UNITED ARAB EMIRATES
MINISTRY OF EDUCATION

عام المجتمع
YEAR OF COMMUNITY

نحن الإمارات
WE THE UAE 2031

2025-2026

Inspire Chemistry

UAE Edition
Grade 11-12 General
Student Edition
Volume 1

McGraw Hill

McGraw-Hill Education

Chemistry

United Arab Emirates Edition

Grade 11-12 General

Volume 1

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Chemistry
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Source ISBN-13 9780021381159
Source ISBN-10 0021381151

Published by McGraw-Hill Education (UK) Limited
Unit 4, Foundation Park,
Roxborough Way,
Maidenhead SL6 3UD
T: +44 (0) 1628 502500
Website: www.mheducation.co.uk

Published by arrangement with McGraw Hill LLC

ISBN 978-1-39-896261-3
ISBN 1-39-896261-9

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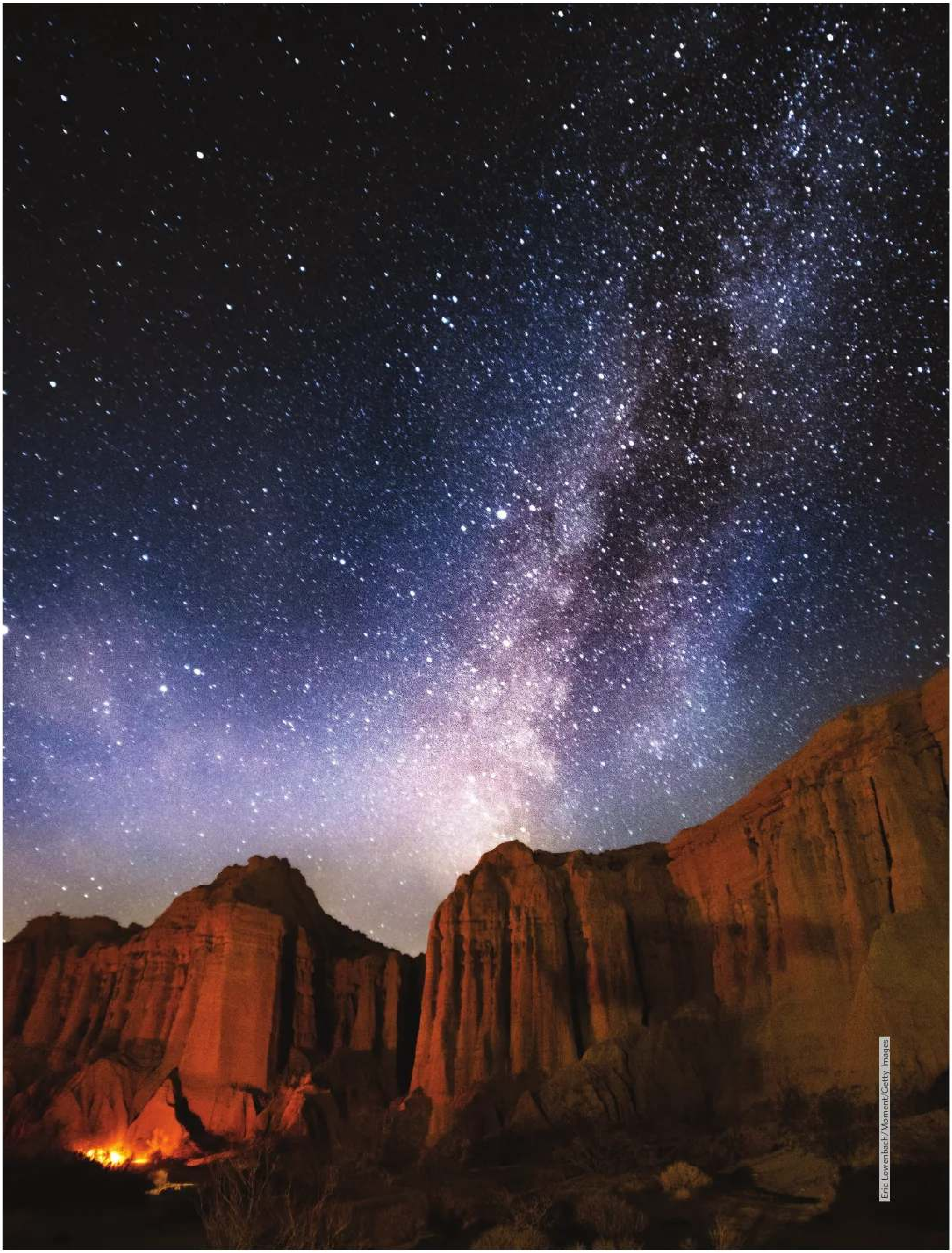

Eric Lowenbach/Moment/Getty Images

MODULE 4

ELECTRONS IN ATOMS

ENCOUNTER THE PHENOMENON

How do we know what stars are made of?

GO ONLINE to play a video about how scientists analyze the light from stars.

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about how we know what stars are made of.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.

LESSON 1: Explore & Explain: Failures of the Wave Model

LESSON 2: Explore & Explain: Bohr's Atomic Model

Smithsonian

Additional Resources

LESSON 1
LIGHT AND QUANTIZED ENERGY

FOCUS QUESTION
What is light made of?

The Atom and Unanswered Questions

After discovering three subatomic particles in the early 1900s, scientists continued their quest to understand atomic structure and the arrangement of electrons within atoms.

Rutherford proposed that all of an atom's positive charge and virtually all of its mass are concentrated in a nucleus that is surrounded by fast-moving electrons. The model did not explain how the atom's electrons are arranged in the space around the nucleus. Nor did it address the question of why the negatively charged electrons are not pulled into the atom's positively charged nucleus. Rutherford's nuclear model did not begin to account for the differences and similarities in chemical behavior among the various elements.

For example, consider the elements lithium, sodium, and potassium, which are found in different periods on the periodic table but have similar chemical behaviors. All three elements appear metallic in nature, and their atoms react vigorously with water to liberate hydrogen gas. In fact, as shown in **Figure 1**, both sodium and potassium react so violently that the hydrogen gas can ignite and even explode.

Lithium Sodium Potassium

Matt Meadows/McGraw-Hill Education

Figure 1 Different elements can have similar reactions with water.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE
Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE
GO ONLINE to find these activities and more resources.

Applying Practice: Wave Characteristics
HS-PS4-1. Use mathematical representations to support a claim regarding relationships among the frequency, wavelength, and speed of waves traveling in various media.

CCC **Identify Crosscutting Concepts**
Create a table of the crosscutting concepts and fill in examples you find as you read.

In the early 1900s, scientists began to unravel the puzzle of chemical behavior. They observed that certain elements emitted visible light when heated in a flame. Analysis of the emitted light revealed that an element's chemical behavior is related to the arrangement of the electrons in its atoms. To understand this relationship and the nature of atomic structure, it will be helpful to first understand the nature of light.

The Wave Nature of Light

Visible light is a type of **electromagnetic radiation**—a form of energy that exhibits wavelike behavior as it travels through space. It can be modeled as a wave of changing electric and magnetic fields. Other examples of electromagnetic radiation include microwaves, X rays, and television and radio waves.

Characteristics of waves

All waves can be described by several characteristics, a few of which might be familiar to you from everyday experience. You might have seen concentric waves when dropping an object into water, as shown in **Figure 2a**.

The **wavelength** (represented by λ, the Greek letter lambda) is the shortest distance between equivalent points on a continuous wave. For example, in **Figure 2b,** the wavelength is measured from crest to crest or from trough to trough. Wavelength is usually expressed in meters, centimeters, or nanometers ($1\ nm = 1 \times 10^{-9}\ m$).

The **frequency** (represented by ν, the Greek letter nu) is the number of waves that pass a given point per second. One hertz (Hz), the SI unit of frequency, equals one wave per second. In calculations, frequency is expressed with units of waves per second, (1/s) or (s^{-1}); the term waves is understood. A particular frequency can be expressed in the following ways: $652\ Hz = 652\ waves/second = 652/s = 652\ s^{-1}$.

The **amplitude** of a wave is the wave's height from the origin to a crest, or from the origin to a trough, as illustrated in **Figure 2b.** Wavelength and frequency do not affect the amplitude of a wave.

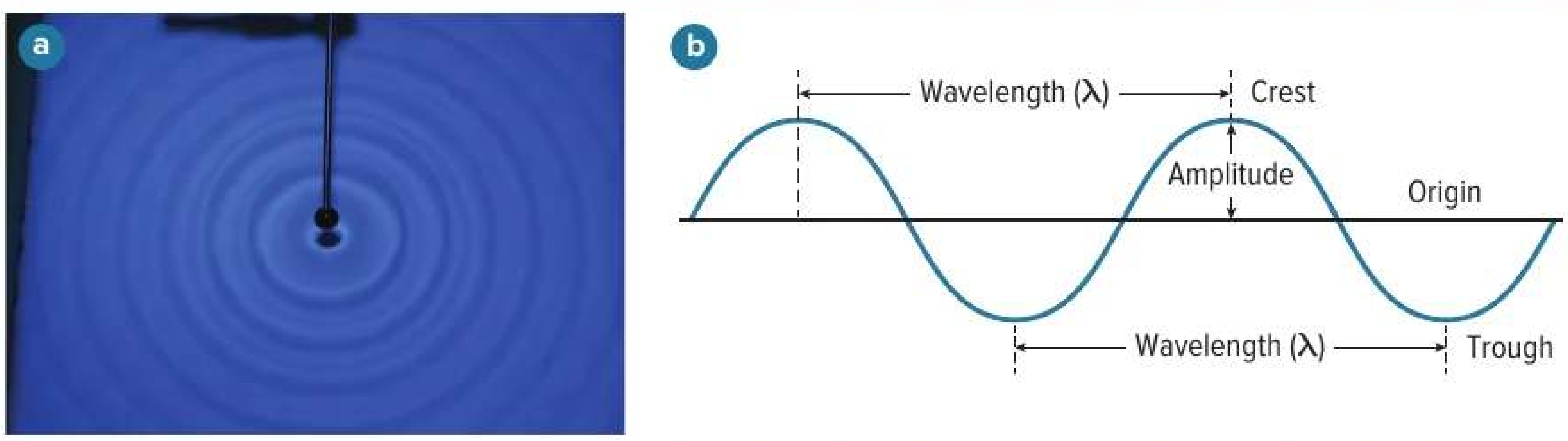


Figure 2 **a.** The concentric waves in the water show the characteristic properties of all waves. **b.** Amplitude, wavelength, and frequency are the main characteristics of waves.

Identify *a crest, a trough, and one wavelength in the photo.*

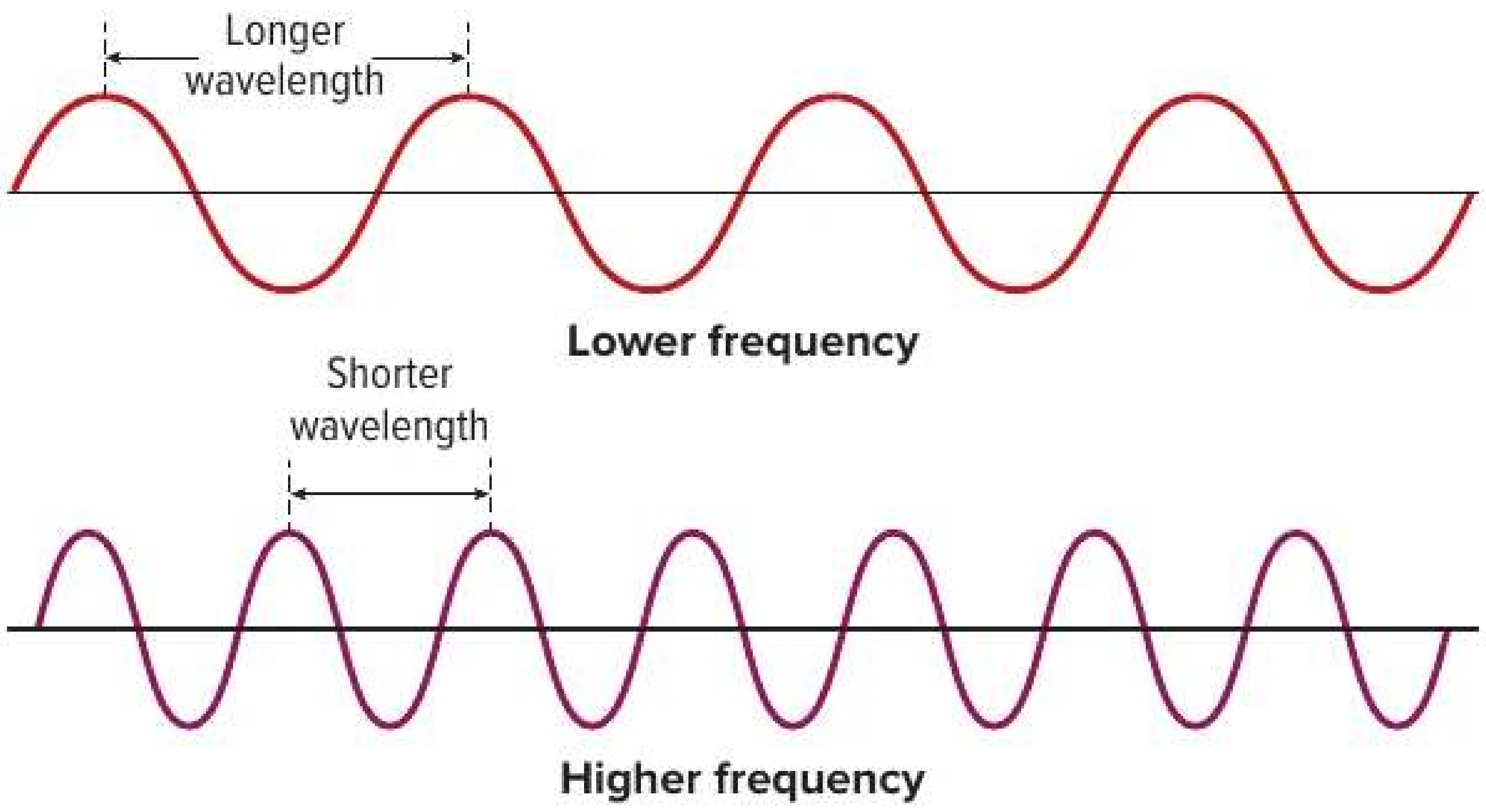


Figure 3 These waves illustrate the relationship between wavelength and frequency. As frequency increases, wavelength decreases.

Infer *Does frequency or wavelength affect amplitude?*

All electromagnetic waves, including visible light, travel at a speed of 3.00×10^8 m/s in a vacuum. Because the speed of light is such an important and universal value, it is given its own symbol, c. The speed of light is the product of its wavelength (λ) and its frequency (ν).

Electromagnetic Wave Relationship

$$c = \lambda\nu$$

c is the speed of light in a vacuum.
λ is the wavelength.
ν is the frequency.

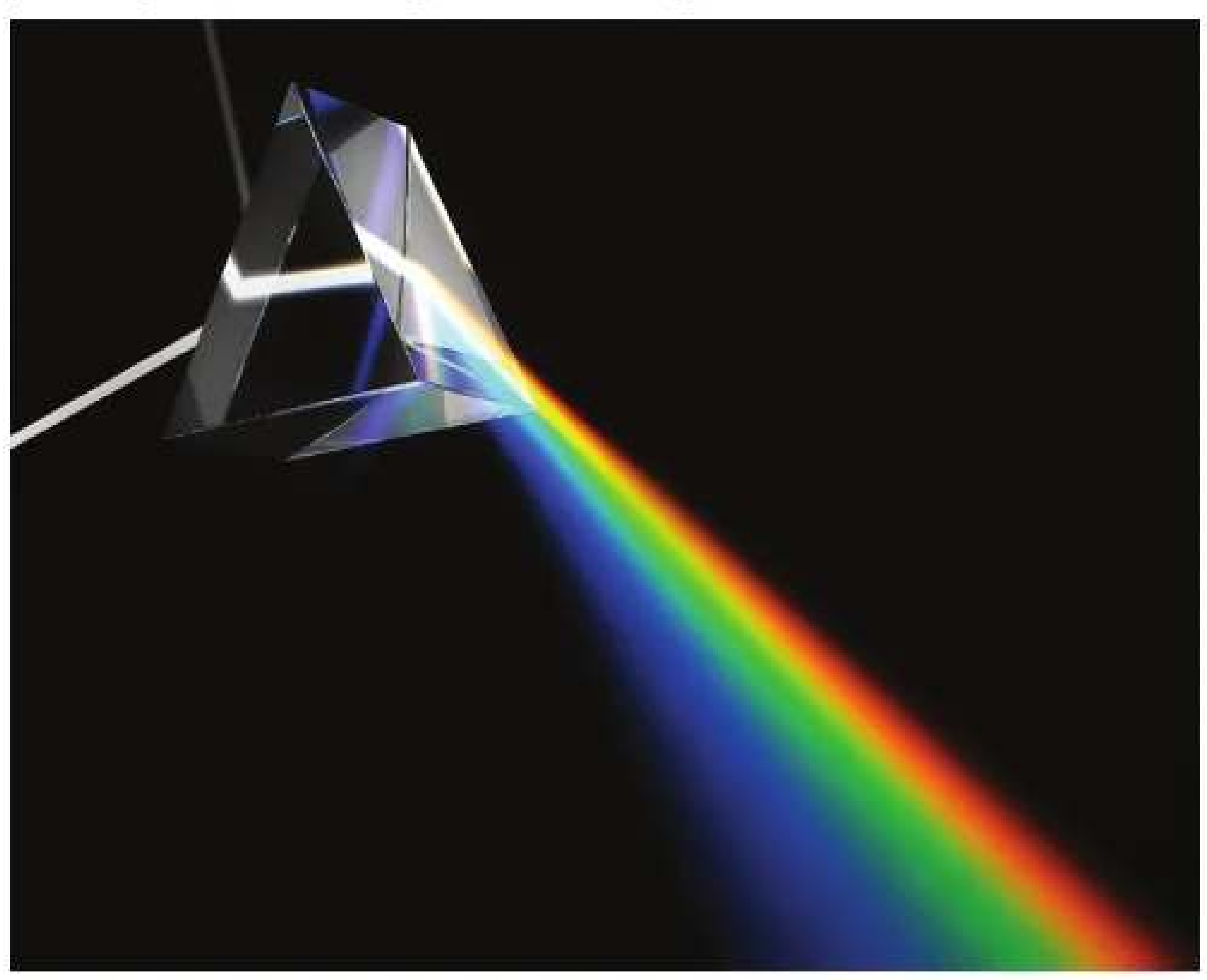
The speed of light in a vacuum is equal to the product of the wavelength and the frequency.

Although the speed of all electromagnetic waves in a vacuum is the same, waves can have different wavelengths and frequencies. As you can see from the equation above, wavelength and frequency are inversely related; in other words, as one quantity increases, the other decreases. To better understand this relationship, examine the two waves illustrated in **Figure 3**. Although both waves travel at the speed of light, you can see that the red wave has a longer wavelength and lower frequency than the violet wave.

Electromagnetic spectrum

Sunlight, which is one example of white light, contains a nearly continuous range of wavelengths and frequencies. White light passing through a prism separates into a continuous spectrum of colors similar to the spectrum in **Figure 4**. These are the colors of the visible spectrum. The spectrum is called continuous because each point of it corresponds to a unique wavelength and frequency.

Figure 4 When white light passes through a prism, it is separated into a continuous spectrum of its different components—red, orange, yellow, green, blue, indigo, and violet light.



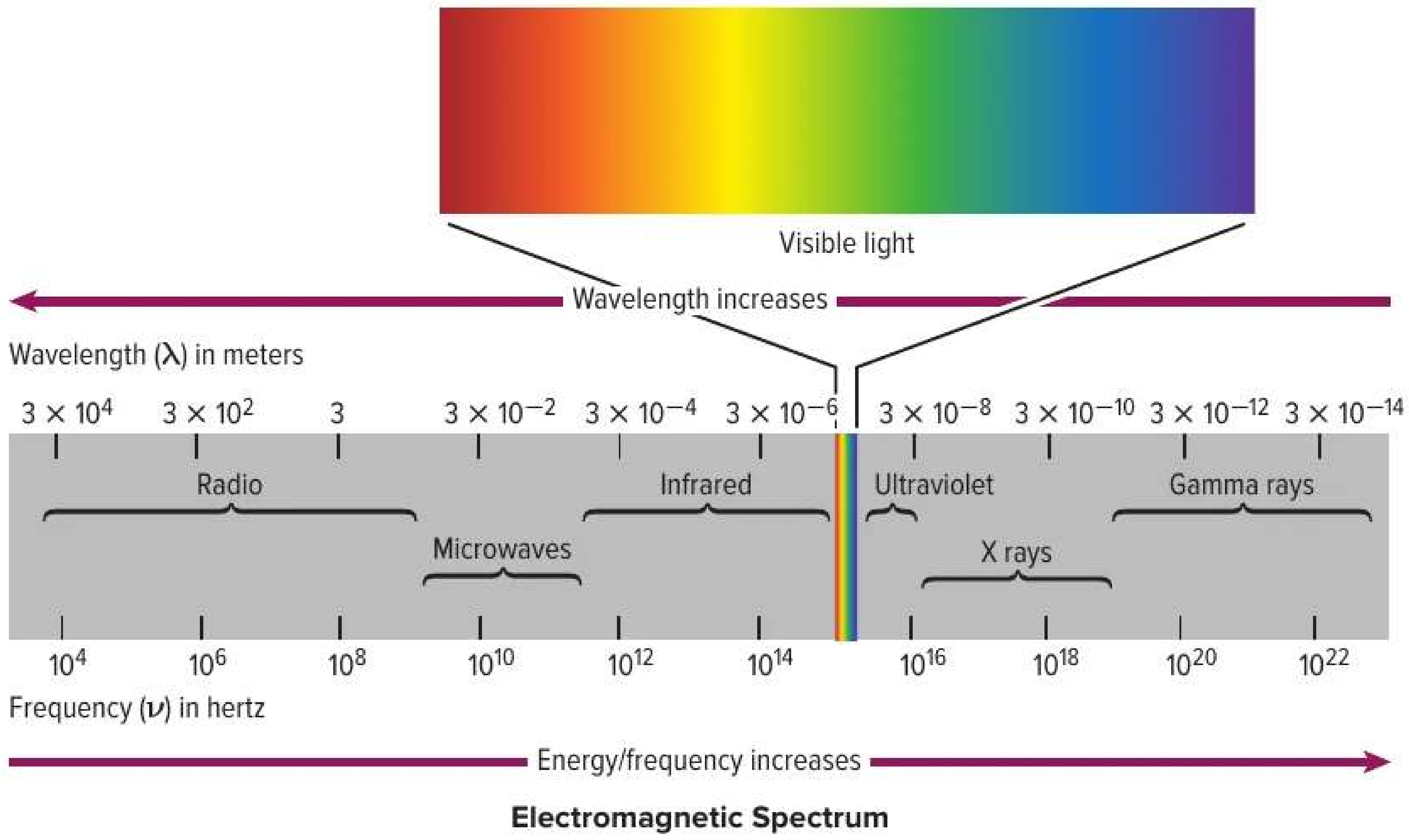


Figure 5 The electromagnetic spectrum covers a wide range of frequencies. The visible-light section of the spectrum is very narrow. As frequency and energy increase, wavelength decreases.

The visible spectrum of light, shown in **Figure 4**, comprises only a small portion of the complete electromagnetic spectrum. The complete electromagnetic spectrum is illustrated in **Figure 5**. The **electromagnetic spectrum,** also called the EM spectrum, includes all forms of electromagnetic radiation, with the only differences in the types of radiation being their frequencies and wavelengths.

Note in **Figure 4** that the bend varies with the wavelengths as they pass through the prism, resulting in the sequence of the colors red, orange, yellow, green, blue, indigo, and violet. In examining the energy of the radiation shown in **Figure 5**, note that energy increases with increasing frequency. Thus, looking back at **Figure 3**, the violet light, with its greater frequency, has more energy than the red light. This relationship between frequency and energy will be explained in the next lesson. The wavelength and frequency of a wave are related to one another by the speed of travel of the wave, which depends on the type of wave and the medium through which it is passing. For light waves, you can use the formula $c = \lambda\nu$ to calculate the wavelength or frequency of any wave.

PHYSICS Connection Electromagnetic radiation from diverse origins constantly bombards us. In addition to the radiation from the Sun, technology such as radio and TV signals, phone relay stations, lightbulbs, medical X-ray equipment, and particle accelerators also produce radiation. Natural sources on Earth, such as lightning, natural radioactivity, and even the glow of fireflies, also contribute. Our knowledge of the universe is based on electromagnetic radiation emitted by distant objects and detected with instruments on Earth.

Get It?

Explain how wavelength and frequency of a wave are related.

EXAMPLE Problem 1

CALCULATING WAVELENGTH OF AN EM WAVE Microwaves are used to cook food and transmit information. What is the wavelength of a microwave that has a frequency of 3.44×10^9 Hz?

1 ANALYZE THE PROBLEM

You are given the frequency of a microwave. You also know that because microwaves are part of the electromagnetic spectrum, their speeds, frequencies, and wavelengths are related by the formula $c = \lambda\nu$. The value of c is a known constant. First, solve the equation for wavelength, then substitute the known values and solve.

Known	Unknown
$\nu = 3.44 \times 10^9$Hz	λ = ? m
$c = 3.00 \times 10^8$ m/s	

2 SOLVE FOR THE UNKNOWN

Solve the equation relating the speed, frequency, and wavelength of an electromagnetic wave for wavelength (λ).

$c = \lambda\nu$ State the electromagnetic wave relationship.

$\lambda = c/\nu$ Solve for λ.

$$\lambda = \frac{3.00 \times 10^8\,\text{m/s}}{3.44 \times 10^9\,\text{Hz}}$$ Substitute $c = 3.00 \times 10^8$ m/s and $\nu = 3.44 \times 10^9$ Hz.

Note that hertz is equivalent to 1/s or s^{-1}.

$$\lambda = \frac{3.00 \times 10^8\,\text{m}/\cancel{\text{s}}}{3.44 \times 10^9\,\cancel{\text{s}^{-1}}}$$ Divide numbers and units.

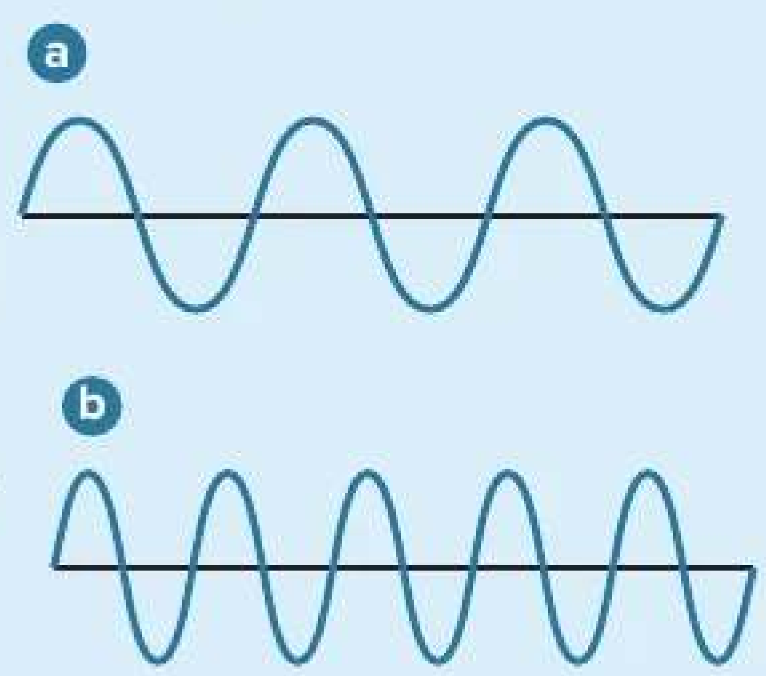
$\lambda = 8.72 \times 10^{-2}$m

3 EVALUATE THE ANSWER

The answer is correctly expressed in a unit of wavelength (m). Both of the known values in the problem are expressed with three significant figures, so the answer should have three significant figures, which it does. The value for the wavelength is within the wavelength range for microwaves shown in **Figure 5.**

PRACTICE Problems ADDITIONAL PRACTICE

1. Objects get their colors from reflecting only certain wavelengths when hit with white light. Light reflected from a green leaf is found to have a wavelength of 4.90×10^{-7} m. What is the frequency of the light?
2. When light or longer wavelength electromagnetic radiation is absorbed in matter, it is generally converted into thermal energy (heat). Shorter wavelength electromagnetic radiation (ultraviolet, X-rays, gamma rays) can ionize atoms and cause damage to living cells. X-rays can penetrate body tissues and are widely used to diagnose and treat disorders of internal body structures. What is the frequency of an X-ray with a wavelength of 1.15×10^{-10} m?
3. After careful analysis, an electromagnetic wave is found to have a frequency of 7.8×10^6 Hz. What is the speed of the wave?
4. **CHALLENGE** While an FM radio station broadcasts at a frequency of 94.7 MHz, an AM station broadcasts at a frequency of 820 kHz. What are the wavelengths of the two broadcasts? Which of the drawings on the right corresponds to the FM station? To the AM station?



The Particle Nature of Light

While considering light as a wave explains much of its everyday behavior, it fails to adequately describe important aspects of light's interactions with matter. The wave model of light cannot explain why heated objects emit only certain frequencies of light at a given temperature, or why some metals emit electrons when light of a specific frequency shines on them. Scientists realized that a new model or a revision of the wave model of light was needed to address these phenomena.

The quantum concept

When objects are heated, they emit glowing light. **Figure 6** illustrates this phenomenon with iron. A piece of iron appears dark gray at room temperature, glows red when heated sufficiently, and turns orange, then bluish in color at even higher temperatures. As the iron gets hotter, it has more energy and emits different colors of light. These different colors correspond to different frequencies and wavelengths.

The wave model could not explain the emission of these different wavelengths. In 1900, German physicist Max Planck (1858–1947) began searching for an explanation of this phenomenon. His study led him to a startling conclusion: matter can gain or lose energy only in small, specific amounts called quanta. A **quantum** is the minimum amount of energy that can be gained or lost by an atom.

Planck proposed that the energy emitted by hot objects was quantized. He also showed that there is a direct relationship between the energy of a quantum and the frequency of emitted radiation.

Figure 6 The wavelength of the light emitted by heated metal, such as the iron above, depends on the temperature. At room temperature, iron is gray. When heated, it first turns red, then glowing orange.

Identify *the color of the piece of iron with the greatest kinetic energy.*

Energy of a Quantum

$$E_{\text{quantum}} = h\nu$$

E_{quantum} represents energy.
h is Planck's constant.
ν represents frequency.

The energy of a quantum is given by the product of Planck's constant and the frequency.

Planck's constant, h, has a value of 6.626×10^{-34} J·s, where J is the symbol for joule, the SI unit of energy. The equation shows that the energy of radiation increases as the radiation's frequency, ν, increases. According to Planck's theory, for a given frequency, ν, matter can emit or absorb energy only in whole-number multiples of $h\nu$; that is, $1h\nu$, $2h\nu$, $3h\nu$, and so on.

A useful analogy is that of a child building a wall with wooden blocks. The child can add to or take away from the wall only in increments of whole numbers of blocks. Similarly, matter can have only certain amounts of energy—quantities of energy between these values do not exist.

Planck and other physicists of the time thought the concept of quantized energy was revolutionary, and some found it disturbing. Prior experience had led scientists to think that energy could be absorbed and emitted in continually varying quantities, with no minimum limit to the amount. For example, think about heating a cup of water in a microwave oven. It seems that you can add any amount of thermal energy to the water by regulating the power and duration of the microwaves. Instead, the water's temperature increases in infinitesimal steps as its molecules absorb quanta of energy. Because these steps are so small, the temperature seems to rise in a continuous, rather than a stepwise, manner.

Real-World Chemistry

The Photoelectric Effect

SOLAR ENERGY is sometimes used to power road signs. Photovoltaic cells use the photoelectric effect to convert the energy of light into electric energy.

The photoelectric effect

Scientists also knew that the wave model of light could not explain a phenomenon called the photoelectric effect. In the **photoelectric effect,** electrons, called photoelectrons, are emitted from a metal's surface when light at or above a certain frequency shines on the surface, as shown in **Figure 7**.

The wave model predicts that given enough time, even low-energy, low-frequency light would accumulate and supply enough energy to eject photoelectrons from a metal. In reality, a metal will not eject photoelectrons below a specific frequency of incident light. For example, no matter how intensely or how long it shines, light with a frequency less than 1.14×10^{15} Hz does not eject photoelectrons from silver. But even dim light with a frequency equal to or greater than 1.14×10^{15} Hz ejects photoelectrons from silver.

Get It?

Describe the photoelectric effect.

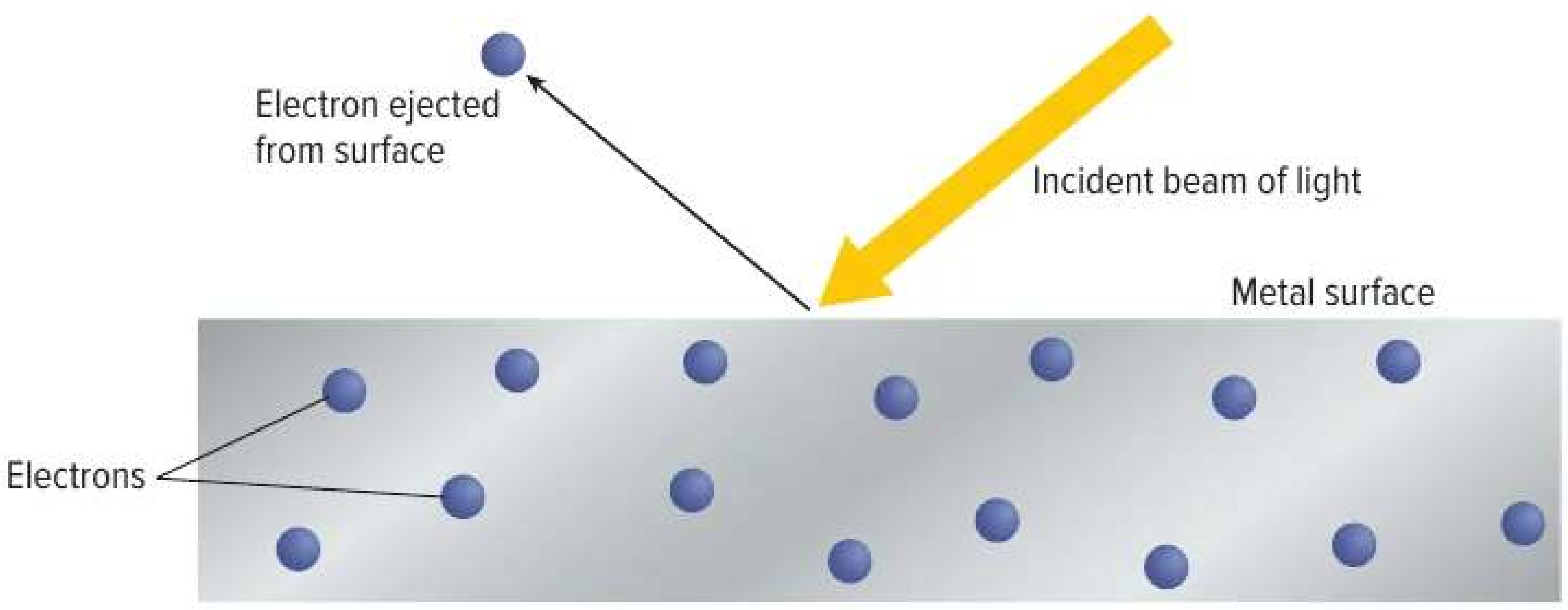


Figure 7 The photoelectric effect occurs when light of a certain frequency strikes a metal surface and ejects electrons. When the intensity of the light increases, the number of electrons ejected increases. When the frequency (energy) of the light increases, the energy of the ejected electrons increases.

Andrew Fox/Corbis/Getty Images

Light's dual nature

To explain the photoelectric effect, Albert Einstein proposed in 1905 that light has a dual nature. A beam of light has wavelike and particlelike properties. It can be thought of, or modeled, as a beam of bundles of energy called photons. A **photon** is a massless particle that carries a quantum of energy. Extending Planck's idea of quantized energy, Einstein calculated that a photon's energy depends on its frequency.

Energy of a Photon

$$E_{\text{photon}} = h\nu$$

E_{photon} represents energy.
h is Planck's constant.
ν represents frequency.

The energy of a photon is given by the product of Planck's constant and the frequency.

Einstein also proposed that the energy of a photon must have a certain threshold value to cause the ejection of a photoelectron from the surface of the metal. Thus, even small numbers of photons with energy above the threshold value will cause the photoelectric effect. Einstein won the Nobel Prize in Physics in 1921 for this work.

EXAMPLE Problem 2

CALCULATE THE ENERGY OF A PHOTON Every object gets its color by reflecting a certain portion of incident light. The color is determined by the wavelength of the reflected photons, thus by their energy. What is the energy of a photon from the violet portion of the Sun's light if it has a frequency of $7.230 \times 10^{14}\ \text{s}^{-1}$?

1 ANALYZE THE PROBLEM

Known	Unknown
$\nu = 7.230 \times 10^{14}\ \text{s}^{-1}$	$E_{\text{photon}} = ?\ \text{J}$
$h = 6.626 \times 10^{-34}\ \text{J·s}$	

2 SOLVE FOR THE UNKNOWN

$E_{\text{photon}} = h\nu$ — State the equation for the energy of a photon.

$E_{\text{photon}} = (6.626 \times 10^{-34}\ \text{J·s})(7.230 \times 10^{14}\ \text{s}^{-1})$ — Substitute $h = 6.626 \times 10^{-34}\ \text{J·s}$ and $\nu = 7.230 \times 10^{14}\ \text{s}^{-1}$.

$E_{\text{photon}} = 4.791 \times 10^{-19}\ \text{J}$ — Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

As expected, the energy of a single photon of light is extremely small. The unit is joules, an energy unit, and there are four significant figures.

PRACTICE Problems

ADDITIONAL PRACTICE

5. Calculate the energy possessed by a single photon of each of the following types of electromagnetic radiation.

a. $6.32 \times 10^{20}\ \text{s}^{-1}$ **b.** $9.50 \times 10^{13}\ \text{Hz}$ **c.** $1.05 \times 10^{16}\ \text{s}^{-1}$

6. The blue color in some fireworks occurs when copper(I) chloride is heated to approximately 1500 K and emits blue light of wavelength 4.50×10^{2} nm. How much energy does one photon of this light carry?

7. **CHALLENGE** The microwaves used to heat food have a wavelength of 0.125 m. What is the energy of one photon of the microwave radiation?

Atomic Emission Spectra

Have you ever wondered how light is produced in the glowing tubes of neon signs? This is another phenomenon that cannot be explained by the wave model of light. The light of a neon sign is produced by passing electricity through a tube filled with neon gas. Neon atoms in the tube absorb energy and become excited. These excited atoms return to their stable state by emitting light to release that energy. If the light emitted by the neon is passed through a glass prism, neon's atomic emission spectrum is produced.

The **atomic emission spectrum** of an element is the set of frequencies of the electromagnetic waves emitted by atoms of the element. **Figure 8** shows the purple-pink glow produced by excited hydrogen atoms and the visible portion of hydrogen's emission spectrum responsible for producing the glow. Note that an atomic emission spectrum is not a continuous spectrum. Rather, it consists of several individual lines of color corresponding to the frequencies of radiation emitted by the atoms.

Each element's atomic emission spectrum is unique and can be used to identify an element. For example, when a platinum wire is dipped into a strontium nitrate solution and then held in a burner flame, the strontium atoms emit a characteristic red color.

Get It?
Explain how an emission spectrum is produced.

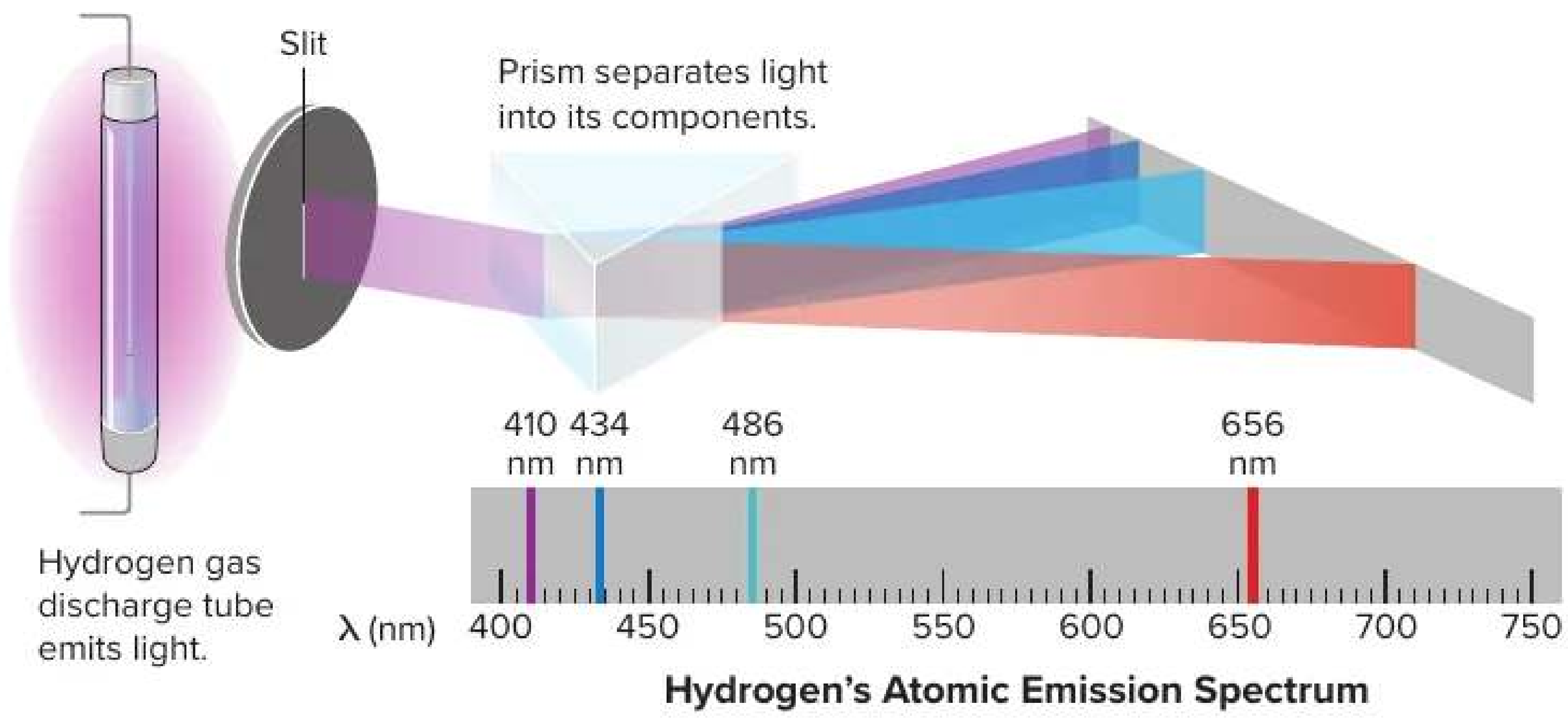


Figure 8 The purple light emitted by hydrogen can be separated into its different components using a prism. Hydrogen has an atomic emission spectrum that comprises four lines of different wavelengths.

Determine *Which line has the highest energy?*

STEM CAREER Connection

Astrochemist

Do you like chemistry, planetary science, chemical biology, physics, astronomy, and computational science? A career in astrochemistry may be the career for you. Astrochemists use telescopes, satellites, and space vehicles to collect spectroscopic data and analyze it. In this career, knowledge from several scientific disciplines is used to analyze and model the data collected.

ACADEMIC VOCABULARY

phenomenon

an observable fact or event *During rainstorms, electric currents often pass from the sky to Earth—a phenomenon called lightning.*

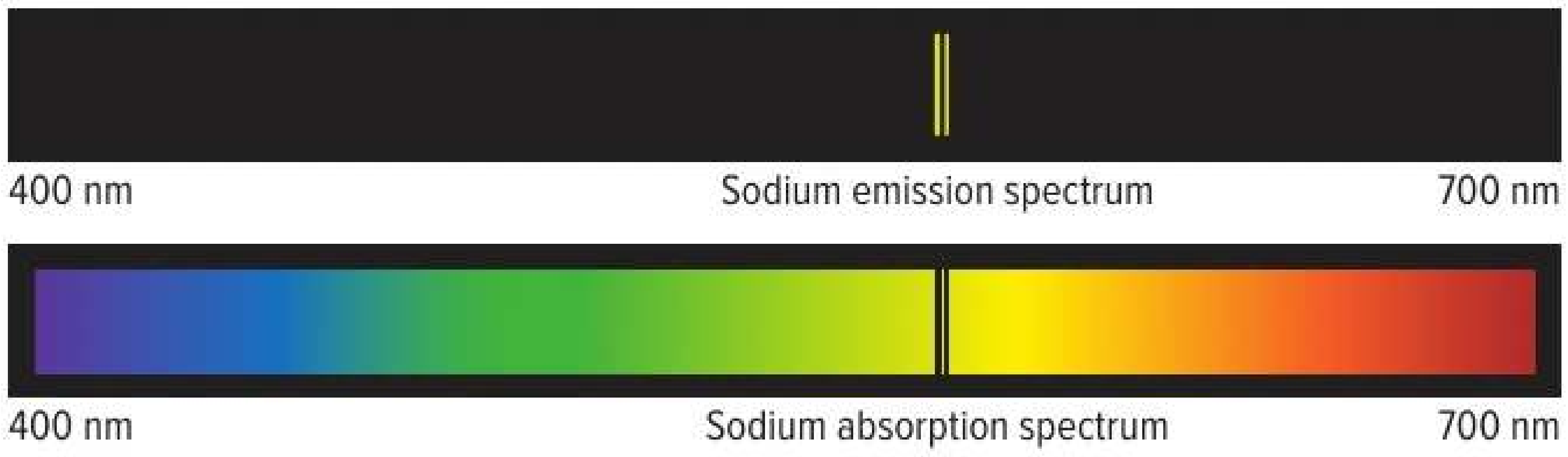


Figure 9 When excited sodium atoms return to a less excited state, they emit light at certain frequencies, producing an emission spectrum. When a continuous spectrum of light passes through sodium gas, atoms in the gas absorb light at those same frequencies, producing an absorption spectrum with dark spectral lines.

ASTRONOMY Connection Astronomers use atomic spectra to determine the composition of the outer layers of stars. When a continuous spectrum of light from within a star passes through the outer layers of the star, atoms in the outer layers absorb light at certain frequencies, producing an absorption spectrum. The lines in the absorption spectrum reveal what elements are in the outer layers of the star because the frequencies absorbed in an element's absorption spectrum are the same as those emitted in the element's emission spectrum, as shown for sodium in **Figure 9.**

Check Your Progress

Summary

- All waves are defined by their wavelengths, frequencies, amplitudes, and speeds.
- In a vacuum, all electromagnetic waves travel at the speed of light.
- All electromagnetic waves have both wave and particle properties.
- Matter emits and absorbs energy in quanta.
- White light produces a continuous spectrum. An element's emission spectrum consists of a series of lines of individual colors.

Demonstrate Understanding

8. **Describe** the relationship between changing electric and magnetic fields and particles.
9. **Compare and contrast** continuous spectrum and emission spectrum.
10. **Discuss** the way in which Einstein utilized Planck's quantum concept to explain the photoelectric effect.
11. **Calculate** Heating 235 g of water from 22.6°C to 94.4°C in a microwave oven requires 7.06×10^4 J of energy. If the microwave frequency is 2.88×10^{10} s^{-1}, how many quanta are required to supply the 7.06×10^4 J?
12. **Interpret Scientific Illustrations** Use **Figure 5** and your knowledge of electromagnetic radiation to match the numbered items with the lettered items. The numbered items may be used more than once or not at all.

a. longest wavelength	1. gamma rays
b. highest frequency	2. ultraviolet light
c. greatest energy	3. radio waves

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LESSON 2
QUANTUM THEORY AND THE ATOM

FOCUS QUESTION

Why does every element produce a unique atomic emission spectrum?

Bohr's Model of the Atom

The dual wave-particle model of light accounted for several previously unexplainable phenomena, but scientists still did not understand the relationships among atomic structure, electrons, and atomic emission spectra. Recall that hydrogen's atomic emission spectrum is discontinuous; that is, it is made up of only certain frequencies of light. Why are the atomic emission spectra of elements discontinuous rather than continuous? Niels Bohr, a Danish physicist working in Rutherford's laboratory in 1913, proposed a quantum model for the hydrogen atom that seemed to answer this question. Bohr's model also correctly predicted the frequencies of the lines in hydrogen's atomic emission spectrum.

Energy states of hydrogen

Building on Planck's and Einstein's concepts of quantized energy, Bohr proposed that the hydrogen atom has only certain allowable energy states, as illustrated in **Figure 10**. The lowest allowable energy state of an atom is called its **ground state.** When an atom gains energy, it is said to be in an excited state.

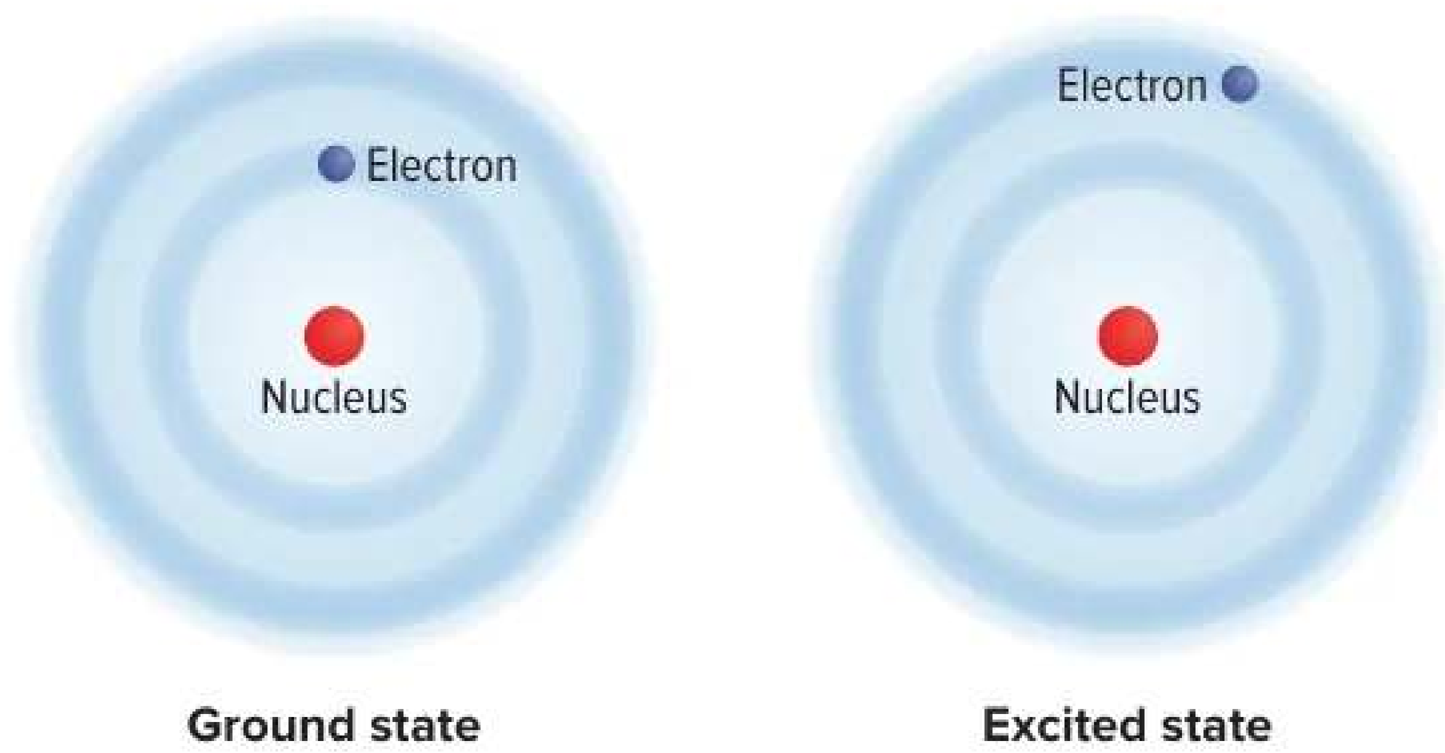


Figure 10 The figure shows an atom that has one electron. Note that the illustration is not to scale. In its ground state, the electron is associated with the lowest energy level. When the atom is in an excited state, the electron is associated with a higher energy level.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Laboratory: The Photoelectric Effect

Use mathematical and computational thinking to observe patterns in the stability and instability of physical systems.

Inquiry into Chemistry: Design Atomic Models

Plan and carry out an investigation to create a model of the charged subsections of an atom.

Table 1 Bohr's Description of the Hydrogen Atom

Bohr's Atomic Orbit	Quantum Number	Orbit Radius (nm)	Corresponding Atomic Energy Level	Relative Energy
First	$n = 1$	0.0529	1	E_1
Second	$n = 2$	0.212	2	$E_2 = 4E_1$
Third	$n = 3$	0.476	3	$E_3 = 9E_1$
Fourth	$n = 4$	0.846	4	$E_4 = 16E_1$
Fifth	$n = 5$	1.32	5	$E_5 = 25E_1$
Sixth	$n = 6$	1.90	6	$E_6 = 36E_1$
Seventh	$n = 7$	2.59	7	$E_7 = 49E_1$

Bohr suggested that the electron in a hydrogen atom moves around the nucleus in only certain allowed circular orbits. The smaller the electron's orbit, the lower the atom's energy state, or energy level. Conversely, the larger the electron's orbit, the higher the atom's energy state, or energy level. Bohr assigned a number, n, called a **quantum number,** to each orbit. He also calculated the radius of each orbit. **Table 1** shows data for the first seven energy levels of a hydrogen atom according to Bohr's model.

