.
6. Use electronegativity values and VSEPR theory to determine whether carbon tetrachloride, $\text{CCl}_4$, and ammonia, $\text{NH}_3$, are polar or nonpolar molecules.
7. Why is table salt, $\text{NaCl(s)}$, soluble in ethanol, $\text{CH}_3\text{OH(l)}$, but not soluble in gasoline, $\text{C}_8\text{H}_{18(l)}$?
8. What is the difference between a weak acid and a dilute solution of a strong acid?
9. Identify the two conjugate acid–base pairs in the following acid–base equilibrium.
   $\text{HCOOH(aq)} + \text{CN}^-(\text{aq}) \rightleftharpoons \text{HCOO}^-(\text{aq}) + \text{HCN(aq)}$
10. Describe the components of a buffer and the role each plays in helping maintain a constant pH.
11. Determine the pH of each of the following solutions, and state whether each is acidic, basic, or neutral:
   (a) tonic water: $[\text{H}_3\text{O}^+\text{(aq)}] = 3.1 \times 10^{-9}\text{ mol/L}$
   (b) wine: $[\text{H}_3\text{O}^+\text{(aq)}] = 2.5 \times 10^{-3}\text{ mol/L}$
12. Distinguish between ionic and polar covalent bonds.
13. Distinguish between table sugar dissolving in water and table salt dissolving in water.
14. Identify each of the following substances as either hydrophilic or hydrophobic:
   (a) $\text{C}_8\text{H}_{18(l)}$  (b) $\text{C}_2\text{H}_5\text{OH(l)}$  (c) $\text{CCl}_4(l)$
15. What property of water accounts for each of the following observations?
   (a) A steel sewing needle floats on water but a large steel nail sinks.
   (b) Dogs pant on a hot summer day.
   (c) Water creeps up the walls in a flooded room.
   (d) Hands are usually washed in water.

Applying Inquiry Skills

16. Water and varsol (a paint thinner) are immiscible liquids. Describe an experiment using these two materials that would determine whether polystyrene foam is polar or nonpolar.

Making Connections

17. (a) Describe three technological uses of radioisotopes.
   (b) Provide three possible reasons for the high costs generally associated with these uses of radioisotopes.