The hydrogen line spectrum

Bohr suggested that a hydrogen atom is in the ground state when its single electron is in the $n = 1$ orbit, also called the first energy level. In the ground state, the atom does not radiate energy. When energy is added from an outside source, the electron moves to a higher-energy orbit, putting the atom in an excited state. When the atom is in an excited state, the electron can drop from the higher-energy orbit to a lower-energy orbit, as shown in **Figure 11.**

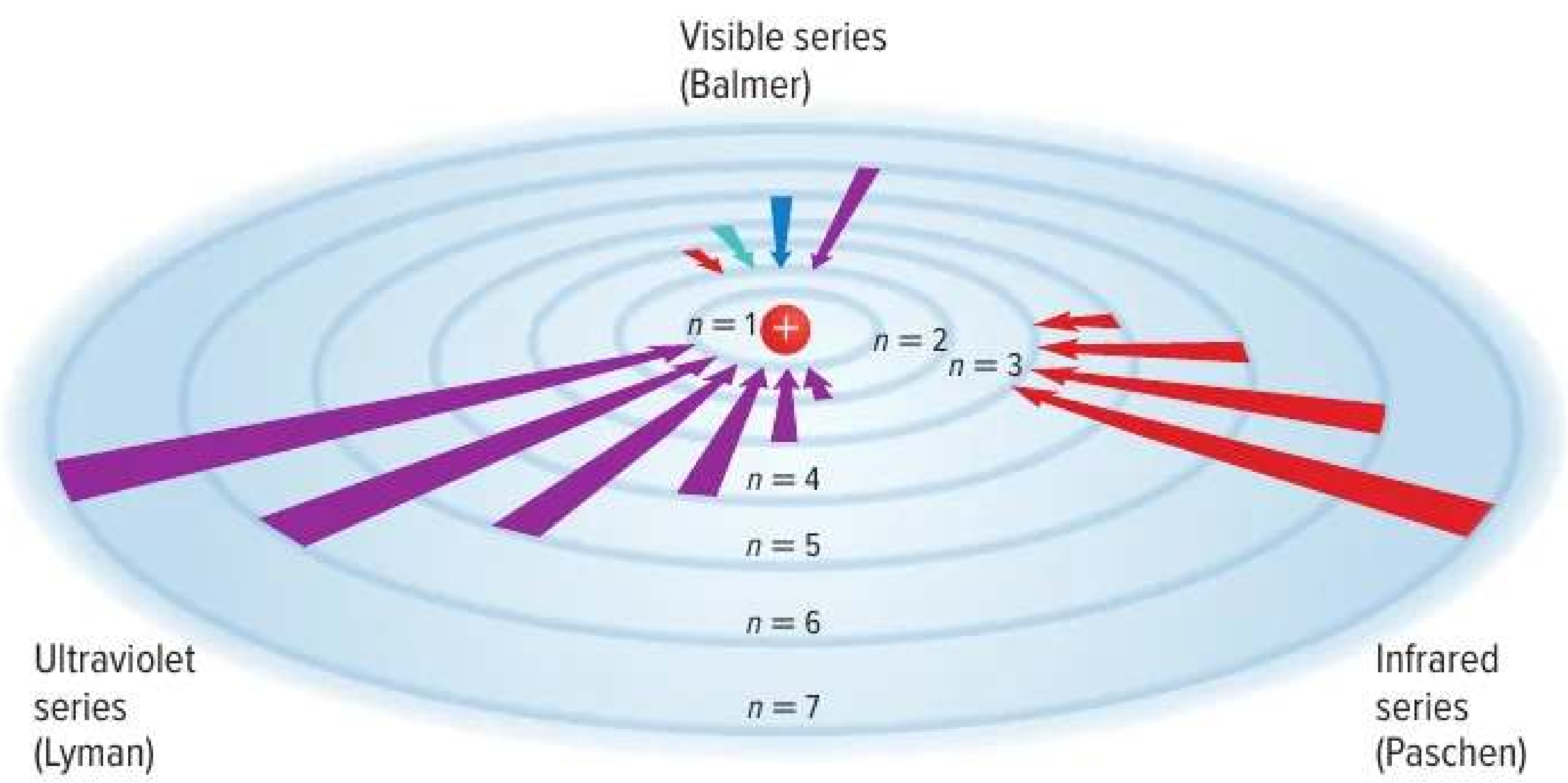


Figure 11 When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon is emitted. The ultraviolet (Lyman), visible (Balmer), and infrared (Paschen) series correspond to electrons dropping to $n = 1$, $n = 2$, and $n = 3$, respectively.

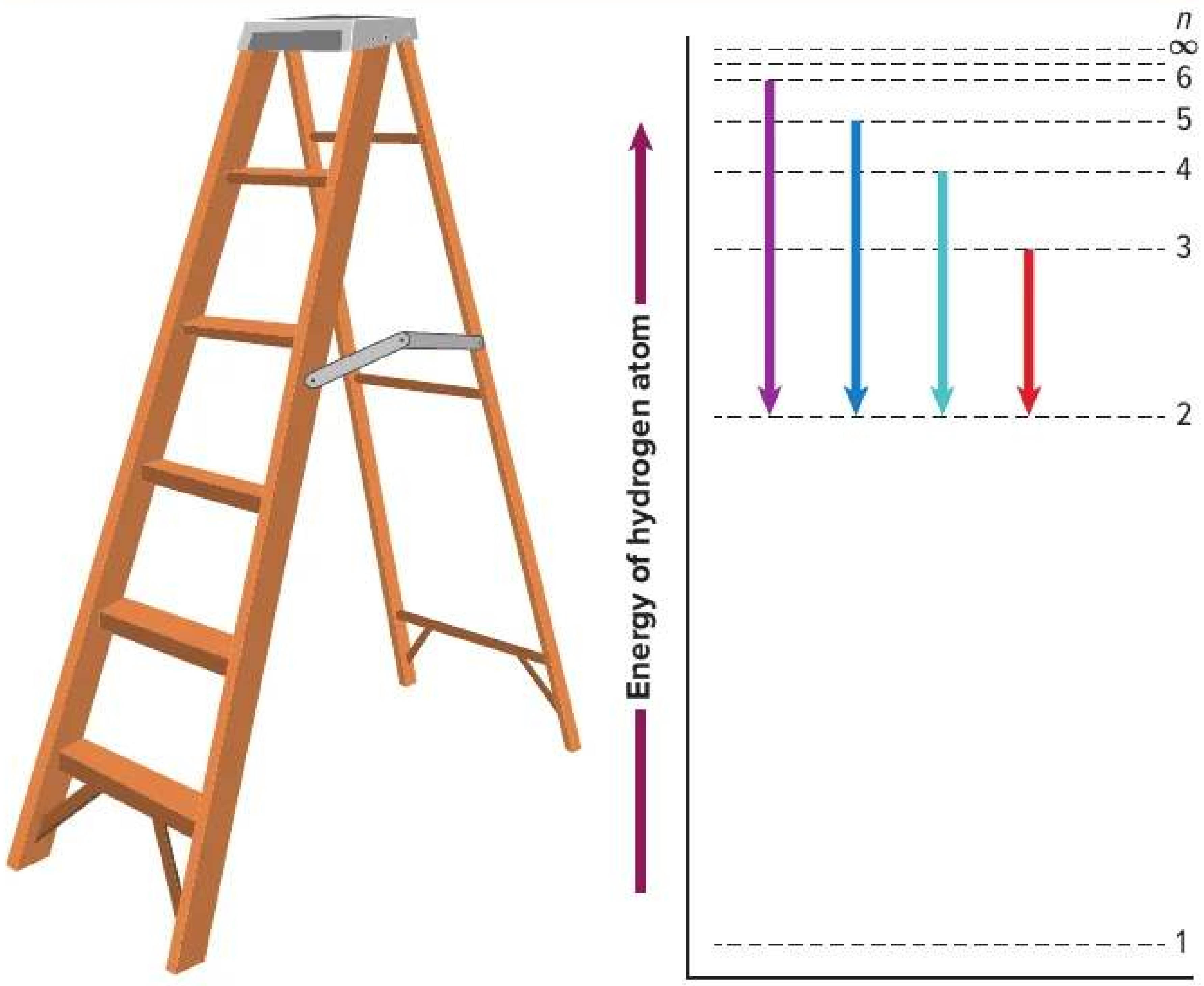


Figure 12 Only certain energy levels are allowed. The energy levels are similar to the rungs of a ladder. The four visible lines correspond to electrons dropping from a higher n to the orbit $n = 2$. As n increases, the hydrogen atom's energy levels are closer to each other.

As a result of this transition, the atom emits a photon corresponding to the energy difference between the two levels.

$$\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}} = E_{\text{photon}} = h\nu$$

Because only certain atomic energies are possible, only certain frequencies of electromagnetic radiation can be emitted.

You might compare hydrogen's atomic energy states to rungs on a ladder, as shown in **Figure 12**. A person can climb up or down the ladder only from rung to rung. Similarly, the hydrogen atom's electron can move only from one allowable orbit to another, and therefore, can emit or absorb only certain amounts of energy, corresponding to the energy difference between the two orbits. Unlike rungs on a ladder, however, the hydrogen atom's energy levels are not evenly spaced.

Figure 12 also illustrates the four electron transitions that account for visible lines in hydrogen's atomic emission spectrum, shown in **Figure 8**. Electron transitions from higher-energy orbits to the second orbit account for all of hydrogen's visible lines, which form the Balmer series. Other electron transitions have been measured that are not visible, such as the Lyman series (ultraviolet), in which electrons drop into the $n = 1$ orbit, and the Paschen series (infrared), in which electrons drop into the $n = 3$ orbit.

Get It?

Explain why different colors of light result from electron behavior in the atom.

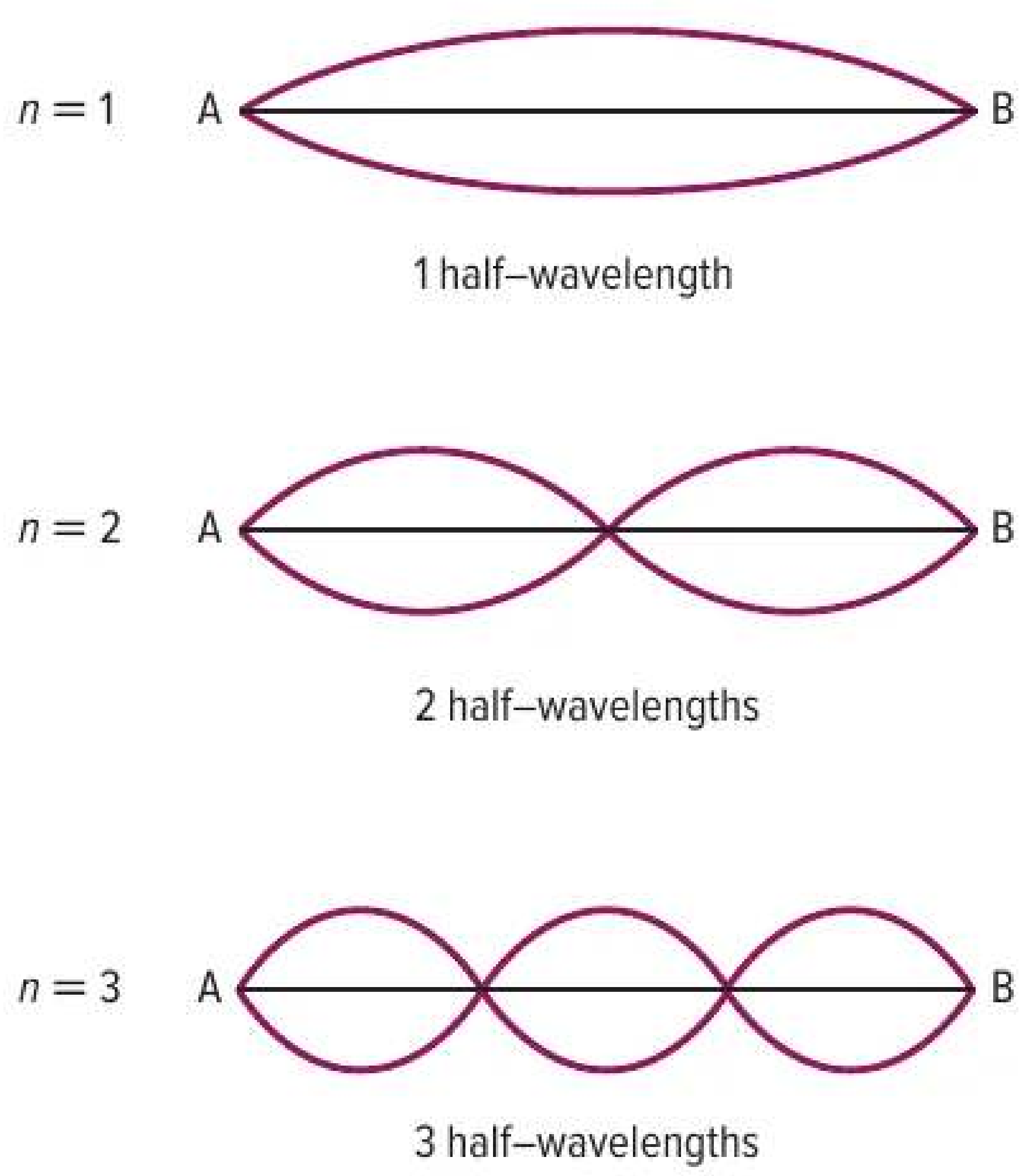
The limits of Bohr's model

Bohr's model explained hydrogen's observed spectral lines. However, the model failed to explain the spectrum of any other element. Moreover, Bohr's model did not fully account for the chemical behavior of atoms. In fact, although Bohr's idea of quantized energy levels laid the groundwork for atomic models to come, later experiments demonstrated that the Bohr model was fundamentally incorrect. The movements of electrons in atoms are not completely understood even now; however, substantial evidence indicates that electrons do not move around the nucleus in circular orbits.

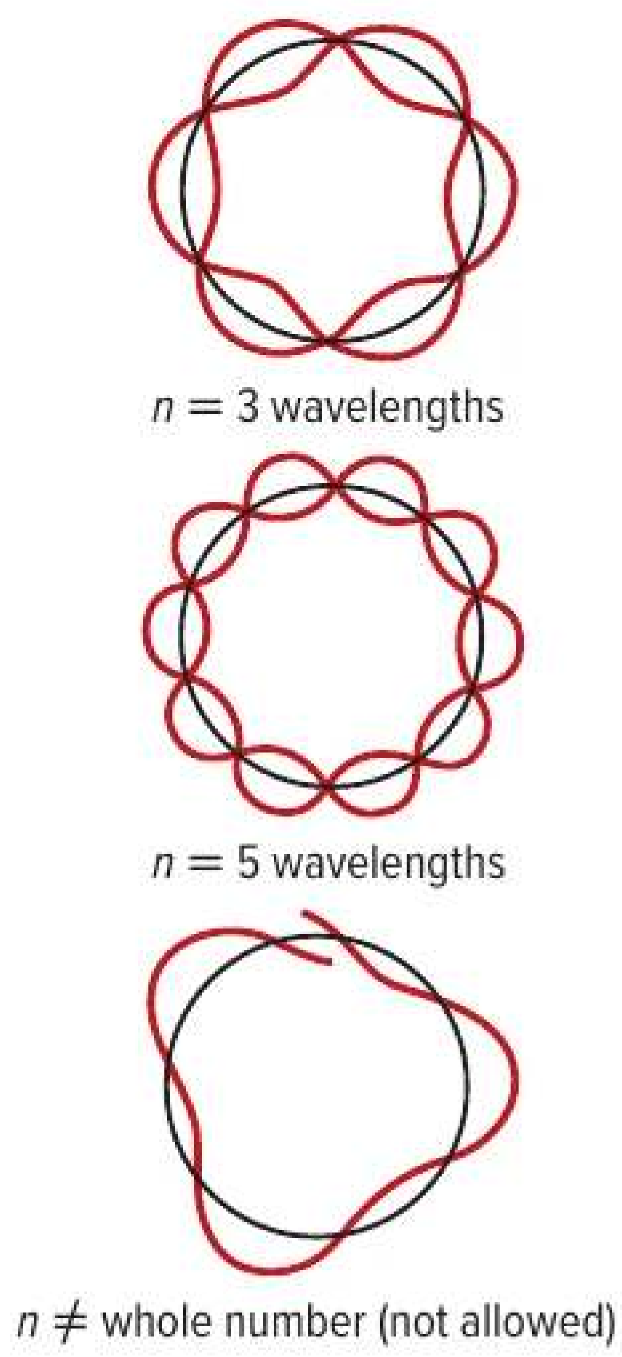
The Quantum Mechanical Model of the Atom

Scientists in the mid-1920s, convinced that the Bohr atomic model was incorrect, formulated new and innovative explanations of how electrons are arranged in atoms. In 1924, a French graduate student in physics named Louis de Broglie (1892–1987) proposed a new idea, shown in **Figure 13** and discussed on the following page.

Figure 13 **a.** The string on the harp vibrates between two fixed endpoints. **b.** The vibrations of a string between the two fixed endpoints labeled A and B are limited to multiples of half-wavelengths. **c.** Electrons on circular orbits can only have whole numbers of wavelengths.



b **Vibrating harp string**



c **Orbiting electron**

Brian Kanof/McGraw-Hill Education

Electrons as waves

De Broglie had been thinking that Bohr's quantized electron orbits had characteristics similar to those of waves. For example, as **Figures 13a** and **13b** show, only multiples of half-wavelengths are possible on a plucked harp string because the string is fixed at both ends. Similarly, de Broglie saw that only whole numbers of wavelengths are allowed in a circular orbit of fixed radius, as shown in **Figure 13c**.

De Broglie also reflected on the fact that light—at one time thought to be strictly a wave phenomenon—has both wave and particle characteristics. These thoughts led de Broglie to pose a new question: If waves can have particlelike behavior, could the opposite also be true? That is, can particles of matter, including electrons, behave like waves?

De Broglie knew that if an electron has wavelike motion and is restricted to circular orbits of fixed radius, only certain wavelengths, frequencies, and energies are possible. Developing his idea, de Broglie derived the following equation, called the **de Broglie equation.**

Particle Electromagnetic-Wave Relationship

$$\lambda = \frac{h}{m\nu}$$

λ represents wavelength.
h is Planck's constant.
m represents mass of the particle.
ν represents velocity.

The wavelength of a particle is the ratio of Planck's constant and the product of the particle's mass and its velocity.

The de Broglie equation predicts that all moving particles have wave characteristics. Note that the equation includes Planck's constant. Planck's constant is an exceedingly small number, 6.626×10^{-34} J·s, which helps explain why it is difficult or impossible to observe the wave characteristics of objects at the scale of everyday experience. For example, an automobile moving at 25 m/s and having a mass of 910 kg has a wavelength of 2.9×10^{-38} m, far too small to be seen or detected. By comparison, an electron moving at the same speed has the easily measured wavelength of 2.9×10^{-5} m. Subsequent experiments have proven that electrons and other moving particles do indeed have wave characteristics.

Get It?
Identify which variables in the de Broglie equation represent wavelike properties.

The Heisenberg uncertainty principle

Step by step, scientists such as Rutherford, Bohr, and de Broglie had been unraveling the mysteries of the atom. However, a conclusion reached by the German theoretical physicist Werner Heisenberg (1901–1976) proved to have profound implications for atomic models.

Heisenberg showed that it is impossible to take any measurement of an object without disturbing the object. Imagine trying to locate a hovering, helium-filled balloon in a darkened room. If you wave your hand about, you can locate the balloon's position when you touch it. However, when you touch the balloon, you transfer energy to it and change its position. You could also detect the balloon's position by turning on a flashlight. Using this method, photons of light reflected from the balloon would reach your eyes and reveal the balloon's location. Because the balloon is a macroscopic object, the effect of the rebounding photons on its position is very small and not observable.

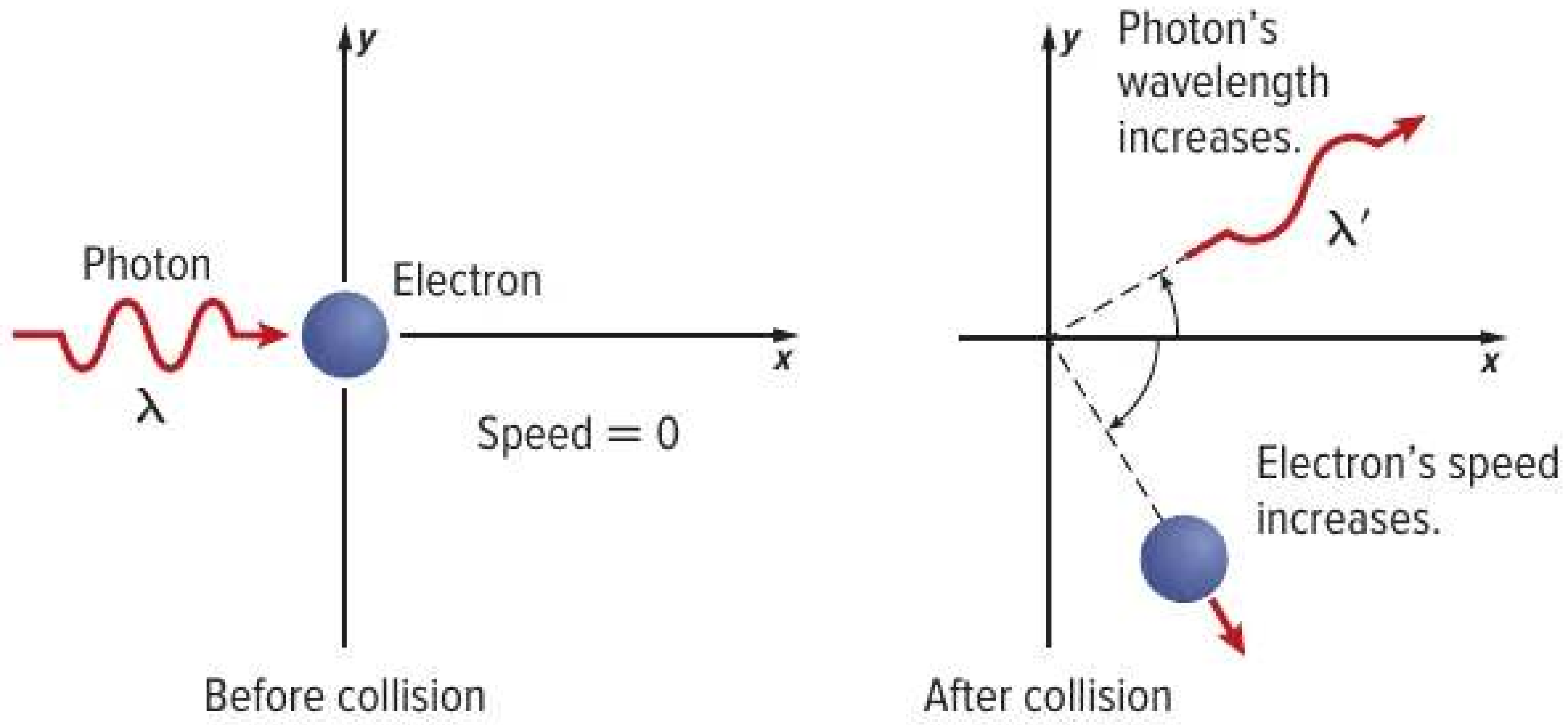


Figure 14 When a photon interacts with an electron at rest, both the velocity and the position of the electron are modified. This illustrates the Heisenberg uncertainty principle. It is impossible to know at the same time the position and the velocity of a particle.

Explain *Why has the photon's energy changed?*

Now imagine trying to determine an electron's location by "bumping" it with a high-energy photon. Because such a photon has about the same energy as an electron, the interaction between the two particles changes both the wavelength of the photon and the position and velocity of the electron, as shown in **Figure 14**. In other words, the act of observing the electron produces a significant, unavoidable uncertainty in the position and motion of the electron. Heisenberg's analysis of interactions, such as those between photons and electrons, led him to his historic conclusion. The **Heisenberg uncertainty principle** states that it is fundamentally impossible to know precisely both the velocity and position of a particle at the same time.

Although scientists of the time found Heisenberg's principle difficult to accept, it has been proven to describe the fundamental limitations of what can be observed. The interaction of a photon with a macroscopic object such as a helium-filled balloon has so little effect on the balloon that the uncertainty in its position is too small to measure. But that is not the case with an electron moving at 6×10^6 m/s near an atomic nucleus. The uncertainty of the electron's position is at least 10^{-9} m, about 10 times greater than the diameter of the entire atom.

The Heisenberg uncertainty principle also means that it is impossible to assign fixed paths for electrons like the circular orbits in Bohr's model. The only quantity that can be known is the probability for an electron to occupy a certain region around the nucleus.

Get It?

Identify the only quantity of an electron's orbit that can be determined.

CCC CROSSCUTTING CONCEPTS

Cause and Effect What empirical evidence did scientists have that supports the claim that electrons have both particle and wave properties?

The Schrödinger wave equation

In 1926, Austrian physicist Erwin Schrödinger (1887–1961) furthered the wave-particle theory proposed by de Broglie. Schrödinger derived an equation that treated the hydrogen atom's electron as a wave. Schrödinger's new model for the hydrogen atom seemed to apply equally well to atoms of other elements—an area in which Bohr's model failed. The atomic model in which electrons are treated as waves is called the wave mechanical model of the atom or the **quantum mechanical model of the atom.** Like Bohr's model, the quantum mechanical model limits an electron's energy to certain values. However, unlike Bohr's model, the quantum mechanical model makes no attempt to describe the electron's path around the nucleus.

Get It?

Compare and contrast Bohr's model and the quantum mechanical model.

Electron's probable location

The Schrödinger wave equation is too complex to be considered here. However, each solution to the equation is known as a wave function, which is related to the probability of finding the electron within a particular volume of space around the nucleus. The wave function predicts a three-dimensional region around the nucleus, called an **atomic orbital,** which describes the electron's probable location. An atomic orbital is like a fuzzy cloud in which the density at a given point is proportional to the probability of finding the electron at that point.

Figure 15a illustrates the probability map that describes the electron in the atom's lowest energy state. The probability map can be thought of as a time-exposure photograph of the electron moving around the nucleus, in which each dot represents the electron's location at an instant in time. The high density of dots near the nucleus indicates the electron's most probable location. However, it is also possible that the electron might be found at a considerable distance from the nucleus.

Get It?

Describe where electrons are located in an atom.

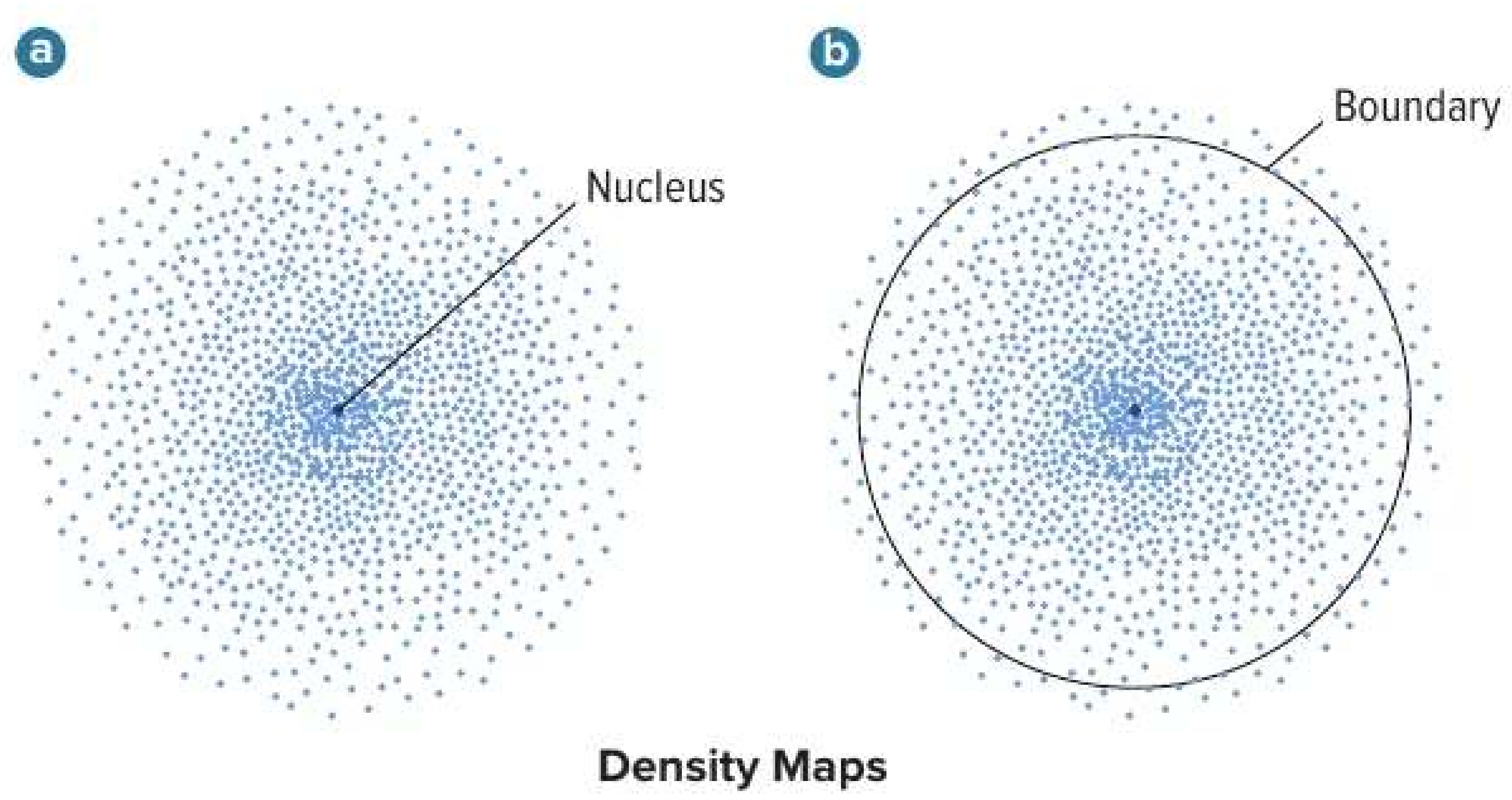


Figure 15 The density map represents the probability of finding an electron at a given position around the nucleus. **a.** The higher density of points near the nucleus shows that the electron is more likely to be found close to the nucleus. **b.** At any given time, there is a 90% probability of finding the electron within the circular region shown. This surface is sometimes chosen to represent the boundary of the atom. In this illustration, the circle corresponds to a projection of the 3-dimensional sphere that contains the electrons.

Hydrogen's Atomic Orbitals

Because the boundary of an atomic orbital is fuzzy, the orbital does not have an exact defined size. To overcome the inherent uncertainty about the electron's location, chemists arbitrarily draw an orbital's surface to contain 90% of the electron's total probability distribution. This means that the probability of finding the electron within the boundary is 0.9 and the probability of finding it outside the boundary is 0.1. In other words, it is more likely to find the electron close to the nucleus and within the volume defined by the boundary, than to find it outside the volume. The circle shown in **Figure 15b** encloses 90% of the lowest-energy orbital of hydrogen.

Principal quantum number

Recall that the Bohr atomic model assigns quantum numbers to electron orbits. Similarly, the quantum mechanical model assigns four quantum numbers to atomic orbitals. The first one is the **principal quantum number** (n) and indicates the relative size and energy of atomic orbitals. As n increases, the orbital becomes larger, the electron spends more time farther from the nucleus, and the atom's energy increases. Therefore, n specifies the atom's major energy levels. Each major energy level is called a **principal energy level.** An atom's lowest principal energy level is assigned a principal quantum number of 1. When the hydrogen atom's single electron occupies an orbital with $n = 1$, the atom is in its ground state. Up to 7 energy levels have been detected for the hydrogen atom, giving n values ranging from 1 to 7.

Energy sublevels

Principal energy levels contain **energy sublevels.** Principal energy level 1 consists of a single sublevel, principal energy level 2 consists of two sublevels, principal energy level 3 consists of three sublevels, and so on. To better understand the relationship between the atom's energy levels and sublevels, picture the seats in a wedge-shaped section of a theater, as shown in **Figure 16**. As you move away from the stage, the rows become higher and contain more seats. Similarly, the number of energy sublevels in a principal energy level increases as n increases.

Get It?

Explain the relationship between energy levels and sublevels.

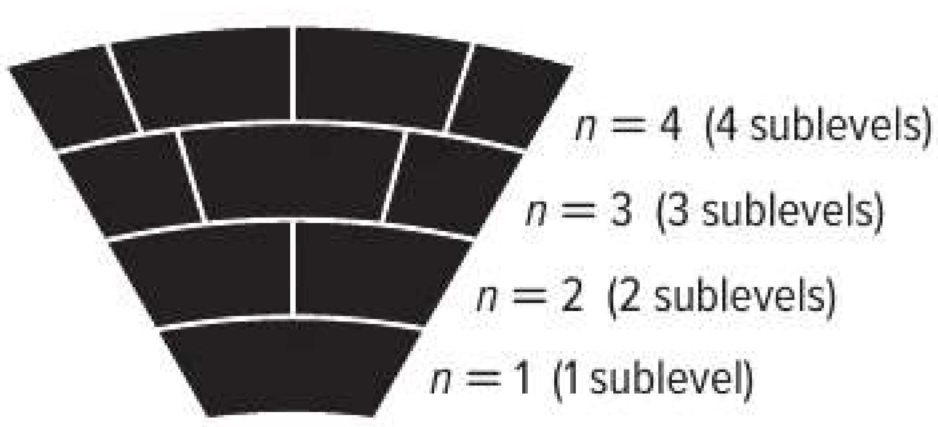


Figure 16 Energy levels can be thought of as rows of seats in a theater. The rows that are higher up and farther from the stage contain more seats. Similarly, energy levels related to orbitals farther from the nucleus contain more sublevels.

Shapes of orbitals

Sublevels are labeled *s, p, d,* or f according to the shapes of the atom's orbitals. All s orbitals are spherical, and all p orbitals are dumbbell-shaped; however, not all d or f orbitals have the same shape. Each orbital can contain, at most, two electrons. The single sublevel in principal energy level 1 corresponds to a spherical orbital called the 1s orbital. The two sublevels in principal energy level 2 are designated 2s and 2p. The 2s sublevel corresponds to the 2s orbital, which is spherical like the 1s orbital but larger in size, as shown in **Figure 17a**. The 2p sublevel corresponds to three dumbbell-shaped p orbitals designated $2p_x$, $2p_y$, and $2p_z$. The subscripts *x*, *y*, and *z* merely designate the orientations of p orbitals along the *x*, *y*, and *z* coordinate axes, as shown in **Figure 17b**. Each of the p orbitals related to an energy sublevel has the same energy.

Get It?

Describe the shapes of *s* and *p* orbitals.

Principal energy level 3 consists of three sublevels designated 3s, 3p, and 3d. Each d sublevel relates to five orbitals of equal energy. Four of the d orbitals have identical shapes but different orientations along the *x*, *y*, and *z* coordinate axes. However, the fifth orbital, d_{z^2}, is shaped and oriented differently than the other four. The shapes and orientations of the five d orbitals are illustrated in **Figure 17c**. The fourth principal energy level ($n = 4$) contains a fourth sublevel, called the 4f sublevel, which relates to seven f orbitals of equal energy. The f orbitals have complex, multilobed shapes.

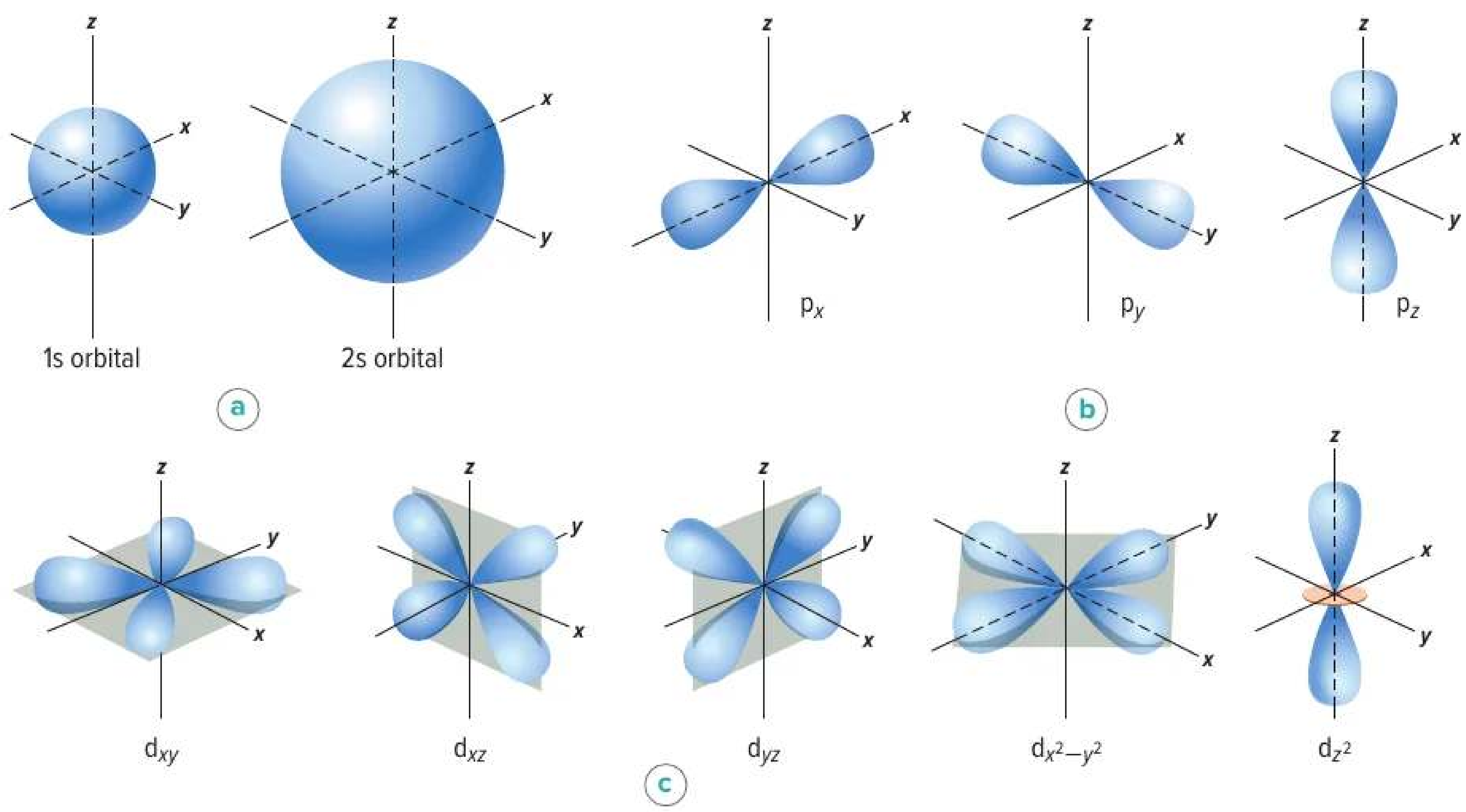


Figure 17 The shapes of atomic orbitals describe the probable distribution of electrons in energy sublevels.
a. All s orbitals are spherical, and their size increases with increasing principal quantum number.
b. The three p orbitals are dumbbell-shaped and are oriented along the three perpendicular *x*, *y*, and *z* axes.
c. Four of the five d orbitals have the same shape but lie in different planes. The d_{z^2} orbital has its own unique shape.

Table 2 Hydrogen's First Four Principal Energy Levels

Principal Quantum Number (n)	Sublevels (Types of Orbitals) Present	Number of Orbitals Related to Sublevel	Total Number of Orbitals Related to Principal Energy Level (n^2)
1	s	1	1
2	s p	1 3	4
3	s p d	1 3 5	9
4	s p d f	1 3 5 7	16

Hydrogen's first four principal energy levels, sublevels, and related atomic orbitals are summarized in **Table 2**. Note that the number of orbitals related to each sublevel is always an odd number, and that the maximum number of orbitals related to each principal energy level equals n^2.

At any given time, the electron in a hydrogen atom can occupy just one orbital. You can think of the other orbitals as unoccupied spaces—spaces available should the atom's energy increase or decrease.

Check Your Progress

Summary

- Bohr's atomic model attributes hydrogen's emission spectrum to electrons dropping from higher-energy to lower-energy orbits.
- The de Broglie equation relates a particle's wavelength to its mass, its velocity, and Planck's constant.
- The quantum mechanical model assumes that electrons have wave properties.
- Electrons occupy three-dimensional regions of space called atomic orbitals.

Demonstrate Understanding

13. **Explain** the reason, according to Bohr's atomic model, why atomic emission spectra contain only certain frequencies of light.
14. **Differentiate** between the wavelength of visible light and the wavelength of a moving soccer ball.
15. **Explain** why the location of an electron in an atom is uncertain using the Heisenberg uncertainty principle. How is the location of electrons in atoms defined?
16. **Compare and contrast** Bohr's model and the quantum mechanical model of the atom.
17. **Enumerate** the sublevels contained in the hydrogen atom's first four energy levels. What orbitals are related to each s sublevel and each p sublevel?
18. **Calculate** Use the information in **Table 1** to calculate how many times larger the hydrogen atom's seventh Bohr radius is than its first Bohr radius.

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LESSON 3
ELECTRON CONFIGURATION

FOCUS QUESTION

How are electrons arranged in atoms?

Ground-State Electron Configuration

The arrangement of electrons in an atom is called the atom's **electron configuration.** Because low-energy systems are more stable than high-energy systems, electrons in an atom tend to assume the arrangement that gives the atom the lowest energy possible. The most stable, lowest-energy arrangement of the electrons is called the element's ground-state electron configuration.

Three rules, or principles—the aufbau principle, the Pauli exclusion principle, and Hund's rule—define how electrons can be arranged in an atom's orbitals.

The aufbau principle

The **aufbau principle** states that each electron occupies the lowest energy orbital available. Therefore, your first step in determining an element's ground-state electron configuration is learning the sequence of atomic orbitals from lowest energy to highest energy. This sequence, known as an aufbau diagram, is shown in **Figure 18**. In the diagram, each box represents an atomic orbital.

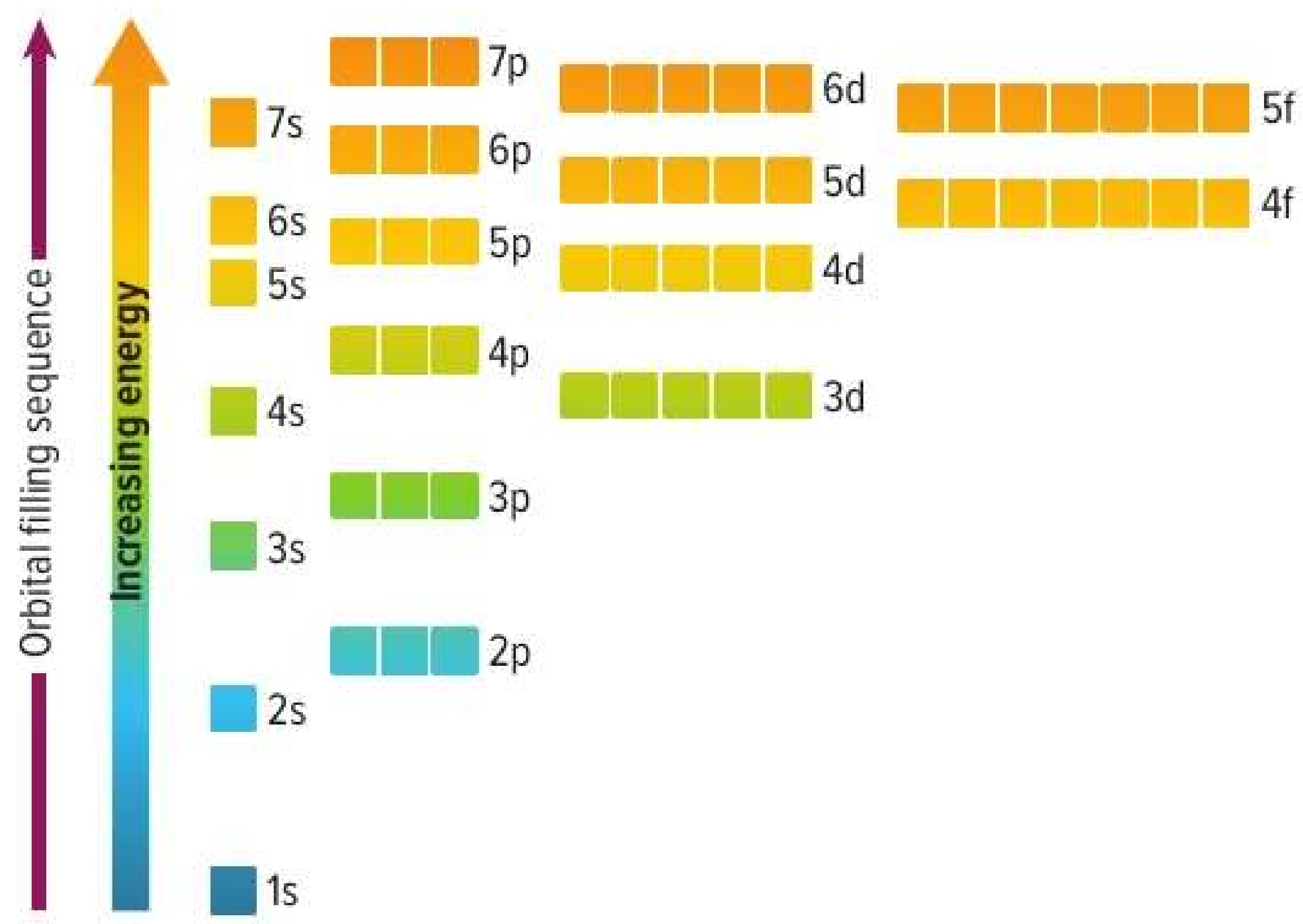


Figure 18 The aufbau diagram shows the energy of each sublevel relative to the energy of other sublevels.

Determine *Which sublevel has the greater energy, 4d or 5p?*

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Virtual Investigation: Electron Configuration

Obtain, evaluate, and communicate information on the patterns present in the periodic table that translate into patterns of electron states.

Laboratory: Electron Charge-to-Mass Ratio

Analyze and interpret data to determine the proportion of charge to mass for an electron.

Table 3 Features of the Aufbau Diagram

Feature	Example
All orbitals related to an energy sublevel are of equal energy.	All three 2p orbitals are of equal energy.
In a multi-electron atom, the energy sublevels within a principal energy level have different energies.	The three 2p orbitals are of higher energy than the 2s orbital.
In order of increasing energy, the sequence of energy sublevels within a principal energy level is s, p, d, and f.	If $n = 4$, then the sequence of energy sublevels is 4s, 4p, 4d, and 4f.
Orbitals related to energy sublevels within one principal energy level can overlap orbitals related to energy sublevels within another principal level.	The orbital related to the atom's 4s sublevel has a lower energy than the five orbitals related to the 3d sublevel.

Table 3 summarizes several features of the aufbau diagram. Although the aufbau principle describes the sequence in which orbitals are filled with electrons, it is important to know that atoms are not built up electron by electron.

The Pauli exclusion principle

Every electron has an associated spin, similar to the way a top spins on its point. Like a top, an electron is able to spin in only one of two directions. The **Pauli exclusion principle,** proposed by Austrian physicist Wolfgang Pauli (1900–1958), states that a maximum of two electrons can occupy a single atomic orbital, but only if the electrons have opposite spins.

Electrons in orbitals can be represented by arrows in boxes. An arrow pointing up $\boxed{\uparrow}$ represents the electron spinning in one direction, and an arrow pointing down $\boxed{\downarrow}$ represents the electron spinning in the opposite direction. An empty box $\boxed{}$ represents an unoccupied orbital, a box containing a single up arrow $\boxed{\uparrow}$ represents an orbital with one electron, and a box containing both up and down arrows $\boxed{\uparrow\downarrow}$ represents a filled orbital containing a pair of electrons with opposite spins.

Hund's rule

The fact that negatively charged electrons repel each other affects the distribution of electrons in equal-energy orbitals. **Hund's rule** states that single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. For example, the boxes below show the sequence in which six electrons occupy the three 2p orbitals. One electron enters each of the orbitals before a second electron enters any of the orbitals.

1. $\boxed{\uparrow}\boxed{}\boxed{}$ **2.** $\boxed{\uparrow}\boxed{\uparrow}\boxed{}$ **3.** $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

4. $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$ **5.** $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$ **6.** $\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$

Get It?

State the three rules that define how electrons are arranged in atoms.

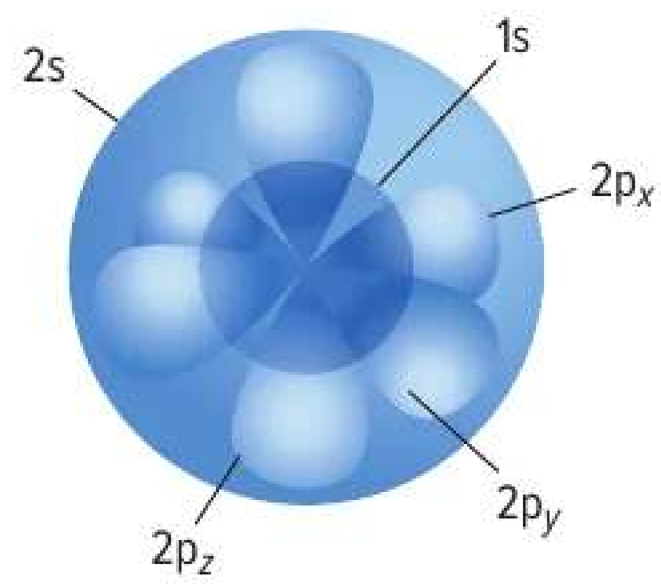
Electron Arrangement

You can represent an atom's electron configuration using one of two convenient methods: orbital diagrams or electron configuration notation.

Orbital diagrams

As mentioned earlier, electrons in orbitals can be represented by arrows in boxes. Each box is labeled with the principal quantum number and sublevel associated with the orbital. For example, the orbital diagram for a ground-state carbon atom, shown below, contains two electrons in the 1s orbital, two electrons in the 2s orbital, and one electron in two of three separate 2p orbitals. The third 2p orbital remains unoccupied.

1s	2s	2p		
↑↓	↑↓	↑	↑	



Neon atom

1s	2s	2p		
↑↓	↑↓	↑↓	↑↓	↑↓

Figure 19 The 1s, 2s, and 2p orbitals of a neon atom overlap.

Determine *how many electrons a neon atom has.*

Electron configuration notation

The electron configuration notation designates the principal energy level and energy sublevel associated with each of the atom's orbitals and includes a superscript representing the number of electrons in the orbital. For example, the electron configuration notation of a ground-state carbon atom is written $1s^22s^22p^2$.

Orbital diagrams and electron configuration notations for the elements in periods one and two of the periodic table are shown in **Table 4. Figure 19** illustrates how the 1s, 2s, $2p_x$, $2p_y$, and $2p_z$ orbitals illustrated earlier in **Figure 17** overlap in the neon atom.

Table 4 Electron Configurations and Orbital Diagrams for Elements 1–10

Element	Atomic Number	Orbital Diagram 1s	2s	$2p_x$	$2p_y$	$2p_z$	Electron Configuration Notation
Hydrogen	1	↑					$1s^1$
Helium	2	↑↓					$1s^2$
Lithium	3	↑↓	↑				$1s^2\ 2s^1$
Beryllium	4	↑↓	↑↓				$1s^2\ 2s^2$
Boron	5	↑↓	↑↓	↑			$1s^2\ 2s^2\ 2p^1$
Carbon	6	↑↓	↑↓	↑	↑		$1s^2\ 2s^2\ 2p^2$
Nitrogen	7	↑↓	↑↓	↑	↑	↑	$1s^2\ 2s^2\ 2p^3$
Oxygen	8	↑↓	↑↓	↑↓	↑	↑	$1s^2\ 2s^2\ 2p^4$
Fluorine	9	↑↓	↑↓	↑↓	↑↓	↑	$1s^2\ 2s^2\ 2p^5$
Neon	10	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2\ 2s^2\ 2p^6$

Note that the electron configuration notation does not usually show the orbital distributions of electrons related to a sublevel. It is understood that a designation such as nitrogen's $2p^3$ represents the orbital occupancy $2p_x^1 2p_y^1 2p_z^1$.

For sodium, the first ten electrons occupy 1s, 2s, and 2p orbitals. Then, according to the aufbau sequence, the eleventh electron occupies the 3s orbital. The electron configuration notation and orbital diagram for sodium are written as follows.

$$1s^2 2s^2 2p^6 3s^1$$

1s	2s	2p			3s
$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow}$

Noble-gas notation Noble gases are the elements in the last column of the periodic table. Their outermost energy levels are full, and they are unusually stable. Noble-gas notation uses bracketed symbols to represent the electron configurations of noble gases. For example, [He] represents the electron configuration for helium, $1s^2$, and [Ne] represents the electron configuration for neon, $1s^2 2s^2 2p^6$.

Compare the electron configuration for neon with sodium's configuration above. Note that the inner-level configuration for sodium is identical to the electron configuration for neon. Using noble-gas notation, sodium's electron configuration can be shortened to the form $[Ne]3s^1$. The electron configuration for an element can be represented using the noble-gas notation for the noble gas in the previous period and the electron configuration for the additional orbitals being filled. The complete and abbreviated (using noble-gas notation) electron configurations of the period 3 elements are shown in **Table 5.**

Table 5 Electron Configurations for Elements 11–18

Element	Atomic Number	Complete Electron Configuration	Electron Configuration Using Noble Gas
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	$[Ne]3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$	$[Ne]3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	$[Ne]3s^2 3p^1$
Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[Ne]3s^2 3p^2$
Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	$[Ne]3s^2 3p^3$
Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne]3s^2 3p^4$
Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[Ne]3s^2 3p^5$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne]3s^2 3p^6$ or [Ar]

Get It?

Explain how to write the noble-gas notation for an element. What is the noble-gas notation for calcium?

SCIENCE USAGE V. COMMON USAGE

period

Science usage: a horizontal row of elements in the current periodic table

There are seven periods in the current periodic table.

Common usage: an interval of time determined by some recurring phenomenon

The period of Earth's orbit is one year.

WORD ORIGIN

aufbau

comes from the German word *aufbauen*, which means to *build up* or *arrange*

Exceptions to predicted configurations

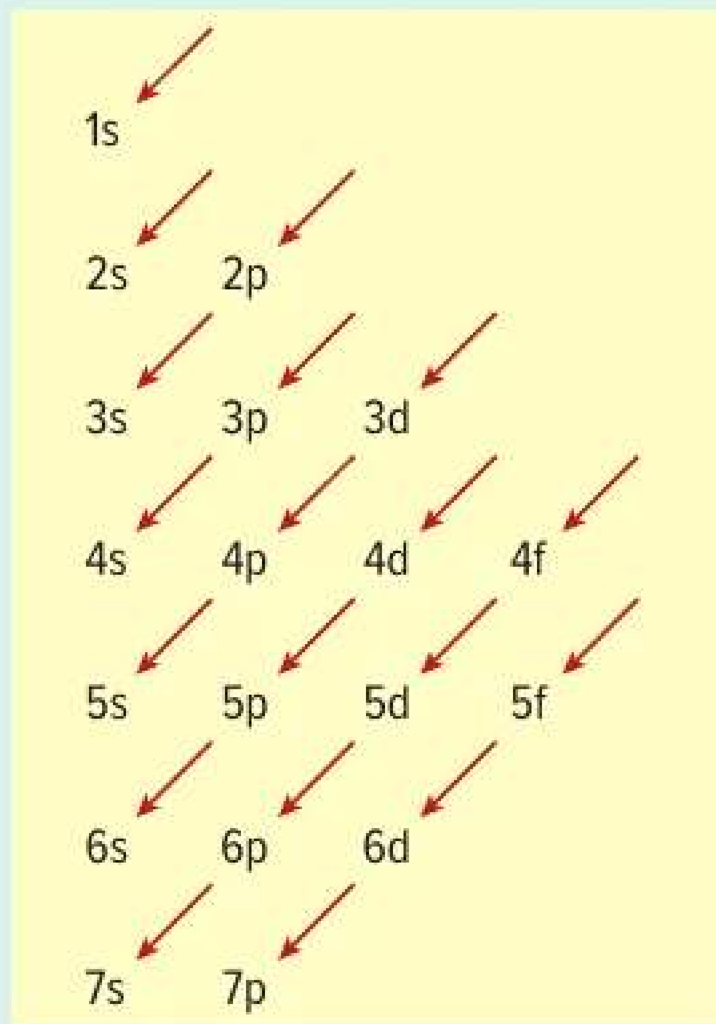
You can use the aufbau diagram to write correct ground-state electron configurations for all elements up to and including vanadium, atomic number 23. However, if you were to proceed in this manner, your configurations for chromium, $[Ar]4s^2 3d^4$, and copper, $[Ar]4s^2 3d^9$, would be incorrect. The correct configurations for these two elements are $[Ar]4s^1 3d^5$ for chromium and $[Ar]4s^1 3d^{10}$ for copper. The electron configurations for these two elements, as well as those of several other elements, illustrate the increased stability of half-filled and filled sets of s and d orbitals.

PROBLEM-SOLVING STRATEGY

Filling Atomic Orbitals

By drawing a sublevel diagram and following the arrows, you can write the ground-state electron configuration for any element.

1. Sketch the sublevel diagram on a blank piece of paper.
2. Determine the number of electrons in one atom of the element for which you are writing the electron configuration. The number of electrons in a neutral atom equals the element's atomic number.
3. Starting with 1s, write the aufbau sequence of atomic orbitals by following the diagonal arrows from the top of the sublevel diagram to the bottom. When you complete one line of arrows, move to the right, to the beginning of the next line of arrows. As you proceed, add superscripts indicating the numbers of electrons in each set of atomic orbitals. Continue only until you have sufficient atomic orbitals to accommodate the total number of electrons in one atom of the element.
4. Apply noble-gas notation.



The sublevel diagram shows the order in which the orbitals are usually filled.

Apply the Strategy

Write the ground-state electron configuration for zirconium.

PRACTICE Problems

ADDITIONAL PRACTICE

19. Write ground-state electron configurations for the following elements.
    - **a.** bromine (Br)
    - **b.** strontium (Sr)
    - **c.** antimony (Sb)
    - **d.** rhenium (Re)
    - **e.** terbium (Tb)
    - **f.** titanium (Ti)
20. A chlorine atom in its ground state has a total of seven electrons in orbitals related to the atom's third energy level. How many of the seven electrons occupy p orbitals? How many of the 17 electrons in a chlorine atom occupy p orbitals?
21. When a sulfur atom reacts with other atoms, electrons in the atom's third energy level are involved. How many such electrons does a sulfur atom have?
22. An element has the ground-state electron configuration $[Kr]5s^2 4d^{10} 5p^1$. It is part of some semiconductors and used in various alloys. What element is it?
23. **CHALLENGE** In its ground state, an atom of an element has two electrons in all orbitals related to the atom's highest energy level for which $n = 6$. Using noble-gas notation, write the electron configuration for this element, and identify the element.

Valence Electrons

Only certain electrons, called valence electrons, determine the chemical properties of an element. **Valence electrons** are defined as electrons in the atom's outermost orbitals—generally those orbitals associated with the atom's highest principal energy level. For example, a sulfur atom contains 16 electrons, only six of which occupy the outermost 3s and 3p orbitals, as shown by sulfur's electron configuration, $[Ne]3s^2 3p^4$. Sulfur has six valence electrons. Similarly, although a cesium atom has 55 electrons, it has just one valence electron, the 6s electron shown in cesium's electron configuration, $[Xe]6s^1$.

Get It?

Cite Evidence How do the properties of electrons influence the properties of elements?

Electron-dot structures

Because valence electrons are involved in forming chemical bonds, chemists often represent them visually using a simple shorthand method, called electron-dot structure. An atom's **electron-dot structure** consists of the element's symbol, which represents the atomic nucleus and inner-level electrons, surrounded by dots representing all of the atom's valence electrons. In writing an atom's electron-dot structure, dots representing valence electrons are placed one at a time on the four sides of the symbol (they may be placed in any sequence) and then paired up until all are shown. **Table 6** shows examples for the second period.

Table 6 Electron Configurations and Dot Structures

Element	Atomic Number	Electron Configuration	Electron-Dot Structure
Lithium	3	$1s^2 2s^1$	Li·
Beryllium	4	$1s^2 2s^2$	·Be·
Boron	5	$1s^2 2s^2 2p^1$	·Ḃ·
Carbon	6	$1s^2 2s^2 2p^2$	·Ċ̣·
Nitrogen	7	$1s^2 2s^2 2p^3$	·N̤̈·
Oxygen	8	$1s^2 2s^2 2p^4$	:Ö̤·
Fluorine	9	$1s^2 2s^2 2p^5$	:F̤̈·
Neon	10	$1s^2 2s^2 2p^6$	:N̤̈e:

EXAMPLE Problem 3

ELECTRON-DOT STRUCTURES Some toothpastes contain stannous fluoride, a compound of tin and fluorine. What is tin's electron-dot structure?

1 ANALYZE THE PROBLEM

Consult the periodic table to determine the total number of electrons in a tin atom. Write out tin's electron configuration, and determine its number of valence electrons. Then use the rules for electron-dot structures to draw the electron-dot structure for tin.

2 SOLVE FOR THE UNKNOWN

Tin has an atomic number of 50. Thus, a tin atom has 50 electrons.

$[Kr]5s^2 4d^{10} 5p^2$ — Write out tin's electron configuration using noble-gas notation. The closest noble gas is Kr.

The two 5s and the two 5p electrons (the electrons in the orbitals related to the atom's highest principal energy level) represent tin's four valence electrons. Draw the four valence electrons around tin's chemical symbol (Sn) to show tin's electron-dot structure.

·Ṩn·

EXAMPLE Problem 3 (continued)

3 EVALUATE THE ANSWER

The correct symbol for tin (Sn) has been used, and the rules for drawing electron-dot structures have been correctly applied.

PRACTICE Problems ADDITIONAL PRACTICE

24. Draw electron-dot structures for atoms of the following elements.

a. magnesium **b.** thallium **c.** xenon

25. An atom of an element has a total of 13 electrons. What is the element, and how many electrons are shown in its electron-dot structure?

26. **CHALLENGE** This element exists in the solid state at room temperature and at normal atmospheric pressure and is found in emerald gemstones. It is known to be one of the following elements: carbon, germanium, sulfur, cesium, beryllium, or argon. Identify the element based on the electron-dot structure at right.

Check Your Progress

Summary

- The arrangement of electrons in an atom is called the atom's electron configuration.
- Electron configurations are defined by the aufbau principle, the Pauli exclusion principle, and Hund's rule.
- An element's valence electrons determine the chemical properties of the element.
- Electron configurations can be represented using orbital diagrams, electron configuration notation, and electron-dot structures.

Demonstrate Understanding

27. **Apply** the Pauli exclusion principle, the aufbau principle, and Hund's rule to write the electron configuration and draw the orbital diagram for each of the following elements.

    a. silicon b. fluorine c. calcium d. krypton

28. **Define** *valence electron*.

29. **Illustrate** and describe the sequence in which ten electrons occupy the five orbitals related to an atom's d sublevel.

30. **Extend** the aufbau sequence through an element that has not yet been identified, but whose atoms would completely fill 7p orbitals. How many electrons would such an atom have? Write its electron configuration using noble-gas notation for the previous noble gas, radon.

31. **Interpret Scientific Illustrations** Which is the correct electron-dot structure for an atom of selenium? Explain.

    a.  b.  c.  d. 

LEARNSMART Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

SCIENTIFIC BREAKTHROUGHS

Batteries of the Future: Super Charged!

Batteries are fundamental to modern technology. Since the invention of the lithium ion battery, researchers have been looking for a better energy source for cars, smart phones, and computers.

Lithium ion batteries are used in electronics.

The New Wave of Battery Power

In most batteries available in devices today, the electrolyte is liquid. However, some researchers are working to develop safer batteries with water-based, air, or solid electrolytes.

The U.S. Army Research Laboratory, in collaboration with the University of Maryland, is shaking things up with new technology that uses a saltwater electrolyte. The researchers say this eliminates the fire and explosion risk associated with some non-aqueous lithium ion batteries, which is especially a concern for military personnel in combat situations. The technology needs to be perfected before it is made commercially available, but so far, it is the first battery of its kind to reach the 4.0 volt mark essential for many electronics.

Lithium-air batteries are another promising possibility for the future. It is projected they could provide three times as much power for a given weight compared with lithium-ion batteries. Researchers across the country are working to determine which electrolyte material, such as lithium iodide, will be most efficient in these batteries to make them cheaper and more powerful.

Solid state batteries use polymer electrolytes to eliminate the liquid electrolyte entirely. This creates a safer, fire-resistant, more powerful, and rechargeable energy source. Solid state batteries are an important advancement to the electric automotive industry, which is currently limited by the range of the best lithium ion batteries.

Given the pervasiveness of battery-powered electronics, new faster, stronger, and safer options could be the catalyst for the next big breakthrough in battery technology.

MAKE AND DEFEND A CLAIM

Research one type of battery described in this feature. Write a report summarizing why you think this battery will or will not be successful in the marketplace.

MODULE 4
STUDY GUIDE

GO ONLINE to study with your Science Notebook

Lesson 1 LIGHT AND QUANTIZED ENERGY

- All waves are defined by their wavelengths, frequencies, amplitudes, and speeds.

$$c = \lambda\nu$$

- In a vacuum, all electromagnetic waves travel at the speed of light.
- All electromagnetic waves have both wave and particle properties.
- Matter emits and absorbs energy in quanta.

$$E_{\text{quantum}} = h\nu$$

- White light produces a continuous spectrum. An element's emission spectrum consists of a series of lines of individual colors.

- electromagnetic radiation
- wavelength
- frequency
- amplitude
- electromagnetic spectrum
- quantum
- Planck's constant
- photoelectric effect
- photon
- atomic emission spectrum

Lesson 2 QUANTUM THEORY AND THE ATOM

- Bohr's atomic model attributes hydrogen's emission spectrum to electrons dropping from higher-energy to lower-energy orbits.

$$\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}} = E_{\text{photon}} = h\nu$$

- The de Broglie equation relates a particle's wavelength to its mass, its velocity, and Planck's constant.

$$\lambda = h/m\nu$$

- The quantum mechanical model assumes that electrons have wave properties.
- Electrons occupy three-dimensional regions of space called atomic orbitals.

- ground state
- quantum number
- de Broglie equation
- Heisenberg uncertainty principle
- quantum mechanical model of the atom
- atomic orbital
- principal quantum number
- principal energy level
- energy sublevel

Lesson 3 ELECTRON CONFIGURATION

- The arrangement of electrons in an atom is called the atom's electron configuration.
- Electron configurations are defined by the aufbau principle, the Pauli exclusion principle, and Hund's rule.
- An element's valence electrons determine the chemical properties of the element.
- Electron configurations can be represented using orbital diagrams, electron configuration notation, and electron-dot structures.

- electron configuration
- aufbau principle
- Pauli exclusion principle
- Hund's rule
- valence electron
- electron-dot structure

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

How do we know what stars are made of?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

SEP Data Analysis Lab

What electron transitions account for the Balmer series?

Hydrogen's emission spectrum comprises three series of lines. Some wavelengths are ultraviolet (Lyman series) and infrared (Paschen series). Visible wavelengths comprise the Balmer series. The Bohr atomic model attributes these spectral lines to transitions from higher-energy states with electron orbits in which $n = n_i$ to lower-energy states with smaller electron orbits in which $n = n_f$.

CER Analyze and Interpret Data

Some hydrogen balmer lines are designated H_α (6562 Å), H_β (4861 Å), H_γ (4340 Å), and H_δ (4101 Å). Each wavelength (λ) is related to an electron transition within a hydrogen atom by the following equation, in which $1.09678 \times 10^7\ m^{-1}$ is known as the Rydberg constant.

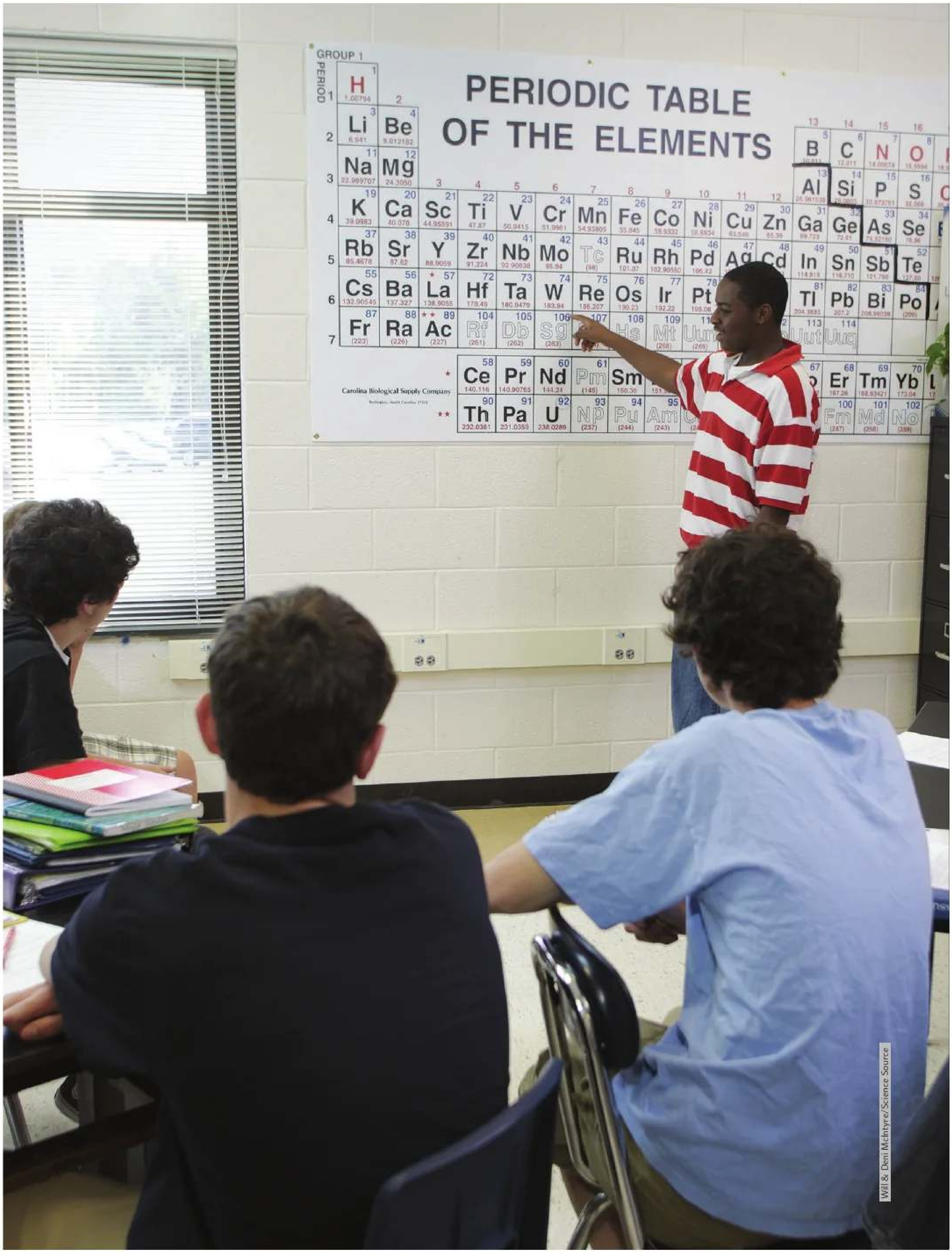
$$\frac{1}{\lambda} = 1.09678 \times 10^7 \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)\ m^{-1}$$

For hydrogen's Balmer series, electron orbit transitions occur from larger orbits to the $n = 2$ orbit; that is, $n_f = 2$.

CER Analyze and Interpret Data

1. **Calculate the wavelengths for the following electron orbit transitions.**
   a. $n_i = 3; n_f = 2$ **c.** $n_i = 5; n_f = 2$
   b. $n_i = 4; n_f = 2$ **d.** $n_i = 6; n_f = 2$
2. **Claim, Evidence, Reasoning** Relate the Balmer-series wavelengths you calculated in Question 1 to those determined experimentally. Allowing for experimental error and calculation uncertainty, do the wavelengths match? Explain your answer. One angstrom (Å) equals 10^{-10} m.
3. **Apply** the formula $E = hc/\lambda$ to determine the energy per quantum for each of the orbit transitions in Question 1.

Eric Lowenbach/Moment/Getty Images



Will & Deni McIntyre/Science Source

MODULE 5

THE PERIODIC TABLE AND PERIODIC LAW

ENCOUNTER THE PHENOMENON

What can we learn from the periodic table?

GO ONLINE to play a video about Mendeleev's approach to organizing the elements.

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

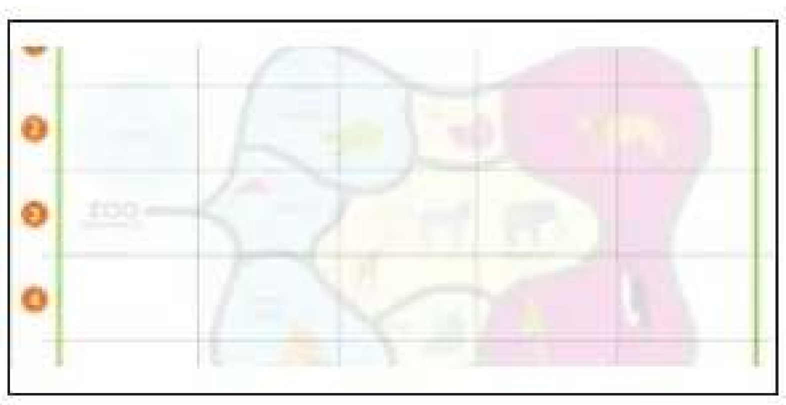
CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about what we can learn from the periodic table.

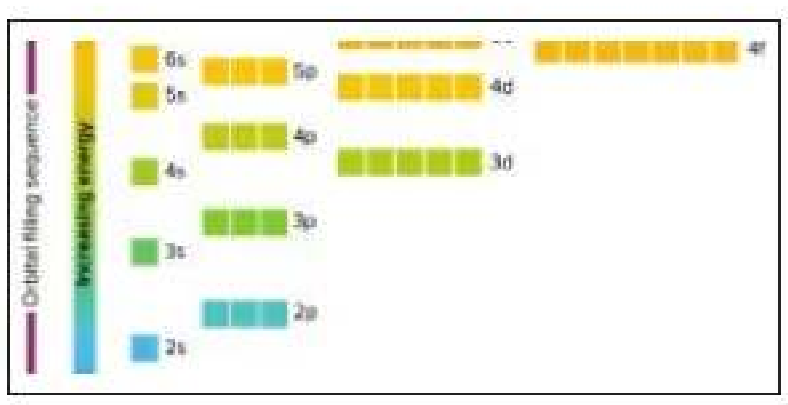
Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.



LESSON 1: Explore & Explain: Groups and Periods



LESSON 2: Explore & Explain: Electron Configuration and the Periodic Table

Smithsonian

Additional Resources

LESSON 1
DEVELOPMENT OF THE MODERN PERIODIC TABLE

FOCUS QUESTION

How are elements organized in the periodic table?

Development of the Periodic Table

In the late 1700s, French scientist Antoine Lavoisier (1743–1794) compiled a list of all elements that were known at the time. The list, shown in **Table 1**, contained 33 elements organized in four categories. Many of these elements, such as silver, gold, carbon, and oxygen, have been known since prehistoric times.

The 1800s brought a large increase in the number of known elements. The advent of electricity, which was used to break down compounds into their components, and the development of the spectrometer, which was used to identify the newly isolated elements, played major roles in the advancement of chemistry. The industrial revolution of the mid-1800s also played a major role, which led to the development of many new chemistry-based industries, such as the manufacture of petrochemicals, soaps, dyes, and fertilizers. By 1870, there were over 60 known elements.

Along with the discovery of new elements came volumes of new scientific data related to the elements and their compounds. Chemists of the time were overwhelmed with learning the properties of so many new elements and compounds. What chemists needed was a tool for organizing the many facts associated with the elements.

Table 1 Lavoisier's Table of Simple Substances (Old English Names)

Gases	light, heat, dephlogisticated air, phlogisticated gas, inflammable air
Metals	antimony, silver, arsenic, bismuth, cobalt, copper, tin, iron, manganese, mercury, molybdena, nickel, gold, platina, lead, tungsten, zinc
Nonmetals	sulphur, phosphorus, pure charcoal, radical muriatique*, radical fluorique*, radical boracique*
Earths	chalk, magnesia, barote, clay, siliceous earth

*no English name

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Virtual Investigation: Periodic Properties
Analyze and interpret the data in the periodic table of elements for patterns of organization and the properties of matter.

ChemLAB: Investigate Descriptive Chemistry
Analyze and interpret data to determine patterns of properties in representative elements.

Organizing the elements

A significant step toward developing a tool for organizing the elements and the large amount of data about their properties came in 1860, when chemists agreed upon a method for accurately determining the atomic masses of the elements. Until this time, different chemists used different mass values in their work, making the results of one chemist's work hard to reproduce by another.

With newly agreed-upon atomic masses for the elements, the search for relationships between atomic mass and elemental properties, and a way to organize the elements, began in earnest.

John Newlands

In 1864, English chemist John Newlands (1837–1898) proposed an organizational scheme for the elements. He noticed that when the elements were arranged by increasing atomic mass, their properties repeated every eighth element. A pattern such as this is called periodic because it repeats in a specific manner. Newlands named the periodic relationship that he observed in chemical properties the law of octaves, after the musical octave in which notes repeat every eighth tone.

Figure 1 shows how Newlands organized 14 of the elements known in the mid-1860s. Acceptance of the law of octaves was hampered because the law did not work for all of the known elements. Also, the use of the word octave was harshly criticized by fellow scientists, who thought that the musical analogy was unscientific. While his law was not generally accepted, the passage of a few years would show that Newlands was basically correct: the properties of elements do repeat in a periodic way.

Meyer and Mendeleev

In 1869, German chemist Lothar Meyer (1830–1895) and Russian chemist Dmitri Mendeleev (1834–1907) each demonstrated a connection between atomic mass and the properties of elements. Mendeleev, however, is generally given more credit than Meyer because he published his organizational scheme first.

Like Newlands several years earlier, Mendeleev noticed that when the elements were ordered by increasing atomic mass, there was a periodic pattern in their properties.

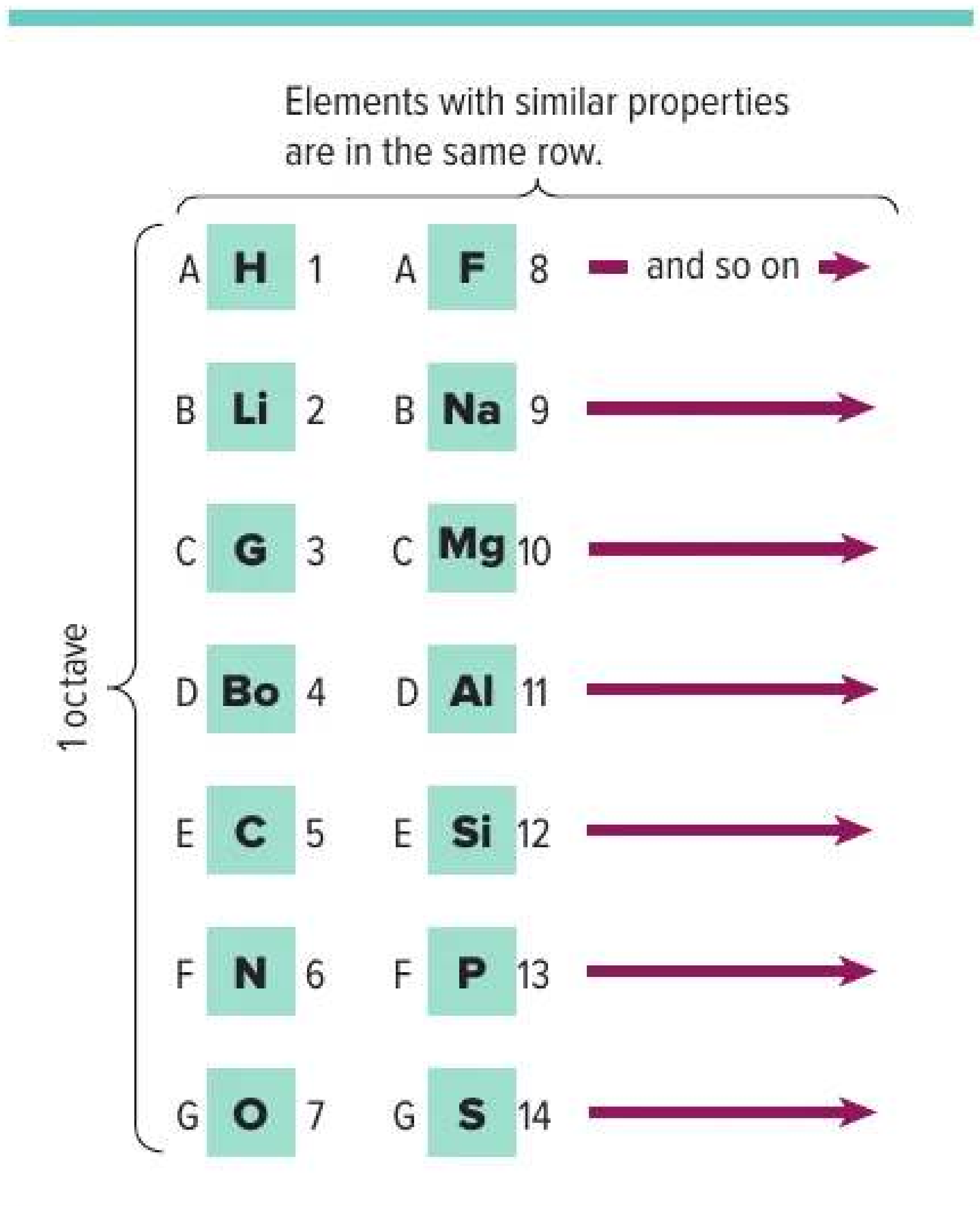
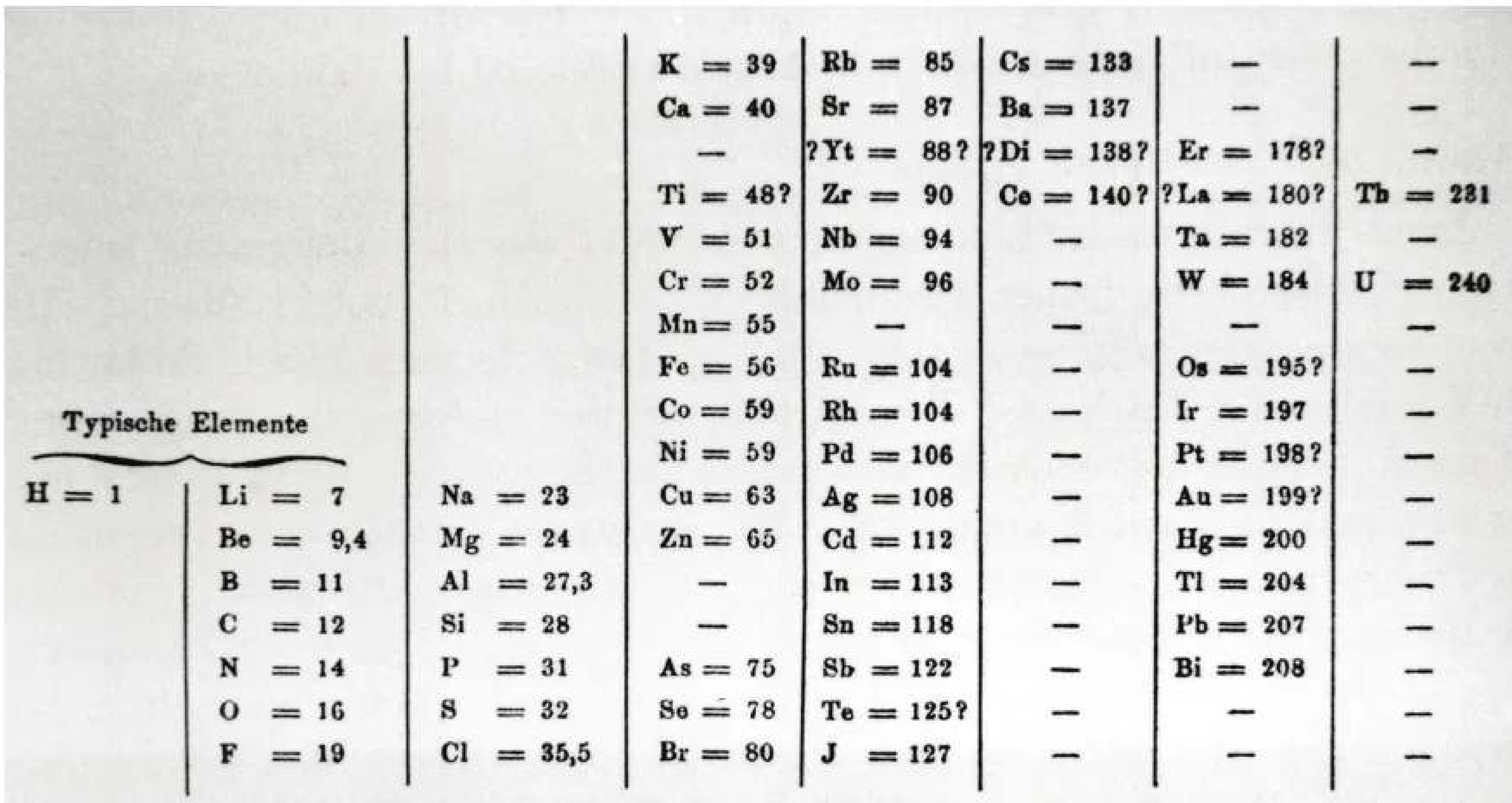


Figure 1 John Newlands noticed that the properties of elements repeated every eighth element in the same way musical notes repeat every eighth note and form octaves.

Get It?
Describe the pattern that both Newlands and Mendeleev noticed about the properties of the elements.

By arranging the elements in order of increasing atomic mass into columns with similar properties, Mendeleev organized the elements into a periodic table.

Mendeleev's table, shown in **Figure 2**, became widely accepted because he predicted the existence and properties of undiscovered elements that were later found. Mendeleev left blank spaces in the table where he thought the undiscovered elements should go. By noting trends in the properties of known elements, he was able to predict the properties of the yet-to-be-discovered elements scandium, gallium, and germanium.



Typische Elemente

			K = 39	Rb = 85	Cs = 133	—	—
			Ca = 40	Sr = 87	Ba = 137	—	—
			—	?Yt = 88?	?Di = 138?	Er = 178?	—
			Ti = 48?	Zr = 90	Ce = 140?	?La = 180?	Th = 231
			V = 51	Nb = 94	—	Ta = 182	—
			Cr = 52	Mo = 96	—	W = 184	U = 240
			Mn = 55	—	—	—	—
			Fe = 56	Ru = 104	—	Os = 195?	—
			Co = 59	Rh = 104	—	Ir = 197	—
			Ni = 59	Pd = 106	—	Pt = 198?	—
H = 1	Li = 7	Na = 23	Cu = 63	Ag = 108	—	Au = 199?	—
	Be = 9,4	Mg = 24	Zn = 65	Cd = 112	—	Hg = 200	—
	B = 11	Al = 27,3	—	In = 113	—	Tl = 204	—
	C = 12	Si = 28	—	Sn = 118	—	Pb = 207	—
	N = 14	P = 31	As = 75	Sb = 122	—	Bi = 208	—
	O = 16	S = 32	Se = 78	Te = 125?	—	—	—
	F = 19	Cl = 35,5	Br = 80	J = 127	—	—	—

Figure 2 In the first version of his table, published in 1869, Mendeleev arranged elements with similar chemical properties horizontally. He left empty spaces for elements that were not yet discovered.

Moseley

Mendeleev's table, however, was not completely correct. After several new elements were discovered and the atomic masses of the known elements were more accurately determined, it became apparent that several elements in his table were not in the correct order. Arranging the elements by mass resulted in several elements being placed in groups of elements with differing properties. The reason for this problem was determined in 1913 by English chemist Henry Moseley (1887–1915). Moseley discovered that atoms of each element contain a unique number of protons in their nuclei—the number of protons being equal to the atom's atomic number. By arranging the elements in order of increasing atomic number, the problems with the order of the elements in the periodic table were solved and a clear periodic pattern of properties resulted.

The statement that there is a periodic repetition of chemical and physical properties of the elements when they are arranged by increasing atomic number is called the **periodic law.**

Get It?

Compare and contrast the ways in which Mendeleev and Moseley organized the elements.

Table 2 summarizes the contributions of Newlands, Meyer, Mendeleev, and Moseley to the development of the periodic table.

The periodic table brought order to seemingly unrelated facts and became a significant tool for chemists. It is a useful reference for understanding and predicting the properties of elements and for organizing knowledge of atomic structure.

Table 2 Contributions to the Classification of Elements

John Newlands (1837–1898)
• arranged elements by increasing atomic mass • noticed the repetition of properties every eighth element • created the law of octaves
Lothar Meyer (1830–1895)
• demonstrated a connection between atomic mass and elements' properties • arranged the elements in order of increasing atomic mass
Dmitri Mendeleev (1834–1907)
• demonstrated a connection between atomic mass and elements' properties • arranged the elements in order of increasing atomic mass • predicted the existence and properties of undiscovered elements
Henry Moseley (1887–1915)
• discovered that atoms contain a unique number of protons called the atomic number • arranged elements in order of increasing atomic number, which resulted in a periodic pattern of properties

The Modern Periodic Table

The modern periodic table consists of boxes, each containing an element name, symbol, atomic number, and atomic mass. A typical box from the table is shown in **Figure 3**.

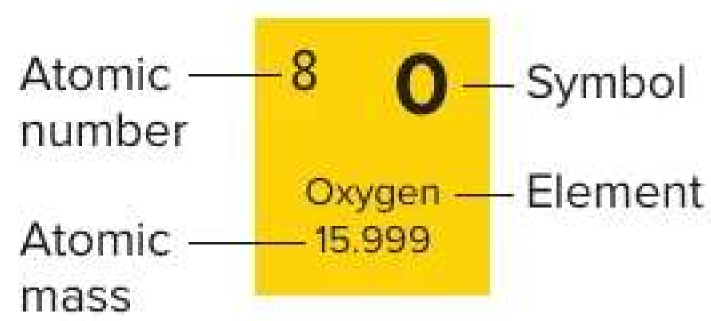


Figure 3 A typical box from the periodic table contains the element's name, its chemical symbol, its atomic number, and its atomic mass.

The table orders elements horizontally by the number of protons in an atom's nucleus, and places those with similar chemical properties in columns. The columns are called **groups** or families. The rows are called **periods.**

The periodic table is shown in **Figure 4** on the next page and on the inside back cover of your textbook. Becoming familiar with the periodic table will help you understand how the properties of different elements relate to one another.

CCC CROSSCUTTING CONCEPTS

Patterns Different patterns can be observed in the periodic table. The patterns organize and can predict the properties of elements. Compare and contrast the periods and groups of the table, shown in **Figure 4,** based on their atomic number and atomic mass. Create a graphic organizer or other simple visual that will help you and your classmates remember the patterns.

WORD ORIGIN

periodic

comes from the Greek word *periodos*, meaning *way around*, *circuit*

Figure 4 **Periodic Table of the Elements**

Atomic number: 1 — Symbol: H — Element: Hydrogen — Atomic mass: 1.008

Metal
Metalloid
Nonmetal
Synthetic

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H Hydrogen 1.008																	2 He Helium 4.003
2	3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
3	11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.066	17 Cl Chlorine 35.453	18 Ar Argon 39.948
4	19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.847	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.39	31 Ga Gallium 69.723	32 Ge Germanium 72.61	33 As Arsenic 74.922	34 Se Selenium 78.971	35 Br Bromine 79.904	36 Kr Krypton 83.80
5	37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium (98)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.411	49 In Indium 114.82	50 Sn Tin 118.710	51 Sb Antimony 121.757	52 Te Tellurium 127.60	53 I Iodine 126.904	54 Xe Xenon 131.290
6	55 Cs Cesium 132.905	56 Ba Barium 137.327	57 La Lanthanum 138.905	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.217	78 Pt Platinum 195.08	79 Au Gold 196.967	80 Hg Mercury 200.59	81 Tl Thallium 204.383	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium 208.982	85 At Astatine 209.987	86 Rn Radon 222.018
7	87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium * (267)	105 Db Dubnium * (270)	106 Sg Seaborgium * (269)	107 Bh Bohrium * (270)	108 Hs Hassium * (277)	109 Mt Meitnerium * (278)	110 Ds Darmstadtium * (281)	111 Rg Roentgenium * (281)	112 Cn Copernicium * (285)	113 Nh Nihonium * (286)	114 Fl Flerovium * (289)	115 Mc Moscovium * (289)	116 Lv Livermorium * (293)	117 Ts Tennessine * (294)	118 Og Oganesson * (294)

The number in parentheses is the mass number of the longest-lived isotope for that element.

*Properties are largely predicted.

Lanthanide series	58 Ce Cerium 140.115	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.242	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.965	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.50	67 Ho Holmium 164.930	68 Er Erbium 167.259	69 Tm Thulium 168.934	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.967
Actinide series	90 Th Thorium 232.038	91 Pa Protactinium 231.036	92 U Uranium 238.029	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium * (252)	100 Fm Fermium * (257)	101 Md Mendelevium * (258)	102 No Nobelium * (259)	103 Lr Lawrencium * (262)

Groups and periods

Beginning with hydrogen in period 1, there are a total of seven periods. Each group is numbered 1 through 18. For example, period 4 contains potassium and calcium. Oxygen is in group 16. The elements in groups 1, 2, and 13 to 18 possess a wide range of chemical and physical properties. For this reason, they are often referred to as the main group, or **representative elements.** The elements in groups 3 to 12 are referred to as the **transition elements.** Elements are also classified as metals, nonmetals, and metalloids.

Metals

Elements that are generally shiny when smooth and clean, solid at room temperature, and good conductors of heat and electricity are called **metals.** Most metals are also malleable and ductile, meaning that they can be pounded into thin sheets and drawn into wires, respectively, as shown in **Figure 5**.

Most representative elements and all transition elements are metals. If you look at boron (B) in column 13, you will see a heavy stairstep line that zigzags down to astatine (At) at the bottom of group 17. This stairstep line is a visual divider between the metals and the nonmetals on the table. In the periodic table shown in **Figure 4** metals are represented by the blue boxes.

Alkali Metals Except for hydrogen, all of the elements on the left side of the table are metals. The group 1 elements (except for hydrogen) are known as the **alkali metals.** Because they are so reactive, alkali metals usually exist as compounds with other elements. Two familiar alkali metals are sodium (Na), one of the components of salt, and lithium (Li), often used in batteries.

Figure 5 Copper, like most metals, is ductile and conducts electricity well. For these reasons copper is used for electrical wiring.

SCIENCE USAGE V. COMMON USAGE

conductor

Science usage: a substance or body capable of transmitting electricity, heat, or sound

Copper is a good conductor of heat.

Common usage: a person who conducts an orchestra, chorus, or other group of musical performers

The new conductor helped the orchestra perform at its best.

Alkaline Earth Metals The **alkaline earth metals** are in group 2. They are also highly reactive. Calcium (Ca) and magnesium (Mg), two minerals important for your health, are examples of alkaline earth metals. Because magnesium is strong and relatively light, it is used in applications in which strength and low mass are important, as shown in **Figure 6**.

Figure 6 Because magnesium is light and strong, it is often used in the production of safety devices such as these caribiners used by climbers.

Transition and Inner Transition Metals The transition elements are divided into **transition metals** and **inner transition metals.** The two sets of inner transition metals, known as the **lanthanide series** and **actinide series,** are located along the bottom of the periodic table. The rest of the elements in groups 3 to 12 make up the transition metals. Elements from the lanthanide series are used extensively as phosphors, substances that emit light when struck by electrons. Because it is strong and light, the transition metal titanium is used to make frames for bicycles and eyeglasses.

Nonmetals

BIOLOGY Connection Nonmetals occupy the upper-right side of the periodic table. They are represented by the yellow boxes in **Figure 4**. **Nonmetals** are elements that are generally gases or brittle, dull-looking solids. They are poor conductors of heat and electricity. The only nonmetal that is a liquid at room temperature is bromine (Br). The most abundant element in the human body is the nonmetal oxygen, which constitutes 65% of the body mass.

Group 17 comprises highly reactive elements that are known as **halogens.** Like the group 1 and group 2 elements, the halogens are often part of compounds. Compounds made with the halogen fluorine (F) are commonly added to toothpaste and drinking water to prevent tooth decay. The extremely unreactive group 18 elements are commonly called the **noble gases.** They are used in applications where their unreactivity is an advantage, such as in lasers, a variety of light bulbs, and neon signs.

Metalloids

The elements in the green boxes bordering the stairstep line in **Figure 4** are called metalloids, or semimetals.

Metalloids have physical and chemical properties of both metals and nonmetals. Silicon (Si) and germanium (Ge) are two important metalloids used extensively in computer chips and solar cells. Silicon is also used to make prosthetics or in lifelike applications, as shown in **Figure 7**.

Figure 7 Scientists developing submarine technology created this robot that looks and swims like a real fish. Its body is made of a silicon resin that softens in water.

This introduction to the periodic table touches only the surface of its durable explanatory power. You can refer to the Elements Handbook at the end of your textbook to learn more about the elements and their various groups.

Check Your Progress

Summary

- The elements were first organized by increasing atomic mass, which led to inconsistencies. Later, they were organized by increasing atomic number.
- The periodic law states that when the elements are arranged by increasing atomic number, there is a periodic repetition of their chemical and physical properties.
- The periodic table organizes the elements into periods (rows) and groups or families (columns); elements with similar properties are in the same group.
- Elements are classified as metals, nonmetals, or metalloids.

Demonstrate Understanding

1. **Describe** the development of the periodic table.
2. **Sketch** a simplified version of the periodic table, and indicate the location of metals, nonmetals, and metalloids.
3. **Describe** the general characteristics of metals, nonmetals, and metalloids.
4. **Identify** each of the following as a representative element or a transition element.
   a. lithium (Li)
   b. platinum (Pt)
   c. promethium (Pm)
   d. carbon (C)
5. **Compare** For each of the given elements, list two other elements with similar chemical properties.
   a. iodine (I) b. barium (Ba) c. iron (Fe)
6. **Compare** According to the periodic table, which two elements have an atomic mass less than twice their atomic number?
7. **Interpret Data** A company plans to make an electronic device. They need to use an element that has chemical behavior similar to that of silicon (Si) and lead (Pb). The element must have an atomic mass greater than that of sulfur (S) but less than that of cadmium (Cd). Use the periodic table to predict which element the company will use.

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LESSON 2
CLASSIFICATION OF THE ELEMENTS

FOCUS QUESTION

Why do elements in the same group have similar properties?

Organizing the Elements by Electron Configuration

As you learned previously, electron configuration determines the chemical properties of an element. Writing out electron configurations using the aufbau diagram can be tedious. Fortunately, you can determine an atom's electron configuration and its number of valence (outermost) electrons from its position on the periodic table. The repeating patterns of the table reflect patterns of outer electron states. The electron configurations for some of the group 1 elements are listed in **Table 3**. All four configurations have a single electron in their outermost orbitals.

Valence electrons

Recall that electrons in the highest principal energy level of an atom are called valence electrons. Each of the group 1 elements has one valence electron. The group 1 elements have similar chemical properties because they have the same number of valence electrons. This is one of the most important relationships in chemistry: atoms in the same group have similar chemical properties because they have the same number of valence electrons. Each group 1 element has a valence electron configuration of s^1. Each group 2 element has a valence electron configuration of s^2. Groups 1, 2, and 13 to 18 all have their own valence electron configurations.

Table 3 Electron Configuration for the Group 1 Elements

Period 1	hydrogen	$1s^1$	$1s^1$
Period 2	lithium	$1s2s^1$	$[He]2s^1$
Period 3	sodium	$1s^22s^22p^63s^1$	$[Ne]3s^1$
Period 4	potassium	$1s^22s^22p^63s^23p64s^1$	$[Ar]4s^1$

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Revisit the Encounter the Phenomenon Question

What information from this lesson can help you answer the module question?

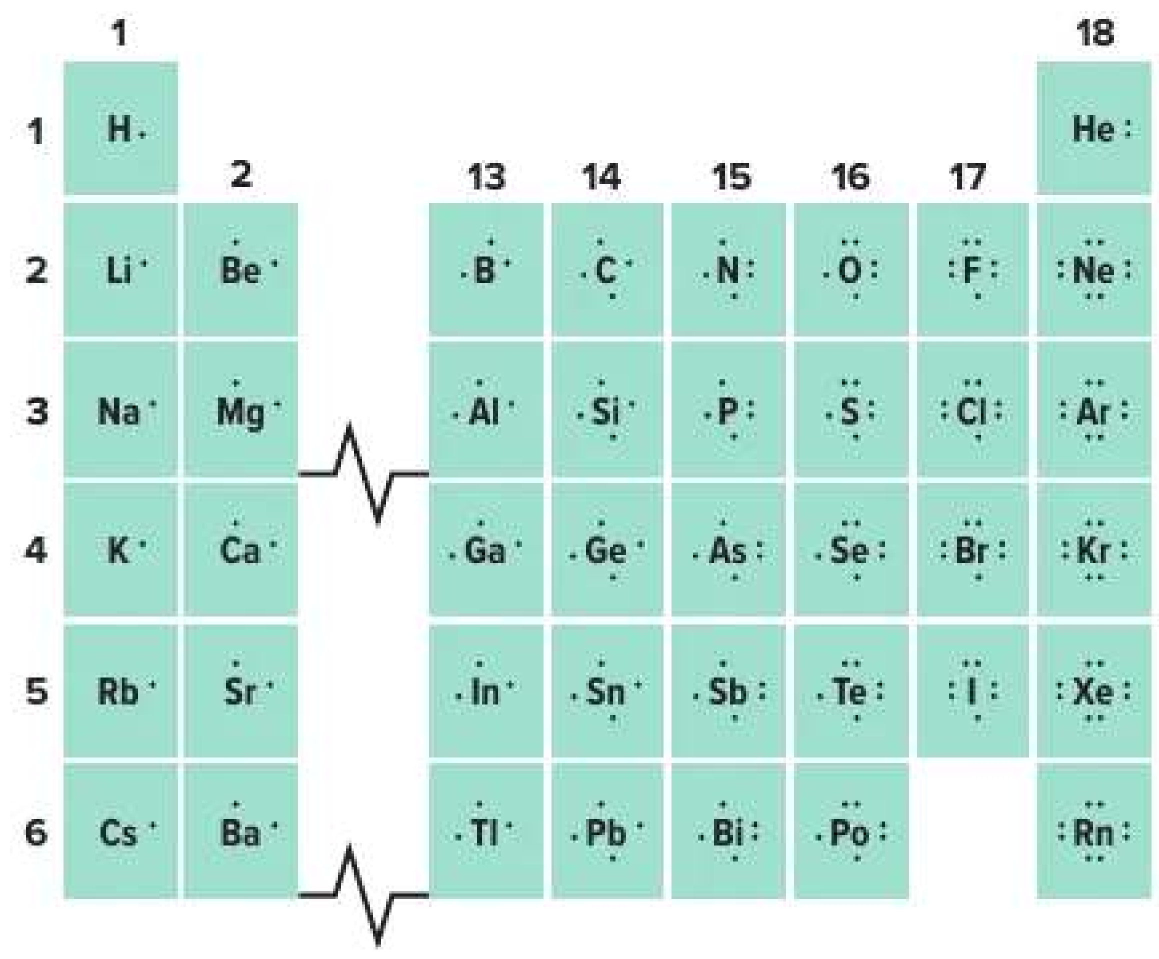

Figure 8 The figure shows the electron-dot structure of most representative elements.

Observe *How does the number of valence electrons vary within a group?*

Valence electrons and period

The energy level of an element's valence electrons indicates the period on the periodic table in which it is found. For example, lithium's valence electron is in the second energy level and lithium is found in period 2. Now look at gallium, with its electron configuration of $[Ar]4s^23d^{10}4p^1$. Gallium's valence electrons are in the fourth energy level, and gallium is found in the fourth period.

Valence electrons of the representative elements

Elements in group 1 have one valence electron; group 2 elements have two valence electrons. Group 13 elements have three valence electrons, group 14 elements have four, and so on. The noble gases in group 18 each have eight valence electrons, with the exception of helium, which has only two valence electrons. **Figure 8** shows how the electron-dot structures you studied previously illustrate the connection between group number and number of valence electrons. Notice that the number of valence electrons for the elements in groups 13 to 18 is ten less than their group number.

The s-, p-, d-, and f-Block Elements

The periodic table has columns and rows of varying sizes. The reason behind the table's odd shape becomes clear if it is divided into sections, or blocks, representing the atom's energy sublevel being filled with valence electrons. Because there are four different energy sublevels (s, p, d, and f), the periodic table is divided into four distinct blocks, as shown in **Figure 9** on the next page.

s-Block elements

The s-block consists of groups 1 and 2, and the element helium. Group 1 elements have partially filled s orbitals containing one valence electron and electron configurations ending in s^1. Group 2 elements have completely filled s orbitals containing two valence electrons and electron configurations ending in s^2. Because s orbitals hold two electrons at most, the s-block spans two groups.

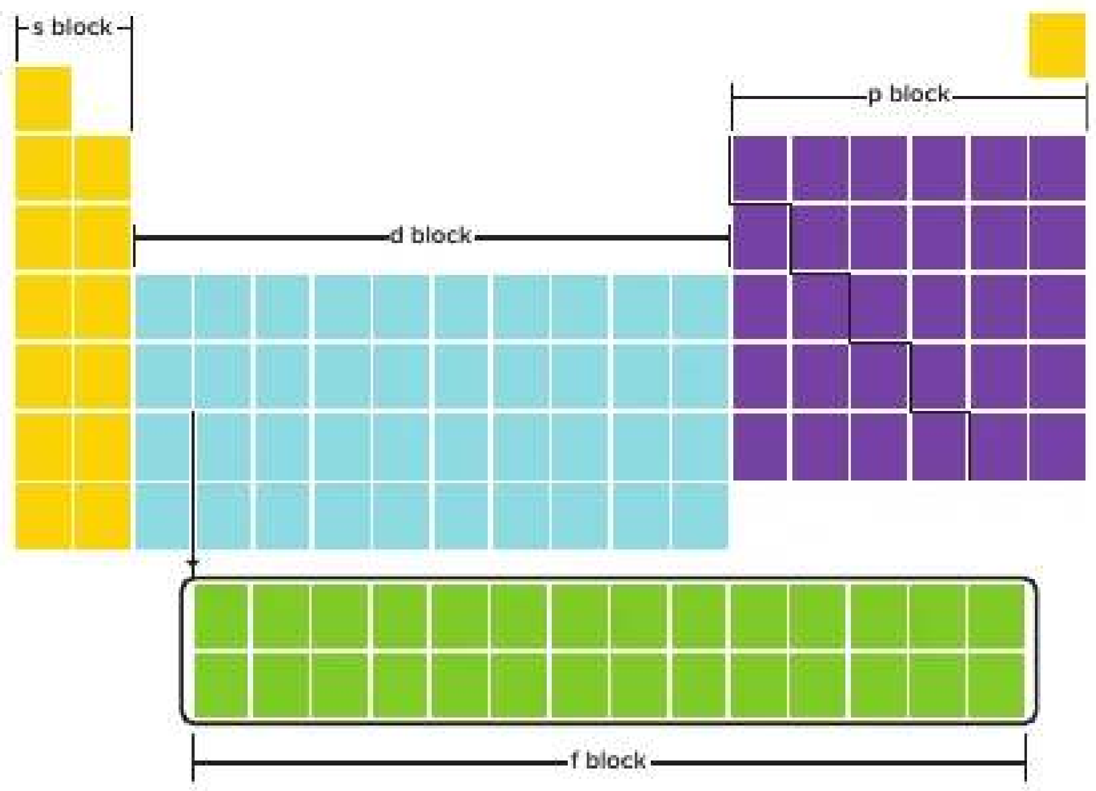

Figure 9 The periodic table is divided into four blocks—s, p, d, and f.

Analyze *What is the relationship between the maximum number of electrons an energy sublevel can hold and the number of columns in that block on the diagram?*

p-Block elements

After the s sublevel is filled, the valence electrons next occupy the p sublevel. The p-block is comprised of groups 13 through 18 and contains elements with filled or partially filled p orbitals. There are no p-block elements in period 1 because the p sublevel does not exist for the first principal energy level (n = 1). The first p-block element is boron (B), which is in the second period. The p-block spans six groups because the three p orbitals can hold a maximum of six electrons.

The group 18 elements, which are called the noble gases, are unique members of the p-block. The atoms of these elements are so stable that they undergo virtually no chemical reactions. The electron configurations of the first four noble gas elements are shown in **Table 4**. Here, both the s and p orbitals corresponding to the period's principal energy level are completely filled. This arrangement of electrons results in an unusually stable atomic structure. Together, the s- and p-blocks comprise the representative elements.

d-Block elements

The d-block contains the transition metals and is the largest of the blocks. With some exceptions, d-block elements are characterized by a filled outermost s orbital of energy level n, and filled or partially filled d orbitals of energy level $n - 1$.

Table 4 **Noble Gas Electron Configuration**

Period	Principal Energy Level	Element	Electron Configuration
1	$n = 1$	helium	$1s^2$
2	$n = 2$	neon	$[He]2s2^2p^6$
3	$n = 3$	argon	$[Ne]3s^23p^6$
4	$n = 4$	krypton	$[Ar]4s^23d^{10}4p^6$

ACADEMIC VOCABULARY

structure

something made up of more-or-less interdependent elements or parts

Many scientists were involved in the discovery of the structure of the atom.

As you move across a period, electrons fill the d orbitals. For example, scandium (Sc), the first d-block element, has an electron configuration of $[Ar]4s^2 3d^1$. Titanium (Ti), the next element on the table, has an electron configuration of $[Ar]4s^2 3d^2$. Note that titanium's filled outermost s orbital has an energy level of $n = 4$, while the d orbital, which is partially filled, has an energy level of $n = 3$.

As you learned previously, the aufbau principle states that the 4s orbital has a lower energy level than the 3d orbital. Therefore, the 4s orbital is filled before the 3d orbital. The five d orbitals can hold a total of 10 electrons; thus, the d-block spans 10 groups on the periodic table.

f-Block elements

The f-block contains the inner transition metals. Its elements are characterized by a filled, or partially filled outermost s orbital, and filled or partially filled 4f and 5f orbitals.

The electrons of the f sublevel do not fill their orbitals in a predictable manner. Because there are seven f orbitals holding up to a maximum of 14 electrons, the f-block spans 14 columns of the periodic table.

EXAMPLE Problem 1

ELECTRON CONFIGURATION AND THE PERIODIC TABLE Strontium, which is used to produce red fireworks, has an electron configuration of $[Kr]5s^2$. Without using the periodic table, determine the group, period, and block of strontium.

1 ANALYZE THE PROBLEM

You are given the electron configuration of strontium.

Known	Unknown
Electron configuration = $[Kr]5s^2$	Group = ?
	Period = ?
	Block = ?

2 SOLVE FOR THE UNKNOWN

The s^2 indicates that strontium's valence electrons fill the s sublevel. Thus, strontium is in **group 2** of the **s-block**.

For representative elements, the number of valence electrons can indicate the group number.

The 5 in $5s^2$ indicates that strontium is in **period 5**.

The number of the highest energy level indicates the period number.

3 EVALUATE THE ANSWER

The relationships between electron configuration and position on the periodic table have been correctly applied.

STEM CAREER Connection

Physical Chemist

Does the idea of using computers and sophisticated laboratory instruments to model, simulate, and analyze experimental results appeal to you? Are you someone who enjoys developing new theories? Physical chemists are interested in how matter behaves at the molecular and atomic level. They have a strong interest and background in chemistry, physics, and math.

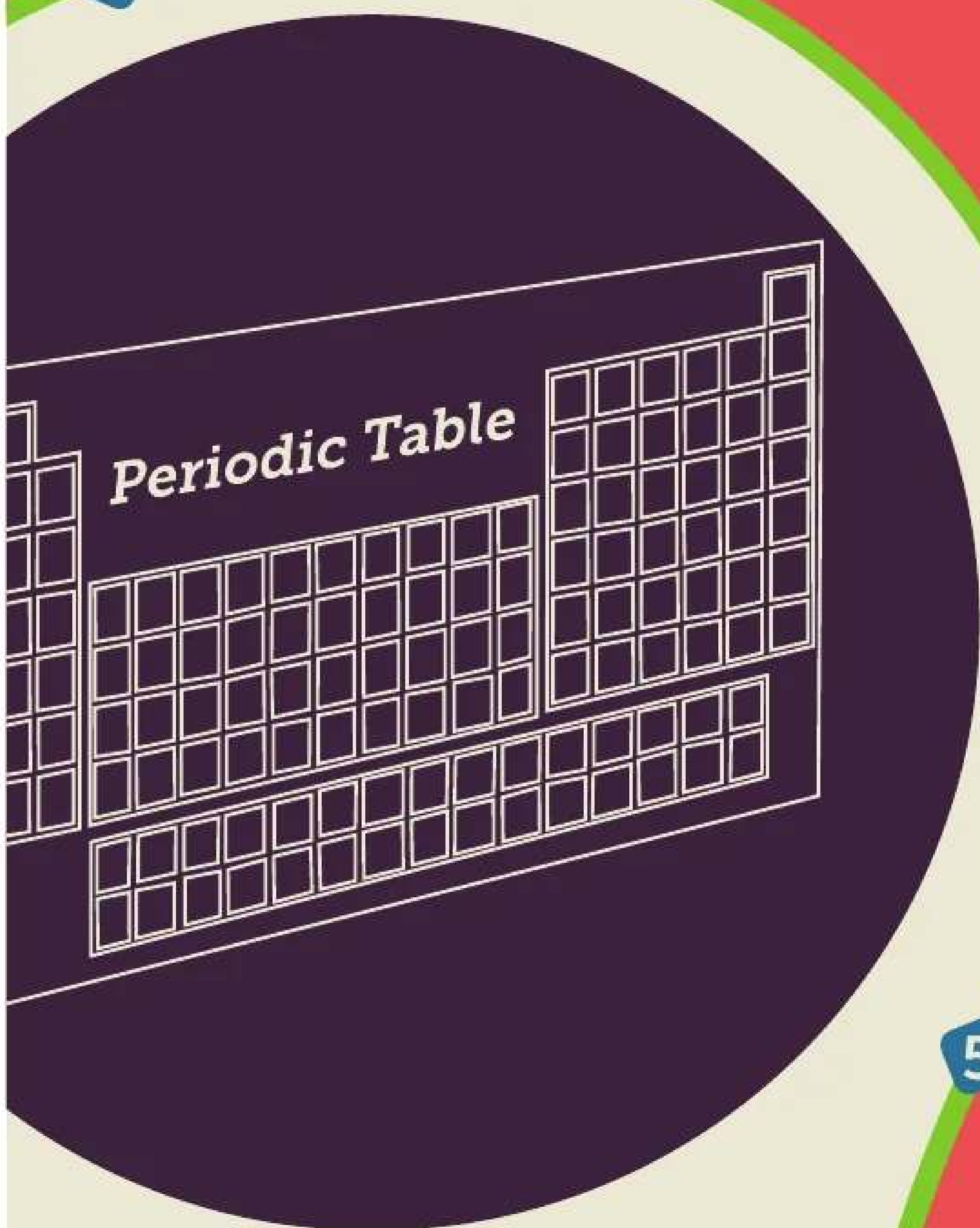

Figure 10

History of the Periodic Table

The modern periodic table is the result of the work of many scientists over the centuries who studied elements and discovered periodic patterns in their properties.

1 **1789** Antoine Lavoisier defines the chemical element, develops a list of all known elements, and distinguishes between metals and nonmetals.

2 **1814** An academic journal publishes Jons Jacob Berzelius's paper proposing letter symbols for the known elements.

3 **1869** Lothar Meyer and Dmitri Mendeleev independently develop tables based on element characteristics and predict the properties of unknown elements.

4 **1894-1900** The noble gases—argon, helium, krypton, neon, xenon, and radon—become a new group in the periodic table.

5 **1913** Henry Moseley determines the atomic number of known elements and establishes that element properties vary periodically with atomic number.

6 **1940** Synthesized elements with an atomic number larger than 92 become part of a new block of the periodic table called the actinides.

7 **1969** Researchers at the University of California, Berkeley synthesize the first element heavier than the actinides. It has a half-life of 4.7 seconds and is named rutherfordium.

8 **2016** The syntheses of elements 113, 115, 117, and 118 are confirmed. This completes the seventh period of the periodic table.

PRACTICE Problems ADDITIONAL PRACTICE

8. Without using the periodic table, determine the group, period, and block of an atom with the following electron configurations.

a. $[Ne]3s^2$ **b.** $[He]2s^2$ **c.** $[Kr]5s^24d^{10}5p^5$

9. What are the symbols for the elements with the following valence electron configurations?

a. s^2d^1 **b.** s^2p^3 **c.** s^2p^6

10. CHALLENGE Write the electron configuration of the following elements.

a. the group 2 element in the fourth period

b. the group 12 element in the fourth period

c. the noble gas in the fifth period

d. the group 16 element in the second period

The development of the periodic table took many years, but like all scientific knowledge, it is open to change. As new elements are synthetized, identified, and named, and as new data about elements arise from experimentation, the periodic table is updated.

Refer to **Figure 10** on the previous page to learn more about the history of the periodic table and the work of the many scientists who contributed to its development. The periodic table is an essential tool in understanding and exploring chemistry and you will use it throughout your study of the subject.

Check Your Progress

Summary

- The periodic table has four blocks (s, p, d, f).
- Elements within a group have similar chemical properties.
- The group number for elements in groups 1 and 2 equals the element's number of valence electrons.
- The energy level of an atom's valence electrons equals its period number.

Demonstrate Understanding

11. **Explain** what determines the blocks in the periodic table.
12. **Determine** in which block of the periodic table are the elements having the following valence electron configurations.

    a. s^2p^4 c. s^2d^1

    b. s^1 d. s^2p^1
13. **Infer** Xenon, a nonreactive gas used in strobe lights, is a poor conductor of heat and electricity. Would you expect xenon to be a metal, a nonmetal, or a metalloid? Where would you expect it to be on the periodic table? Explain.
14. **Explain** why elements within a group have similar chemical properties.
15. **Model** Make a simplified sketch of the periodic table, and label the s-, p-, d-, and f-blocks.

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LESSON 3
PERIODIC TRENDS

FOCUS QUESTION

How can you use the periodic table to predict an element's properties?

Atomic Radius

Many properties of the elements tend to change in a predictable way, known as a trend, as you move across a period or down a group. Atomic size is one such periodic trend. The sizes of atoms are influenced by electron configuration.

Recall that the electron cloud surrounding a nucleus does not have a clearly defined edge. The outer limit of an electron cloud is defined as the spherical surface within which there is a 90% probability of finding an electron. However, this surface does not exist in a physical way, as the outer surface of a golf ball does. Atomic size is defined by how closely an atom lies to a neighboring atom. Because the nature of the neighboring atom can vary from one substance to another, the size of the atom itself also tends to vary somewhat from substance to substance.

For metals such as sodium, the atomic radius is defined as half the distance between adjacent nuclei in a crystal of the element, as shown in **Figure 11**. For elements that commonly occur as molecules, such as many nonmetals, the atomic radius is defined as half the distance between nuclei of identical atoms that are chemically bonded together. The atomic radius of a nonmetal diatomic hydrogen molecule (H_2) is shown in **Figure 11**.

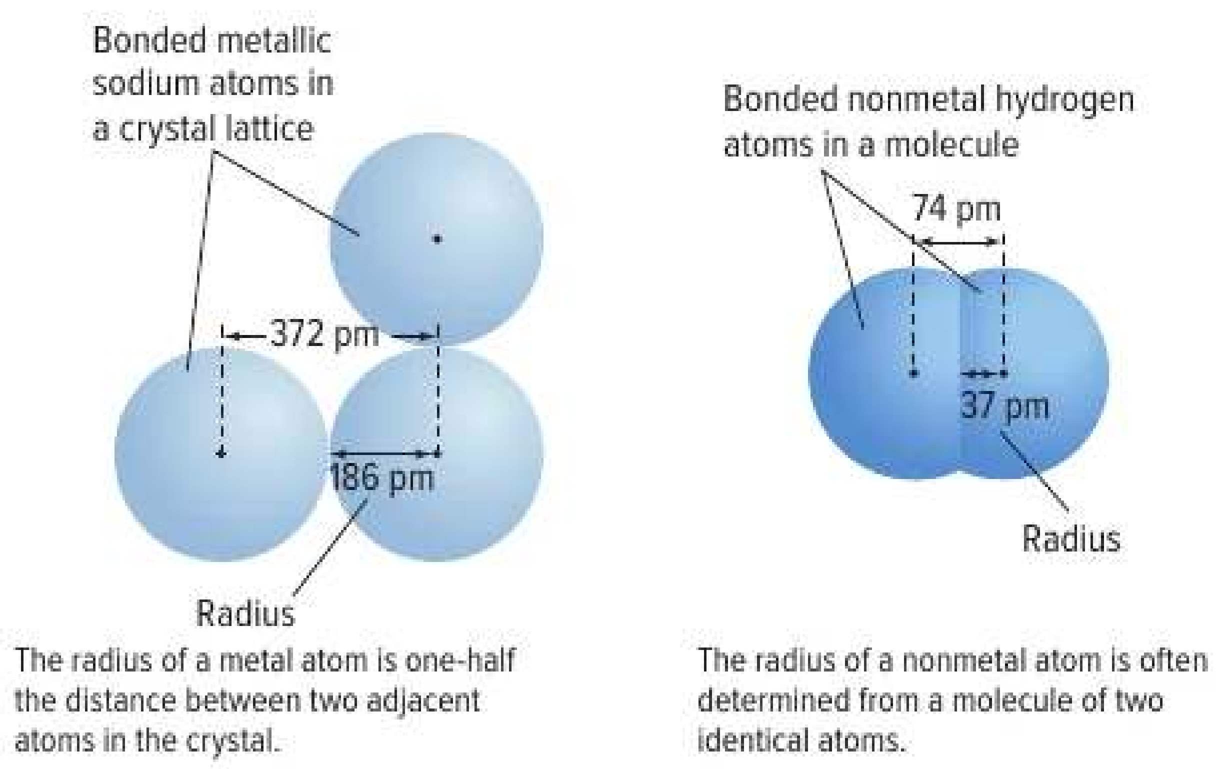

Figure 11 Atomic radii depend on the type of bonds that atoms form.

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COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practice: Electron Patterns of Atoms

HS-PS1-1. Use the periodic table as a model to predict the relative properties of elements based on patterns of electrons in the outermost energy level of atoms.

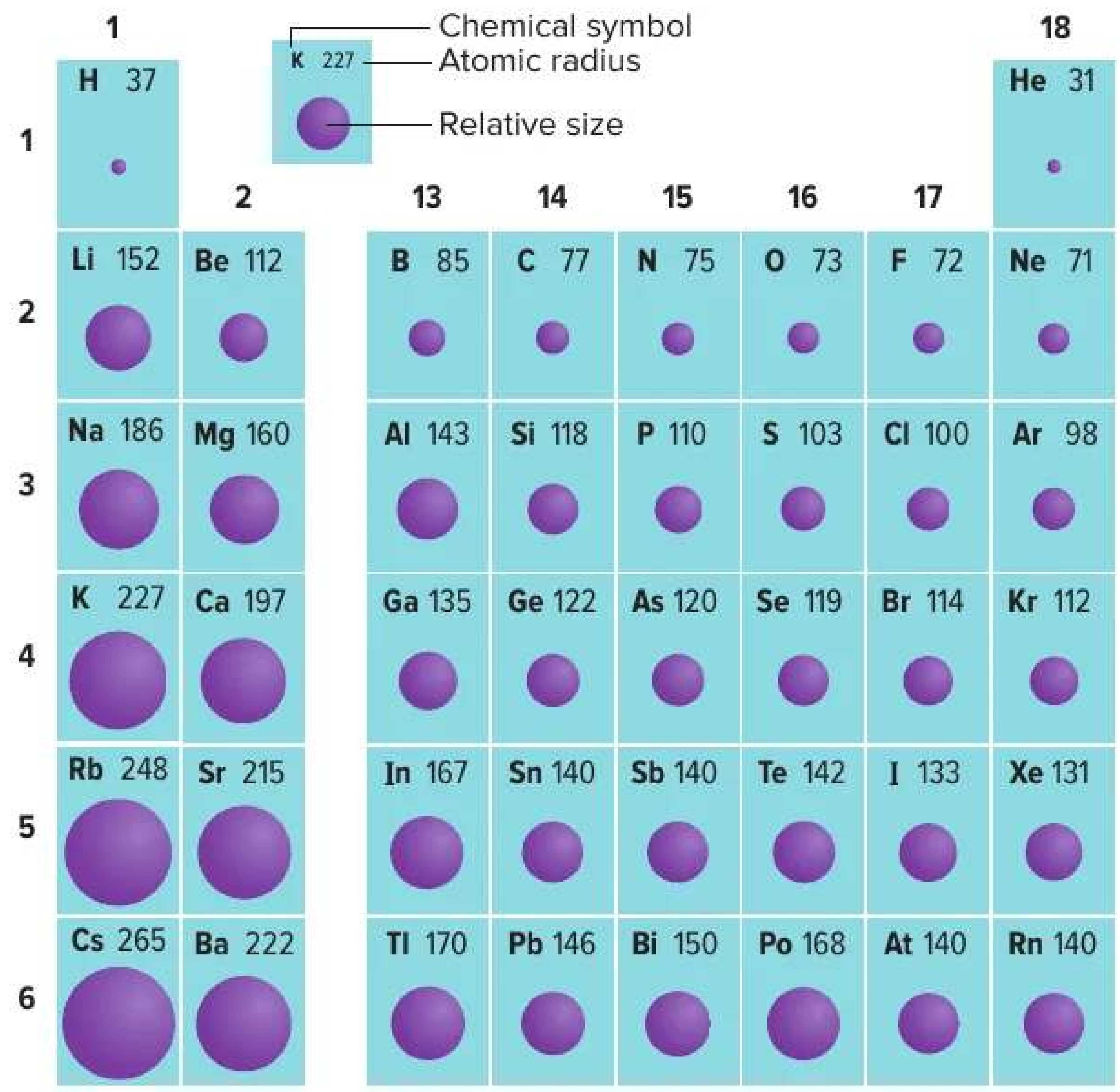


Figure 12 The atomic radii of the representative elements, given in picometers (10^{-12} m), vary as you move from left to right within a period and down a group.

Infer *why the atomic radii increase as you move down a group.*

Trends within periods

In general, there is a decrease in atomic radii as you move from left to right across a period. This trend is illustrated in **Figure 12**. It is caused by the increasing positive charge in the nucleus and the fact that the principal energy level within a period remains the same. Each successive element has one additional proton and electron, and each additional electron is added to orbitals corresponding to the same principal energy level. Moving across a period, no additional electrons come between the valence electrons and the nucleus. Thus, the valence electrons are not shielded from the increased nuclear charge, which pulls the outermost electrons closer to the nucleus.

Get It?

Discuss how the fact that the principal energy level remains the same within a period explains the decrease in the atomic radii across a period.

Trends within groups

Atomic radii generally increase as you move down a group on the periodic table. The nuclear charge increases, and electrons are added to orbitals corresponding to successively higher principal energy levels. However, the increased nuclear charge does not pull the outer electrons toward the nucleus to make the atom smaller as you might expect. Why does the increased nuclear charge not make the atom smaller?

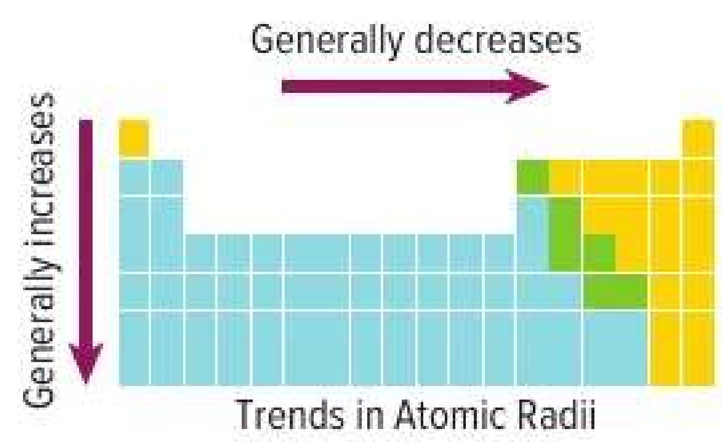


Figure 13 Atomic radii generally decrease from left to right in a period and generally increase as you move down a group.

Moving down a group, the outermost orbital increases in size along with the increasing principal energy level; thus, the atom becomes larger. The larger orbital means that the outer electrons are farther from the nucleus. This increased distance offsets the pull of the increased nuclear charge. Also, as additional orbitals between the nucleus and the outer electrons are occupied, these electrons shield the outer electrons from the nucleus. **Figure 13** summarizes the group and period trends.

EXAMPLE Problem 2

ELECTRON CONFIGURATION AND THE PERIODIC TABLE Which has the largest atomic radius: carbon (C), fluorine (F), beryllium (Be), or lithium (Li)? Answer without referring to **Figure 12**. Explain your answer in terms of trends in atomic radii.

1 ANALYZE THE PROBLEM

You are given four elements. First, determine the groups and periods the elements occupy. Then apply the general trends in atomic radii to determine which has the largest atomic radius.

2 SOLVE FOR THE UNKNOWN

From the periodic table, all the elements are found to be in period 2. *Determine the periods.*

Ordering the elements from left-to-right across the period yields: Li, Be, C, and F.

The first element in period 2, lithium, has the largest radius. *Apply the trend of decreasing radii across a period.*

3 EVALUATE THE ANSWER

The period trend in atomic radii has been correctly applied. Checking radii values in **Figure 12** verifies the answer.

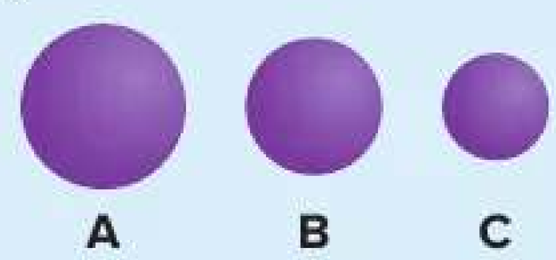
PRACTICE Problems

ADDITIONAL PRACTICE

Answer the following questions using your knowledge of group and period trends in atomic radii. Do not use the atomic radii values in Figure 12 to answer the questions.

16. Which has the largest atomic radius: magnesium (Mg), silicon (Si), sulfur (S), or sodium (Na)? The smallest?

17. The figure on the right shows helium, krypton, and radon. Which one is krypton? How can you tell?



18. Can you determine which of two unknown elements has the larger radius if the only known information is that the atomic number of one of the elements is 20 greater than the other? Explain.

19. CHALLENGE Determine which element in each pair has the largest atomic radius:

a. the element in period 2, group 1; or the element in period 3, group 18

b. the element in period 5, group 2; or the element in period 3, group 16

c. the element in period 3, group 14; or the element in period 6, group 15

d. the element in period 4, group 18; or the element in period 2, group 16

Ionic Radius

Atoms can gain or lose one or more electrons to form ions. Because electrons are negatively charged, atoms that gain or lose electrons acquire a net charge. Thus, an **ion** is an atom or a bonded group of atoms that has a positive or negative charge.

You will learn more about ions later, but for now, consider how the formation of an ion affects the size of an atom.

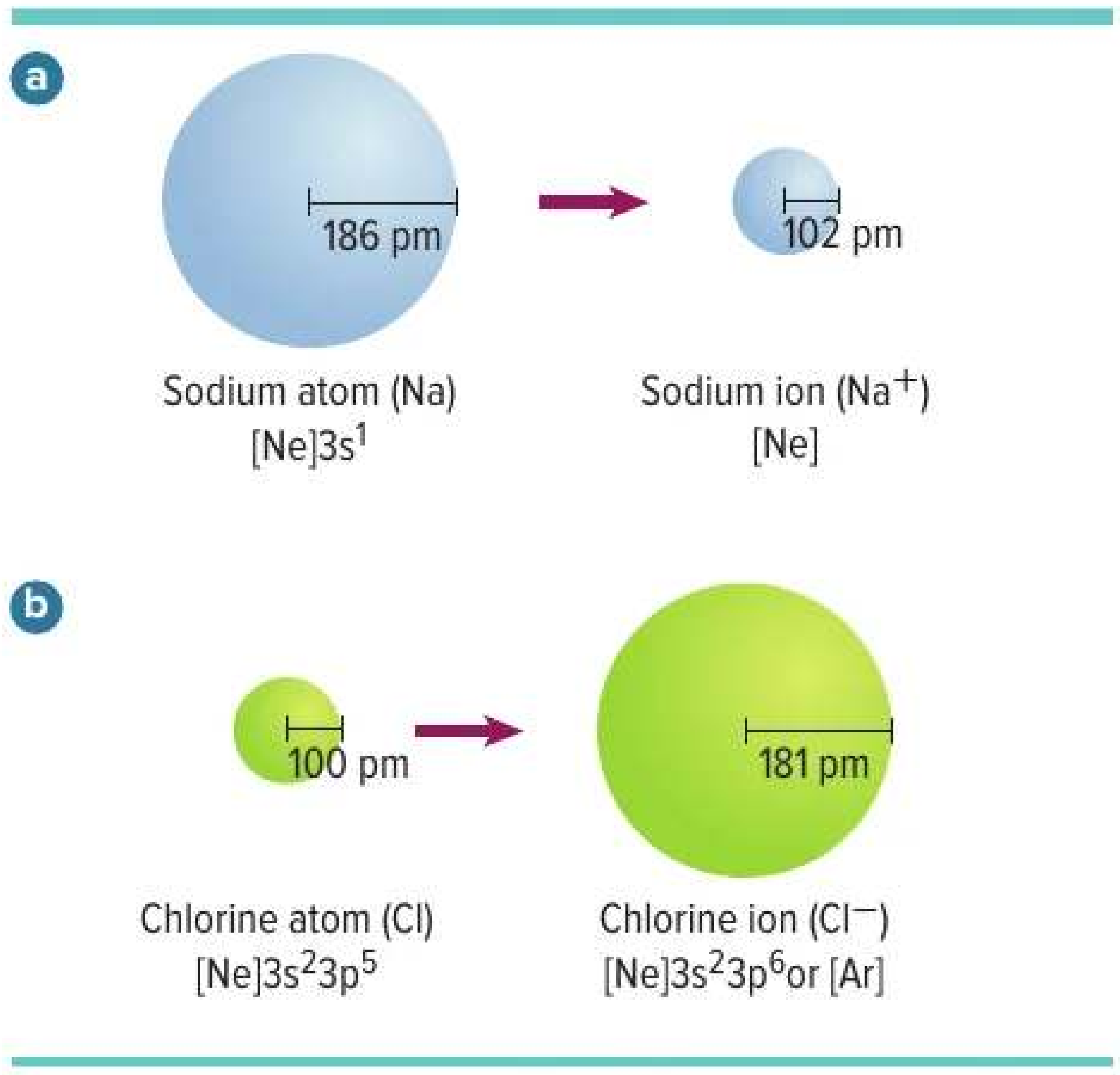


Figure 14 The size of atoms varies greatly when they form ions.

a. Positive ions are smaller than the neutral atoms from which they form.

b. Negative ions are larger than the neutral atoms from which they form.

Losing electrons

When atoms lose electrons and form positively charged ions, they always become smaller. The reason is twofold. The electron lost from the atom will almost always be a valence electron. The loss of a valence electron can leave a completely empty outer orbital, which results in a smaller radius. Furthermore, the electrostatic repulsion between the now-fewer number of remaining electrons decreases. As a result, they experience a greater nuclear charge allowing these remaining electrons to be pulled closer to the positively charged nucleus.

Figure 14a illustrates how the radius of sodium decreases when sodium atoms form positive ions. The outer orbital of the sodium atom is unoccupied in the sodium ion, so the sodium ion is much smaller than the sodium atom.

Gaining electrons

When atoms gain electrons and form negatively charged ions, they become larger. The addition of an electron to an atom increases the electrostatic repulsion between the atom's outer electrons, forcing them to move farther apart. The increased distance between the outer electrons results in a larger radius.

Figure 14b shows how the radius of chlorine increases when chlorine atoms form negative ions. Adding an electron to a chlorine atom increases the electrostatic repulsion among its valence electrons. The increased repulsion causes the electrons to move farther apart and results in the radius of a chloride ion being almost twice as large as that of a chlorine atom.

Get It?

Explain why a lithium ion is smaller than a lithium atom.

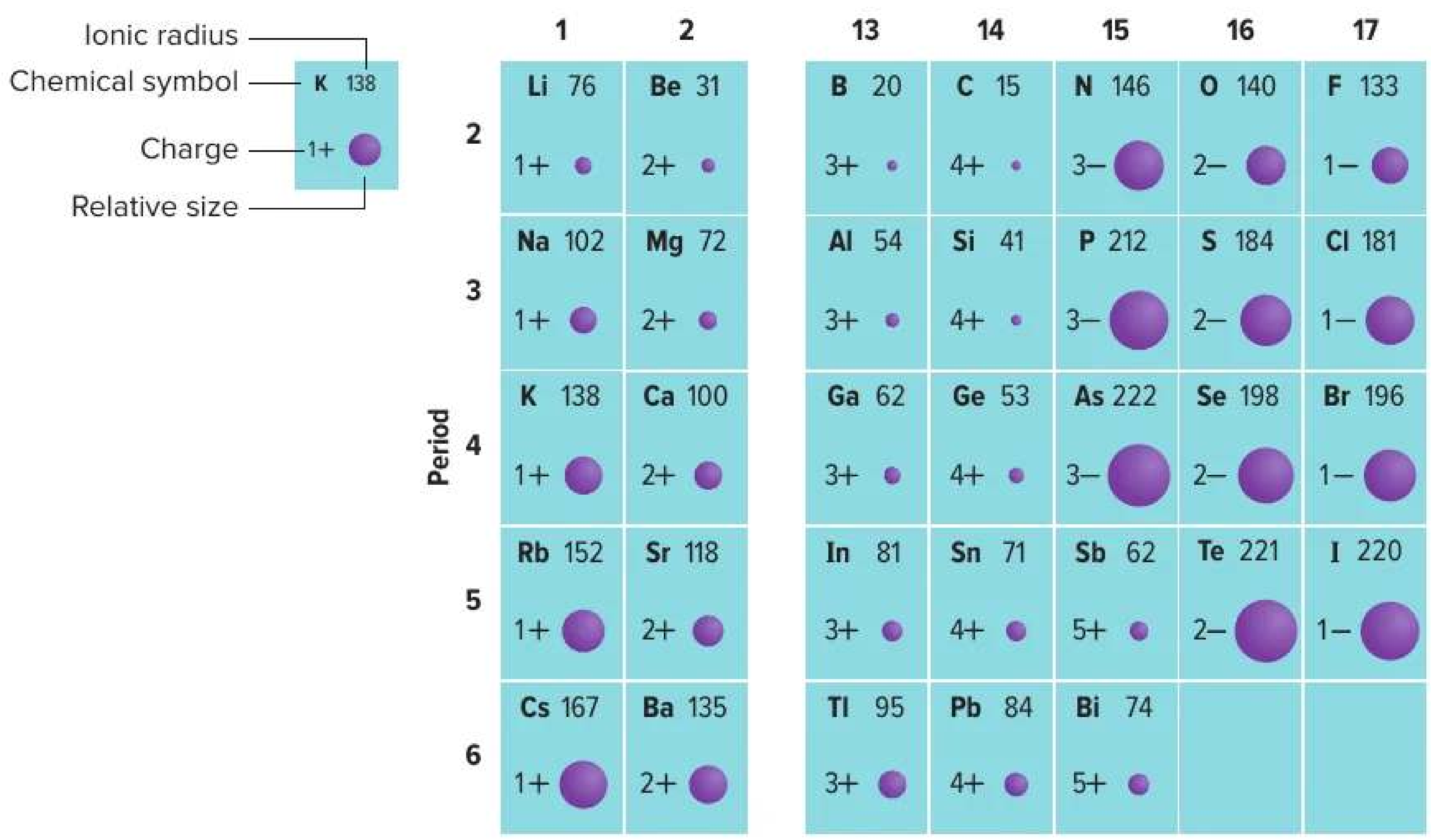


Figure 15 The ionic radii of most of the representative elements are shown in picometers (10^{-12} m).
Explain *why the ionic radii increase for both positive and negative ions as you move down a group.*

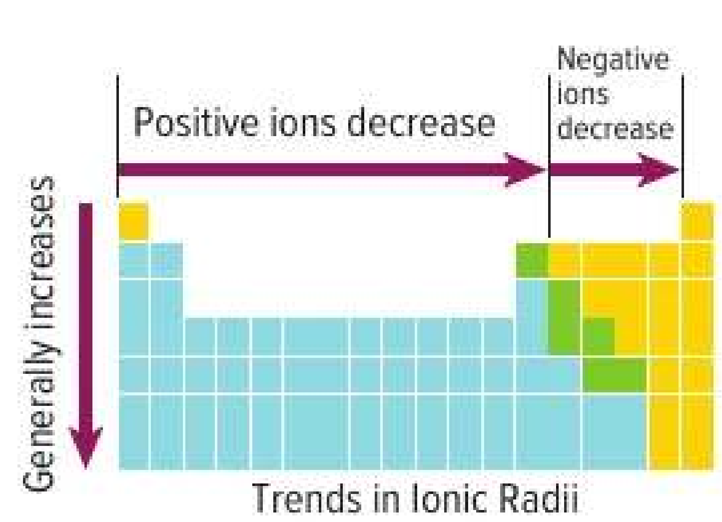
Trends within periods

The ionic radii of most of the representative elements are shown in **Figure 15**. Note that elements on the left side of the table form smaller positive ions, and elements on the right side of the table form larger negative ions.

In general, as you move from left to right across a period, the size of the positive ions gradually decreases. Then, beginning in group 15 or 16, the size of the much-larger negative ions also gradually decreases.

Trends within groups

As you move down a group, an ion's outer electrons are in orbitals corresponding to higher principal energy levels, resulting in a gradual increase in ionic size. Thus, the ionic radii of both positive and negative ions increase as you move down a group. The group and period trends in ionic radii are summarized in **Figure 16**.



Get It?

Explain why calcium has a greater atomic radius than magnesium.

Figure 16 The diagram summarizes the general trends in ionic radii.

Ionization Energy

To form a positive ion, an electron must be removed from a neutral atom. This requires energy. The energy is needed to overcome the attraction between the positive charge of the nucleus and the negative charge of the electron.

Ionization energy is defined as the energy required to remove an electron from a gaseous atom. For example, 8.64×10^{-19} J is required to remove an electron from a gaseous lithium atom. The energy required to remove the first outermost electron from an atom is called the first ionization energy. The first ionization energy of lithium equals 8.64×10^{-19} J. The loss of the electron results in the formation of a Li^+ ion. The first ionization energies of the elements in periods 1 through 5 are plotted on the graph in **Figure 17**.

Get It?

Define *ionization energy.*

Think of ionization energy as an indication of how strongly an atom's nucleus holds onto its valence electrons. A high ionization energy value indicates the atom has a strong hold on its electrons. Atoms with large ionization energy values are less likely to form positive ions. Likewise, a low ionization energy value indicates an atom loses an outer electron easily. Such atoms are likely to form positive ions. Lithium's low ionization energy, for example, is important for its use in lithium-ion computer backup batteries, where the ability to lose electrons easily makes a battery that can quickly provide a large amount of electrical power.

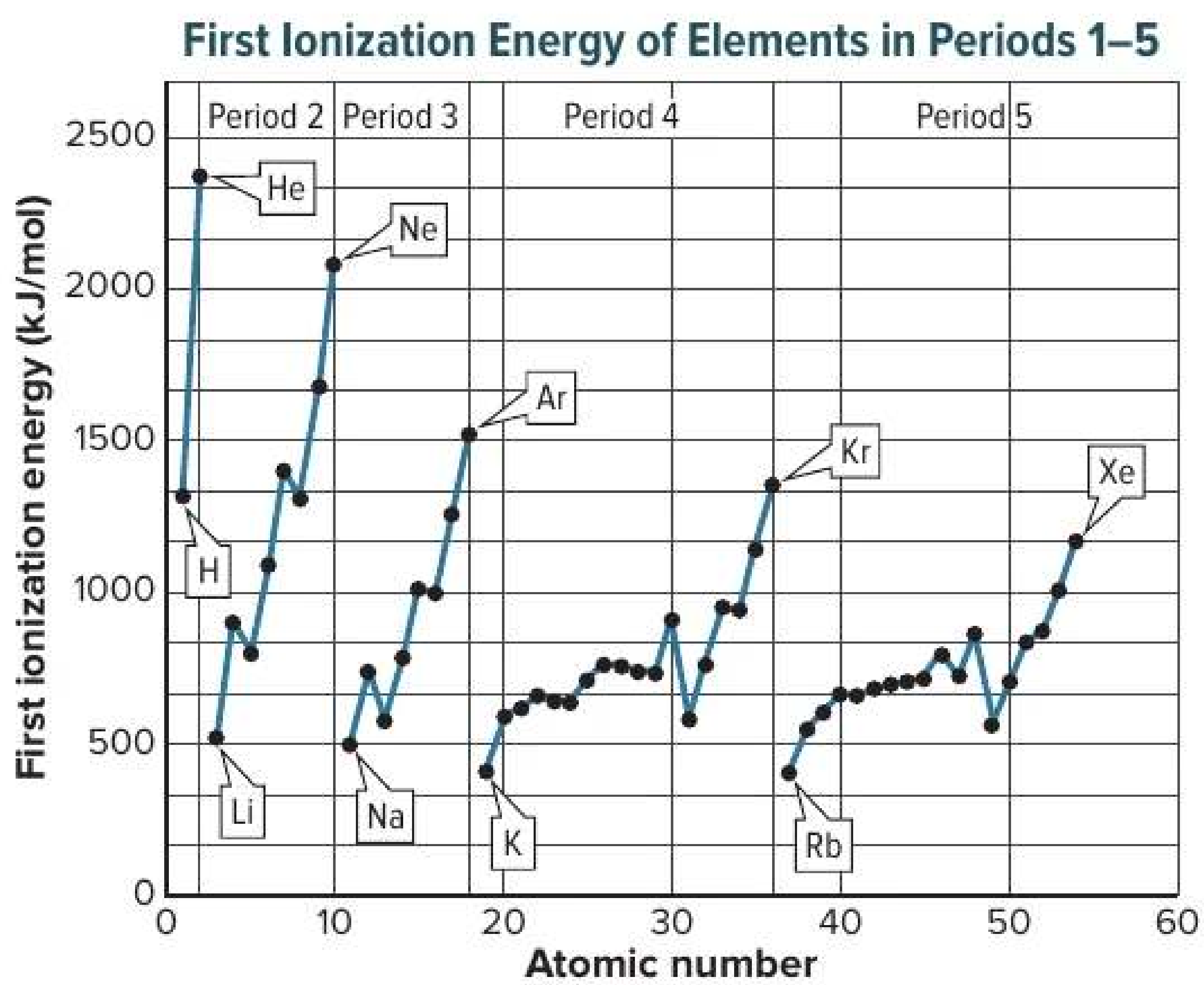


Figure 17 The first ionization energies for elements in periods 1 through 5 are shown as a function of the atomic number.

Describe *how ionization energy and atomic number are related as shown on this scatter plot.*

Each set of connected points on the graph in **Figure 17** represents the elements in a period. The group 1 metals have low ionization energies. Thus, group 1 metals (Li, Na, K, Rb) are likely to form positive ions. The group 18 elements (He, Ne, Ar, Kr, Xe) have high ionization energies and are unlikely to form ions. The stable electron configuration of gases of group 18 greatly limits their reactivity.

Removing more than one electron

After removing the first electron from an atom, it is possible to remove additional electrons. The amount of energy required to remove a second electron from a 1+ ion is called the second ionization energy, the amount of energy required to remove a third electron from a 2+ ion is called the third ionization energy, and so on. **Table 5** lists the first through ninth ionization energies for elements in period 2.

Reading across **Table 5** from left to right, you will see that the energy required for each successive ionization always increases. However, the increase in energy does not occur smoothly. Note that for each element there is an ionization for which the required energy increases dramatically. For example, the second ionization energy of lithium (7300 kJ/mol) is much greater than its first ionization energy (520 kJ/mol). This means that a lithium atom is likely to lose its first valence electron but extremely unlikely to lose its second.

Get It?
Infer how many electrons carbon is likely to lose.

If you examine the **Table 5**, you will notice that the ionization at which the large increase in energy occurs is related to the atom's number of valence electrons. The element lithium has one valence electron and the increase occurs after the first ionization energy. Lithium easily forms the common lithium 1+ ion but is unlikely to form a lithium 2+ ion. The increase in ionization energy shows that atoms hold onto their inner core electrons much more strongly than they hold onto their valence (outermost) electrons.

Table 5 Successive Ionization Energies for the Period 2 Elements

Element	Valence Electrons	Ionization Energy (kJ/mol)*								
		1st	2nd	3rd	4th	5th	6th	7th	8th	9th
Li	1	520	7300	11,810						
Be	2	900	1760	14,850	21,010					
B	3	800	2430	3660	25,020	32,820				
C	4	1090	2350	4620	6220	37,830	47,280			
N	5	1400	2860	4580	7480	9440	53,270	64,360		
O	6	1310	3390	5300	7470	10,980	13,330	71,870	84,080	
F	7	1680	3370	6050	8410	11,020	15,160	17,870	92,040	106,430
Ne	8	2080	3950	6120	9370	12,180	15,240	20,000	23,070	115,380

*mol is an abbreviation for mole, a quantity of matter

Trends within periods

As shown in **Figure 17** and by the values in **Table 5**, first ionization energies generally increase as you move from left to right across a period. The increased nuclear charge of each successive element produces an increased hold on the valence electrons.

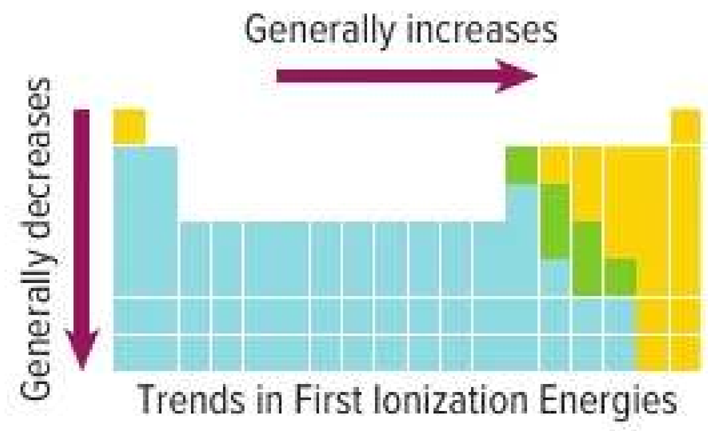


Figure 18 Ionization energies generally increase from left to right in a period and generally decrease as you move down a group.

Trends within groups

First ionization energies generally decrease as you move down a group. This decrease in energy occurs because atomic size increases as you move down the group. Less energy is required to remove the valence electrons farther from the nucleus. **Figure 18** summarizes the group and period trends in first ionization energies.

Octet rule

When a sodium atom loses its single valence electron to form a 1+ sodium ion, its electron configuration changes as shown below.

Sodium atom $1s^22s^22p^63s^1$ Sodium ion $1s^22s^22p^6$

Note that the sodium ion has the same electron configuration as neon ($1s^22s^22p^6$), a noble gas. This observation leads to one of the most important principles in chemistry, the **octet rule.** The octet rule states that atoms tend to gain, lose, or share electrons in order to acquire a full set of eight valence electrons. This reinforces what you learned earlier, that the electron configuration of filled s and p orbitals of the same energy level (consisting of eight valence electrons) is unusually stable. Note that the first-period elements are an exception to the rule, as they are complete with only two valence electrons.

Real-World Chemistry

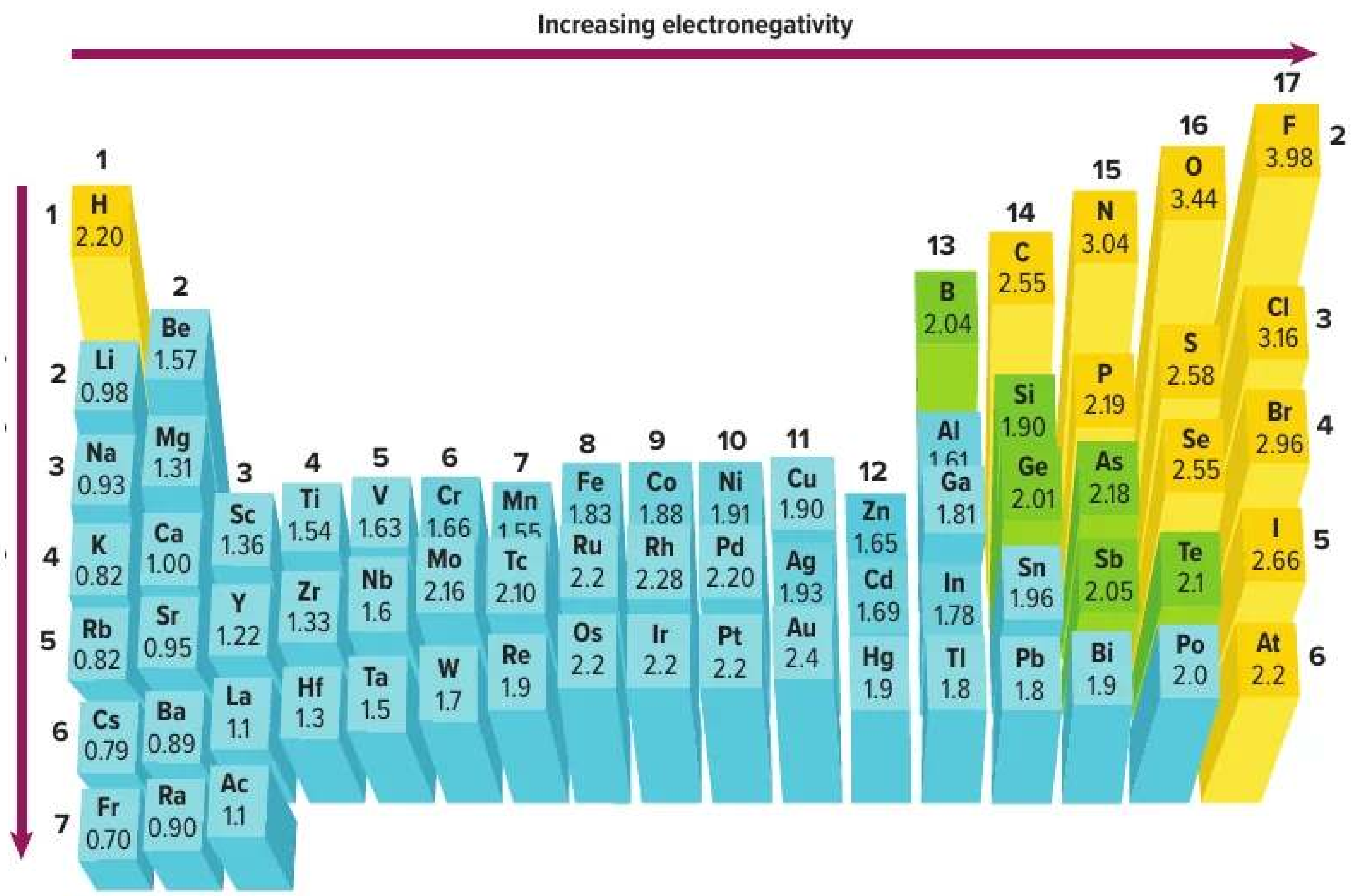
Ionization Energy

SCUBA DIVING The increased pressure that scuba divers experience far below the water's surface can cause too much oxygen to enter their blood, which would result in confusion and nausea. To avoid this, divers sometimes use a gas mixture called heliox—oxygen diluted with helium. Helium's high ionization energy ensures that it will not react chemically in the bloodstream.

Electronegativity

The **electronegativity** of an element indicates the relative ability of its atoms to attract electrons in a chemical bond. As shown in **Figure 19**, on the next page, electronegativity generally decreases as you move down a group. **Figure 19** also indicates that electronegativity generally increases as you move from left to right across a period. Fluorine is the most electronegative element, with a value of 3.98, meaning it attracts electrons more strongly than any other element in a chemical bond. Cesium and francium are the least electronegative elements, with values of 0.79 and 0.70, repectively. In a chemical bond, the atom with the greater electronegativity more strongly attracts the bond's electrons. Note that because the noble gases form very few compounds, they do not have electronegativity values.

Jag_cz/Shutterstock.com



Electronegativity Values in Paulings

Figure 19 The electronegativity values for most of the elements are shown. The values are given in Paulings, a unit named after American scientist Linus Pauling (1901–1994).

Infer *why electronegativity values are not listed for the noble gases.*

Check Your Progress

Summary

- Atomic and ionic radii decrease from left to right across a period, and increase as you move down a group.
- Ionization energies generally increase from left to right across a period, and decrease down a group.
- The octet rule states that atoms gain, lose, or share electrons to acquire a full set of eight valence electrons.
- Electronegativity generally increases from left to right across a period, and decreases down a group.

Demonstrate Understanding

20. **Explain** how the period and group trends in atomic radii are related to electron configuration.
21. **Indicate** whether fluorine or bromine has a larger value for each of the following properties.
    a. electronegativity
    b. ionic radius
    c. atomic radius
    d. ionization energy
22. **Explain** why it takes more energy to remove the second electron from a lithium atom than it does to remove the fourth electron from a carbon atom.
23. **Calculate** Determine the differences in electronegativity, ionic radius, atomic radius, and first ionization energy for oxygen and beryllium.
24. **Make and Use Graphs** Graph the atomic radii of the representative elements in periods 2, 3, and 4 versus their atomic numbers. Connect the points of elements in each period, so that there are three separate curves on the graph. Summarize the trends in atomic radii shown on your graph. Explain.

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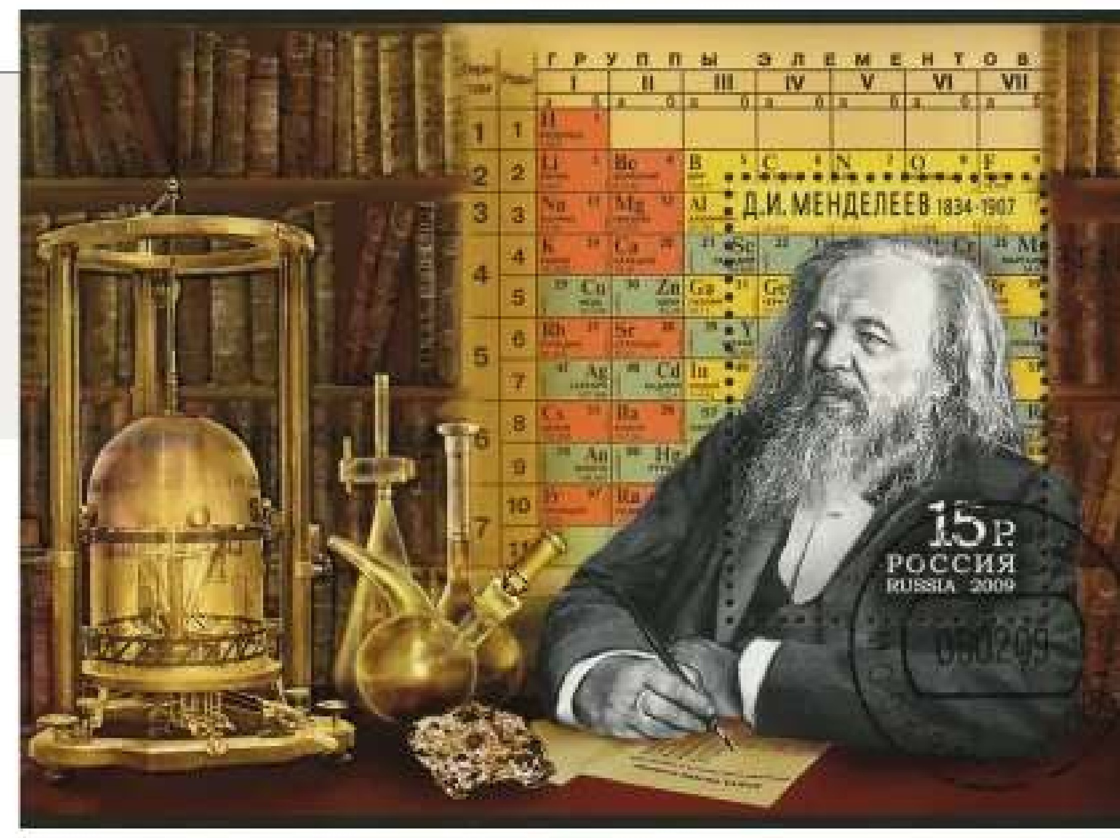
NATURE OF SCIENCE

The Evolving Periodic Table

Chemists have used the periodic table since its development in the late 1860s, but it has evolved over the years, and it is still evolving today.



This Russian stamp commemorates Dmitri Mendeleev, who first published the periodic table in 1869.

The Adaptable Periodic Table

The design of the periodic table that we use today was developed in the 19th century, before the discovery of all of the naturally occurring elements, including the noble gases, and before the synthesis of elements. Initially, the elements were organized on the periodic table by atomic mass. This caused some inconsistencies. Once the atomic number was used to align the elements into rows and properties were used to organize the elements into columns, the modern periodic table was born.

The modern periodic table demonstrates the elegance of the nature of science. The early periodic table evolved to incorporate new information. For example, a new column was added when the noble gases were discovered. Period 7 is now completely filled after the recently discovered elements 113, 115, 117, and 118 were added.

Newly synthesized elements must now be added to another row. If and when these new elements are discovered, period 8 will be added to the periodic table.

When the periodic table was being developed chemists did not understand why these groups of elements had similar properties, but they recognized the periodic trend in these properties. Today, students and chemists understand that elements in a group have similar properties and the same number of valence electrons. Our understanding of the atom evolved along with our understanding of what elements in a group on the periodic table have in common in terms of valence electrons, reactivity, and properties.

While the nature of atoms and elements were not understood at the time of the periodic table's development, the original table was well designed. As new information was discovered about atoms and elements, the periodic table evolved and incorporated the new information resulting in a table that is just as useful today as it was when it was first developed.

APPLY SCIENTIFIC PRINCIPLES AND EVIDENCE

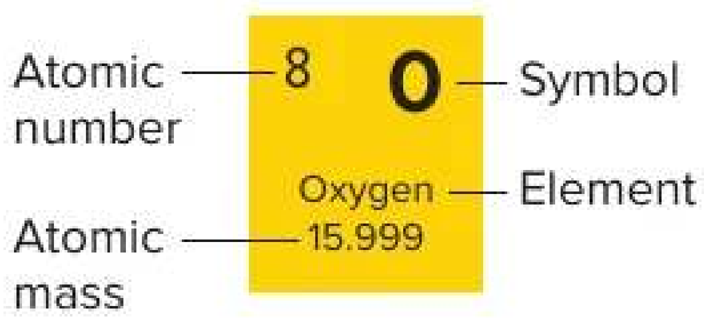
Explain how a discovery about an element or elements was incorporated into the modern periodic table.

MODULE 5
STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 DEVELOPMENT OF THE MODERN PERIODIC TABLE

- The elements were first organized by increasing atomic mass, which led to inconsistencies. Later, they were organized by increasing atomic number.
- The periodic law states that when the elements are arranged by increasing atomic number, there is a periodic repetition of their chemical and physical properties.
- The periodic table organizes the elements into periods (rows) and groups or families (columns); elements with similar properties are in the same group.
- Elements are classified as metals, nonmetals, or metalloids.



- periodic law
- group
- period
- representative element
- transition element
- metal
- alkali metal
- alkaline earth metal
- transition metal
- inner transition metal
- lanthanide series
- actinide series
- nonmetal
- halogen
- noble gas
- metalloid

Lesson 2 CLASSIFICATION OF THE ELEMENTS

- The periodic table has four blocks (s, p, d, f).
- Elements within a group have similar chemical properties.
- The group number for elements in groups 1 and 2 equals the element's number of valence electrons.
- The energy level of an atom's valence electrons equals its period number.

Lesson 3 PERIODIC TRENDS

- Atomic and ionic radii decrease from left to right across a period, and increase as you move down a group.
- Ionization energies generally increase from left to right across a period, and decrease down a group.
- The octet rule states that atoms gain, lose, or share electrons to acquire a full set of eight valence electrons.
- Electronegativity generally increases from left to right across a period, and decreases down a group.

- ion
- ionization energy
- octet rule
- electronegativity

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

What can we learn from the periodic table?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will apply your evidence from this module and complete your project.

GO FURTHER

SEP Data Analysis Lab

Can you predict the properties of an element?

Francium was discovered in 1939, but its existence was predicted by Mendeleev in the 1870s. It is the least stable of the first 101 elements, with a half-life of just 22 minutes for its most stable isotope. Use the properties of other alkali metals, shown in the table, to predict some of francium's properties.

Alkali Metals Data

Element	Melting Point (°C)	Boiling Point (°C)	Radius (pm)
Lithium	180.5	1342	152
Sodium	97.8	883	186
Potassium	63.4	759	227
Rubidium	39.3	688	248
Cesium	28.4	671	265
Francium	?	?	?

CER Analyze and Interpret Data

Use the given information about the known properties of the alkali metals to devise a method for predicting the corresponding properties of francium.

1. **Claim, Evidence, Reasoning** Devise an approach that clearly displays the trends for each of the properties given in the table and allows you to extrapolate a value for francium. Use the periodic law as a guide.
2. **Claim, Evidence, Reasoning** Predict whether francium is a solid, a liquid, or a gas. How can you support your prediction?
3. **Infer** which column of data presents the greatest possible error in making a prediction. Explain.
4. **Determine** why producing 1 million francium atoms per second is not enough to make measurements, such as density or melting point.

Will & Deni McIntyre/Science Source


Mark A. Schneider/Science Source

MODULE 6

IONIC COMPOUNDS AND METALS

ENCOUNTER THE PHENOMENON

Why do some crystals form cubes?

GO ONLINE to play a video about mineral crystals that have formed in a cave.

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about why some crystals form cubes.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.

LESSON 2: Explore & Explain: Ionic Bond Formation

LESSON 2: Explore & Explain: Properties of Ionic Compounds

Smithsonian

Additional Resources

LESSON 1
ION FORMATION

FOCUS QUESTION
Why do elements form compounds?

Valence Electrons and Chemical Bonds

Imagine going on a scuba dive, diving below the ocean's surface and observing the awe-inspiring world below. You might explore the colorful and exotic organisms teeming around a coral reef, such as the one shown in **Figure 1.** The reef is formed from a compound called calcium carbonate, which is just one of thousands of compounds found on Earth. How do so many compounds form from the relatively few elements known to exist? The answer to this question involves the electron structure of atoms and the nature of the forces between atoms.

In previous chapters, you learned that elements within a group on the periodic table have similar properties. Many of these properties depend on the number of valence electrons the atom has. These valence electrons are involved in the formation of chemical bonds between two atoms. A **chemical bond** is the force that holds two atoms together.

Figure 1 As carbon dioxide dissolves in ocean water, carbonate ions are produced. Coral polyps capture these carbonate ions, producing crystals of calcium carbonate, which they secrete as an exoskeleton. Over time, the coral reef forms. A coral reef is a complex habitat that supports coral, algae, mollusks, echinoderms, and a variety of fishes.

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3D THINKING | DCI Disciplinary Core Ideas | CCC Crosscutting Concepts | SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Review the News

Obtain information from a current news story about ion formation. Evaluate your source and communicate your findings to your class.

Chemical bonds can form by the attraction between the positive nucleus of one atom and the negative electrons of another atom, or by the attraction between positive ions and negative ions. This chapter discusses chemical bonds formed by ions, atoms that have acquired a positive or negative charge. You will learn about bonds that form from the sharing of electrons in a later chapter.

Valence electrons

Recall that an electron-dot structure is a type of diagram used to keep track of valence electrons. Electron-dot structures are especially helpful when used to illustrate the formation of chemical bonds. **Table 1** shows several examples of electron-dot structures. For example, carbon, with an electron configuration of $1s^22s^22p^2$, has four valence electrons in the second energy level. These valence electrons are represented by the four dots around the symbol C in the table.

Table 1 Electron-Dot Structures

Group	1	2	13	14	15	16	17	18
Diagram	Li•	•Be•	•B•	•C•	•N•	O	F	Ne

Recall that ionization energy refers to how easily an atom loses an electron. *Electron affinity* (EA) is a term used to describe how much attraction an atom has for electrons. Noble gases, which have high ionization energies and low electron affinities, show a general lack of chemical reactivity. Other elements on the periodic table react with each other, forming numerous compounds. The difference in reactivity is directly related to the valence electrons.

The difference in reactivity involves the octet—the stable arrangement of eight valence electrons in the outer energy level. Unreactive noble gases have electron configurations that have a full outermost energy level. This level is filled with two electrons for helium ($1s^2$) and eight electrons for the other noble gases (ns^2np^6). Elements tend to react to acquire the stable electron structure of a noble gas.

Positive Ion Formation

A positive ion forms when an atom loses one or more valence electrons in order to attain a noble gas configuration. A positively charged ion is called a **cation.** To understand the formation of a positive ion, compare the electron configurations of the noble gas neon (atomic number 10) and the alkali metal sodium (atomic number 11).

Neon atom (Ne) $1s^22s^22p^6$

Sodium atom (Na) $1s^22s^22p^63s^1$

Note that the sodium atom has one 3s valence electron; it differs from the noble gas neon by that single valence electron. When sodium loses this outer valence electron, the resulting electron configuration is identical to that of neon.

Figure 2 shows how a sodium atom loses its valence electron to become a sodium cation. By losing an electron, the sodium atom acquires the stable outer electron configuration of neon. It is important to understand that although sodium now has the electron configuration of neon, it is not neon. It is a sodium ion with a single positive charge. The 11 protons that establish the character of sodium still remain within its nucleus.

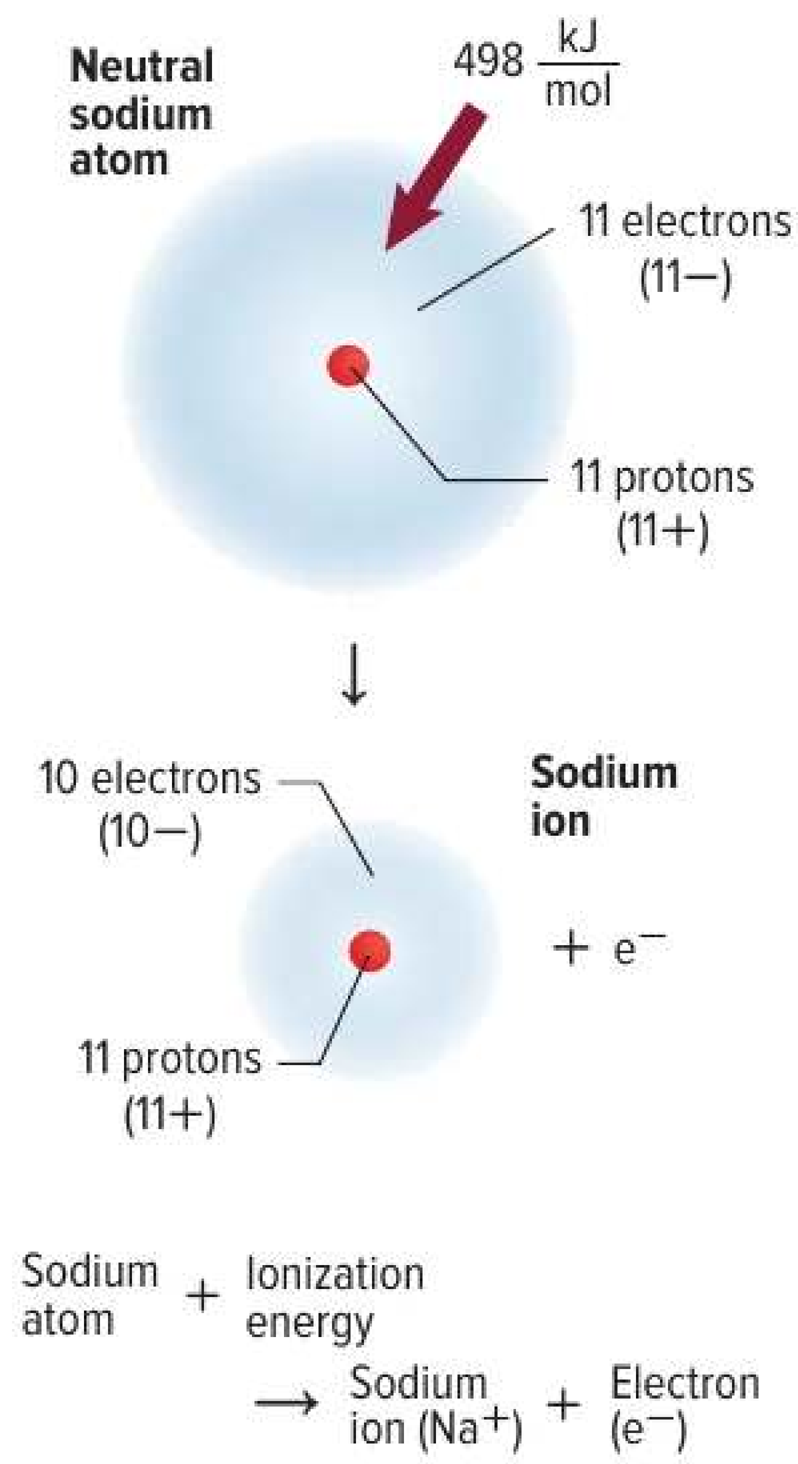

Figure 2 In the formation of a positive ion, a neutral atom loses one or more valence electrons. The atom is neutral because it contains equal numbers of protons and electrons; the ion, however, contains more protons than electrons and has a positive charge.

Analyze *Does the removal of an electron from a neutral atom require energy or release energy?*

Get It?

Identify the number of electrons in the outermost energy level that are associated with maximum stability.

Metal ions

Metal atoms are reactive because they lose valence electrons easily. The group 1 and 2 metals are the most reactive metals on the periodic table. For example, potassium and magnesium, group 1 and 2 elements, respectively, form K^+ and Mg^{2+} ions. Some group 13 atoms also form ions. The ions formed by metal atoms in groups 1, 2, and 13 are summarized in **Table 2**.

Transition metal ions

Recall that, in general, transition metals have an outer energy level of ns^2. Going from left to right across a period, atoms of each element fill an inner d sublevel. When forming positive ions, transition metals commonly lose their two valence electrons, forming 2+ ions. However, it is also possible for d electrons to be lost. Thus, transition metals also commonly form ions of 3+ or greater, depending on the number of d electrons in the electron structure. It is difficult to predict the number of electrons that will be lost. For example, iron (Fe) forms both Fe^{2+} and Fe^{3+} ions. A useful rule of thumb for these metals is that they form ions with a 2+ or a 3+ charge.

Get It?

Explain in your own words why transition metals can form ions with 2+ or 3+ charges.

Table 2 Group 1, 2, and 13 Ions

Group	Configuration	Charge of Ion Formed
1	[noble gas]ns^1	1+ when the s^1 electron is lost
2	[noble gas]ns^2	2+ when the s^2 electrons are lost
13	[noble gas]ns^2np^1	3+ when the s^2p^1 electrons are lost

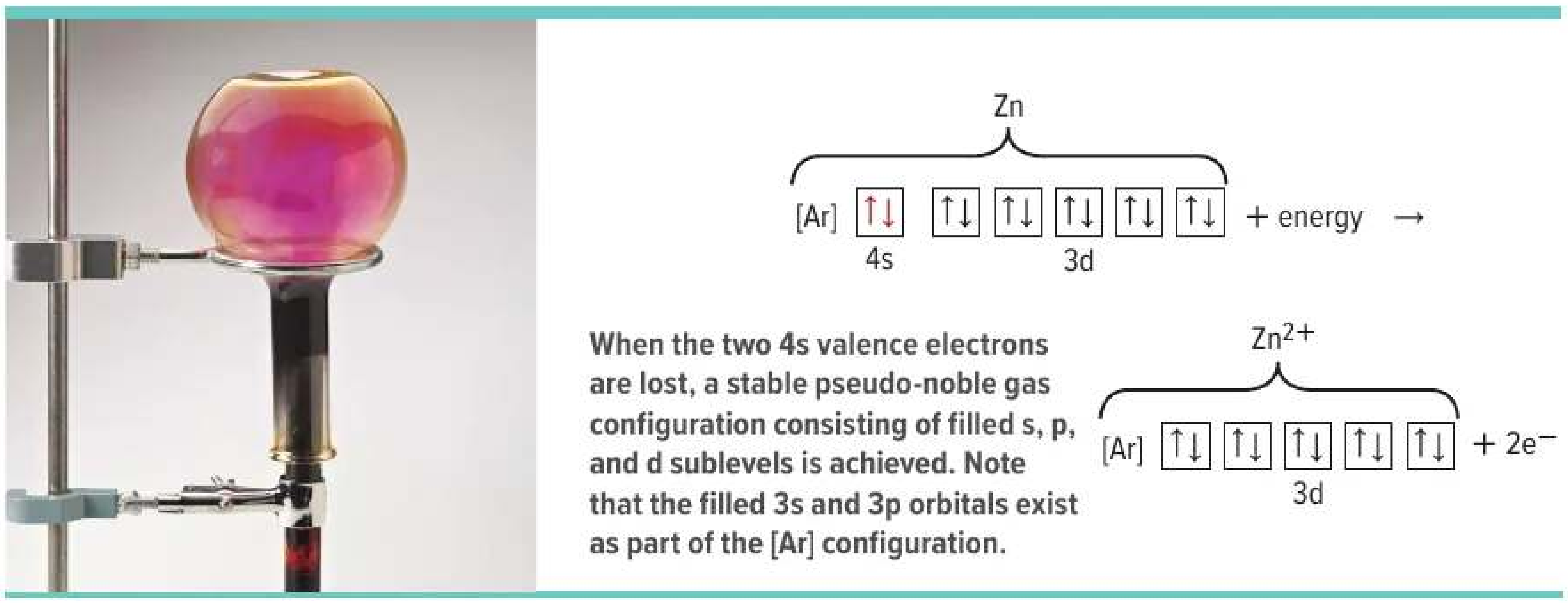

Figure 3 When zinc reacts with iodine, the heat of the reaction causes solid iodine to sublimate into a purple vapor. At the bottom of the tube, ZnI_2 is formed containing Zn^{2+} ions with a pseudo-noble gas configuration.

Pseudo-noble gas configurations

Although the formation of an octet is the most stable electron configuration, other electron configurations can also provide some stability. For example, elements in groups 11–14 lose electrons to form an outer energy level containing full s, p, and d sublevels. These relatively stable electron arrangements are referred to as pseudo-noble gas configurations. The zinc atom has the electron configuration of $1s^22s^22p^63s^23p^64s^23d^{10}$, in **Figure 3.** When forming an ion, the zinc atom loses the two 4s electrons in the outer energy level, and the stable configuration of $1s^22s^22p^63s^23p^63d^{10}$ results in a pseudo-noble gas configuration.

Negative Ion Formation

Nonmetals, which are located on the right side of the periodic table, easily gain electrons to attain a stable outer electron configuration. Examine **Figure 4**. To attain a noble-gas configuration, chlorine gains one electron, forming an ion with a 1– charge. After gaining the electron, the chloride ion has the electron configuration of an argon atom.

Chlorine atom (Cl)	$1s^22s^22p^63s^23p^5$
Argon atom (Ar)	$1s^22s^22p^63s^23p^6$
Chloride ion (Cl^-)	$1s^22s^22p^63s^23p^6$

An **anion** is a negatively charged ion. To designate an anion, the ending *-ide* is added to the root name of the element. Thus, a chlorine atom becomes a chloride anion.

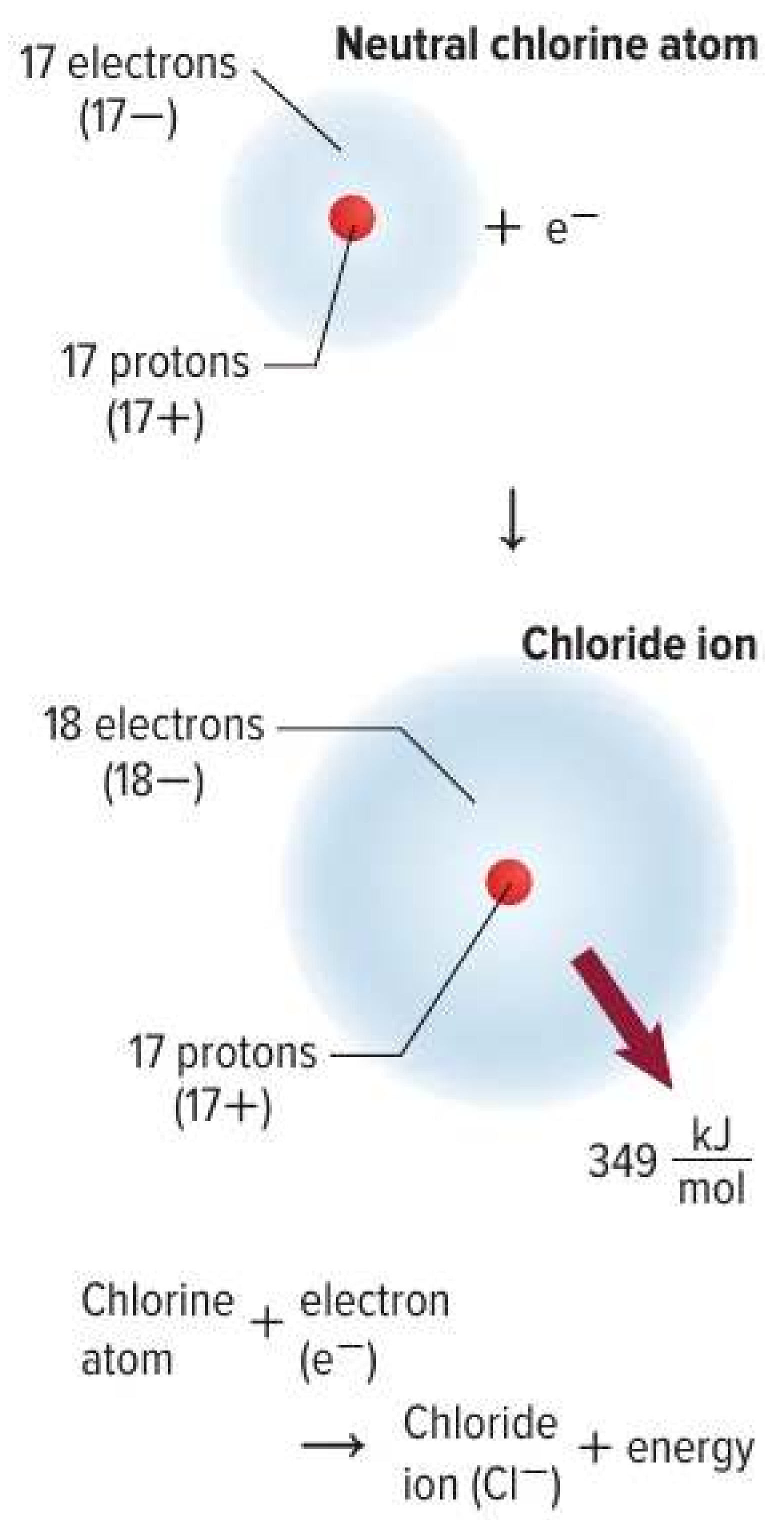

Figure 4 During the formation of the negative chloride ion, a neutral atom gains one electron. The process releases 349 kJ/mol of energy.

Compare *How do the energy changes accompanying positive ion and negative ion formation compare?*

Table 3 Group 15–17 Ions

Group	Configuration	Charge of Ion Formed
15	[noble gas]ns^2np^3	3– when three electrons are gained
16	[noble gas]ns^2np^4	2– when two electrons are gained
17	[noble gas]ns^2np^5	1– when one electron is gained

Nonmetal ions

As shown in **Table 3**, nonmetals gain the number of electrons that, when added to their valence electrons, equals 8. For example, consider phosphorus, with five valence electrons. To form a stable octet, the atom gains three electrons and forms a phosphide ion with a 3– charge. Likewise, oxygen, with six valence electrons, gains two electrons and forms an oxide ion with a 2– charge.

Some nonmetals can lose or gain other numbers of electrons to form an octet. For example, in addition to gaining three electrons, phosphorus can lose five. However, in general, group 15 elements gain three electrons, group 16 elements gain two, and group 17 elements gain one to achieve an octet.

Check Your Progress

Summary

- A chemical bond is the force that holds two atoms together.
- Some atoms form ions to gain stability. A stable configuration involves a complete outer energy level, usually consisting of eight valence electrons.
- Ions are formed by the loss or gain of valence electrons.
- The number of protons remains unchanged during ion formation.

Demonstrate Understanding

1. **Relate** the properties of atoms, their position in the periodic table, and their number of valence electrons to their chemical reactivity.
2. **Describe** two different causes of the force of attraction in a chemical bond.
3. **Apply** Why are all of the elements in group 18 relatively unreactive, whereas those in group 17 are very reactive?
4. **Summarize** ionic bond formation by correctly pairing these terms: cation, anion, electron gain, and electron loss.
5. **Apply** Write out the electron configuration for each atom. Then, predict the change that must occur in each to achieve a noble-gas configuration.
   a. nitrogen b. sulfur c. barium d. lithium
6. **Model** Draw models to represent the formation of the positive calcium ion and the negative bromide ion.

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LESSON 2

IONIC BONDS AND IONIC COMPOUNDS

FOCUS QUESTION

How are the ions in ionic compounds arranged?

Formation of an Ionic Bond

What do the reactions shown in **Figure 5** have in common? In both cases, elements react with each other to form a compound. **Figure 5a** shows the reaction between the elements sodium and chlorine. During this reaction, a sodium atom transfers its valence electron to a chlorine atom and becomes a positive ion. The chlorine atom accepts the electron into its outer energy level and becomes a negative ion. The oppositely charged ions attract each other, forming the compound sodium chloride. The electrostatic force that holds oppositely charged particles together in an ionic compound is referred to as an **ionic bond.** Compounds that contain ionic bonds are **ionic compounds.** If ionic bonds occur between metals and the nonmetal oxygen, oxides form. Most other ionic compounds are called salts.

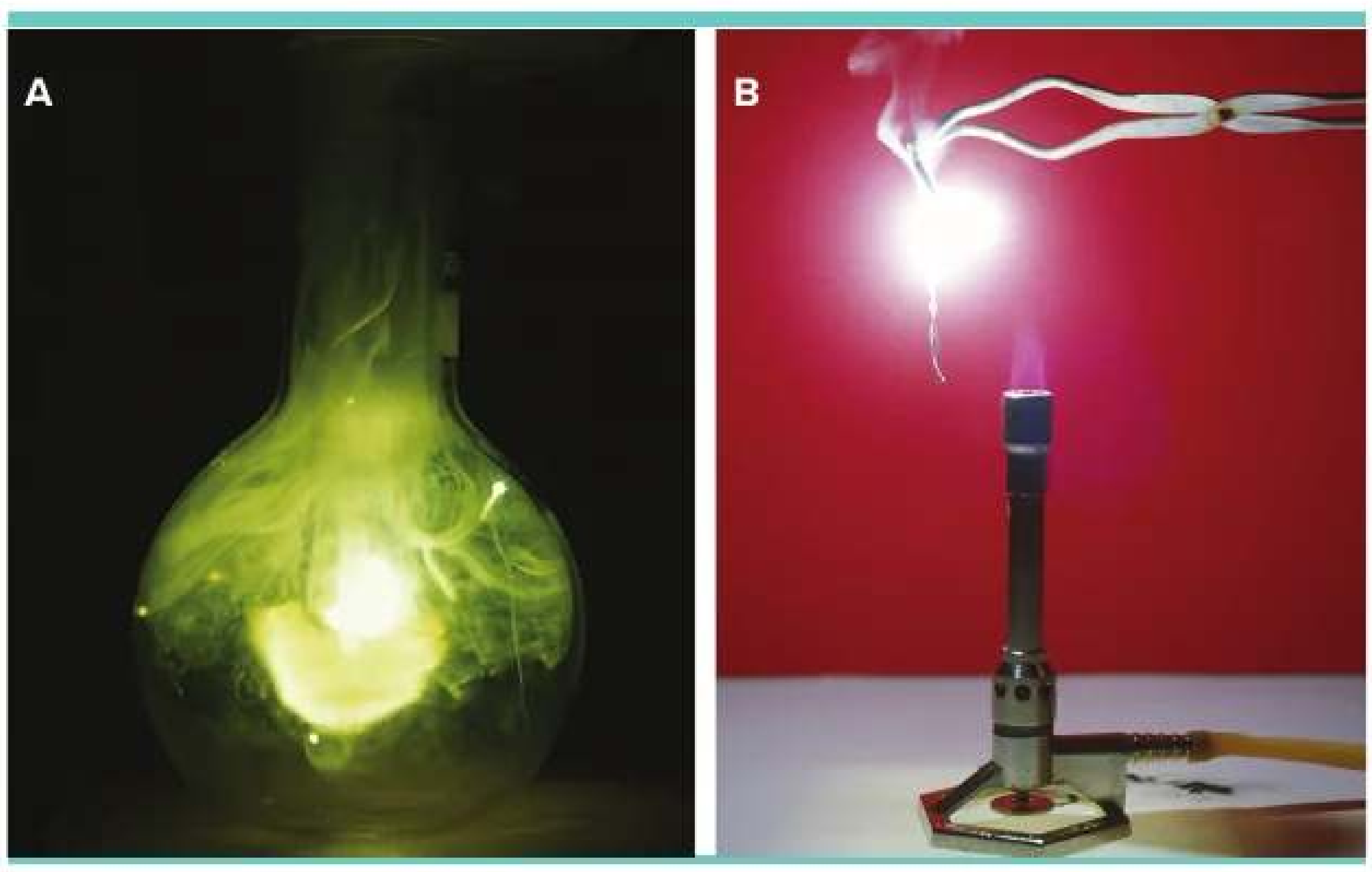


Figure 5 Each of these chemical reactions produces an ionic compound while releasing a large amount of energy. **a.** The reaction between elemental sodium and chlorine gas produces a white crystalline solid. **b.** When a ribbon of magnesium metal burns in air, it forms the ionic compound magnesium oxide.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

ChemLAB: Synthesize an Ionic Compound

Carry out an investigation to determine if a compound has ionic bonds based on physical properties.

Probeware Lab: Conductivity

Obtain, evaluate, and communicate information to discover the function of an electric current to determine the properties of a dissolved solid.

Binary ionic compounds

Thousands of compounds contain ionic bonds. Many ionic compounds are binary, which means that they contain only two different elements. Binary ionic compounds contain a metallic cation and a nonmetallic anion. Sodium chloride (NaCl) is a binary compound because it contains two different elements, sodium and chlorine. Magnesium oxide (MgO), the reaction product shown in **Figure 5b** on the previous page, is also a binary ionic compound.

Table 4 summarizes several ways in which the formation of an ionic compound such as sodium chloride can be represented.

Table 4 Formation of Sodium Chloride

Chemical Equation

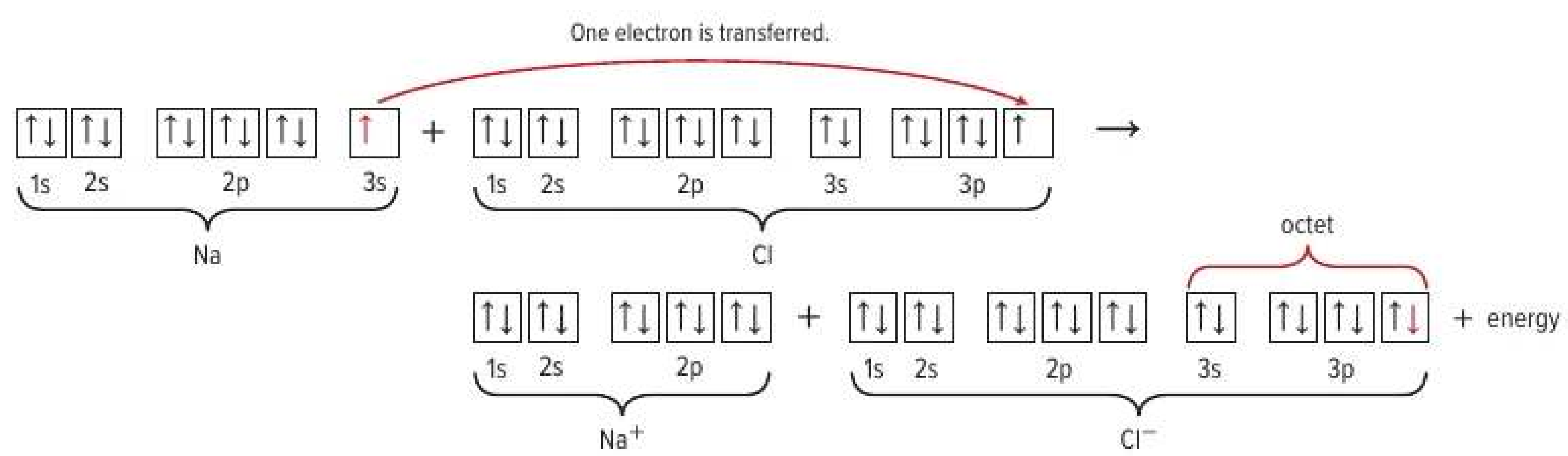
$$Na + Cl \rightarrow Na^+ + Cl^- + \text{energy}$$

Electron Configurations

One electron is transferred.

$$\underbrace{[Ne]3s^1}_{Na} + \underbrace{[Ne]3s^2 3p^5}_{Cl} \rightarrow \underbrace{[Ne]}_{Na^+} + \underbrace{[Ar]}_{Cl^-} + \text{energy}$$

Orbital Notation

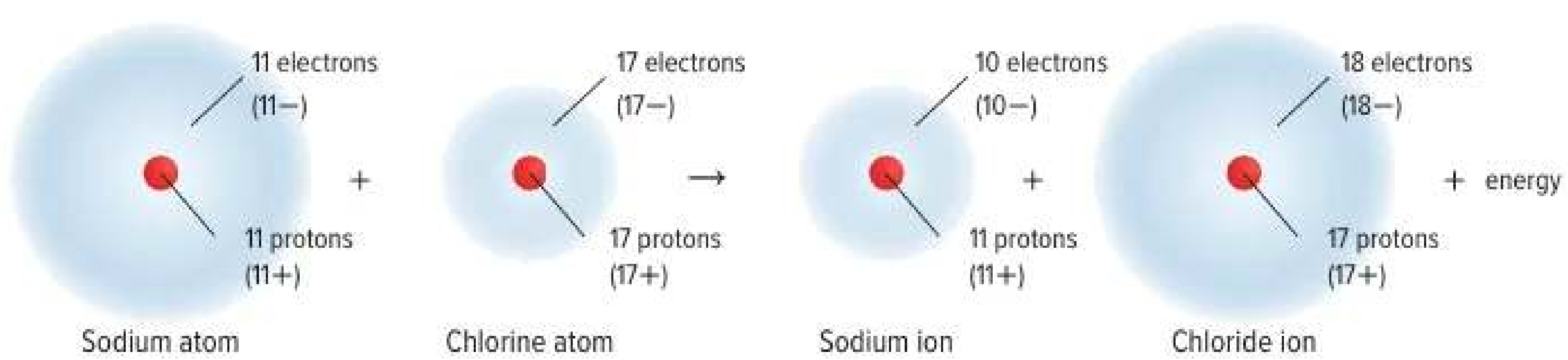


Electron-Dot Structures

One electron is transferred.

$$Na\cdot + \cdot\ddot{Cl}: \rightarrow [Na]^+ + [:\ddot{Cl}:]^- + \text{energy}$$

Atomic Models



What role does ionic charge play in the formation of ionic compounds? To answer this question, consider how calcium fluoride forms. Calcium has the electron configuration $[Ar]4s^2$, and needs to lose two electrons to attain the stable configuration of argon. Fluorine has the configuration $[He]2s^22p^5$, and must gain one electron to attain the stable configuration of neon. Because the number of electrons lost and gained must be equal, two fluorine atoms are needed to accept the two electrons lost from the calcium atom. As you can see, the overall charge of one unit of calcium fluoride (CaF_2) is zero.

$$1\ \cancel{\text{Ca ion}}\left(\frac{2+}{\cancel{\text{Ca ion}}}\right) + 2\ \cancel{\text{F ions}}\left(\frac{1-}{\cancel{\text{F ion}}}\right) = (1)(2+) + (2)(1-) = 0$$

Next, consider aluminum oxide, the whitish coating that forms on aluminum chairs. To acquire a noble-gas configuration, each aluminum atom loses three electrons and each oxygen atom gains two electrons. Thus, three oxygen atoms are needed to accept the six electrons lost by two aluminum atoms. The neutral compound formed is aluminum oxide (Al_2O_3).

$$2\ \cancel{\text{Al ions}}\left(\frac{3+}{\cancel{\text{Al ion}}}\right) + 3\ \cancel{\text{O ions}}\left(\frac{2-}{\cancel{\text{O ion}}}\right) = 2(3+) + 3(2-) = 0$$

Real-World Chemistry

Ionic Compounds

MINERAL SUPPLEMENTS To function properly, your body requires a daily intake of many different minerals. To ensure they are getting what they need, many people take a daily multivitamin and a mineral supplement. The minerals in these supplements come from a variety of ionic compounds. In fact, the majority of minerals found in mineral supplements come from ground-up rocks.

PRACTICE Problems ADDITIONAL PRACTICE

Explain how an ionic compound forms from these elements.

7. sodium and nitrogen

8. lithium and oxygen

9. strontium and fluorine

10. aluminum and sulfur

11. CHALLENGE Explain how elements in the two groups shown on the periodic table at the right combine to form an ionic compound.

Properties of Ionic Compounds

The chemical bonds in a compound determine many of its properties and applications. For ionic compounds, electrical forces in the ionic bonds produce unique physical structures, unlike those of other compounds. The physical structures of ionic compounds also contribute to their bulk physical properties.

Physical structure

In an ionic compound, large numbers of positive ions and negative ions exist together in a ratio determined by the number of electrons transferred from the metal atom to the nonmetal atom. These ions are packed into a regular repeating pattern that balances the forces of attraction and repulsion between the ions.

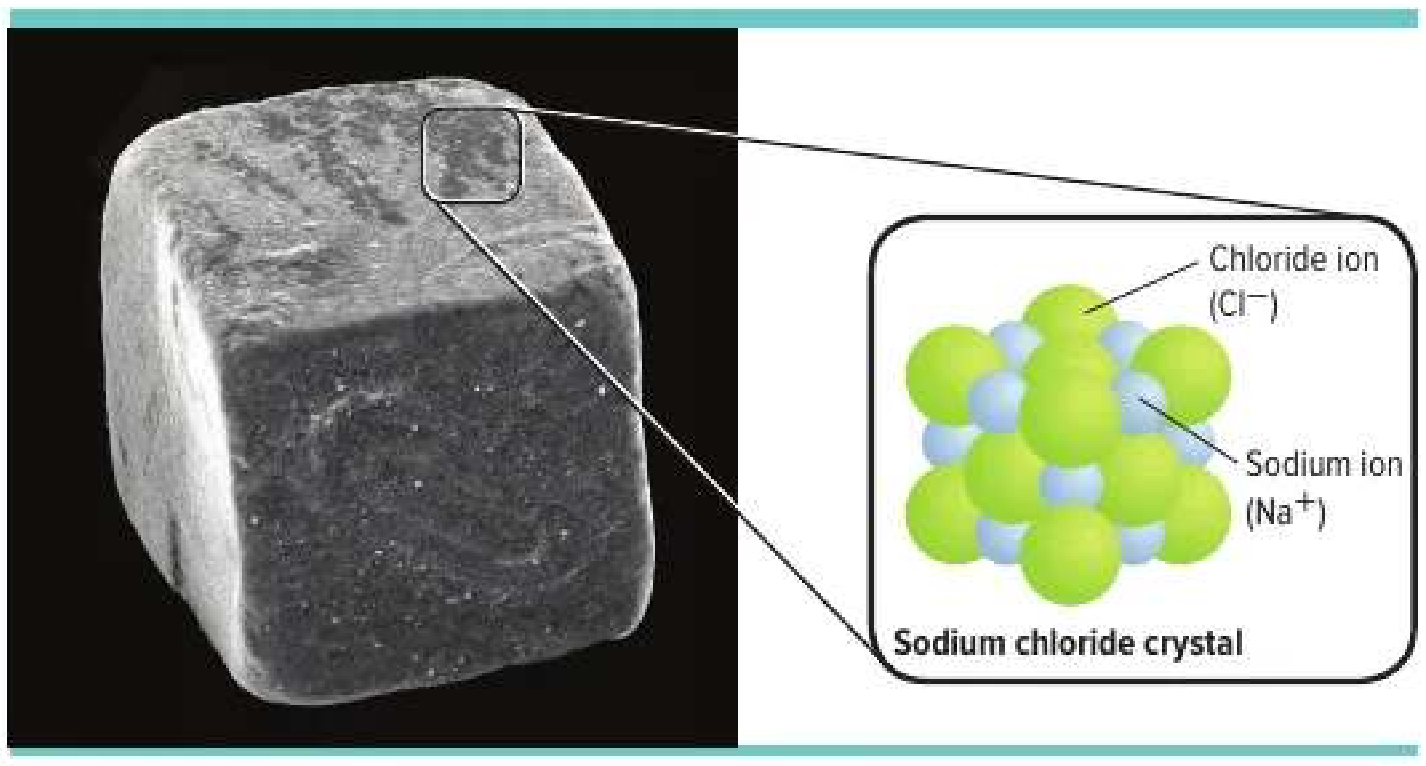


Figure 6 When viewed with a scanning electron microscope, the cubic shape of a sodium chloride crystal is visible. The structure of the crystal is highly ordered.

Compare *the shape of the crystal at the atomic scale to the shape of the crystal at the bulk scale.*

Examine the pattern of the ions in the sodium chloride crystal shown in **Figure 6**. Note the highly organized nature of an ionic crystal—the consistent spacing of the ions and the uniform pattern formed by them. Although the ion sizes are not the same, each sodium ion in the crystal is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. What shape would you expect a large crystal of this compound to be? As shown in **Figure 6**, the one-to-one ratio of sodium and chloride ions produces a highly ordered cubic crystal. As in all ionic compounds, in NaCl, no single unit consisting of only one sodium ion and one chloride ion is formed. Instead, large numbers of sodium ions and chloride ions exist together. If you can, obtain a magnifying lens and use it to examine some crystals of table salt (NaCl). What is the shape of these small salt crystals?

Get It?

Explain what determines the ratio of positive ions to negative ions in an ionic mineral crystal.

The strong attractions among the positive ions and the negative ions in an ionic compound result in the formation of a **crystal lattice.** A crystal lattice is a three-dimensional geometric arrangement of particles. In a crystal lattice, each positive ion is surrounded by negative ions, and each negative ion is surrounded by positive ions. Ionic crystals vary in shape due to the sizes and relative numbers of the ions bonded, as shown by the minerals in **Figure 7**.

Aragonite ($CaCO_3$) Barite ($BaSO_4$) Beryl ($Be_3Al_2Si_6O_{18}$)

Figure 7 Aragonite ($CaCO_3$), barite ($BaSO_4$), and beryl ($Be_3Al_2Si_6O_{18}$) are examples of minerals that are ionic compounds. The ions that form them are bonded together in a crystal lattice. Differences in ion size and charge result in different ionic crystal shapes, a topic that will be discussed later.

EARTH SCIENCE Connection The minerals in **Figure 7,** on the previous page, are just a few of the types studied by mineralogists, scientists who study minerals. They make use of several classification schemes to organize the thousands of known minerals. Color, crystal structure, hardness, chemical, magnetic, and electric properties, and numerous other characteristics are used to classify minerals. The types of anions minerals contain can also be used to identify them. For example, more than one-third of all known minerals are silicates, which are minerals that contain an anion that is a combination of silicon and oxygen. Halides contain fluoride, chloride, bromide, or iodide ions. Other mineral classes include boron-containing anions known as borates and carbon-oxygen containing anions known as carbonates.

Get It?

Identify In **Figure 7,** which mineral is a silicate, and which mineral is a carbonate? Can you classify any of the minerals in the video about crystal formation in caves?

Physical properties

Melting point, boiling point, and hardness are bulk physical properties of matter that are determined by the strength of electrical forces between particles that make up the matter. Because ionic bonds are relatively strong, ionic crystals require a large amount of energy to be broken apart. Thus, ionic crystals have high melting points and high boiling points, as shown in **Table 5**. Many crystals, including gemstones, have brilliant colors. These colors are due to the presence of transition metals in the crystal lattices.

Table 5 Melting and Boiling Points of Some Ionic Compounds

Compound	Melting Point (°C)	Boiling Point (°C)
NaI	660	1304
KBr	734	1435
NaBr	747	1390
$CaCl_2$	782	>1600
NaCl	801	1413
MgO	2852	3600

Ionic crystals are also hard, rigid, brittle solids due to the strong attraction between electric charges that holds the ions in place. When an external force is applied to the crystal—a force strong enough to overcome the attractive forces holding the ions in position within the crystal—the crystal cracks or breaks apart, as shown in **Figure 8**. The crystal breaks apart because the applied force repositions the like-charged ions next to each other; the resulting repulsion between electric forces breaks apart the crystal.

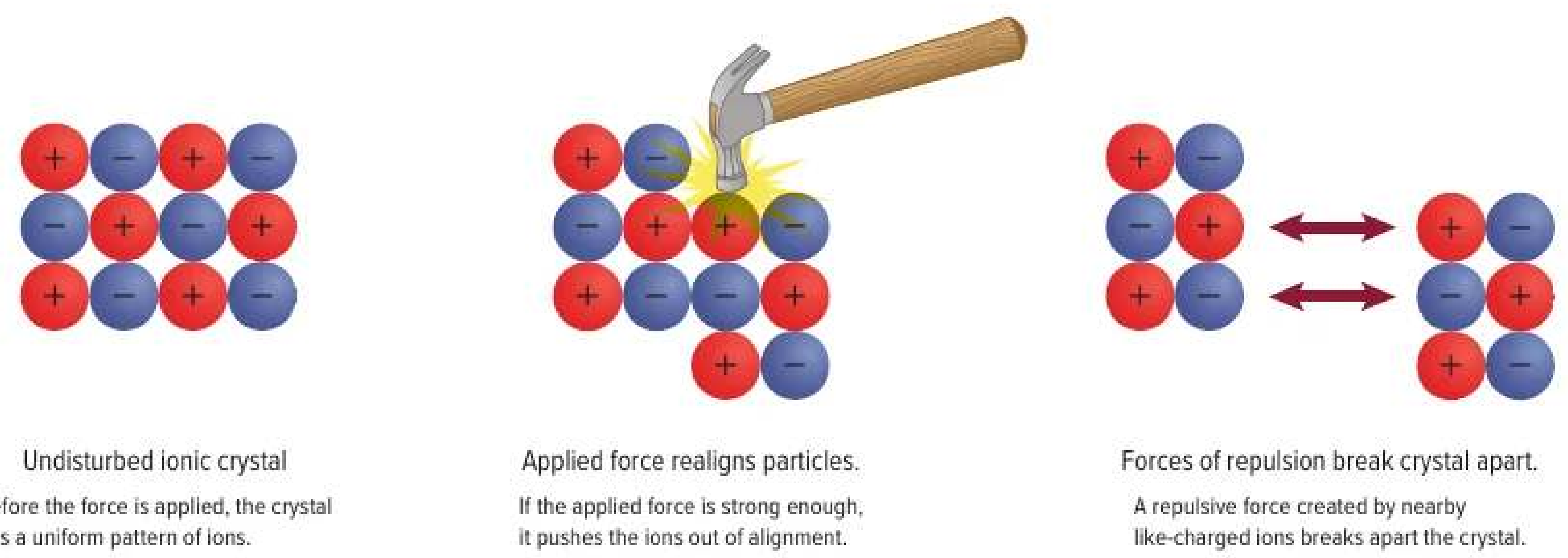


Figure 8 Strong attractive forces hold the ions in place until a force strong enough to overcome the attraction is applied.

Another property—the ability of a material to conduct electricity—depends on the availability of freely moving charged particles. Ions are charged particles, so whether they are free to move determines whether an ionic compound conducts electricity. In the solid state, the ions in an ionic compound are locked into fixed positions by strong attractive forces. As a result, ionic solids do not conduct electricity.

The situation changes dramatically, however, when an ionic solid melts to become a liquid or is dissolved in solution. The ions—previously locked in position—are now free to move and conduct an electric current. Both ionic compounds in solution and in the liquid state are excellent conductors of electricity. An ionic compound whose aqueous solution conducts an electric current is called an **electrolyte.**

Energy and the Ionic Bond

During every chemical reaction, energy is either absorbed or released. If energy is absorbed during a chemical reaction, the reaction is endothermic. If energy is released, it is exothermic. The formation of ionic compounds from positive ions and negative ions is always exothermic. The attraction of the positive ion for the negative ions close to it forms a more stable system that is lower in energy than the individual ions. If the amount of energy released during bond formation is reabsorbed, the bonds holding the positive ions and negative ions together will break apart.

Lattice energy

Because the ions in an ionic compound are arranged in a crystal lattice, the energy required to separate 1 mol of the ions of an ionic compound is referred to as the **lattice energy.** The strength of the electrical forces holding ions in place is reflected by the lattice energy. The greater the lattice energy, the stronger the force of attraction.

Lattice energy is directly related to the size of the ions bonded. Smaller ions form compounds with more closely spaced ionic charges. Because the electrostatic force of attraction between opposite charges increases as the distance between the charges decreases, smaller ions produce stronger attractions and greater lattice energies. For example, the lattice energy of a lithium compound is greater than that of a potassium compound with the same anion because a lithium ion is smaller than a potassium ion.

Get It?

Explain the relationship between lattice energy and the size of the ions in an ionic compound.

SCIENCE USAGE v. COMMON USAGE

conduct

Science usage: to transmit light, heat, sound, or electricity

The material did not conduct electricity well.

Common usage: to guide or lead

It was the manager's job to conduct the training session.

The value of lattice energy is also affected by the charge of the ion. The ionic bond formed from the attraction of ions with larger positive or negative charges generally has a greater lattice energy. The lattice energy of MgO is almost four times greater than that of NaF because the charge of the ions in MgO is greater than the charge of the ions in NaF. The lattice energy of $SrCl_2$ is between the lattice energies of MgO and NaF because $SrCl_2$ contains ions with both higher and lower charges.

Table 6 Lattice Energies of Some Ionic Compounds

Compound	Lattice Energy (kJ/mol)	Compound	Lattice Energy (kJ/mol)
KI	632	KF	808
KBr	671	AgCl	910
RbF	774	NaF	910
NaI	682	LiF	1030
NaBr	732	$SrCl_2$	2142
NaCl	769	MgO	3795

Table 6 shows the lattice energies of some ionic compounds. Examine the lattice energies of RbF and KF. Because K+ has a smaller ionic radius than Rb+, KF has a greater lattice energy than RbF. This confirms that lattice energy is related to ion size. Notice the lattice energies of $SrCl_2$ and AgCl. How do they show the relationship between lattice energy and the charge of the ions involved?

Check Your Progress

Summary

- Ionic compounds contain ionic bonds formed by the attraction of oppositely charged ions.
- Ions in an ionic compound are arranged in a repeating pattern known as a crystal lattice.
- Ionic compound properties are related to ionic bond strength.
- Ionic compounds conduct an electric current in the liquid phase and in aqueous solution.
- Lattice energy is the energy needed to remove 1 mol of ions from its lattice.

Demonstrate Understanding

12. **Explain** how an ionic compound made up of charged particles can be electrically neutral.
13. **Describe** the endothermic and exothermic energy changes associated with ionic bond formation, and relate these to stability.
14. **Identify** three physical properties of ionic compounds that are associated with ionic bonds, and relate them to bond strength.
15. **Explain** how ions form bonds, and describe the structure of the resulting compound.
16. **Relate** lattice energy to ionic-bond strength.
17. **Apply** Use electron configurations, orbital notation, and electron-dot structures to represent the formation of an ionic compound from the metal strontium and the nonmetal chlorine.
18. **Design** a concept map that shows the relationships among ionic bond strength, physical properties of ionic compounds, lattice energy, and stability.

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

LESSON 3

NAMES AND FORMULAS FOR IONIC COMPOUNDS

FOCUS QUESTION

What are the names and formulas of ionic compounds?

Formulas for Ionic Compounds

Because chemists around the world need to be able to communicate with one another, they have developed a set of rules for naming compounds. Using this standardized naming system, you can write a chemical formula from a compound's name and name a compound given its chemical formula.

Recall that an ionic compound is made up of ions arranged in a repeating pattern. The chemical formula for an ionic compound, called a **formula unit,** represents the simplest ratio of the ions involved. For example, the formula unit of magnesium chloride is $MgCl_2$ because the magnesium and chloride ions exist in a 1:2 ratio.

The overall charge of a formula unit is zero because the formula unit represents the entire crystal, which is electrically neutral. The formula unit for $MgCl_2$ contains one Mg^{2+} ion and two Cl^- ions, for a total charge of zero.

Monatomic ions

Binary ionic compounds are composed of positively charged monatomic ions of a metal and negatively charged monatomic ions of a nonmetal. A **monatomic ion** is a one-atom ion, such as Mg^{2+} or Br^-. **Table 7** indicates the charges of common monatomic ions according to their location on the periodic table. What is the formula for the beryllium ion? The iodide ion? The nitride ion?

Table 7 Common Monatomic Ions

Group	Atoms that Commonly Form Ions	Charge of Ions
1	H, Li, Na, K, Rb, Cs	1+
2	Be, Mg, Ca, Sr, Ba	2+
15	N, P, As	3−
16	O, S, Se, Te	2−
17	F, Cl, Br, I	1−

Transition metals, which are in groups 3 through 12, and metals in groups 13 and 14 are not included in **Table 7** because of the variance in ionic charges of atoms in the groups. Most transition metals and metals in groups 13 and 14 can form several different positive ions.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Review the News

Obtain information from a current news story about recent additions to the periodic table. Evaluate your source and communicate your findings to your class.

Oxidation numbers

The oxidation number of a monatomic ion, also called its oxidation state, is equal to the net charge of the ion. As shown in **Table 8**, most transition metals and group 13 and 14 metals have more than one possible ionic charge. Note that the ionic charges given in the table are the most common ones, but not the only ones possible.

The oxidation number of an element in an ionic compound equals the number of electrons transferred from the atom to form the ion. For example, a sodium atom transfers one electron to a chlorine atom to form sodium chloride. This results in Na^+ and Cl^-. Thus, the oxidation number of sodium in the compound is 1+ because one electron was transferred from the sodium atom. Because an electron is transferred to the chlorine atom, its oxidation number is 1–.

Table 8 Monatomic Metal Ions

Group	Common Ions
3	Sc^{3+}, Y^{3+}, La^{3+}
4	Ti^{2+}, Ti^{3+}
5	V^{2+}, V^{3+}
6	Cr^{2+}, Cr^{3+}
7	Mn^{2+}, Mn^{3+}, Tc^{2+}
8	Fe^{2+}, Fe^{3+}
9	Co^{2+}, Co^{3+}
10	Ni^{2+}, Pd^{2+}, Pt^{2+}, Pt^{4+}
11	Cu^{+}, Cu^{2+}, Ag^{+}, Au^{+}, Au^{3+}
12	Zn^{2+}, Cd^{2+}, Hg_2^{2+}, Hg^{2+}
13	Al^{3+}, Ga^{2+}, Ga^{3+}, In^{+}, In^{2+}, In^{3+}, Tl^{+}, Tl^{3+}
14	Sn^{2+}, Sn^{4+}, Pb^{2+}, Pb^{4+}

Formulas for binary ionic compounds

In the chemical formula for any ionic compound, the symbol of the cation is always written first, followed by the symbol of the anion. Subscripts, which are small numbers to the lower right of a symbol, represent the number of ions of each element in an ionic compound. If no subscript is written, it is assumed to be one. You can use oxidation numbers to write formulas for ionic compounds. Recall that ionic compounds have no charge. If you add the oxidation number of each ion multiplied by the number of these ions in a formula unit, the total must be zero.

Suppose you need to determine the formula for one formula unit of the compound that contains sodium and fluoride ions. Start by writing the symbol and charge for each ion: Na^+ and F^-. The ratio of ions in a formula unit of the compound must show that the number of electrons lost by the metal equals the number of electrons gained by the nonmetal. This occurs when one sodium atom transfers one electron to the fluorine atom; the formula unit is NaF.

Get It?

Relate the charge of an ion to its oxidation number.

STEM CAREER Connection

Pharmacist

Would you like using chemistry to help people treat disease? Pharmacists prepare and dispense medications to patients and provide expertise on the safe and proper use of medications. They may work in retail pharmacies or in hospitals. Pharmacists also conduct health screenings, give immunizations such as flu shots, and advise patients on general health topics such as diet, exercise, and managing stress.

ACADEMIC VOCABULARY

transfer

to cause to pass from one to another

Carlos had to transfer to a new school when his parents moved to a new neighborhood.

EXAMPLE Problem 1

FORMULA FOR AN IONIC COMPOUND Determine the formula for the ionic compound formed from potassium and oxygen.

1 ANALYZE THE PROBLEM

You are given that potassium and oxygen ions form an ionic compound, the formula for the compound is the unknown. First, write out the symbol and oxidation number for each ion involved in the compound. Potassium, from group 1, forms 1+ ions, and oxygen, from group 16, forms 2− ions.

K^+ O^{2-}

Because the charges are not the same, you need to determine the subscripts to use to indicate the ratio of positive ions to negative ions.

2 SOLVE FOR THE UNKNOWN

A potassium atom loses one electron. while an oxygen atom gains two electrons. If combined in a one-to-one ratio, the number of electrons lost by potassium will not balance the number of electrons gained by oxygen. Thus, two potassium ions are needed for each oxide ion. The formula is **K_2O.**

3 EVALUATE THE ANSWER

The overall charge of the compound is zero.

$$2\ \cancel{\text{K ions}}\left(\frac{1+}{\cancel{\text{K ion}}}\right) + 1\ \cancel{\text{O ion}}\left(\frac{2-}{\cancel{\text{O ion}}}\right) = 2(1+) + 1(2-) = 0$$

EXAMPLE Problem 2

FORMULA FOR AN IONIC COMPOUND Determine the formula for the compound formed from aluminum ions and sulfide ions.

1 ANALYZE THE PROBLEM

You are given that aluminum and sulfur form an ionic compound; the formula for the ionic compound is the unknown. First, determine the charges of each ion. Aluminum, from group 13, forms 3+ ions, and sulfur, from group 16, forms 2— ions.

Al^{3+} S^{2-}

Each aluminum atom loses three electrons, while each sulfur atom gains two electrons. The number of electrons lost must equal the number of electrons gained.

2 SOLVE FOR THE UNKNOWN

The smallest number that can be divided evenly by both 2 and 3 is 6. Therefore, six electrons are transferred. Three sulfur atoms accept the six electrons lost by two aluminum atoms. The correct formula, **Al_2S_3,** shows two aluminum ions bonded to three sulfide ions.

3 EVALUATE THE ANSWER

The overall charge of one formula unit of this compound is zero.

$$2\ \cancel{\text{Al ions}}\left(\frac{3+}{\cancel{\text{Al ion}}}\right) + 3\ \cancel{\text{S ions}}\left(\frac{2-}{\cancel{\text{S ion}}}\right) = 2(3+) + 3(2-) = 0$$

PRACTICE Problems ADDITIONAL PRACTICE

Write formulas for the ionic compounds formed by the following ions. Use units as a guide to your solutions.

19. potassium and iodide
20. magnesium and chloride
21. aluminum and bromide
22. cesium and nitride
23. **CHALLENGE** Write the general formula for the ionic compound formed by elements from the two groups shown on the periodic table at the right.

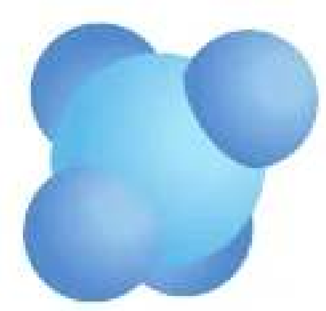
Formulas for polyatomic ionic compounds

Many ionic compounds contain **polyatomic ions,** which are ions made up of more than one atom. **Table 9** and **Figure 9** list some common polyatomic ions. Also see **Table R-5** in the Student Resources. A polyatomic ion acts as an individual ion in a compound and its charge applies to the entire group of atoms. Thus, the formula for a polyatomic compound follows the same rules used for a binary compound.

Because a polyatomic ion exists as a unit, never change subscripts of the atoms within the ion. If more than one polyatomic ion is needed, place parentheses around the ion and write the appropriate subscript outside the parentheses. For example, consider the compound formed from the ammonium ion (NH_4^+) and the oxide ion (O^{2-}). To balance the charges, the compound must have two ammonium ions for each oxide ion. To add a subscript to ammonium, enclose it in parentheses, then add the subscript. The correct formula is $(NH_4)_2O$.

Table 9 Common Polyatomic Ions

Ion	Name	Ion	Name
NH_4^+	ammonium	IO_4^-	periodate
NO_2^-	nitrite	$C_2H_3O_2^-$	acetate
NO_3^-	nitrate	$H_2PO_4^-$	dihydrogen phosphate
OH^-	hydroxide	CO_3^{2-}	carbonate
CN^-	cyanide	SO_3^{2-}	sulfite
MnO_4^-	permanganate	SO_4^{2-}	sulfate
HCO_3^-	hydrogen carbonate	$S_2O_3^{2-}$	thiosulfate
ClO^-	hypochlorite	O_2^{2-}	peroxide
ClO_2^-	chlorite	CrO_4^{2-}	chromate
ClO_3^-	chlorate	$Cr_2O_7^{2-}$	dichromate
ClO_4^-	perchlorate	HPO_4^{2-}	hydrogen phosphate
BrO_3^-	bromate	PO_4^{3-}	phosphate
IO_3^-	iodate	AsO_4^{3-}	arsenate


Ammonium ion (NH_4^+)

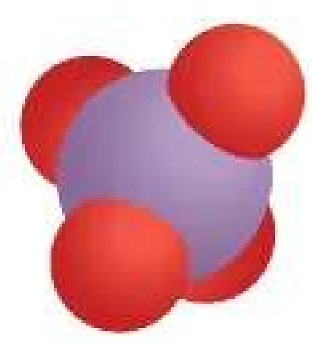

Phosphate ion (PO_4^{3-})

Figure 9 Ammonium and phosphate ions are polyatomic; that is, they are made up of more than one atom. Each polyatomic ion, however, acts as a single unit and has one charge.

Identify *What are the charges of the ammonium ion and phosphate ion, respectively?*

EXAMPLE Problem 3

FORMULA FOR A POLYATOMIC IONIC COMPOUND A compound formed by calcium ions and phosphate ions is often used In fertilizers. Write the compound's formula.

1 ANALYZE THE PROBLEM

You know that calcium and phosphate ions form an ionic compound; the formula for the compound is the unknown. First, write each ion along with its charge. Calcium, from group 2, forms 2+ ions, and the polyatomic phosphate acts as a single unit with a 3− charge.

$$Ca^{2+} \qquad PO_4^{3-}$$

Each calcium atom loses two electrons, while each polyatomic phosphate group gains three electrons. The number of electrons lost must equal the number of electrons gained.

2 SOLVE FOR THE UNKNOWN

The smallest number evenly divisible by both charges is 6. Thus, a total of six electrons are transferred. The negative charge from two phosphate ions equals the positive charge from three calcium ions. In the formula, place the polyatomic ion in parentheses and add a subscript to the outside. The correct formula for the compound is $Ca_3(PO_4)_2$.

3 EVALUATE THE ANSWER

The overall charge of one formula unit of calcium phosphate is zero.

$$3\ \cancel{\text{Ca ions}}\left(\frac{2+}{\cancel{\text{Ca ion}}}\right) + 2\ \cancel{\text{PO}_4\text{ ions}}\left(\frac{3-}{\cancel{\text{PO}_4\text{ ion}}}\right) = 3(2+) + 2(3-) = 0$$

PRACTICE Problems ADDITIONAL PRACTICE

Write formulas for ionic compounds composed of the following ions.

Use units as a guide to your solutions.

24. sodium and nitrate

25. calcium and chlorate

26. aluminum and carbonate

27. **CHALLENGE** Write the formula for an ionic compound formed by ions from a group 2 element and polyatomic ions composed of only carbon and oxygen.

Names for Ions and Ionic Compounds

Scientists use a systematic approach when naming ionic compounds. Because ionic compounds have both cations and anions, the naming system accounts for both of these ions.

Naming an oxyanion

An **oxyanion** is a polyatomic ion composed of an element, usually a nonmetal, bonded to one or more oxygen atoms. More than one oxyanion exists for some nonmetals, such as nitrogen and sulfur. These ions are easily named using the rules in **Table 10**.

Table 10 Oxyanion Naming Conventions for Sulfur and Nitrogen

• Identify the ion with the greatest number of oxygen atoms. This ion is named using the root of the nonmetal and the suffix *-ate*.				
• Identify the ion with fewer oxygen atoms. This ion is named using the root of the nonmetal and the suffix *-ite*.				
Examples:	NO_3^- nitr*ate*	NO_2^- nitr*ite*	SO_4^{2-} sulf*ate*	SO_3^{2-} sulf*ite*

As shown in **Table 11**, chlorine forms four oxyanions that are named according to the number of oxygen atoms present. Names of similar oxyanions formed by other halogens follow the rules used for chlorine. For example, bromine forms the bromate ion (BrO_3^-), and iodine forms the periodate ion (IO_4^-) and the iodate ion (IO_3^-).

Table 11 Oxyanion Naming Conventions for Chlorine

• The oxyanion with the greatest number of oxygen atoms is named using the prefix *per-*, the root of the nonmetal, and the suffix *-ate*.
• The oxyanion with one fewer oxygen atom is named using the root of the nonmetal and the suffix *-ate*.
• The oxyanion with two fewer oxygen atoms is named using the root of the nonmetal and the suffix *-ite*.
• The oxyanion with three fewer oxygen atoms is named using the prefix hypo-, the root of the nonmetal, and the suffix *-ite*.
Examples: ClO_4^- perchlorate ClO_3^- chlorate ClO_2^- chlorite ClO^- hypochlorite

Naming ionic compounds

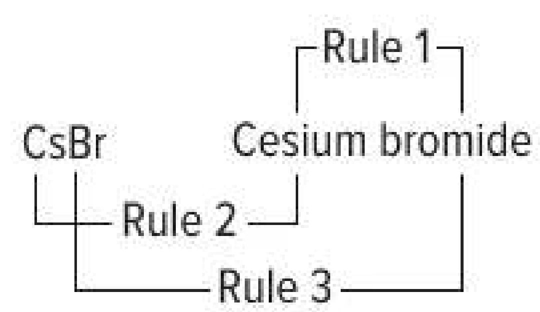
Chemical nomenclature is a systematic way of naming compounds. Now that you are familiar with chemical formulas, you can use the following five rules to name ionic compounds.

1. Name the cation followed by the anion. Remember that the cation is always written first in the formula.

2. For monatomic cations, use the element name.

3. For monatomic anions, use the root of the element name plus the suffix *-ide*.

Example:



4. To distinguish between multiple oxidation numbers of the same element, the name of the chemical formula must indicate the oxidation number of the cation. The oxidation number is written as a Roman numeral in parentheses after the name of the cation.

Note: This rule applies to the transition metals and metals on the right side of the periodic table, which often have more than one oxidation number. See **Table 8** earlier in this lesson. It does not apply to group 1 and group 2 cations, as they have only one oxidation number.

Examples: Fe^{2+} and O_2^- ions form FeO, known as iron(II) oxide.
Fe^{3+} and O_2^- ions form Fe_2O_3, known as iron(III) oxide.

5. When the compound contains a polyatomic ion, simply use the name of the polyatomic ion in place of the anion or cation.

Examples: The name for NaOH is sodium hydroxide.
The name for $(NH_4)_2S$ is ammonium sulfide.

PRACTICE Problems ADDITIONAL PRACTICE

Interpret the formula representations and name these compounds.

28. NaBr

29. $CaCl_2$

30. KOH

31. $Cu(NO_3)_2$

32. Ag_2CrO_4

33. CHALLENGE The ionic compound NH_4ClO_4 is a key reactant used in solid rocket boosters, such as those that powered the Space Shuttle into orbit. Name this compound.

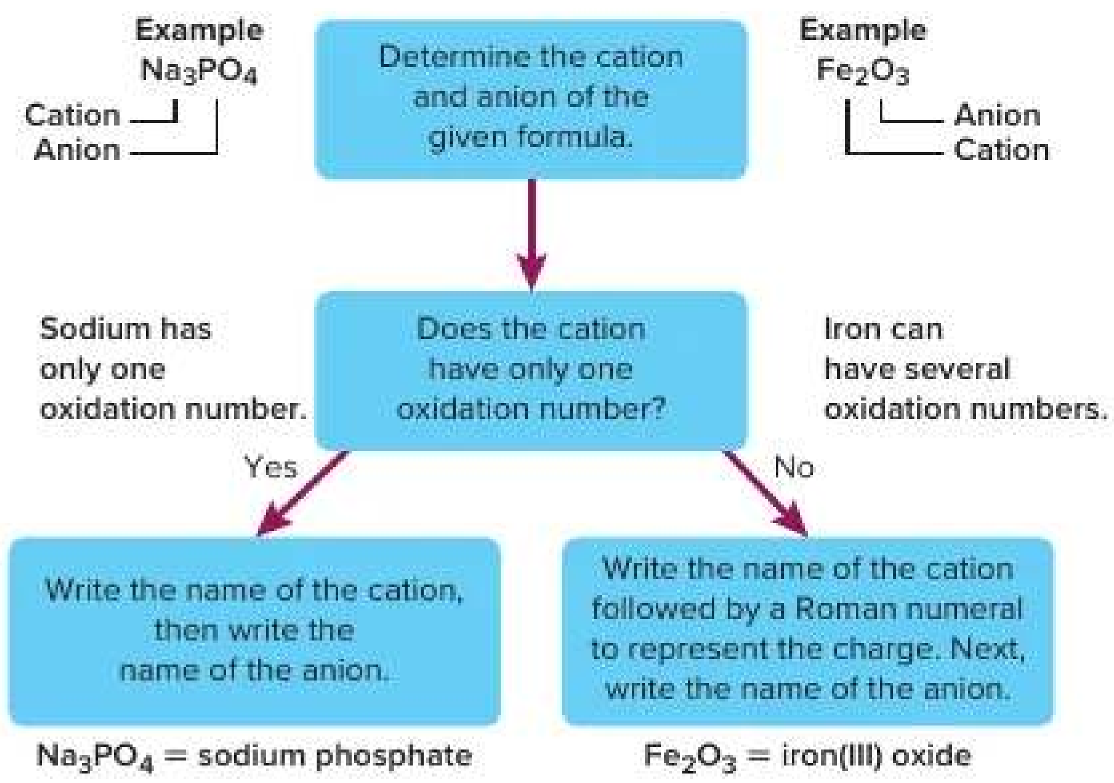
PROBLEM-SOLVING STRATEGY

Naming Ionic Compounds

Naming ionic compounds is easy if you follow this naming-convention flowchart.

Apply the Strategy

Name the compounds KOH and Ag_2CrO_4 using this flowchart.



Naming ionic compounds is important in communicating the cation and anion in a crystalline solid or aqueous solution. The Problem-Solving Strategy reviews the steps used in naming ionic compounds if the formula is known. How might you change the diagram to help you write the formulas for ionic compounds if you know their names?

Check Your Progress

Summary

- A formula unit gives the ratio of cations to anions in the ionic compound.
- A monatomic ion is formed from one atom. The charge of a monatomic ion is equal to its oxidation number.
- Roman numerals indicate the oxidation number of cations having multiple possible oxidation states.
- Polyatomic ions consist of more than one atom and act as a single unit.
- To indicate more than one polyatomic ion in a chemical formula, place parentheses around the polyatomic ion and use a subscript.

Demonstrate Understanding

34. **State** the order in which the ions associated with a compound composed of potassium and bromine would be written in the chemical formula and the compound name.
35. **Determine** the difference between a monatomic ion and a polyatomic ion, and give an example of each.
36. **Apply** Ion X has a charge of 2+, and ion Y has a charge of 1−. Write the formula unit of the compound formed from the ions.
37. **State** the name and formula for the compound formed from Mg and Cl.
38. **Write** the name and formula for the compound formed from sodium ions and nitrite ions.
39. **Analyze** What subscripts would you most likely use if the following substances formed an ionic compound?
    a. an alkali metal and a halogen
    b. an alkali metal and a nonmetal from group 16
    c. an alkaline earth metal and a halogen
    d. an alkaline earth metal and a nonmetal from group 16

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

LESSON 4

METALLIC BONDS AND THE PROPERTIES OF METALS

FOCUS QUESTION

How do metals and ionic compounds compare and contrast?

Metallic Bonds

Although metals are not ionic, they share several properties with ionic compounds. The bonding in both metals and ionic compounds is based on the attraction of particles with unlike charges. In the solid state, metals often form lattices similar to ionic crystal lattices. In such a lattice, 8 to 12 other metal atoms closely surround each metal atom.

A sea of electrons

In a metallic lattice, metal atoms do not share their valence electrons with neighboring atoms, nor do they lose their valence electrons. Instead, the outer energy levels of the metal atoms overlap, as shown in **Figure 10**. This unique arrangement is described by the **electron sea model,** which proposes that all the metal atoms in a metallic solid contribute their valence electrons to form a "sea" of electrons that surrounds the metal cations in the lattice. The electrons present in the outer energy levels of the bonding metallic atoms are not held by any specific atom and can move easily from one atom to the next. Because they are free to move, they are often referred to as **delocalized electrons.** When the atom's outer electrons move freely throughout the solid, a metallic cation is formed. Each such ion is bonded to the lattice by the sea of valence electrons. A **metallic bond** is the attraction of a metallic cation for delocalized electrons.

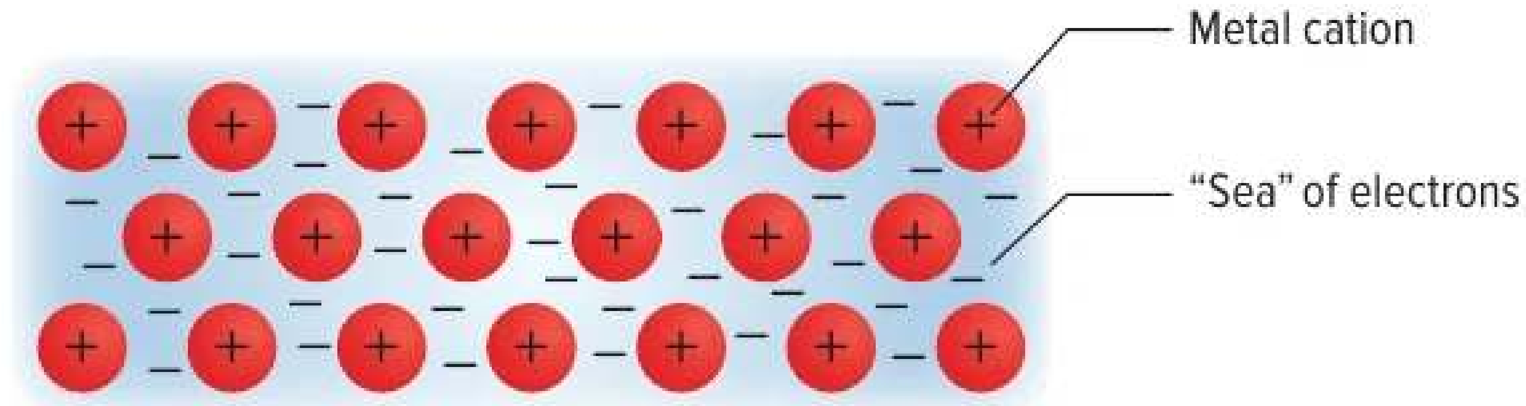


Figure 10 The valence electrons in metals (shown as a **blue cloud** of minus signs) are evenly distributed among the metallic cations (**shown in red**). Attractions between positive cations and the negative "sea" hold the metal atoms together in a lattice.

Explain *Why are electrons in metals known as delocalized electrons?*

3D THINKING | DCI Disciplinary Core Ideas | CCC Crosscutting Concepts | SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Modeling Electrostatic Forces—Ionic and Metallic Bonding

HS-PS2-4. Use mathematical representations of Newton's Law of Gravitation and Coulomb's Law to describe and predict the gravitational and electrostatic forces between objects.

Quick Investigation: Observe Properties

Construct an explanation of the structures and malleable properties of steel.

Properties of metals

The physical properties of metals at the bulk scale can be explained by metallic bonding. These properties provide evidence of the strength of metallic bonds.

Table 12 Melting and Boiling Points

Element	Melting Point (°C)	Boiling Point (°C)
Lithium	180	1342
Tin	232	2602
Aluminum	660	2519
Barium	727	1897
Silver	962	2162
Copper	1085	2562

Melting and boiling points The melting points of metals vary greatly. Mercury is a liquid at room temperature, which makes it useful in scientific instruments such as thermometers and barometers. On the other hand, tungsten has a melting point of 3422°C. Lightbulb filaments are usually made from tungsten, as are certain spacecraft parts. In general, metals have moderately high melting points and high boiling points, as shown in **Table 12**. The melting points are less extreme than the boiling points because the cations and electrons are mobile in a metal, so it does not take much energy for them to be able to move past each other. However, during boiling, atoms must be completely separated from the lattice, which requires much more energy.

Thermal conductivity and electrical conductivity The movement of mobile electrons around positive metallic cations makes metals good conductors. The delocalized electrons move heat from one place to another much more quickly than the electrons in a material that does not contain mobile electrons. Mobile electrons easily move as part of an electric current when an electric potential is applied to a metal. These same delocalized electrons interact with light, absorbing and releasing photons, thereby creating the property of luster in metals.

Malleability, ductility, and durability Metals are malleable, which means they can be hammered into sheets, and they are ductile, which means they can be drawn into wire. **Figure 11** shows how the mobile particles involved in metallic bonding can be pushed or pulled past each other. Metals are generally durable. Although metallic cations are mobile in a metal, they are strongly attracted to the electrons surrounding them and are not easily removed from the metal.

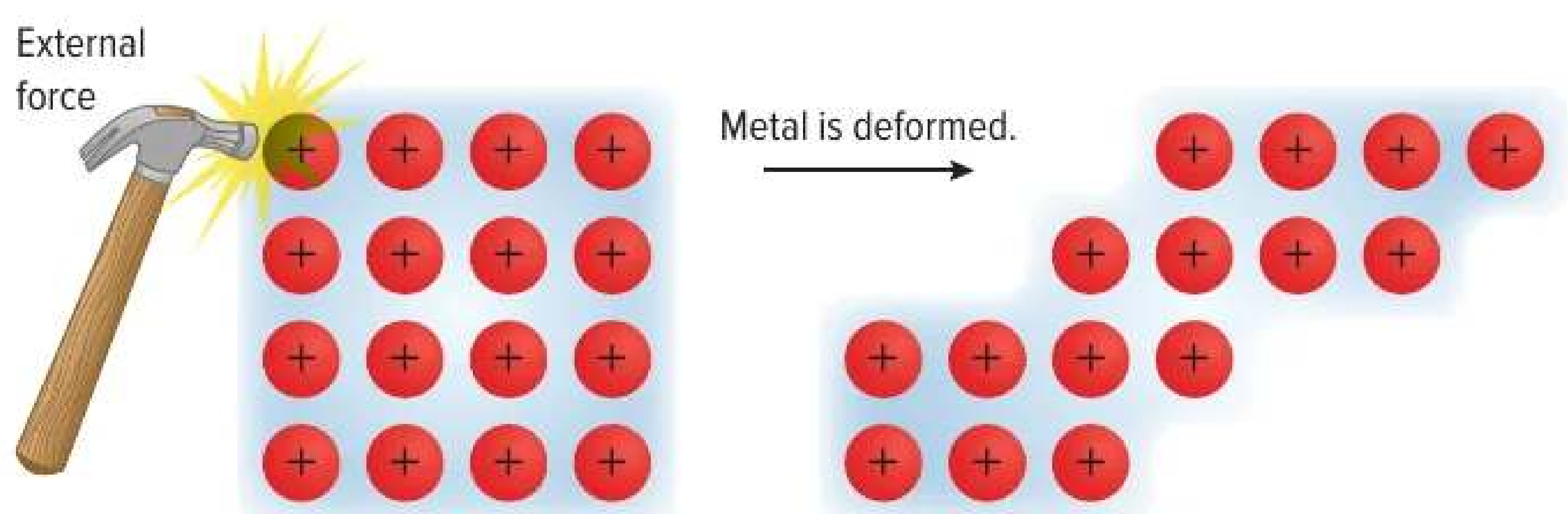


Figure 11 An applied force causes metal ions to move through delocalized electrons, making metals malleable and ductile.

Describe *in your own words what happens to metal ions when a metal is struck with a hammer.*

Hardness and strength The mobile electrons in transition metals consist not only of the two outer s electrons but also of the inner d electrons. As the number of delocalized electrons increases, so do the properties of hardness and strength. For example, strong metallic bonds are found in transition metals such as chromium, iron, and nickel, whereas alkali metals are considered soft because they have only one delocalized electron, ns^1.

Get It?
Contrast the behavior of metals and ionic compounds when each is struck by a hammer.

Metal Alloys

Due to the nature of metallic bonds, it is relatively easy to introduce other elements into the metallic crystal, forming an alloy. An **alloy** is a mixture of elements that has metallic properties. Because of their unique blend of properties, alloys have a wide range of commercial applications. Stainless steel, brass, and cast iron are a few of the many useful alloys. **Table 13** lists some commercially important alloys and their uses.

You likely have items made of alloys in your home, but you might not have realized it. Magnets often are used to attach decorative items or mementos to refrigerator doors. Candlesticks made of brass are common home accessories. Yellow gold jewelry is made of an alloy of copper and gold. Silver gold jewelry often is made of an alloy of gold with silver or palladium. Stainless steel often is used to make kitchen sinks, appliance doors and exposed sides, and tableware.

Table 13 Commercial Alloys

Common Name	Composition	Uses
Alnico	Fe 50%, Al 20%, Ni 20%, Co 10%	magnets
Brass	Cu 67–90%, Zn 10–33%	plumbing, hardware, lighting
Bronze	Cu 70–95%, Zn 1–25%, Sn 1–18%	bearings, bells, medals
Cast iron	Fe 96–97%, C 3–4%	casting
Gold, 10-carat	Au 42%, Ag 12–20%, Cu 37.46%	jewelry
Lead shot	Pb 99.8%, As 0.2%	shotgun shells
Pewter	Sn 70–95%, Sb 5–15%, Pb 0–15%	tableware
Stainless steel	Fe 73–79%, Cr 14–18%, Ni 7–9%	instruments, sinks
Sterling silver	Ag 92.5%, Cu 7.5%	tableware, jewelry

CCC CROSSCUTTING CONCEPTS
Structure and Function Make a list of the properties of metals discussed in this lesson. For each item on the list, explain how that property could be seen as evidence of the structure of metals at the atomic scale.

WORD ORIGIN
alloy
comes from the Latin word *alligare*, which means *to bind*

Figure 12 Bicycle frames are sometimes made of 3/2.5 titanium alloy, an alloy of titanium containing 3% aluminum and 2.5% vanadium.

Properties of alloys

The properties of alloys differ somewhat from the properties of the elements they contain. For example, steel is iron mixed with at least one other element. Some properties of iron are present, but steel has additional properties, such as increased strength. Alloys are classified into two basic types, substitutional alloys and interstitial alloys.

Substitutional alloys In a substitutional alloy, some of the atoms in the original metal are replaced by other metals of similar atomic size. Sterling silver is an example of a substitutional alloy. In sterling silver, copper atoms replace some of the silver atoms in the metallic crystal. The resulting solid has properties of both silver and copper.

Interstitial alloys An interstitial alloy, such as the titanium alloy shown in **Figure 12,** is formed when the small holes (interstices) in a metallic crystal are filled with smaller atoms. The best-known interstitial alloy is carbon steel, in which holes in an iron crystal are filled with carbon atoms. Iron alone is relatively soft and malleable, but the added carbon makes the solid harder, stronger, and less ductile.

Check Your Progress

Summary

- A metallic bond forms when metal cations attract freely moving, delocalized valence electrons.
- In the electron sea model, electrons move through the metallic crystal and are not held by any particular atom.
- The electron sea model explains the physical properties of metallic solids.
- Metal alloys are formed when a metal is mixed with one or more other elements.

Demonstrate Understanding

40. **Contrast** the structures of ionic compounds and metals.
41. **Explain** how the conductivity of electricity and the high boiling points of metals are explained by metallic bonding.
42. **Contrast** the cause of the attraction in ionic bonds and metallic bonds.
43. **Summarize** alloy types by correctly pairing these terms and phrases: substitutional, interstitial, replaced, and filled in.
44. **Design** an experiment that could be used to distinguish between a metallic solid and an ionic solid. Include at least two different methods for comparing the solids. Explain your reasoning.
45. **Model** Draw a model to represent the physical property of metals known as ductility, or the ability to be drawn into a wire. Base your drawing on the electron sea model shown in **Figure 10.**

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

ENGINEERING & TECHNOLOGY

From Salty to Fresh

Desalination is the process of removing dissolved salts from seawater. Desalination technologies are becoming more important as other sources of fresh water are depleted.

Desalination plants such as this one can produce millions of liters of fresh water from seawater every day.

Taking the Salt out of Saltwater

Only 1 percent of all water on Earth is accessible fresh water. Most of Earth's water is in oceans. As fresh water sources are depleted, looking to the oceans as a source of water seems an obvious solution. The challenge for engineers is to develop desalination processes that are efficient, economical, and environmentally friendly.

Thermal distillation is a technique that has been used for thousands of years. In this process, heat is used to evaporate seawater. The water becomes a vapor and the other substances in seawater are left behind as a residue. The evaporated water is collected and condensed by cooling. The condensed water contains no salts or other dissolved impurities and is safe to drink.

A variety of technologies use thermal distillation, ranging from small solar stills that produce fresh water for individuals to huge distillation plants that produce fresh water for cities.

The main problem with thermal distillation is that it requires a lot of energy. Since the mid-1900s, desalination research has focused on another process called reverse osmosis, which uses much less energy than thermal distillation. In reverse osmosis, water is subjected to high pressure to cause it to pass through a semi-permeable membrane. The membrane allows water molecules but not ions to pass through, resulting in salt-free, drinkable water on one side of the membrane, and concentrated seawater on the other.

Reverse osmosis can be used in small devices and large desalination plants. The technology has drawbacks, though, including potential harm to marine life when the concentrated seawater is released back to the ocean. Engineers and scientists continue to improve desalination technologies to ensure a stable, safe source of drinking water for the future.

DEVELOP A MODEL TO ILLUSTRATE

Make diagrams that compare desalination using thermal distillation and reverse osmosis. Be sure to include labels and captions.

MODULE 6
STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 ION FORMATION

- A chemical bond is the force that holds two atoms together.
- Some atoms form ions to gain stability. This stable configuration involves a complete outer energy level, usually consisting of eight valence electrons.
- Ions are formed by the loss or gain of valence electrons.
- The number of protons remains unchanged during ion formation.

- chemical bond
- cation
- anion

Lesson 2 IONIC BONDS AND IONIC COMPOUNDS

- Ionic compounds contain ionic bonds formed by the attraction of oppositely charged ions.
- Ions in an ionic compound are arranged in a repeating pattern known as a crystal lattice.
- Ionic compound properties are related to ionic bond strength.
- Ionic compounds conduct an electric current in the liquid phase and in aqueous solution.
- Lattice energy is the energy needed to remove 1 mol of ions from its lattice.

- ionic bond
- ionic compound
- crystal lattice
- electrolyte
- lattice energy

Lesson 3 NAMES AND FORMULAS FOR IONIC COMPOUNDS

- A formula unit gives the ratio of cations to anions in the ionic compound.
- A monatomic ion is formed from one atom. The charge of a monatomic ion is equal to its oxidation number.
- Roman numerals indicate the oxidation number of cations having multiple possible oxidation states.
- Polyatomic ions consist of more than one atom and act as a single unit.
- To indicate more than one polyatomic ion in a chemical formula, place parentheses around the polyatomic ion and use a subscript.

- formula unit
- monatomic ion
- polyatomic ion
- oxyanion

Lesson 2 METALLIC BONDS AND THE PROPERTIES OF METALS

- A metallic bond forms when metal cations attract freely moving, delocalized valence electrons.
- In the electron sea model, electrons move through the metallic crystal and are not held by any particular atom.
- The electron sea model explains the physical properties of metallic solids.
- Metal alloys are formed when a metal is mixed with one or more other elements.

- electron sea model
- delocalized electron
- metallic bond
- alloy

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

Why do some crystals form cubes?

CER Claim, Evidence, Reasoning

Explain Your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

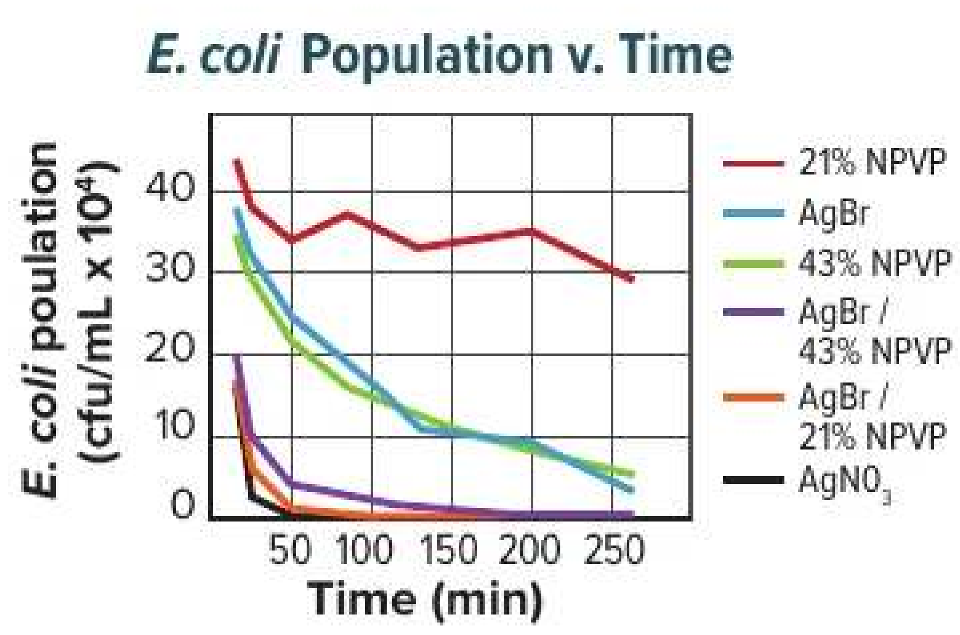
Based on Real Data*

SEP Data Analysis Lab

Can embedding nanoparticles of silver into a polymer give the polymer antimicrobial properties?

Researchers tested the antimicrobial properties of a new composite material—the polymer poly(4-vinyl-N-hexylpyridinium bromide), known as NPVP, which attracts cations. It is known that silver ions from silver bromide and silver nitrate exhibit antimicrobial activity. Silver bromide was embedded into the NPVP polymer. Scientists tested the antimicrobial properties of the composite material. Their results, illustrated in the graph, show the growth of *E. coli* bacteria over a period of approximately four hours. Each line represents the *E. coli* population in response to the introduction of a particular substance.

Data and Observations

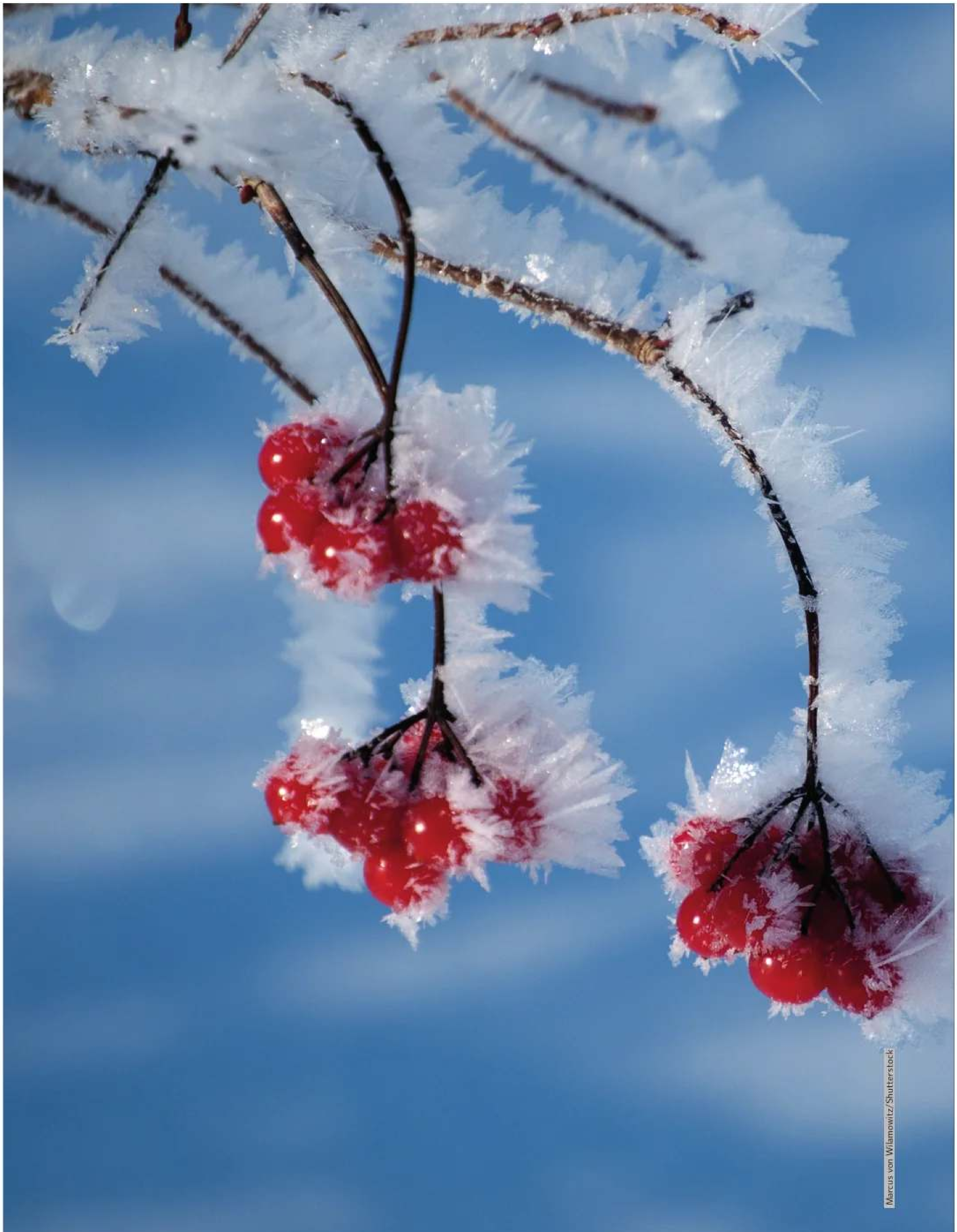


*Data obtained from: Sambhy, V., et al. Published on the Web 7/7/2006. Silver Bromide Nanoparticle/Polymer Composites. Journal of the American Chemical Society.

CER Analyze and Interpret Data

1. **Claim** Does the addition of silver bromide (AgBr) ions to NPVP improve the antimicrobial properties of the composite?
2. **Evidence, Reasoning** Does a composite polymer containing NPVP and silver bromide show antimicrobial properties? Explain your answer.

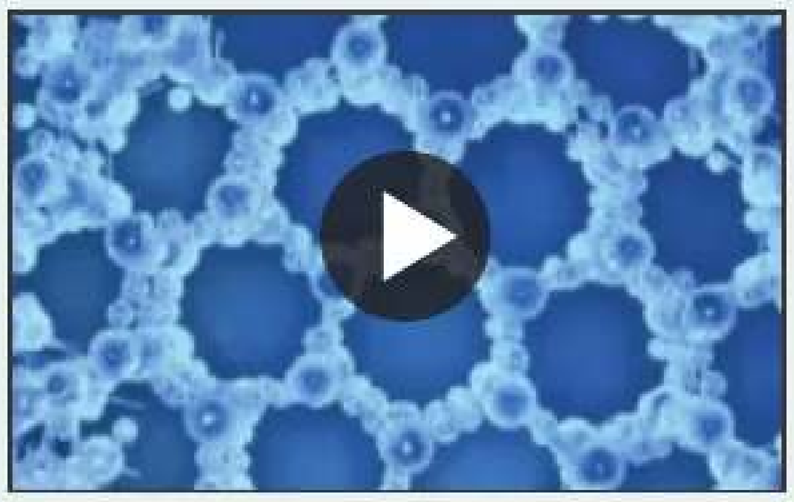
Mark A. Schneider/Science Source


Marcus von Wilamowitz/Shutterstock

MODULE 7
COVALENT BONDING

ENCOUNTER THE PHENOMENON

Why does water expand when it freezes?



GO ONLINE to play a video about covalent bonding in water.

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

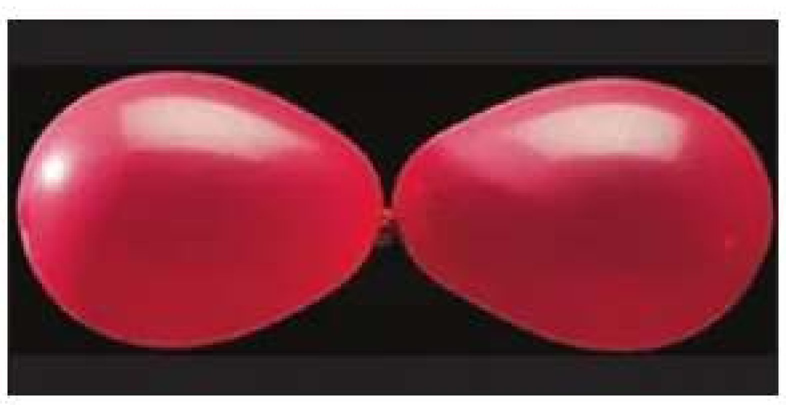
CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about why water expands when it freezes.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.



LESSON 4: Explore & Explain: Hybridization and Molecular Shape

LESSON 5: Explore & Explain: Intermolecular Forces and Properties of Covalent Compounds

Smithsonian

Additional Resources

LESSON 1
THE COVALENT BOND

FOCUS QUESTION

How do atoms bond in covalent molecules?

Why do atoms bond?

To understand why new compounds form, recall what you know about elements that do not tend to form new compounds—the noble gases. You learned that all noble gases have stable electron arrangements. This stable arrangement consists of a full outer energy level and has lower potential energy than other electron arrangements. Because of their stable configurations, noble gases seldom form compounds. Other elements frequently form compounds, such as the hydrogen and oxygen that form the water shown in **Figure 1**.

Gaining stability

The stability of an atom, ion, or compound is related to its energy, with lower energy states being more stable. In ionic bonds, metals and nonmetals gain stability by transferring electrons to form ions. The resulting ions have stable noble-gas electron configurations. In this module, you will learn that valence electron sharing is another way atoms can acquire the stable electron configuration of noble gases, resulting in stable molecules with less energy than the same set of atoms separated.

Figure 1 Each water droplet is made up of water molecules. Each water molecule is made up of two hydrogen atoms and one oxygen atom that have bonded by sharing electrons. The shapes of the drops are due to intermolecular forces acting on the water molecules.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Electron States and Simple Chemical Reactions

HS-PS1-2. Construct and revise an explanation for the outcome of a simple chemical reaction based on the outermost electron states of atoms, trends in the periodic table, and knowledge of the patterns of chemical properties.

Revisit the Encounter the Phenomenon Question

What information from this lesson can help you answer the module question?

What is a covalent bond?

You just read that atoms can share electrons to form stable electron configurations. How does this occur? Are there different ways in which electrons can be shared? How are the properties of these compounds different from those formed by ions? Read on to answer these questions.

Shared electrons

Atoms in nonionic compounds share electrons. The chemical bond that results from sharing valence electrons is a **covalent bond.** A **molecule** is formed when two or more atoms bond covalently. In a covalent bond, the shared electrons are considered to be part of the outer energy levels of both atoms involved. Covalent bonding generally can occur between elements that are near each other on the periodic table. The majority of covalent bonds form between atoms of nonmetallic elements.

Covalent bond formation

Diatomic molecules, such as hydrogen (H_2), nitrogen (N_2), oxygen (O_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2), form when two atoms of each element share electrons. They exist this way because the two-atom molecules are more stable than the individual atoms.

Consider fluorine, which has an electron configuration of $1s^22s^22p^5$. Each fluorine atom has seven valence electrons and needs another electron to form an octet. As two fluorine atoms approach each other, several forces act, as shown in **Figure 2**. Two repulsive forces act on the atoms, one from each atom's like-charged electrons and one from each atom's like-charged protons. A force of attraction also acts, as one atom's protons attract the other atom's electrons. As the fluorine atoms move closer, the attraction of the protons in each nucleus for the other atom's electrons increases until a point of maximum net attraction is achieved. At that point, the two atoms bond covalently, and a molecule forms. Now, each atom has a completed octet because they are each sharing a pair of electrons. If the two nuclei move closer, the repulsion forces increase and exceed the attractive forces.

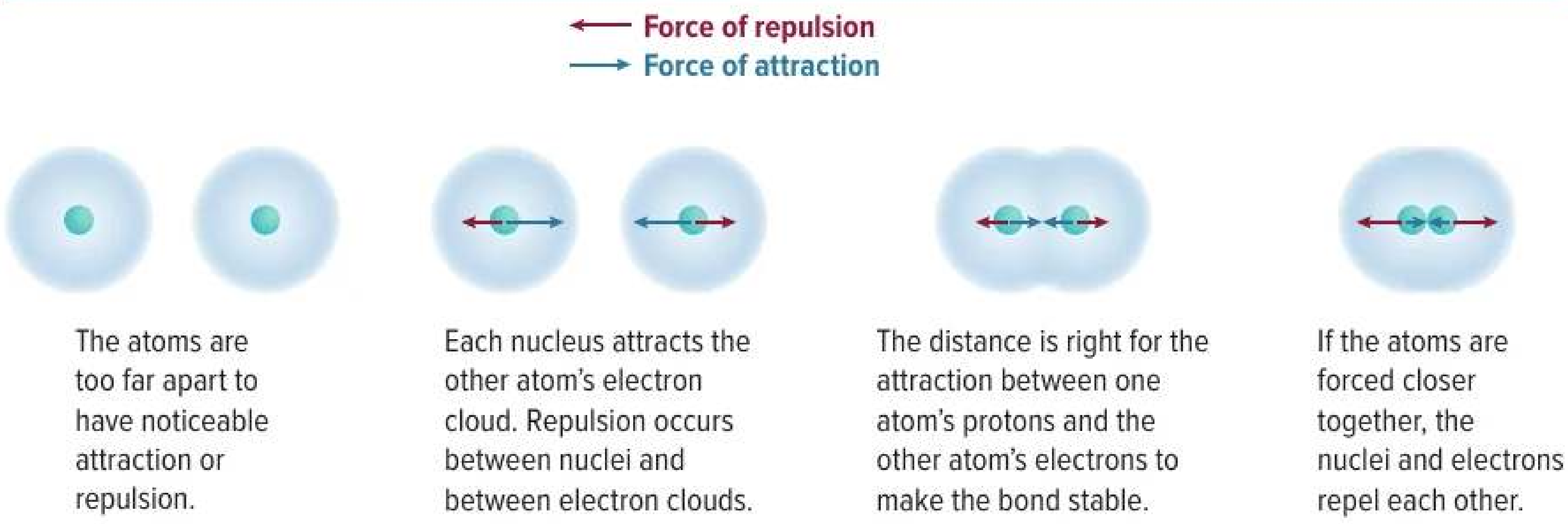


Figure 2 The arrows in this diagram show the net forces of attraction and repulsion acting on two fluorine atoms as they move toward each other. The overall force between two atoms is the result of electron-electron repulsion, nucleus-nucleus repulsion, and nucleus-electron attraction. At the position of maximum net attraction, a covalent bond forms.

Relate *How is the stability of the bond related to the forces acting on the atoms?*

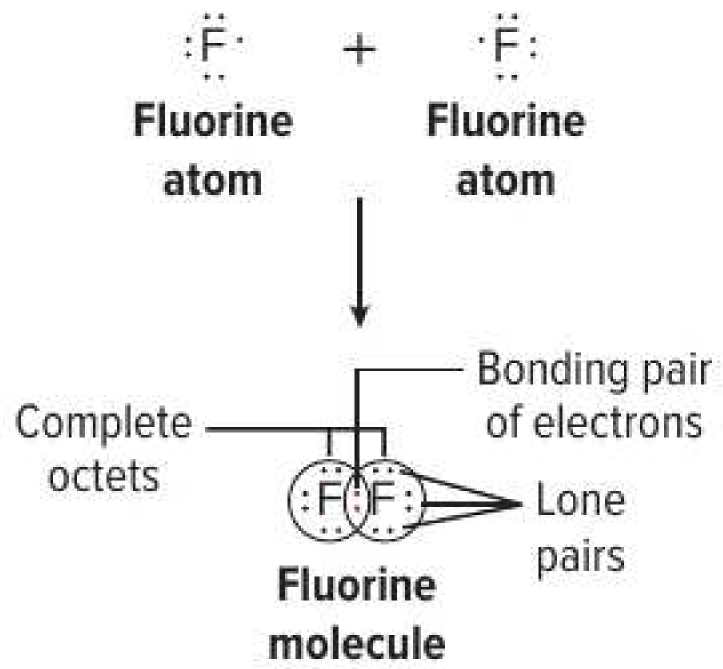


Figure 3 Two fluorine atoms share a pair of electrons to form a covalent bond. Note that the shared electron pair gives each atom a complete octet.

Infer *how the energy of the fluorine atoms changes when they form a covalent bond.*

The most stable arrangement of atoms in a covalent bond exists at some optimal distance between nuclei. At this point, the net attraction is greater than the net repulsion. Fluorine exists as a diatomic molecule because the sharing of one pair of electrons gives each fluorine atom a stable noble-gas configuration. As shown in **Figure 3**, each fluorine atom in the fluorine molecule has one pair of electrons that are covalently bonded (shared) and three pairs of electrons that are unbonded (not shared). Unbonded pairs are also known as lone pairs.

Single Covalent Bonds

When only one pair of electrons is shared, such as in a hydrogen molecule, it is a single covalent bond. The shared electron pair is often referred to as the bonding pair. For a hydrogen molecule, shown in **Figure 4,** each covalently bonded atom equally attracts the pair of shared electrons. Thus, the two shared electrons belong to each atom simultaneously, which gives each hydrogen atom the noble-gas configuration of helium ($1s^2$) and lower energy. The hydrogen molecule is more stable than either hydrogen atom is by itself.

Recall that electron-dot diagrams can be used to show valence electrons of atoms. In a **Lewis structure,** they represent the arrangement of electrons in a molecule. A line or a pair of vertical dots between the symbols of elements represents a single covalent bond in a Lewis structure. For example, a hydrogen molecule is written as H–H or H:H.

Group 17 and single bonds

The halogens—the group 17 elements, such as fluorine—have seven valence electrons. To form an octet, one more electron is needed. Therefore, atoms of group 17 elements form single covalent bonds with atoms of other nonmetals, such as carbon. You have already read that the atoms of some group 17 elements form covalent bonds with identical atoms. For example, fluorine exists as F_2, and chlorine exists as Cl_2. Because atoms from group 17 need only one more electron to form an octet they are very unstable and reactive.

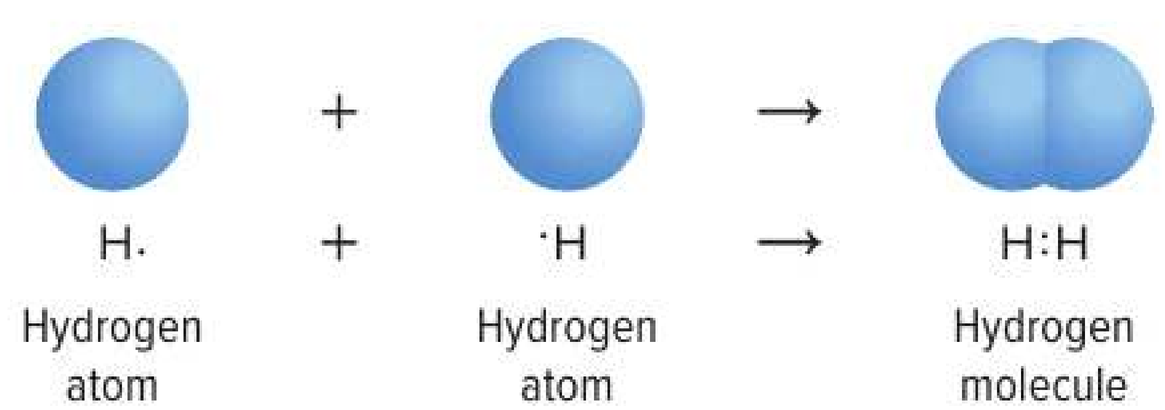


Figure 4 When two hydrogen atoms share a pair of electrons, each hydrogen atom is stable because it has a full outer energy level.

Group 16 and single bonds

An atom of a group 16 element can share two electrons and can form two covalent bonds. Oxygen is a group 16 element with an electron configuration of $1s^2 2s^2 2p^4$. Water is composed of two hydrogen atoms and one oxygen atom. Each hydrogen atom has the noble-gas configuration of helium when it shares one electron with oxygen. Oxygen, in turn, has the noble-gas configuration of neon when it shares one electron with each hydrogen atom. **Figure 5a** shows the Lewis structure for a molecule of water. Notice that the oxygen atom has two single covalent bonds and two unshared pairs of electrons.

Group 15 and single bonds

Group 15 elements form three covalent bonds with atoms of nonmetals. Nitrogen is a group 15 element with the electron configuration of $1s^2 2s^2 2p^3$. Ammonia (NH_3) has three single covalent bonds. Three nitrogen electrons bond with the three hydrogen atoms leaving one pair of unshared electrons on the nitrogen atom. **Figure 5b** shows the Lewis structure for an ammonia molecule. Nitrogen also forms similar compounds with atoms of group 17 elements, such as nitrogen trifluoride (NF_3), nitrogen trichloride (NCl_3), and nitrogen tribromide (NBr_3). Each atom of these group 17 elements and the nitrogen atom share an electron pair.

Group 14 and single bonds

Atoms of group 14 elements form four covalent bonds. A methane molecule (CH_4) forms when one carbon atom bonds with four hydrogen atoms. Carbon, a group 14 element, has an electron configuration of $1s^2 2s^2 2p^2$. With four valence electrons, carbon needs four more electrons for a noble gas configuration. Therefore, when carbon bonds with other atoms, it forms four bonds. Because a hydrogen atom, a group 1 element, has one valence electron, it takes four hydrogen atoms to provide the four electrons needed by a carbon atom. The Lewis structure for methane is shown in **Figure 5c**. Carbon also forms single covalent bonds with other nonmetal atoms, including those in group 17.

Get It?

Describe how a Lewis structure shows a single covalent bond.

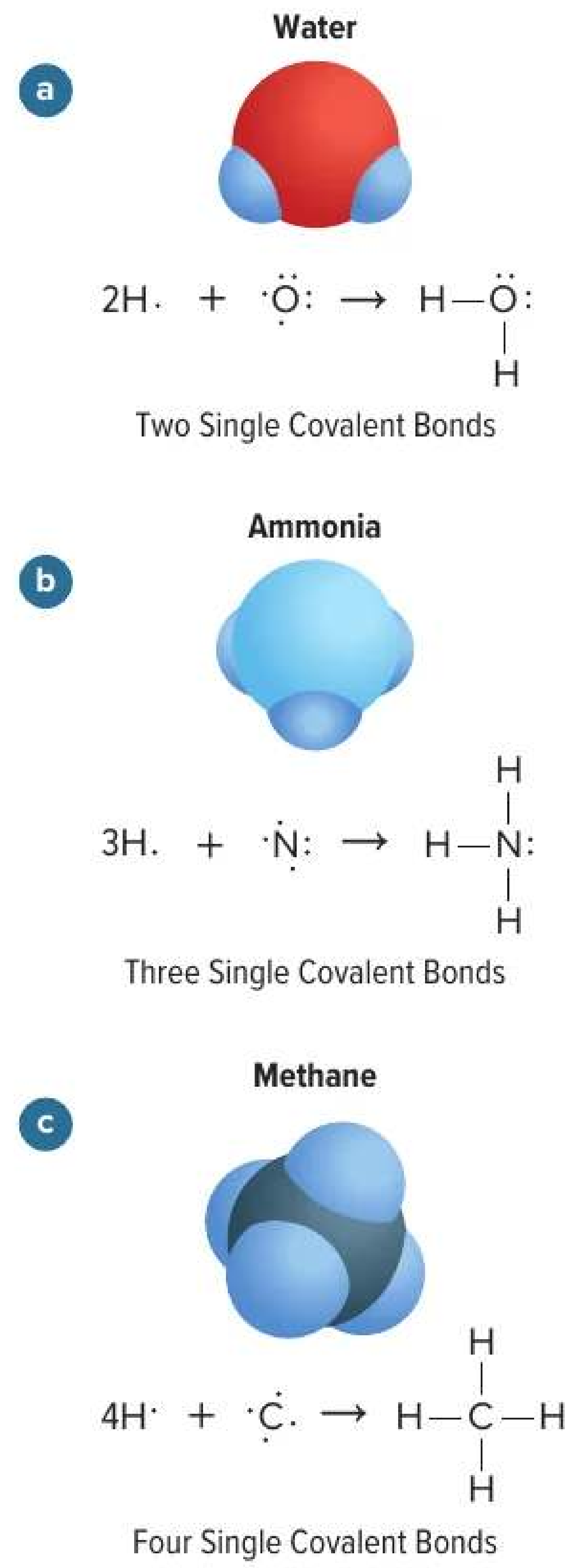


Figure 5 These chemical equations show how atoms share electrons and become stable. As shown by the Lewis structure for each molecule, all atoms in each molecule achieve a full outer energy level.

Describe *For the central atom in each molecule, describe how the octet rule is met.*

Figure 6 The frosted-looking portions of this glass were chemically etched using hydrogen fluoride (HF), a weak acid. Hydrogen fluoride reacts with silica, the major component of glass, and forms gaseous silicon tetrafluoride (SiF_4) and water.

EXAMPLE Problem 1

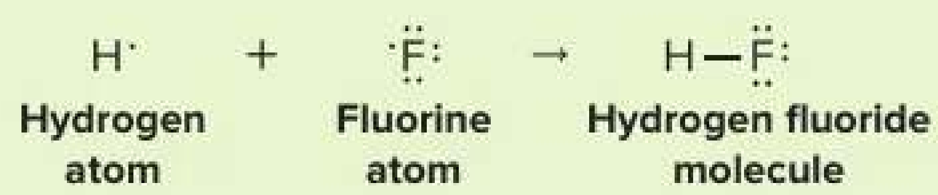
LEWIS STRUCTURE OF A MOLECULE The pattern on the glass shown in **Figure 6** was made by chemically etching its surface with hydrogen fluoride (HF). Draw the Lewis structure for a molecule of hydrogen fluoride.

1 ANALYZE THE PROBLEM

You are given the information that hydrogen and fluorine form the molecule hydrogen fluoride. An atom of hydrogen, a group 1 element, has only one valence electron. It can bond with any nonmetal atom when they share one pair of electrons. An atom of fluorine, a group 17 element, needs one electron to complete its octet. Therefore, a single covalent bond forms when atoms of hydrogen and fluorine bond.

2 SOLVE FOR THE UNKNOWN

To draw a Lewis structure, first draw the electron-dot diagram for each of the atoms. Then, rewrite the chemical symbols and draw a line between them to show the shared pair of electrons. Finally, add dots to show the unshared electron pairs.



3 EVALUATE THE ANSWER

Each atom in the new molecule now has a noble-gas configuration and is stable.

PRACTICE Problems

ADDITIONAL PRACTICE

Draw the Lewis structure for each molecule.

1. PH_3
2. H_2S
3. HCl
4. CCl_4
5. SiH_4
6. **CHALLENGE** Draw a generic Lewis structure for a molecule formed between atoms of group 1 and group 16 elements.

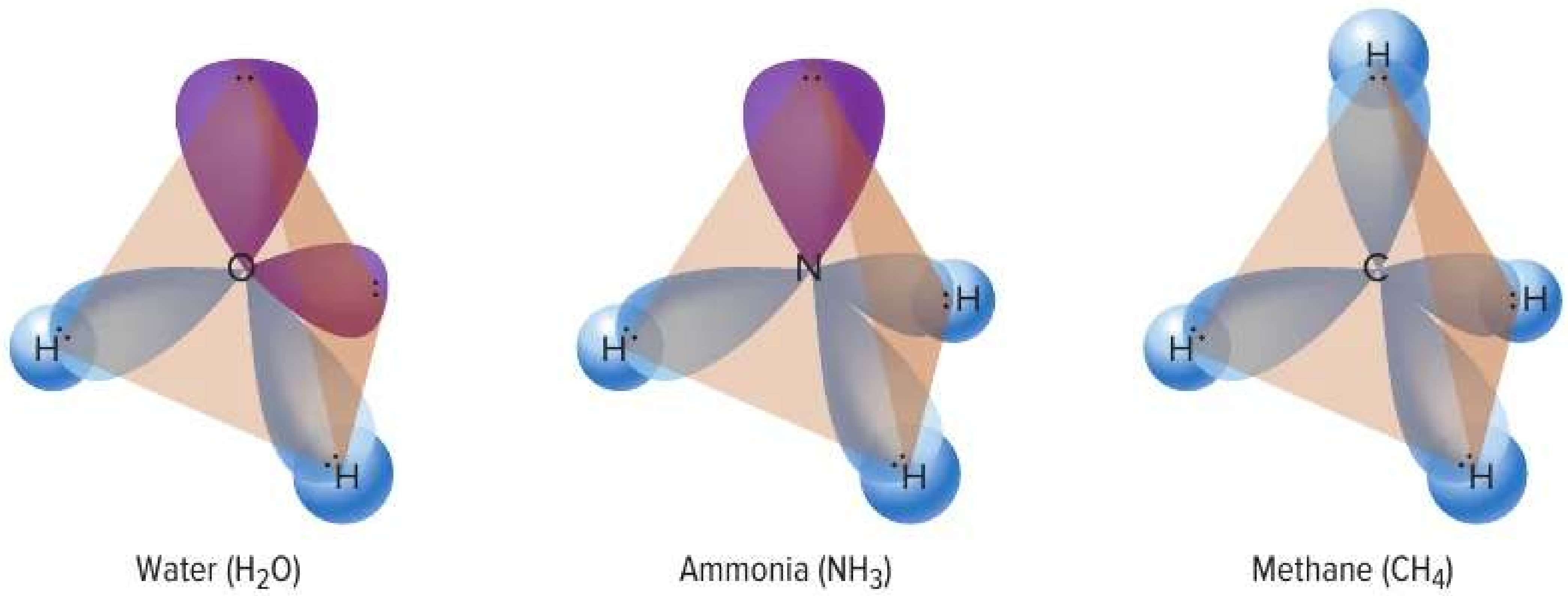


Figure 7 Sigma bonds formed in each of these molecules when the atomic orbital of each hydrogen atom overlapped end-to-end with the orbital of the central atom.

Interpret *Identify the number of sigma bonds in each molecule.*

The sigma bond

Single covalent bonds are also called **sigma bonds,** represented by the Greek letter sigma (σ). A sigma bond occurs when the pair of shared electrons is in an area centered between the two atoms. When two atoms share electrons, their valence atomic orbitals overlap end-to-end, concentrating the electrons in a bonding orbital between the two atoms. A bonding orbital is a localized region where bonding electrons will most likely be found. Sigma bonds can form when an s orbital overlaps with another s orbital or a p orbital, or two p orbitals overlap end-to-end. Water (H_2O), ammonia (NH_3), and methane (CH_4) have sigma bonds, as shown in **Figure 7**.

Get It?

List the orbitals that can form sigma bonds in a covalent compound, such as in DNA molecules.

Multiple Covalent Bonds

In some molecules, atoms have noble-gas configurations when they share more than one pair of electrons with one or more atoms. Sharing multiple pairs of electrons forms multiple covalent bonds. A double covalent bond and a triple covalent bond are examples of multiple bonds. Carbon, nitrogen, oxygen, and sulfur atoms often form multiple bonds with other nonmetals. How do you know if two atoms will form a multiple bond? In general, the number of valence electrons needed to form an octet equals the number of covalent bonds that can form.

ACADEMIC VOCABULARY

overlap

to occupy the same area in part

The two driveways overlap at the street forming a common entrance.

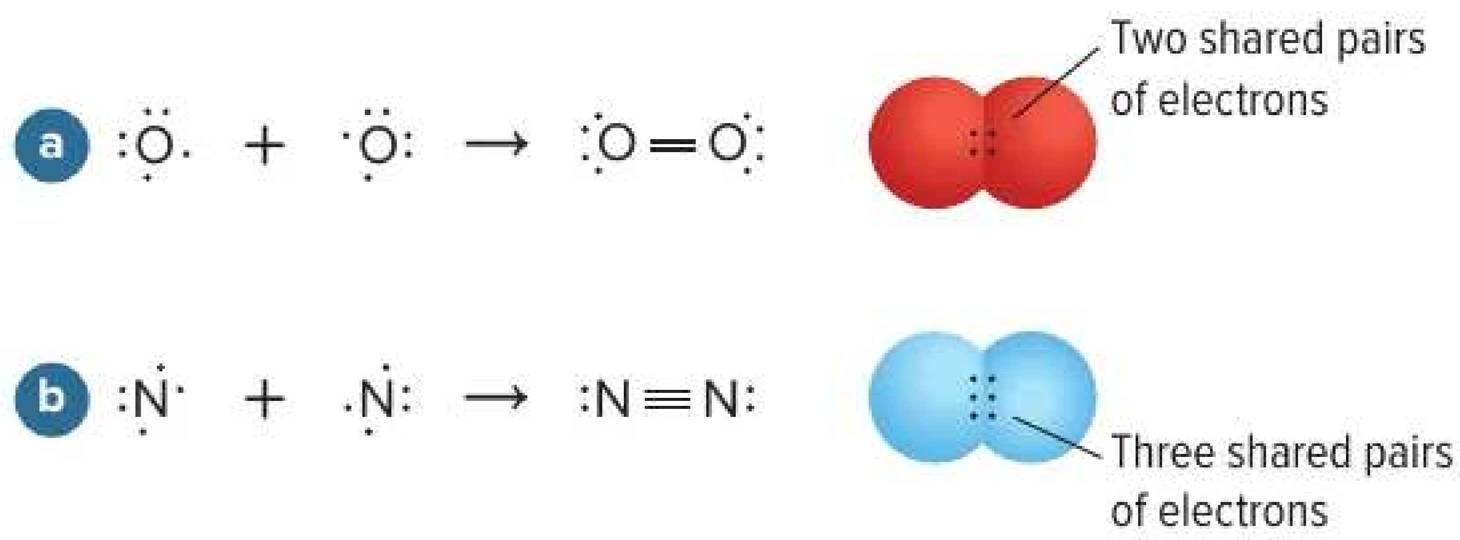


Figure 8 Multiple covalent bonds form when two atoms share more than one pair of electrons.

a. Two oxygen atoms form a double bond.

b. A triple bond forms between two nitrogen atoms.

Double bonds

A double covalent bond forms when two pairs of electrons are shared between two atoms. For example, atoms of the element oxygen only exist as diatomic molecules. Each oxygen atom has six valence electrons and must obtain two additional electrons for a noble-gas configuration, as shown in **Figure 8a**. A double covalent bond forms when each oxygen atom shares two electrons; a total of two pairs of electrons are shared between the two atoms.

Triple bonds

A triple covalent bond forms when three pairs of electrons are shared between two atoms. Diatomic nitrogen (N_2) molecules contain a triple covalent bond. Each nitrogen atom shares three electron pairs, forming a triple bond with the other nitrogen atom as shown in **Figure 8b**.

The pi bond

A multiple covalent bond consists of one sigma bond and at least one pi bond. A **pi bond,** represented by the Greek letter pi (π), forms when parallel orbitals overlap and share electrons. The shared electron pair of a pi bond occupies the space above and below the line that represents where the two atoms are joined together.

It is important to note that molecules having multiple covalent bonds contain both sigma and pi bonds. A double covalent bond, as shown in **Figure 9**, consists of one pi bond and one sigma bond. A triple covalent bond consists of two pi bonds and one sigma bond.

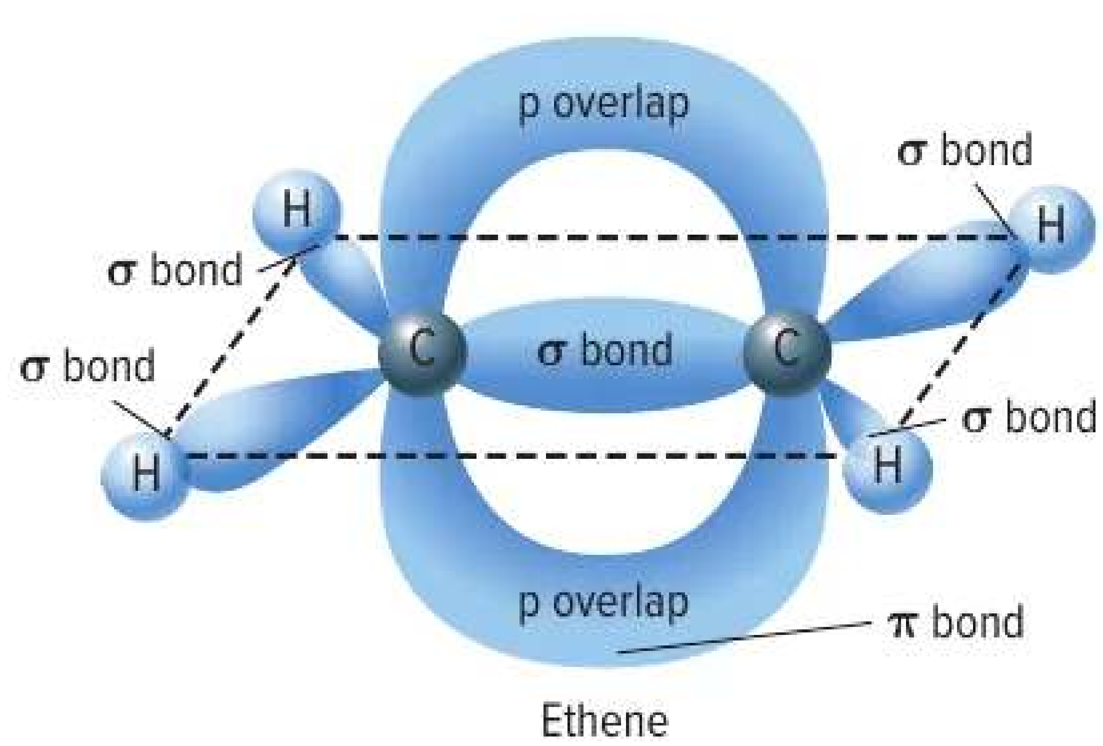


Figure 9 Notice how the multiple bond between the two carbon atoms in ethene (C_2H_4) consists of a sigma bond and a pi bond. The sigma bond is formed by the end-to-end overlap of orbitals directly between the two carbon atoms. The carbon atoms are close enough that the side-by-side p orbitals overlap and form the pi bond. This results in a doughnut-shaped cloud around the sigma bond.

The Strength of Covalent Bonds

Recall that a covalent bond involves attractive and repulsive forces. In a molecule, nuclei and electrons attract each other, but nuclei repel other nuclei, and electrons repel other electrons. When this balance of forces is upset, a covalent bond can be broken. Because covalent bonds differ in strength, some bonds break more easily than others. Several factors influence the strength of covalent bonds.

Bond length

The strength of a covalent bond depends on the distance between the bonded nuclei. The distance between the two bonded nuclei at the position of maximum attraction is called bond length, as shown in **Figure 10**. It is determined by the sizes of the two bonding atoms and how many electron pairs they share. Bond lengths for molecules of fluorine (F_2), oxygen (O_2), and nitrogen (N_2) are listed in **Table 1**. Notice that as the number of shared electron pairs increases, the bond length decreases.

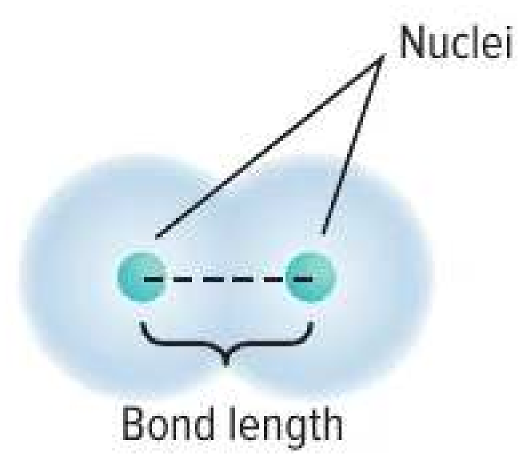


Figure 10 Bond length is the distance from the center of one nucleus to the center of the other nucleus of two bonded atoms.

Bond length and bond strength are also related: the shorter the bond length, the stronger the bond. Therefore, a single bond, such as that in F_2, is weaker than a double bond, such as that in O_2. Likewise, the double bond in O_2 is weaker than the triple bond in N_2.

Table 1 Covalent Bond Type, Bond Length, and Bond-Dissociation Energy

Molecule	Bond Type	Bond Length	Bond-Dissociation Energy
F_2	single covalent	1.43×10^{-10} m	159 kJ/mol
O_2	double covalent	1.21×10^{-10} m	498 kJ/mol
N_2	triple covalent	1.10×10^{-10} m	945 kJ/mol

Bonds and energy

An energy change occurs when a bond between atoms in a stable molecule forms or breaks. The amount of energy needed to break apart the molecule must be at least the amount released during its formation. This relationship between the amount of energy released during bond formation and the amount of energy needed to break covalent bonds is an example of how energy flows through chemical systems. The amount of energy required to break a specific covalent bond is called bond-dissociation energy and is always a positive value. The bond-dissociation energies for the covalent bonds in molecules of fluorine, oxygen, and nitrogen are listed in **Table 1**.

Bond-dissociation energy also indicates the strength of a chemical bond because of the inverse relationship between bond energy and bond length. As indicated in **Table 1**, the smaller the bond length is, the greater the bond-dissociation energy. In addition, the sum of the bond-dissociation energy values for all of the bonds in a molecule is the amount of chemical potential energy in a molecule of that compound.

Get It?

Compare the flow of energy in forming a molecule with the flow of energy when the molecule breaks apart.

The total energy change of a chemical reaction is determined from the energy of the bonds broken and formed. An **endothermic reaction** occurs when a greater amount of energy is required to break the existing bonds in the reactants than is released when the new bonds form in the products. An **exothermic reaction** occurs when more energy is released during product bond formation than is required to break bonds in the reactants. **Figure 11** illustrates a common exothermic reaction. You will study exothermic and endothermic reactions in much greater detail when you study the energy changes in chemical reactions.

Figure 11 Breaking the C–C bonds in charcoal and the O–O bonds in the oxygen in air requires an input of energy. Energy is released as heat and light when bonds form, producing CO_2. Thus, the burning of charcoal is an exothermic reaction.

Check Your Progress

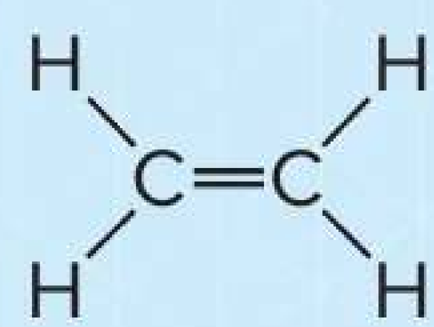
Summary

- Covalent bonds form when atoms share one or more pairs of electrons.
- Sharing one pair, two pairs, and three pairs of electrons forms single, double, and triple covalent bonds, respectively.
- Orbitals overlap directly in sigma bonds. Parallel orbitals overlap in pi bonds. A single covalent bond is a sigma bond but multiple covalent bonds are made of both sigma and pi bonds.
- Bond-dissociation energy is needed to break a covalent bond.

Demonstrate Understanding

7. **Describe** how the energy of a molecule compares to the same set of atoms separated.
8. **Describe** how the stability of an atom relates to the octet rule and the formation of covalent bonds.
9. **Illustrate** the formation of single, double, and triple covalent bonds using Lewis structures.
10. **Compare and contrast** ionic bonds and covalent bonds.
11. **Contrast** sigma bonds and pi bonds.
12. **Apply** Create a graph using the bond-dissociation energy data and the bond-length data in **Table 1**. Describe the relationship between bond length and bond-dissociation energy.
13. **Compare** the bond-dissociation energies needed to break the bonds in the structures below. Then explain how these values relate to the amount of energy needed to form the compounds.

a. H—C≡C—H

b. 

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LESSON 2

NAMING MOLECULES

FOCUS QUESTION

How do you name molecules?

Naming Binary Molecular Compounds

Many molecular compounds have common names, but they also have scientific names that reveal their composition. To write the formulas and names of molecules, you will use processes similar to those described for ionic compounds.

Start with a binary molecular compound. Note that a binary molecular compound is composed only of two nonmetal atoms—not metal atoms or ions. An example is dinitrogen monoxide (N_2O), a gaseous anesthetic that is more commonly known as nitrous oxide or laughing gas. The naming of N_2O is explained in the following rules.

1. The first element in the formula is always named first, using the entire element name. **N is the symbol for *nitrogen*.**
2. The second element in the formula is named using its root and adding the suffix *-ide*. **O is the symbol for oxygen so the second word is *oxide*.**
3. Prefixes are used to indicate the number of atoms of each element that are present in the compound. **Table 2** lists the most common prefixes used. **There are two atoms of nitrogen and one atom of oxygen, so the first word is *dinitrogen* and the second word is *monoxide*.**

Table 2 Prefixes in Covalent Compounds

Number of Atoms	Prefix	Number of Atoms	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

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COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Review the News

Obtain information from a current news story about naming molecules. Evaluate your source and communicate your findings to your class.

There are exceptions to using the prefixes shown in **Table 2**. The first element in the compound name never uses the *mono-* prefix. For example, CO is carbon monoxide, not monocarbon monoxide. Also, if using a prefix results in two consecutive vowels, one of the vowels is usually dropped to avoid an awkward pronunciation. For example, notice that the oxygen atom in CO is called monoxide, not monooxide.

EXAMPLE Problem 2

NAMING BINARY MOLECULAR COMPOUNDS Name the compound P_2O_5, which is used as a drying and dehydrating agent.

1 ANALYZE THE PROBLEM

You are given the formula for a compound. The formula contains the elements and the number of atoms of each element in one molecule of the compound. Because only two different elements are present and both are nonmetals, the compound can be named using the rules for naming binary molecular compounds.

2 SOLVE FOR THE UNKNOWN

First, name the elements involved in the compound.

phosphorus	The first element, represented by P, is phosphorus.
oxide	The second element, represented by O, is oxygen. Add the suffix *-ide* to the root of oxygen, *ox-*.
phosphorus oxide	Combine the names.

Now modify the names to indicate the number of atoms present in a molecule.

diphosphorus pentoxide	From the formula P_2O_5, you know that two phosphorus atoms and five oxygen atoms make up a molecule of the compound. From Table 2, you know that *di-* is the prefix for two and *penta-* is the prefix for five. The *a* in *penta-* is not used because *oxide* begins with a vowel.

3 EVALUATE THE ANSWER

The name diphosphorus pentoxide shows that a molecule of the compound contains two phosphorus atoms and five oxygen atoms, which agrees with the compound's chemical formula, P_2O_5.

PRACTICE Problems

ADDITIONAL PRACTICE

Name each of the binary covalent compounds listed below.

14. CO_2

15. SO_2

16. NF_3

17. CCl_4

18. CHALLENGE What is the formula for diarsenic trioxide?

Common names for some molecular compounds

Have you ever enjoyed an icy, cold glass of dihydrogen monoxide on a hot day? You probably have but you most likely called it by its common name, water. Recall that many ionic compounds have common names in addition to their scientific ones. For example, baking soda is sodium hydrogen carbonate and common table salt is sodium chloride.

Many binary molecular compounds, such as nitrous oxide and water, were discovered and given common names long before the present-day naming system was developed. Other binary covalent compounds that are generally known by their common names rather than their scientific names are ammonia (NH_3), hydrazine (N_2H_4), and nitric oxide (NO).

Get It?

Apply What are the scientific names for ammonia, hydrazine, and nitric oxide?

Naming Acids

Water solutions, also called aqueous solutions, of some molecules are acidic and are named as acids. Acids are important compounds with specific properties. If a compound produces hydrogen ions (H^+) in solution, it is an acid. For example, HCl produces H^+ in solution and is an acid. Two common types of acids exist—binary acids and oxyacids.

Naming binary acids

A binary acid contains hydrogen and one other element. The naming of the common binary acid known as hydrochloric acid is explained in the following rules.

1. The first word has the prefix *hydro-* to name the hydrogen part of the compound. The rest of the first word consists of a form of the root of the second element plus the suffix *-ic*. **HCl (hydrogen and chlorine) becomes *hydrochloric.***
2. The second word is always *acid*. **Thus, HCl in a water solution is called *hydrochloric acid.***

Although the term *binary* indicates exactly two elements, a few acids that contain more than two elements are named according to the rules for naming binary acids.

If no oxygen is present in the formula for the acidic compound, the acid is named in the same way as a binary acid, except that the root of the second part of the name is the root of the polyatomic ion that the acid contains. For example, HCN, which is composed of hydrogen and the cyanide ion, is called *hydrocyanic acid* in solution.

Naming oxyacids

An acid that contains both a hydrogen atom and an oxyanion is referred to as an **oxyacid**. Recall that an oxyanion is a polyatomic ion containing one or more oxygen atoms. The following rules explain how to name nitric acid, an important oxyacid that has a chemical formula of HNO_3.

Table 3 **Naming Oxyacids**

Compound	Oxyanion	Acid Suffix	Acid Name
$HClO_3$	chlorate	-ic	chloric acid
$HClO_2$	chlorite	-ous	chlorous acid
HNO_3	nitrate	-ic	nitric acid
HNO_2	nitrite	-ous	nitrous acid

1. First, identify the oxyanion present. The first word of an oxyacid's name consists of the root of the oxyanion and the prefix *per-* or *hypo-* if it is part of the oxyanion's name. The first word of the oxyacid's name also has a suffix that depends on the oxyanion's suffix. If the oxyanion's name ends with the suffix *-ate*, replace it with the suffix *-ic*. If the name of the oxyanion ends with the suffix *-ite*, replace it with the suffix *-ous*. **NO_3^-, the nitrate ion, becomes *nitric*.**
2. The second word of the name is always *acid*. **HNO_3 (hydrogen and the nitrate ion) becomes *nitric acid*.**

Table 3 shows how the names of several oxyacids follow these rules. Notice that the hydrogen in an oxyacid is not part of the name.

You have learned that naming covalent compounds follows different sets of rules depending on the composition of the compound. **Table 4** summarizes the formulas and names of several covalent compounds. Note that an acid, whether a binary acid or an oxyacid, can have a common name in addition to its compound name.

PRACTICE Problems ADDITIONAL PRACTICE

Name the following acids. Assume each compound is dissolved in water.

19. HI

20. $HClO_3$

21. $HClO_2$

22. H_2SO_4

23. H_2S

24. CHALLENGE What is the formula for periodic acid?

Table 4 **Formulas and Names of Some Covalent Compounds**

Formula	Common Name	Molecular Compound Name
H_2O	water	dihydrogen monoxide
NH_3	ammonia	nitrogen trihydride
N_2H_4	hydrazine	dinitrogen tetrahydride
HCl	muriatic acid	hydrochloric acid
$C_9H_8O_4$	aspirin	2-(acetyloxy)benzoic acid

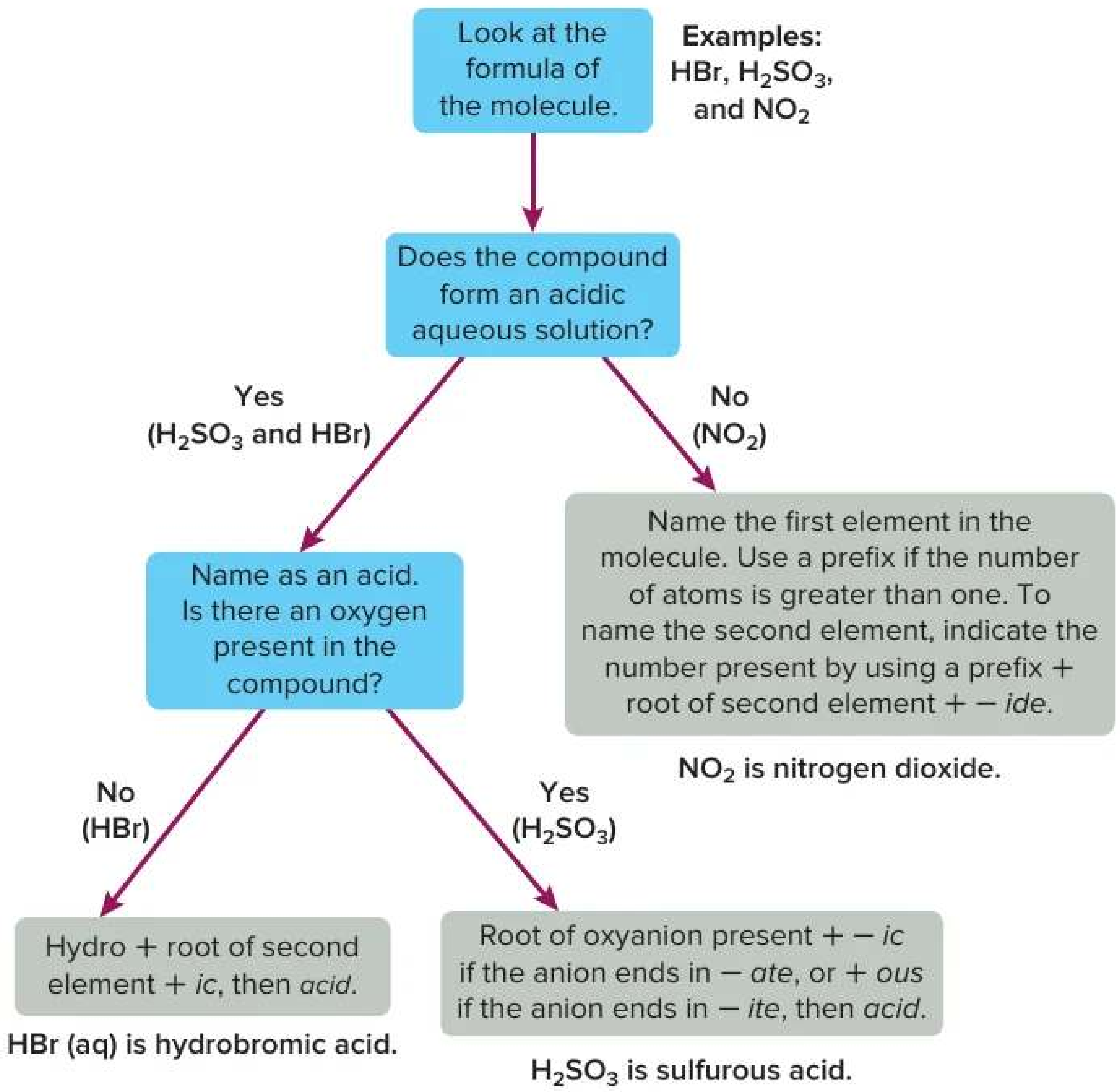

Figure 12 Use this flowchart to name molecular compounds when their formulas are known.

Apply *Which compound in the figure is an oxyacid? Which is a binary acid?*

The flowchart in **Figure 12** can help you determine the name of a molecular covalent compound. To use the chart, start at the top and work downward by reading the text contained in the colored boxes. Apply the text at each step of the flowchart to the formula of the compound that you wish to name.

Writing Formulas from Names

The name of a molecular compound reveals its composition and is important in communicating the nature of the compound. Given the name of any binary molecule, you should be able to write the correct chemical formula. The prefixes used in a name indicate the exact number of each atom present in the molecule and determine the subscripts used in the formula. If you are having trouble writing formulas from the names for binary compounds, you might want to review the naming rules listed on the pages at the beginning of this lesson.

Get It?

Interpret How do you determine the correct subscripts to use in order to write the chemical formula of a binary molecular compound when given the compound's name?

The formula for an acid can also be derived from the name. It is helpful to remember that all binary acids contain hydrogen and one other element. Also, keep in mind that the first part of the name of a binary acid will always use the prefix *hydro-*.

For oxyacids—acids containing oxyanions—you will need to know the names of the common oxyanions. If you need to review the formulas and names of oxyanions, see **Table 9** in the previous module.

PRACTICE Problems ADDITIONAL PRACTICE

Identify the formula for each compound.

25. silver chloride

26. dihydrogen monoxide

27. chlorine trifluoride

28. diphosphorus trioxide

29. disulfur decafluoride

30. CHALLENGE What is the formula for carbonic acid?

Check Your Progress

Summary

- Names of covalent molecular compounds include prefixes for the number of each atom present. The final letter of the prefix is dropped if the element name begins with a vowel.
- Molecules that produce H^+ in solution are acids. Binary acids contain hydrogen and one other element. Oxyacids contain hydrogen and an oxyanion.

Demonstrate Understanding

31. **Summarize** the rules for naming binary molecular compounds.
32. **Define** a binary molecular compound.
33. **Describe** the difference between a binary acid and an oxyacid.
34. **Apply** Using the system of rules for naming binary molecular compounds, describe how you would name the molecule N_2O_4.
35. **Apply** Write the molecular formula for each of these compounds: *iodic acid, disulfur trioxide, dinitrogen monoxide*, and *hydrofluoric acid*.
36. **State** the molecular formula for each compound listed below.
    a. dinitrogen trioxide
    b. nitrogen monoxide
    c. hydrochloric acid
    d. chloric acid
    e. sulfuric acid
    f. sulfurous acid

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LESSON 3

MOLECULAR STRUCTURES

FOCUS QUESTION

How are electrons shared in covalent molecules?

Structural Formulas

You have already studied the structure of ionic compounds—substances formed from ionic bonds. Covalent molecules described in this module have structures that are different from those of ionic compounds. When studying the molecular structures of covalent compounds, various models are used as representations of the molecules.

The molecular formula, which shows the element symbols and numerical subscripts, tells you the type and number of each atom in a molecule. As shown in **Figure 13**, there are several different models that can be used to represent a molecule. Note that in the ball-and-stick and space-filling molecular models, atoms of each specific element are represented by spheres of a representative color, as shown in **Table R-1** in the Student Resources. These colors are used for identifying the atoms if the chemical symbol of the element is not present.

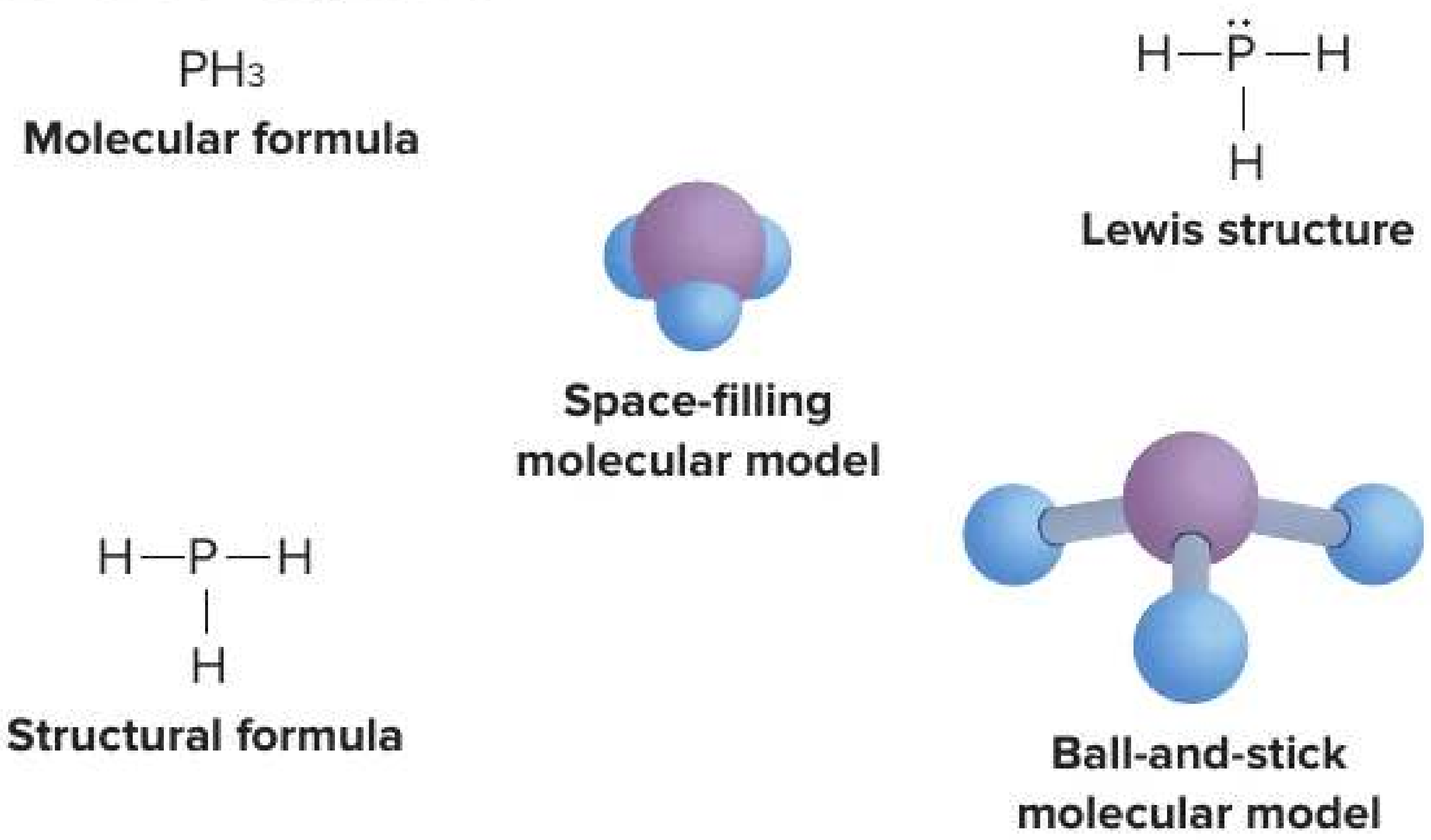


Figure 13 All of these models can be used to show the relative locations of atoms and electrons in the phosphorus trihydride (phosphine) molecule.

Compare and contrast *the types of information contained in each model.*

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

CCC **Identify Crosscutting Concepts**

Create a table of the crosscutting concepts and fill in examples you find as you read.

Revisit the Encounter the Phenomenon Question

What information from this lesson can help you answer the module question?

One of the most useful molecular models is the **structural formula,** which uses letter symbols and bonds to show relative positions of atoms. You can predict the structural formula for many molecules by drawing the Lewis structure. You have already seen some simple examples of Lewis structures, but more involved structures are needed to help you determine the shapes of molecules.

Lewis structures

Although it is fairly easy to draw Lewis structures for most compounds formed by nonmetals, it is a good idea to follow a regular procedure. Whenever you need to draw a Lewis structure, follow the steps outlined in this Problem-Solving Strategy.

PROBLEM-SOLVING STRATEGY

Drawing Lewis Structures

1. Predict the location of certain atoms. The atom that has the least attraction for shared electrons will be the central atom in the molecule. This element is usually the one closer to the left side of the periodic table. The central atom is located in the center of the molecule; all other atoms become terminal atoms.

Hydrogen is always a terminal, or end, atom. Because it can share only one pair of electrons, hydrogen can be connected to only one other atom.

2. Determine the number of electrons available for bonding.

This number is equal to the total number of valence electrons in the atoms that make up the molecule.

3. Determine the number of bonding pairs.

To do this, divide the number of electrons available for bonding by two.

4. Place the bonding pairs.

Place one bonding pair (single bond) between the central atom and each of the terminal atoms.

5. Determine the number of electron pairs remaining.

To do this, subtract the number of pairs used in Step 4 from the total number of bonding pairs determined in Step 3. These remaining pairs include lone pairs as well as pairs used in double and triple bonds. Place lone pairs around each terminal atom (except H atoms) bonded to the central atom to satisfy the octet rule. Any remaining pairs will be assigned to the central atom.

6. Determine whether the central atom satisfies the octet rule.

Is the central atom surrounded by four electron pairs? If not, it does not satisfy the octet rule. To satisfy the octet rule, convert one or two of the lone pairs on the terminal atoms into a double bond or a triple bond between the terminal atom and the central atom. These pairs are still associated with the terminal atom as well as with the central atom. Remember that carbon, nitrogen, oxygen, and sulfur often form double and triple bonds.

Apply the Strategy

Study Example Problems 3 through 5 to see how the steps in the Problem-Solving Strategy are applied.

EXAMPLE Problem 3

LEWIS STRUCTURE FOR A COVALENT COMPOUND WITH SINGLE BONDS Ammonia is a raw material used in the manufacture of many products, including fertilizers, cleaning products, and explosives. Draw the Lewis structure for ammonia (NH_3).

1 ANALYZE THE PROBLEM

Ammonia molecules consist of one nitrogen atom and three hydrogen atoms. Because hydrogen must be a terminal atom, nitrogen is the central atom.

2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons available for bonding.

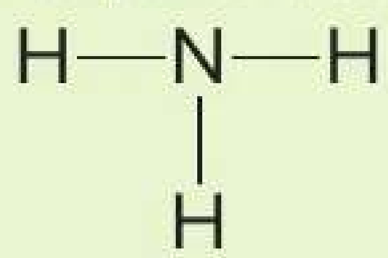
$$1 \text{ N atom} \times \frac{5 \text{ valence electrons}}{1 \text{ N atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{1 \text{ H atom}} = 8 \text{ valence electrons}$$

There are 8 valence electrons available for bonding.

$$\frac{8 \text{ electrons}}{2 \text{ electrons/pair}} = 4 \text{ pairs}$$

Determine the total number of bonding pairs. To do this, divide the number of available electrons by two.

Four pairs of electrons are available for bonding.



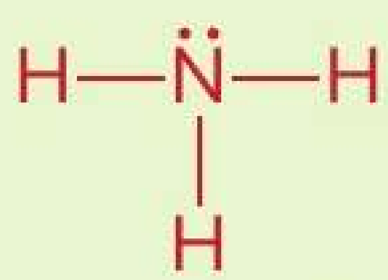
Place a bonding pair (a single bond) between the central nitrogen atom and each terminal hydrogen atom.

Determine the number of bonding pairs remaining.

4 pairs total — 3 pairs used
= 1 pair available

Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.

The remaining pair—a lone pair—must be added to either the terminal atoms or the central atom. Because hydrogen atoms can have only one bond, they have no lone pairs.



Place the remaining lone pair on the central nitrogen atom.

3 EVALUATE THE ANSWER

Each hydrogen atom shares one pair of electrons, as required, and the central nitrogen atom shares three pairs of electrons and has one lone pair, providing a stable octet.

PRACTICE Problems

ADDITIONAL PRACTICE

37. Draw the Lewis structure for BH_3.

38. CHALLENGE A nitrogen trifluoride molecule contains numerous lone pairs. Draw its Lewis structure.

CCC CROSSCUTTING CONCEPTS

System and System Models Refer to the five common molecular models, shown on the first page of this lesson. With a partner, choose 2 compounds listed in question number 36 in the lesson 2 Review It!. Individually, without sharing your work, use 5 index cards to draw each model type to represent each compound, using a total of 10 index cards. Share your cards with your partner and compare your models. If they differ, work together to reach an agreed configuration.

EXAMPLE Problem 4

LEWIS STRUCTURE FOR A COVALENT COMPOUND WITH MULTIPLE BONDS Carbon dioxide is a product of all cellular respiration. Draw the Lewis structure for carbon dioxide (CO_2).

1 ANALYZE THE PROBLEM

The carbon dioxide molecule consists of one carbon atom and two oxygen atoms. Because carbon has less attraction for shared electrons, carbon is the central atom, and the two oxygen atoms are terminal.

2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons available for bonding.

$$1 \text{ C atom} \times \frac{4 \text{ valence electrons}}{1 \text{ C atom}} + 2 \text{ O atoms} \times \frac{6 \text{ valence electrons}}{1 \text{ O atom}} = 16 \text{ valence electrons}$$

There are 16 valence electrons available for bonding.

$$\frac{16 \text{ electrons}}{2 \text{ electrons/pair}} = 8 \text{ pairs}$$

Determine the total number of bonding pairs by dividing the number of available electrons by two.

Eight pairs of electrons are available for bonding.

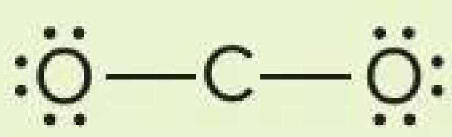
O—C—O

Place a bonding pair (a single bond) between the central carbon atom and each terminal oxygen atom.

Determine the number of electron pairs remaining.

8 pairs total − 2 pairs used
= 6 pairs available

Subtract the number of pairs used in these bonds from the total number of pairs of electrons available.



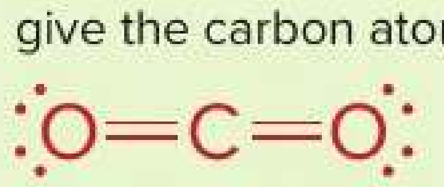
Add three lone pairs to each terminal oxygen atom.

Determine the number of electron pairs remaining.

6 pairs available − 6 pairs used
= 0 pairs available

Subtract the lone pairs from the pairs available.

Examine the incomplete structure above (showing the placement of the lone pairs). Note that the carbon atom does not have an octet and that there are no more electron pairs available. To give the carbon atom an octet, the molecule must form double bonds.



Use a lone pair from each O atom to form a double bond with the C atom.

3 EVALUATE THE ANSWER

Both carbon and oxygen now have an octet, which satisfies the octet rule.

PRACTICE Problems

ADDITIONAL PRACTICE

39. Draw the Lewis structure for ethylene, C_2H_4.

40. CHALLENGE A molecule of carbon disulfide contains both lone pairs and multiple-covalent bonds. Draw its Lewis structure.

Lewis structures for polyatomic ions

To find the total number of electrons available for bonding in a polyatomic ion, first find the number available in the atoms present in the ion. Then, add the ion charge if the ion is negative, or subtract the ion charge if the ion is positive. Compared to the number of valence electrons present in the atoms that make up the ion, more electrons are present if the ion is negatively charged and fewer are present if the ion is positive.

EXAMPLE Problem 5

LEWIS STRUCTURE FOR A POLYATOMIC ION Draw the correct Lewis structure for the polyatomic ion phosphate (PO_4^{3-}).

1 ANALYZE THE PROBLEM

You are given that the phosphate ion consists of one phosphorus atom and four oxygen atoms and has a charge of 3−. Because phosphorus has less attraction for shared electrons than oxygen, phosphorus is the central atom and the four oxygen atoms are terminal atoms.

2 SOLVE FOR THE UNKNOWN

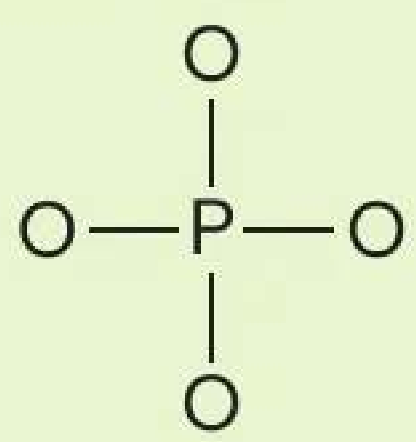
Find the total number of valence electrons available for bonding.

$$1 \text{ P atom} \times \frac{5 \text{ valence electrons}}{1 \text{ P atom}} + 4 \text{ O atoms} \times \frac{6 \text{ valence electrons}}{1 \text{ O atom}}$$

$$+\ 3 \text{ electrons from the negative charge} = 32 \text{ valence electrons}$$

$$\frac{32 \text{ electrons}}{2 \text{ electrons/pair}} = 16 \text{ pairs}$$

Determine the total number of bonding pairs.

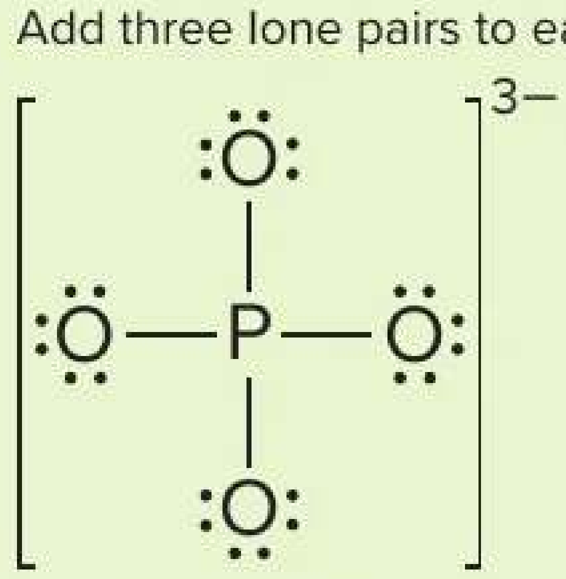


Draw single bonds from each terminal oxygen atom to the central phosphorus atom.

16 pairs total − 4 pairs used = 12 pairs available

Subtract the number of pairs used from the total number of pairs of electrons available.

Add three lone pairs to each terminal oxygen atom. 12 pairs available − 12 lone pairs used = 0



Subtracting the lone pairs used from the pairs available verifies that there are no electron pairs available for the phosphorus atom. The Lewis structure for the phosphate ion is shown.

3 EVALUATE THE ANSWER

All of the atoms have an octet, and the group has a net charge of 3−.

PRACTICE Problems

ADDITIONAL PRACTICE

41. Draw the Lewis structure for the NH_4^+ ion.

42. CHALLENGE The ClO_4^- ion contains numerous lone pairs. Draw its Lewis structure.

Resonance Structures

Using the same sequence of atoms, it is possible to have more than one correct Lewis structure when a molecule or polyatomic ion has both a double bond and a single bond. Consider the polyatomic ion nitrate (NO_3^-), shown in **Figure 14a**. Three equivalent structures can be used to represent the nitrate ion.

Resonance is a condition that occurs when more than one valid Lewis structure can be written for a molecule or ion. The two or more correct Lewis structures that represent a single molecule or ion are referred to as resonance structures. Resonance structures differ only in the position of the electron pairs, never the atom positions. The location of the lone pairs and bonding pairs differs in resonance structures. The molecule O_3 and the polyatomic ions NO_3^-, NO_2^-, SO_3^{2-}, and CO_3^{2-} all exhibit resonance.

It is important to note that each molecule or ion that exhibits resonance behaves as if it has only one structure. Refer to **Figure 14b**. Experimentally measured bond lengths show that the bonds are identical to each other. They are shorter than single bonds but longer than double bonds. The actual bond length is an average of the bonds in the resonance structures.

Real-World Chemistry

Phosphorus and Nitrogen

ALGAL BLOOMS Phosphorus and nitrogen are nutrients required for algae growth. Both can enter lakes and streams from discharges of sewage and industrial waste, and in fertilizer runoff. If these substances build up in a body of water, a rapid growth of algae, known as an algal bloom, can occur, forming a thick layer of green slime over the water's surface. When the algae use up the supply of nutrients, they die and decompose. This process reduces the amount of dissolved oxygen in the water that is available to other aquatic organisms.

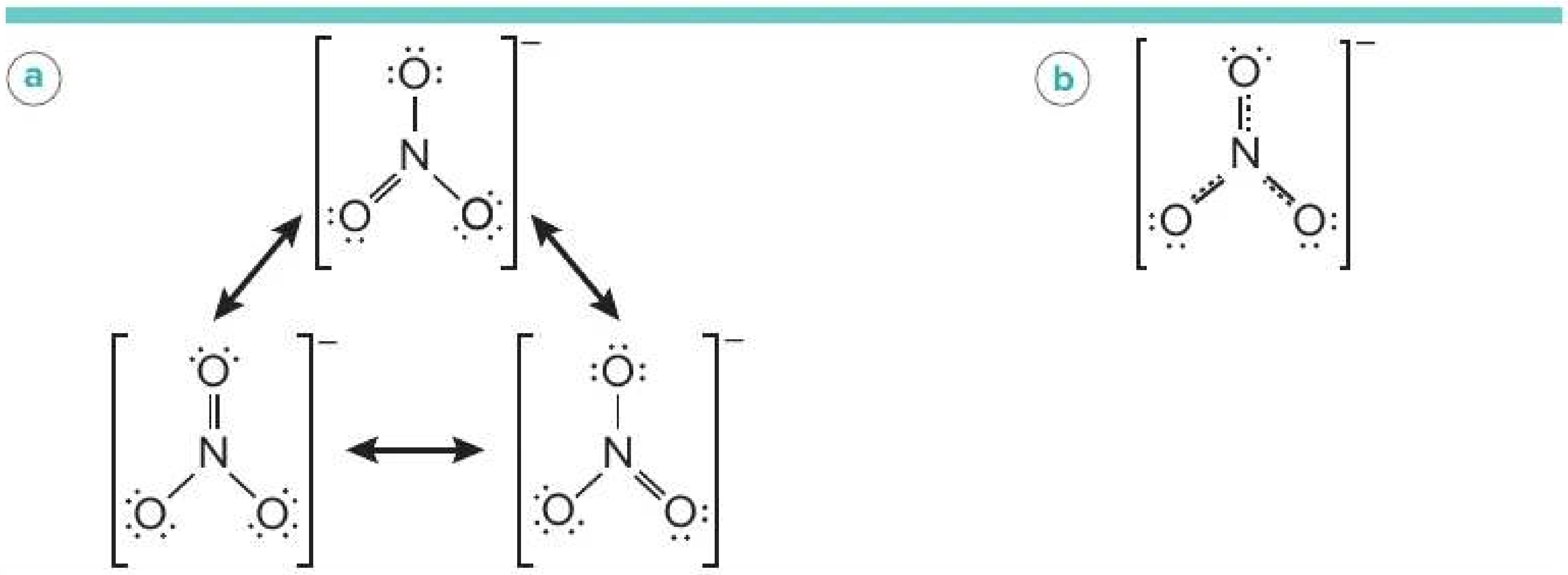


Figure 14 The nitrate ion (NO_3^-) exhibits resonance.

a. These resonance structures differ only in the location of the double bond. The locations of the nitrogen and oxygen atoms stay the same.

b. The actual nitrate ion is like an average of the three resonance structures in **a**. The dotted lines indicate possible locations of the double bond.

Jerome Wexler/Science Source

PRACTICE Problems ADDITIONAL PRACTICE

Draw the Lewis resonance structures for the following molecules.

43. NO_2^- **44.** SO_2 **45.** O_3

46. CHALLENGE Draw the Lewis resonance structure for the ion SO_3^{2-}.

Exceptions to the Octet Rule

Generally, atoms attain an octet when they bond with other atoms. Some molecules and ions, however, do not obey the octet rule. There are several reasons for these exceptions.

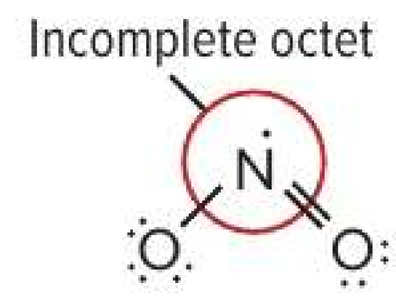
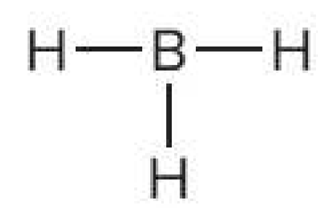


Figure 15 The central nitrogen atom in this NO_2 molecule does not satisfy the octet rule; the nitrogen atom has only seven electrons in its outer energy level.

Odd number of valence electrons

First, a small group of molecules might have an odd number of valence electrons and be unable to form an octet around each atom. For example, NO_2 has five valence electrons from nitrogen and 12 from oxygen, totaling 17, which cannot form an exact number of electron pairs. See **Figure 15**. ClO_2 and NO are other examples of molecules with odd numbers of valence electrons.

Suboctets and coordinate covalent bonds



Another exception to the octet rule is due to a few compounds that form suboctets—stable configurations with fewer than eight electrons present around an atom. This group is relatively rare, and BH_3 is an example. Boron, a group 13 metalloid, forms three covalent bonds with other nonmetallic atoms.The boron atom shares only six electrons—too few to form an octet. Such compounds tend to be reactive and can share an entire pair of electrons donated by another atom.

A **coordinate covalent bond** forms when one atom donates both of the electrons to be shared with an atom or ion that needs two electrons to form a stable electron arrangement with lower potential energy. Refer to **Figure 16**. Atoms or ions with lone pairs often form coordinate covalent bonds with atoms or ions that need two more electrons.

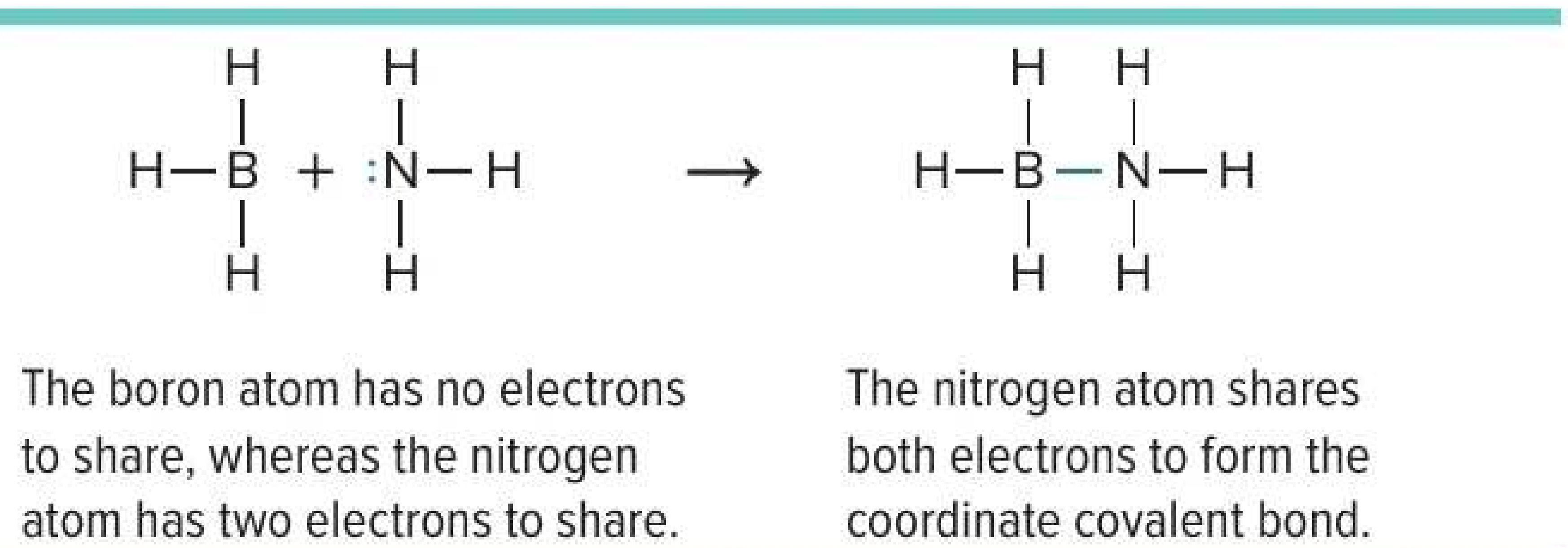


Figure 16 In this reaction between boron trihydride (BH_3) and ammonia (NH_3), the nitrogen atom donates both electrons that are shared by boron and nitrogen, forming a coordinate covalent bond.

Interpret *Does the coordinate covalent bond in the product molecule satisfy the octet rule?*

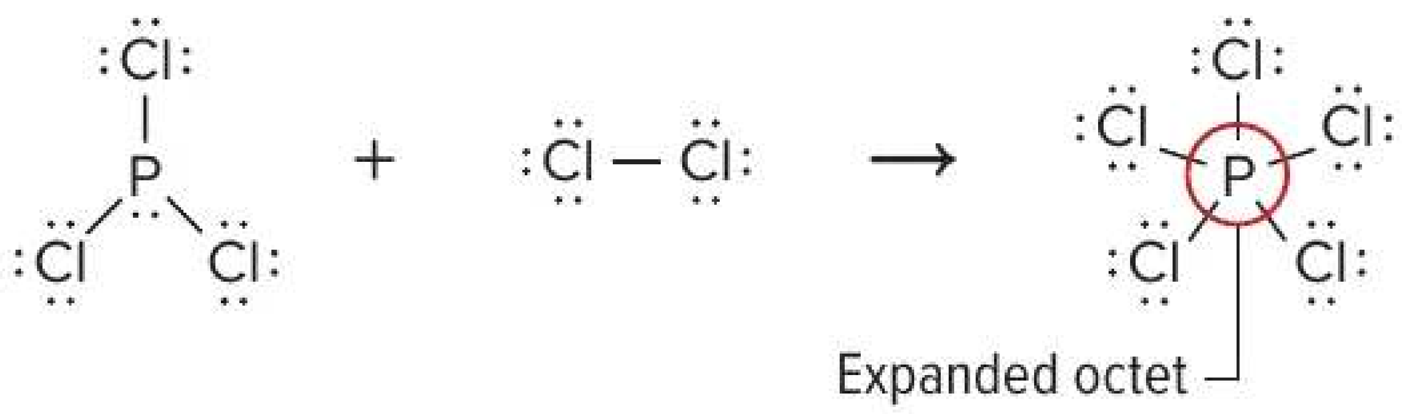


Figure 17 Prior to the reaction of PCl_3 and Cl_2, every reactant atom follows the octet rule. After the reaction, the product, PCl_5, has an expanded octet containing ten electrons.

Expanded octets

The third group of compounds that does not follow the octet rule has central atoms that contain more than eight valence electrons. This electron arrangement is referred to as an expanded octet. An expanded octet can be explained by considering the d orbitals that occur in the energy levels of elements in period three or higher. An example of an expanded octet, shown in **Figure 17**, is the bond formation in the molecule PCl_5. Five bonds are formed with ten electrons shared in one s orbital, three p orbitals, and one d orbital. Another example is the molecule SF_6, which has six bonds sharing 12 electrons in an s orbital, three p orbitals, and two d orbitals. When you draw the Lewis structures for these compounds, either extra lone pairs are added to the central atom or more than four bonding atoms are present in the molecule.

Get It?

Summarize three reasons why some molecules do not conform to the octet rule.

EXAMPLE Problem 6

LEWIS STRUCTURE: EXCEPTION TO THE OCTET RULE Xenon is a noble gas that will form a few compounds with nonmetals that strongly attract electrons. Draw the correct Lewis structure for xenon tetrafluoride (XeF_4).

1 ANALYZE THE PROBLEM

You are given that a molecule of xenon tetrafluoride consists of one xenon atom and four fluorine atoms. Xenon has less attraction for electrons, so it is the central atom.

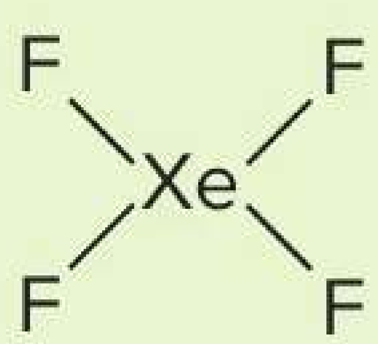
2 SOLVE FOR THE UNKNOWN

First, find the total number of valence electrons.

$$1 \text{ Xe atom} \times \frac{8 \text{ valence electrons}}{1 \text{ Xe atom}} + 4 \text{ F atoms} \times \frac{7 \text{ valence electrons}}{1 \text{ F atom}} = 36 \text{ valence electrons}$$

$$\frac{36 \text{ electrons}}{2 \text{ electrons/pair}} = 18 \text{ pairs}$$

Determine the total number of bonding pairs.



Use four bonding pairs to bond the four F atoms to the central Xe atom.

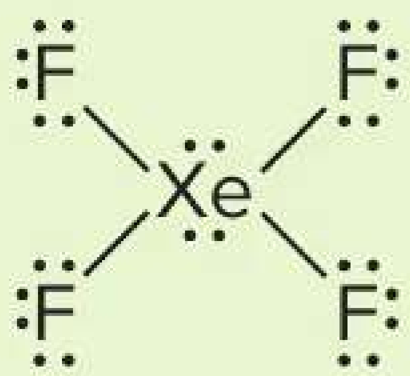
EXAMPLE Problem 6 (continued)

18 pairs available − 4 pairs used
= 14 pairs available

Determine the number of remaining pairs.

$14 \text{ pairs} - 4 \text{ F atoms} \times \frac{3 \text{ pairs}}{1 \text{ F atom}}$
$= 2 \text{ pairs unused}$

Add three pairs to each F atom to obtain an octet. Determine how many pairs remain.



Place the two remaining pairs on the central Xe atom.

3 EVALUATE THE ANSWER

This structure gives xenon 12 total electrons, an expanded octet. Xenon compounds, such as the XeF_4 shown here, are toxic because they are highly reactive.

PRACTICE Problems — ADDITIONAL PRACTICE

Draw the expanded octet Lewis structure for each molecule.

47. ClF_3

48. PCl_5

49. CHALLENGE Draw the Lewis structure for the molecule formed when six fluorine atoms and one sulfur atom bond covalenty.

Check Your Progress

Summary

- Different models can be used to represent molecules.
- Resonance occurs when more than one valid Lewis structure exists for the same molecule.
- Exceptions to the octet rule occur in some molecules.

Demonstrate Understanding

50. **Describe** the information contained in a structural formula.
51. **State** the steps used to draw Lewis structures in your own words.
52. **Summarize** exceptions to the octet rule by correctly pairing these molecules and phrases: odd number of valence electrons, PCl_5, ClO_2, BH_3, expanded octet, less than an octet.
53. **Evaluate** A classmate states that a binary compound having only sigma bonds displays resonance. Could the classmate's statement be true?
54. **Draw** the resonance structures for the dinitrogen oxide (N_2O) molecule.
55. **Draw** the Lewis structures for CN^-, SiF_4, HCO_3^-, and, AsF_6^-.

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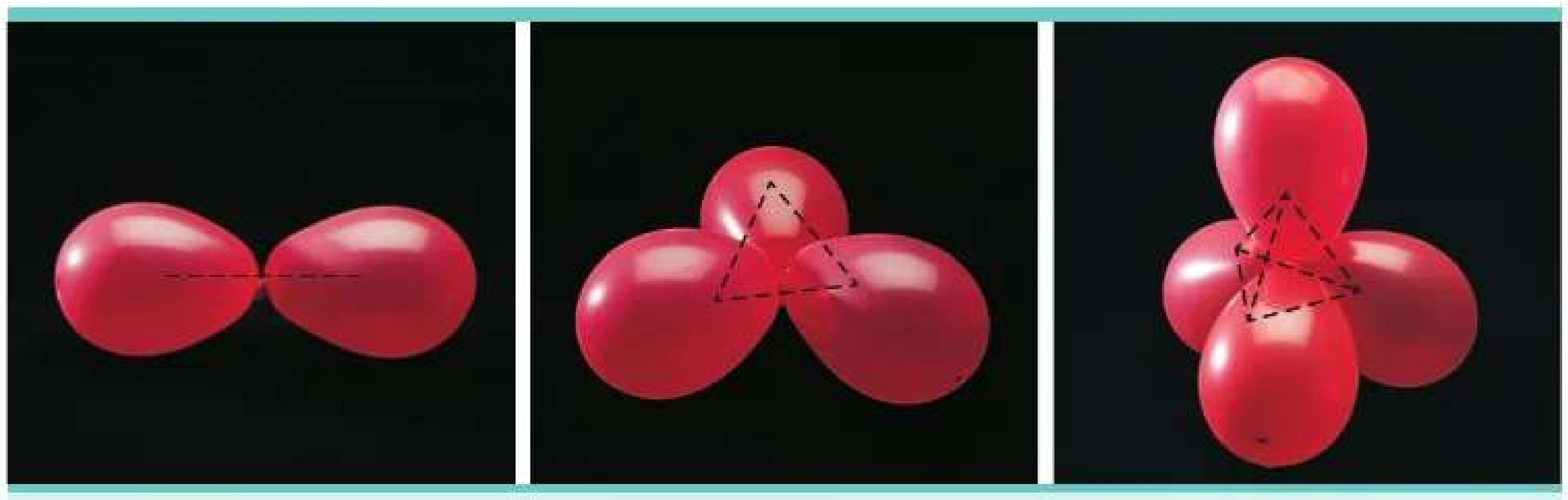
LESSON 4
MOLECULAR SHAPES

FOCUS QUESTION
What shapes do molecules form?

VSEPR Model

Once a Lewis structure is drawn, the molecular geometry, or shape, of a molecule can be determined using the **V**alence **S**hell **E**lectron **P**air **R**epulsion model, or **VSEPR model.** This model is based on an arrangement that minimizes the repulsion of shared and unshared electron pairs around the central atom. Recall that the attraction and repulsion between electric charges at the atomic scale explain the structure, properties, and transformations of matter.

To understand the VSEPR model better, imagine balloons that are inflated to similar sizes and tied together, as shown in **Figure 18**. Each balloon represents an electron-dense region. When a set of balloons is connected at a central point, which represents a central atom, the balloons naturally form a shape that minimizes interactions between the balloons.



Linear Trigonal planar Tetrahedral

Stephen Frisch/McGraw-Hill Education

Figure 18 Electron pairs in a molecule are located as far apart as they can be, just as these balloons are arranged. Two pairs form a linear shape. Three pairs form a trigonal planar shape. Four pairs form a tetrahedral shape.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE
Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE
GO ONLINE to find these activities and more resources.

Applying Practices: Modeling Electrostatic Forces—Covalent Bonding
HS-PS2-4. Use mathematical representations of Newton's Law of Gravitation and Coulomb's Law to describe and predict the gravitational and electrostatic forces between objects.

Virtual Investigation: Molecular Shapes
Develop and use models to discover the structure and function of atomic electric charges.

Bond angle

The electron pairs in a molecule repel one another in a similar way. These forces cause the atoms in a molecule to be positioned at fixed angles relative to one another. The angle formed by two terminal atoms and the central atom is a bond angle. Bond angles predicted by VSEPR are supported by experimental evidence.

Unshared pairs of electrons are also important in determining the shape of the molecule. These electrons occupy a slightly larger orbital than shared electrons. Therefore, shared bonding orbitals are pushed together by unshared pairs.

BIOLOGY Connection The shape of a molecule determines many of its physical and chemical properties. For example, scientists discovered that the shape of food molecules is important to our sense of taste. The surface of your tongue is covered with taste buds, each of which contains from 50 to 100 taste receptor cells. Taste receptor cells can detect five distinct tastes—sweet, bitter, salty, sour, and MSG (monosodium glutamate).

The shapes of food molecules are determined by their chemical structures. When a molecule enters a taste bud, it must have the correct shape for the nerve in each receptor cell to respond and send a message to the brain. The brain then interprets the message as a certain taste. When such molecules bind to sweet receptors, they are sensed as sweet. The greater the number of food molecules that fit a sweet receptor cell, the sweeter the food tastes. Sugars and artificial sweeteners are not the only sweet molecules. Some proteins found in fruits are also sweet molecules. Some common molecular shapes are illustrated in **Table 5** and **Table 6** on the following pages.

Hybridization

Sometimes atomic orbitals fuse to form a new hybridized orbital. **Hybridization** occurs when two things are combined and the result has characteristics of both. For example, during chemical bonding, if electrons come from 2 different atomic orbitals, such as a p or an s, they must rearrange or combine into a hybrid orbital with the same shape and energy level. To understand hybrid orbitals, consider the bonding involved in the methane molecule (CH_4). The hybrid orbitals in a carbon atom are shown in **Figure 19** on the next page in blue. Note, as shown in the electron notation, that although carbon initially has only 2 electrons in its p orbital, a 1 s electron is promoted from the s to the p orbital so that a total of 4 unpaired electrons can be shared. These 4 unpaired electrons are rearranged into four hybrid sp^3 orbitals that can now be used to bond with 4 hydrogen atoms.

WORD ORIGIN

trigonal planar

comes from the Latin words *trigonum*, which means *triangular, and plan-*, which means *flat*

CCC CROSSCUTTING CONCEPTS

Structure and Function Designing new systems requires examining properties of materials, structures of components, and connections of components to solve a problem. Write a blog post for a science web site that applies this statement to developing robots with artificial taste receptors that mimic human ones.

The hybrid orbitals in a carbon atom are shown in **Figure 19**. The hybrid orbital is called an sp^3 orbital because the four hybrid orbitals form from one s orbital and three p orbitals.

The number of atomic orbitals that mix and form the hybrid orbital equals the total number of pairs of electrons, as shown in **Table 5** and **Table 6**. In addition, the number of hybrid orbitals formed equals the number of atomic orbitals mixed. For example, $AlCl_3$ has a total of three pairs of electrons and VSEPR predicts a trigonal planar molecular shape. This shape results when one s and two p orbitals on the central atom, Al, mix and form three identical sp^2 hybrid orbitals.

Lone pairs also occupy hybrid orbitals. Compare the hybrid orbitals of $BeCl_2$ in **Table 5** and H_2O in **Table 6**. Both compounds contain three atoms. Why does an H_2O molecule contain sp^3 orbitals? There are two lone pairs on the central oxygen atom in H_2O. Therefore, there must be four hybrid orbitals—two for bonding and two for the lone pairs.

Recall that multiple covalent bonds consist of one sigma bond and one or more pi bonds. Only the two electrons in the sigma bond occupy hybrid orbitals such as sp and sp^2. The remaining unhybridized p orbitals overlap to form pi bonds. It is important to note that single, double, and triple covalent bonds contain only one hybrid orbital. Thus, CO_2, with two double bonds, forms sp hybrid orbitals.

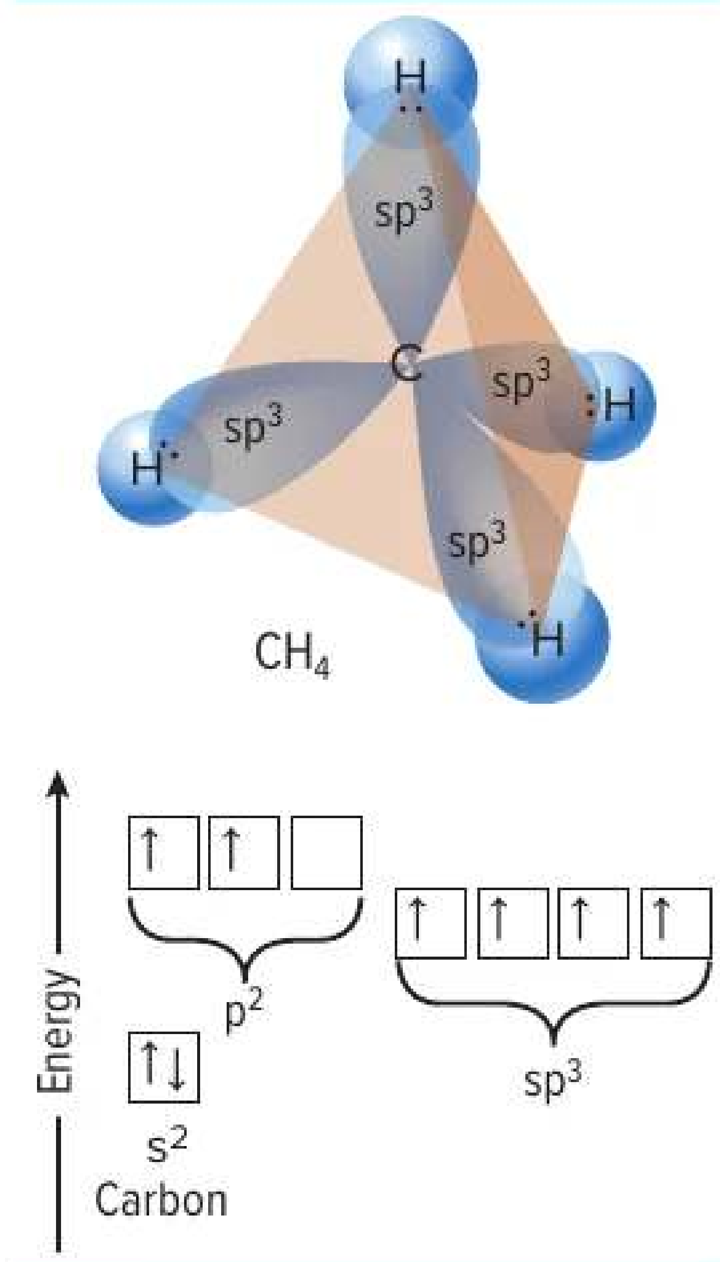


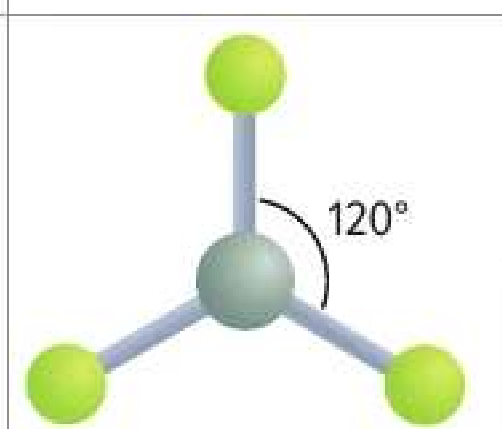
Figure 19 Notice that the hybrid orbitals have an intermediate amount of potential energy when compared with the energy of the original s and p orbitals. According to VSEPR theory, a tetrahedral shape minimizes repulsion between the hybrid orbitals in a CH_4 molecule.

Identify *How many faces does the tetrahedral shape formed by the sp^3 orbitals have?*

Get It?

Describe how repulsion between electric charges at the atomic scale explains the structure of a methane molecule.

Table 5 **Molecular Shapes: 2 or 3 Total Pairs**

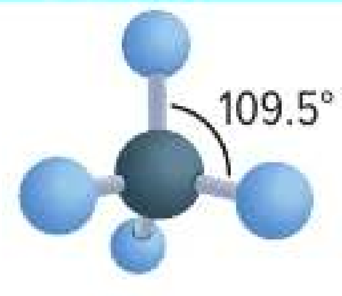
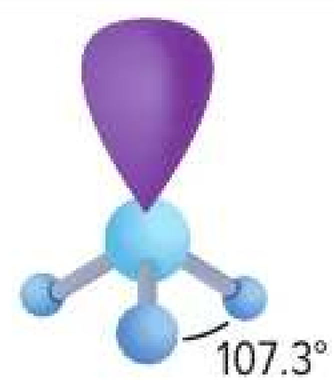
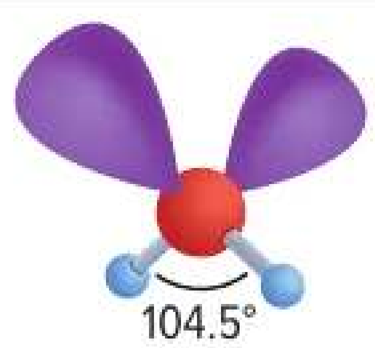
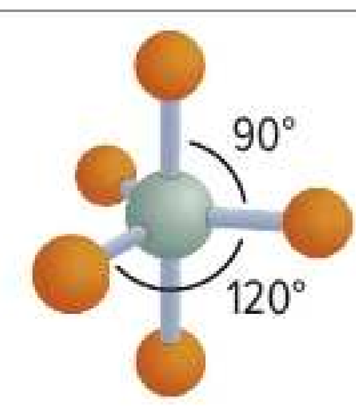
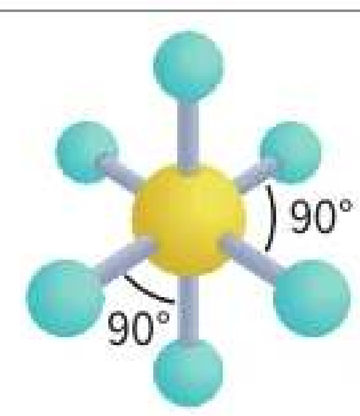
Molecule	Total Pairs	Shared Pairs	Lone Pairs	Hybrid Orbitals	Molecular Shape*
$BeCl_2$	2	2	0	sp	 Linear
$AlCl_3$	3	3	0	sp^2	 Trigonal planar

The $BeCl_2$ molecule contains only two pairs of electrons shared with the central Be atom. These bonding electrons have the maximum separation, a bond angle of 180°, and the molecular shape is linear.

The three bonding electron pairs in $AlCl_3$ have maximum separation in a trigonal planar shape with 120° bond angles.

*Balls represent atoms, sticks represent bonds, and lobes represent lone pairs of electrons.

Table 6 **Molecular Shapes: 4, 5, or 6 Total Pairs**

Molecule	Total Pairs	Shared Pairs	Lone Pairs	Hybrid Orbitals	Molecular Shape*
CH_4	4	4	0	sp^3	 Tetrahedral
NH_3	4	3	1	sp^3	 Trigonal pyramidal
H_2O	4	2	2	sp^3	 Bent
$NbBr_5$	5	5	0	sp^3d	 Trigonal bipyramidal
SF_6	6	6	0	sp^3d^2	 Octahedral

When the central atom in a molecule has four pairs of bonding electrons, as CH_4 does, the shape is tetrahedral. The bond angles are 109.5°.

NH_3 has three single covalent bonds and one lone pair. The lone pair takes up a greater amount of space than the shared pairs. There is stronger repulsion between the lone pair and the bonding pairs than between two bonding pairs. The resulting geometry is trigonal pyramidal, with 107.3° bond angles.

Water has two covalent bonds and two lone pairs. Repulsion between the lone pairs causes the angle to be 104.5°, less than both tetrahedral and trigonal pyramid. As a result, water molecules have a bent shape.

The $NbBr_5$ molecule has five pairs of bonding electrons. The trigonal bipyramidal shape minimizes the repulsion of these shared electron pairs.

As with $NbBr_5$, SF_6 has no unshared electron pairs on the central atom. However, six shared pairs arranged about the central atom result in an octahedral shape.

*Balls represent atoms, sticks represent bonds, and lobes represent lone pairs of electrons.

EXAMPLE Problem 7

FIND THE SHAPE OF A MOLECULE Phosphorus trihydride, a colorless gas, is produced when organic materials, such as fish flesh, rot. What is the shape of a phosphorus trihydride molecule? Predict the bond angle and identify hybrid orbitals.

1 ANALYZE THE PROBLEM

A phosphorus trihydride molecule has three hydrogen atoms bonded to a central phosphorus atom.

EXAMPLE Problem 7 (continued)

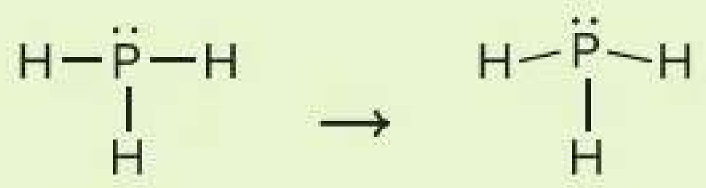
2 SOLVE FOR THE UNKNOWN

Find the total number of valence electrons and the number of electron pairs.

$$1 \text{ P atom} \times \frac{5 \text{ valence electrons}}{1 \text{ P atom}} + 3 \text{ H atoms} \times \frac{1 \text{ valence electron}}{1 \text{ H atom}} = 8 \text{ valence electrons}$$

$$\frac{8 \text{ electrons}}{2 \text{ electrons/pair}} = 4 \text{ pairs}$$

Determine the total number of bonding pairs.



Lewis structure **Molecular shape**

Draw the Lewis structure, using one pair of electrons to bond each H atom to the central P atom and assigning the lone pair to the P atom.

The molecular shape is trigonal pyramidal with a predicted 107° bond angle and sp^3 hybrid orbitals.

3 EVALUATE THE ANSWER

All electron pairs are used and each atom has a stable electron configuration.

PRACTICE Problems ADDITIONAL PRACTICE

Determine the molecular shape, bond angle, and hybrid orbitals for each molecule.

56. BF_3 **57.** OCl_2 **58.** BeF_2 **59.** CF_4

60. CHALLENGE For a NH_4^+ ion, identify its molecular shape, bond angle, and hybrid orbitals.

Check Your Progress

Summary

- VSEPR model theory states that electron pairs repel each other and determine both the shape of and bond angles in a molecule.
- Hybridization explains the observed shapes of molecules by the presence of equivalent hybrid orbitals.

Demonstrate Understanding

61. **Summarize** how the VSEPR model helps explain how electric charges affect bonding and molecular shape in covalent compounds.
62. **Define** the term *bond angle*.
63. **Apply** Use the term *hybridization* to describe the bonds in a methane molecule.
64. **Compare** the size of an orbital that has a shared electron pair with one that has a lone pair.
65. **Identify** the type of hybrid orbitals present and bond angles for a molecule with a tetrahedral shape.
66. **Compare** the molecular shapes and hybrid orbitals of PF_3 and PF_5 molecules. Explain why their shapes differ.
67. **List** in a table the Lewis structure, molecular shape, bond angle, and hybrid orbitals for molecules of CS_2, CH_2O, H_2Se, CCl_2F_2, and NCl_3.

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LESSON 5

ELECTRONEGATIVITY AND POLARITY

FOCUS QUESTION

How does molecular shape affect the way that covalent compounds are held together?

Electronegativity and Bond Character

The type of bond formed during a reaction is related to each atom's attraction for electrons. The version of the periodic table of the elements shown in **Figure 20** lists electronegativity values. Recall that electronegativity indicates the relative ability of an atom to attract electrons for bonding. Because noble gases do not generally form compounds, the noble gases are not shown on this table.

Electronegativity Values for Selected Elements

Metal
Metalloid
Nonmetal

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
1	1 H 2.20																
2	3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98
3	11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
4	19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96
5	37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66
6	55 Cs 0.79	56 Ba 0.89	57 La 1.10	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2
7	87 Fr 0.7	88 Ra 0.9	89 Ac 1.1														

Figure 20 Electronegativity values are derived by comparing an atom's attraction for shared electrons to that of a fluorine atom's attraction for shared electrons. Note, the electonegativity values for the lanthanide and actinide series, which are not shown, range from 1.12 to 1.7.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Applying Practices: Communicate Information about Contact and Noncontact Forces

HS-PS2-6. Communicate scientific and technical information about why the molecular-level structure is important in the functioning of designed materials.

Revisit the Encounter the Phenomenon Question

What information from this lesson can help you answer the module question?

Table 7 EN Difference Between Atoms within a Compound and Bond Character

Electronegativity Difference	Bond Character
> 1.7	mostly ionic
0.4 – 1.7	polar covalent
< 0.4	mostly covalent
0	nonpolar covalent

Electronegativity

The scale of electronegativities—shown in **Figure 20**—allows chemists to evaluate the affinity of specific atoms in a compound for electrons. Electonegativity is a measure of the tendency of an atom to accept an electron. Excluding noble gases, electronegativity increases with increasing atomic number within a period and decreases with increasing atomic number within a group. This shows that some atoms are more capable of attracting electrons than others. The explanation for why this occurs becomes clear when you consider the number of valence electrons an atom has. As you move from left to right on the table, the number of electrons needed to complete the octet rule becomes less, thereby increasing an atom's affinity for electrons in order to reach a more stable state.

Note that flourine has the greatest electronegativity value (3.98), while francium has the least (0.7). Although xenon is a noble gas, it can sometimes form bonds with highly electonegative atoms, such as flourine.

Get It?

Explain how electric charges at the atomic scale relate to the electronegativity of an atom.

Bond character

A chemical bond between atoms of different elements is never completely ionic or covalent. The character of a bond depends on how strongly each of the bonded atoms attracts electrons. As shown in **Table 7**, the character and type of a chemical bond can be predicted using the electronegativity difference between atoms that bond within a compound. Electrons in bonds between identical atoms have an electronegativity difference of zero—meaning that the electrons are equally shared between the two atoms. This type of bond is considered nonpolar covalent, or a pure covalent bond.

On the other hand, because different elements have different electronegativities, the electron pairs in a covalent bond between different atoms are not shared equally. Unequal sharing results in a **polar covalent bond.** When there is a large difference in the electronegativity between bonded atoms, an electron is transferred from one atom to the other, which results in bonding that is primarily ionic.

Bonding is not often clearly ionic or covalent. An electronegativity difference of 1.70 is considered 50 percent covalent and 50 percent ionic. As the difference in electronegativity increases, the bond becomes more ionic in character. As shown in **Table 7,** ionic bonds form when the electronegativity difference is greater than 1.70. However, this cutoff is sometimes inconsistent with experimental observations of two nonmetals bonding together.

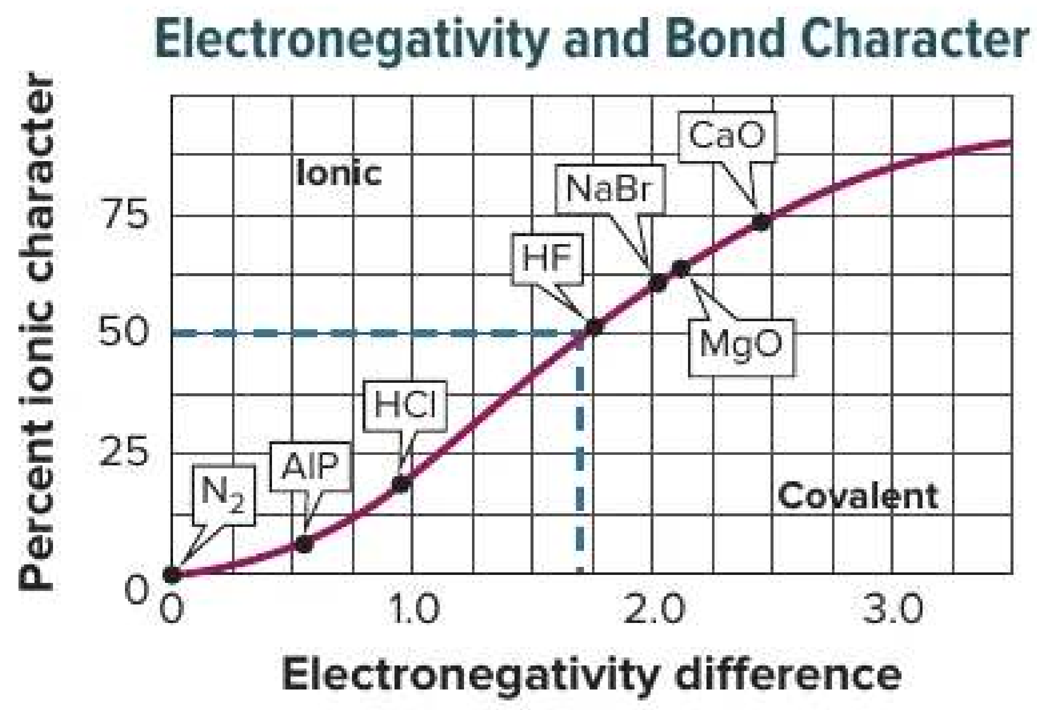


Figure 21 This graph shows that the difference in electronegativity between bonding atoms determines the percent ionic character of the bond. Above 50% ionic character, bonds are mostly ionic.

Figure 21 summarizes the range of chemical bonding between two atoms. What percent ionic character is a bond between two atoms that have an electronegativity difference of 2.00? Where would LiBr be plotted on the graph?

Get It?
Describe how electronegativity relates to polar covalent bonding.

Polar Covalent Bonds

As you just learned, polar covalent bonds form because not all atoms that share electrons attract them equally. A polar covalent bond is similar to a tug-of-war in which the two teams are not of equal strength. Although both sides share the rope, the stronger team pulls more of the rope toward its side. When a polar bond forms, the shared electron pair or pairs are pulled toward one of the atoms. Thus, the electrons spend more time around that atom than the other atom. This results in partial charges at the ends of the bond.

The Greek letter delta (δ) is used to represent a partial charge. In a polar covalent bond, δ^- represents a partial negative charge and δ^+ represents a partial positive charge. As shown in **Figure 22**, δ^- and δ^+ can be added to a molecular model to indicate the polarity of the covalent bond. The more-electronegative atom is at the partially negative end, while the less-electronegative atom is at the partially positive end. The resulting polar bond often is referred to as a dipole (two poles).

Electronegativity	Cl = 3.16
Electronegativity	H = 2.20
Difference	= 0.96

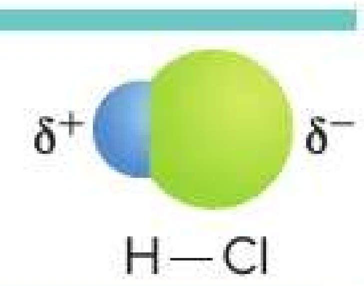


Figure 22 Chlorine's electronegativity is higher than that of hydrogen. Therefore, in a molecule containing hydrogen and chlorine, the shared pair of electrons is with the chlorine atom more often than it is with the hydrogen atom. Symbols are used to indicate the partial charge at each end of the molecule from this unequal sharing of electrons.

Molecular polarity

Covalently bonded molecules are either polar or nonpolar; which type depends on the location and nature of the covalent bonds in the molecule. A distinguishing feature of nonpolar molecules is that they are not attracted by an electric field. Polar molecules, however, are attracted by an electric field. Because polar molecules are dipoles with partially charged ends, they have an uneven electron density. This results in the tendency of polar molecules to align with an electric field.

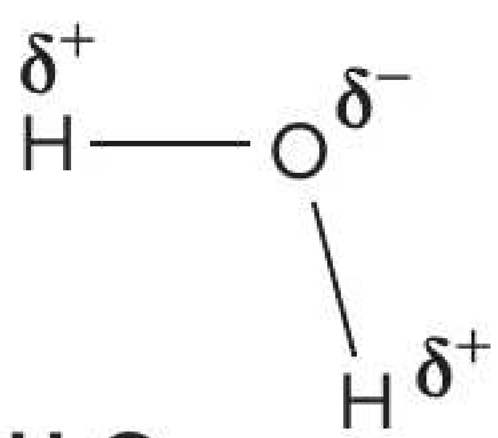
Polarity and molecular shape

You can learn why some molecules are polar and some are not by comparing water (H_2O) and carbon tetrachloride (CCl_4) molecules. Both molecules have polar covalent bonds. According to **Figure 20** earlier in the lesson, the electronegativity difference between a hydrogen atom and an oxygen atom is 1.24. The electronegativity difference between a chlorine atom and a carbon atom is 0.61. Although these electronegativity differences vary, a H—O bond and a C—Cl bond are considered to be polar covalent.

$$\overset{\delta^+}{\text{H}}\text{—}\overset{\delta^-}{\text{O}} \qquad \overset{\delta^+}{\text{C}}\text{—}\overset{\delta^-}{\text{Cl}}$$

According to their molecular formulas, both molecules have more than one polar covalent bond. However, only the water molecule is polar. Why might one molecule with polar covalent bonds be polar, while a second molecule with polar covalent bonds is nonpolar? The answer lies in the shapes of the molecules.

The shape of an H_2O molecule, as determined by VSEPR, is bent because the central oxygen atom has lone pairs of electrons, as shown in **Figure 23**. Because the polar H—O bonds are asymmetric in a water molecule, the molecule has a definite positive end and a definite negative end. Thus, it is polar.



H_2O

Figure 23 The bent shape of a water molecule makes it polar.

STEM CAREER Connection

Surface Chemist

Surface chemistry involves the study of chemical reactions at surfaces and interfaces, usually between two different solid, liquid, or gas phases. Surface chemists are interested in developing new surface treatments, using materials such as polymer films or nanoparticles to solve industrial problems. A bachelor's degree, master's degree, or Ph.D. is needed, as an understanding of chemical synthesis is required.

CCC CROSSCUTTING CONCEPTS

Patterns Use graph paper to construct the graph in **Figure 21**. Calculate the electronegativity differences for the following : PbCl, CH, CO, NaCl, LiI, SiO. Plot the values on the curve to determine how they are bonded. What evidence can you cite to explain the pattern in electronegativity when moving left to right on the periodic table?

A CCl_4 molecule is tetrahedral, and therefore, symmetrical, as shown in **Figure 24**. The electric charge measured at any distance from its center is identical to the charge measured at the same distance to the opposite side. The average center of the negative charge is located on the carbon atom. The positive center is also located on the carbon atom. Because the partial charges are balanced, CCl_4 is a nonpolar molecule. Note that symmetric molecules are usually nonpolar, and molecules that are asymmetric are polar as long as the bond type is polar.

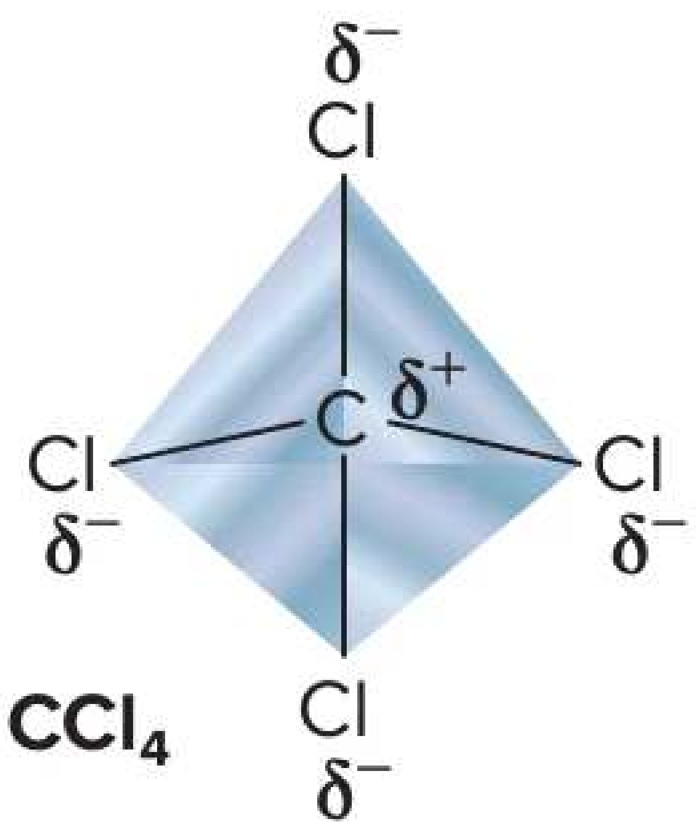


Figure 24 The symmetry of a CCl_4 molecule results in an equal distribution of charge, and the molecule is nonpolar.

Is the molecule of ammonia (NH_3), shown in **Figure 25**, polar? It has a central nitrogen atom and three terminal hydrogen atoms. Its shape is trigonal pyramidal because of the lone pair of electrons present on the nitrogen atom. Using **Figure 20**, you can find that the electronegativity difference of hydrogen and nitrogen is 0.84, making each N–H bond polar covalent. The charge distribution is unequal because the molecule is asymmetric. Thus, the molecule is polar.

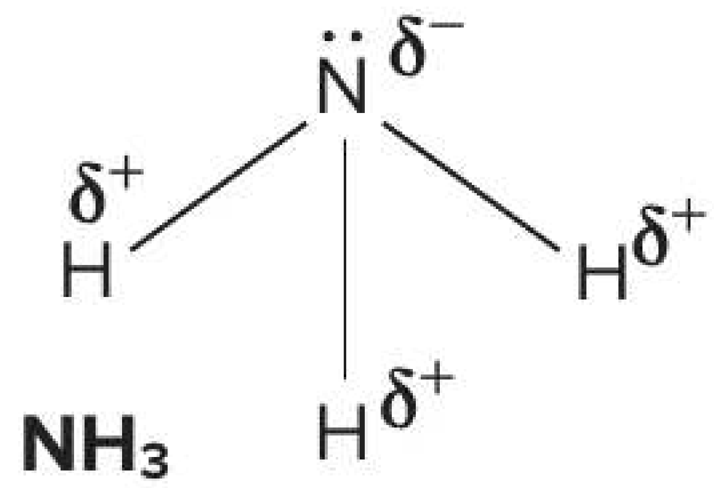


Figure 25 The asymmetric shape of an ammonia molecule results in an unequal charge distribution, and the molecule is polar.

Properties of Covalent Compounds

Table salt, an ionic solid, and table sugar, a covalent solid, are similar in appearance. However, these compounds behave differently when heated. Salt does not melt, but sugar melts at a relatively low temperature. Does the type of bonding in a compound affect its properties?

Intermolecular forces

Differences in properties of atoms are a result of differences in both attractive and repulsive forces. In a covalent compound, the covalent bonds between atoms in molecules are strong, but the attraction forces between molecules are relatively weak. These weak attraction forces are known as intermolecular forces, or van der Waals forces. Intermolecular forces vary in strength but are weaker than the bonds that join atoms in a molecule or those that join ions in an ionic compound.

Get It?
Distinguish covalent bonds from van der Waals forces.

Van der Waals forces, unlike ionic or covalent bonds, are distance-dependent, meaning that as the distance between interacting molecules increases, the van der Waals forces quickly vanish and are no longer present.

Figure 26 Symmetric covalent molecules, such as oil and most petroleum products, are nonpolar. Asymmetric molecules, such as water, are usually polar. As shown in this photo, polar and nonpolar substances usually do not mix.

Infer *Will water alone clean oil from a fabric?*

There are different types of intermolecular forces. Between nonpolar molecules, the force is weak and is called a dispersion force, or induced dipole. The force between oppositely charged ends of two polar molecules is called a dipole-dipole force. The more polar the molecule, the stronger the dipole-dipole force. The third force, a hydrogen bond, is especially strong. It forms between the hydrogen end of one dipole and a fluorine, oxygen, or nitrogen atom on another dipole. You will study intermolecular forces in more detail when you study states of matter.

Get It?

Compare dispersion forces, dipole-dipole forces, and hydrogen bonds.

Solubility of polar molecules

The physical property known as solubility is the ability of a substance to dissolve in another substance. The bond type and the shape of the molecules present determine solubility. Polar molecules and ionic compounds are usually soluble in polar substances, but nonpolar molecules dissolve only in nonpolar substances, as shown in **Figure 26**.

Forces and properties

The properties of covalent molecular compounds are related to the relatively weak intermolecular forces holding the molecules together. These weak forces result in the relatively low melting and boiling points of molecular substances compared with those of ionic substances. That is why, when heated moderately, sugar melts but salt does not.

Weak intermolecular forces also explain why many molecular substances exist as gases or vaporize readily at room temperature. Oxygen (O_2), carbon dioxide (CO_2), and hydrogen sulfide (H_2S) are examples of covalent compounds that are gases at room temperature. Because the hardness of a substance depends on the intermolecular forces between individual molecules, many covalent molecules are relatively soft solids. Paraffin, found in candles and other products, is a common example of a covalent solid.

In the solid phase, molecules align to form a crystal lattice. This molecular lattice is similar to that of an ionic solid, but with less attraction between particles. The structure of the lattice is affected by molecular shape and the type of intermolecular force. Most molecular information has been determined by studying molecular solids.

Covalent Network Solids

There are some solids, often called covalent network solids, that are composed only of atoms interconnected by a network of covalent bonds. Quartz and diamond are two common examples of network solids.

In contrast to molecular solids, network solids are typically brittle, nonconductors of heat or electricity, and extremely hard. Analyzing the structure of a diamond explains some of its properties. In a diamond, each carbon atom is bonded to four other carbon atoms. This tetrahedral arrangement, which is shown in **Figure 27**, forms a strongly bonded crystal system that is extremely hard and has a very high melting point.

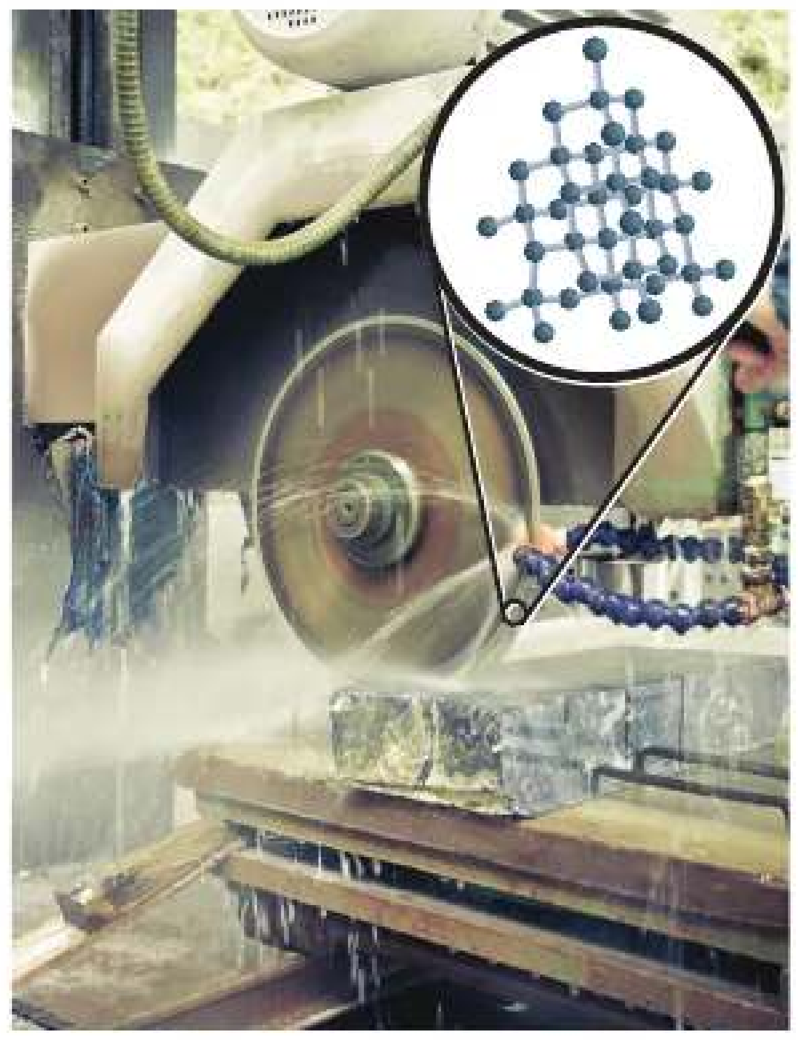


Figure 27 Network solids are often used in cutting tools because of their extreme hardness. Here, a diamond-tipped saw blade cuts through stone.

Check Your Progress

Summary

- Attraction and repulsion of electrical charges at the atomic scale explain the electronega-tivity, bonding, and polar characteristics of compounds.
- The electronegativity difference between atoms in a compound determines the character of a bond between atoms.
- Polar bonds occur when electrons are not shared equally, forming a dipole.
- The spatial arrangement of polar bonds in a molecule determines the overall polarity of a molecule.
- Molecules attract each other by weak intermolecular forces. In a covalent network solid, each atom is covalently bonded to many other atoms.

Demonstrate Understanding

68. **Summarize** how electonegativity relates to bonding and what pattern of electronegativity is observed in elements as you move from left to right on the periodic table.
69. **Describe** a polar covalent bond.
70. **Describe** a polar molecule.
71. **List** three properties of a covalent compound in the solid phase.
72. **Categorize** bond types using electronegativity differences.
73. **Generalize** Describe the general characteristics of covalent network solids.
74. **Predict** the type of bond that will form between the following pair of atoms:
    a. H and S b. C and H c. Na and S
75. **Identify** each molecule as polar or nonpolar: SCl_2, CS_2, and CF_4.
76. **Determine** whether a compound made of hydrogen and sulfur atoms is polar or nonpolar.
77. **Draw** the Lewis structures for the molecules SF_4 and SF_6. Analyze each structure to determine whether the molecule is polar or nonpolar.

LEARNSMART® Go online to follow your personalized learning path to review, practice, and reinforce your understanding.

SCIENCE & SOCIETY

Plastics: The Good, The Bad, and the Ugly

Plastic's properties make it a very versatile material. Its chemical structure makes it strong, lightweight, and easy to produce. However, because plastic doesn't decompose very easily, it's causing some very worrying effects on the environment.

Large patches of plastic waste and microplastics float in Earth's oceans.

Past, Present, and Future

The first completely synthetic plastic, Bakelite, was invented in 1907. By the 1950s, plastic could be found in almost every room in an American house, from plastic-wrapped baked goods in the kitchen to plastic-coated children's toys in the nursery.

Plastic's properties are due to its structure. Plastic is formed when carbon monomers covalently bond in long chains. The resulting material is easily moldable and durable. However, the arrangement of bonds in synthetic plastic polymers makes them difficult to be broken by bacteria and other decomposers. The result is that most plastics aren't biodegradable.

Many plastics are recyclable, but the processes used to recycle plastics are expensive. As a result, much of the world's plastic doesn't make it into the recycling bin, and instead finds its way into the environment, carried by runoff into the ocean.

The problem of plastic pollution isn't limited to unsightly bags and bottles lying on the beach. Once in the environment, plastic can be broken down by sunlight into smaller pieces called microplastics. After they find their way into an ocean ecosystem, microplastics have long-reaching effects on marine life.

Thanks to education and advocacy groups, the public is paying attention to this problem. Some communities have recycling processes that are easier for the public to use, and many are using less plastic. Some scientists are concentrating on creating plant-based plastics that are more biodegradable, while others are studying bacteria and fungi that are able to use plastic for energy. The key to solving the problem with plastics will be to find a balance between making plastics work for both people and the environment.

EVALUATE A DESIGN SOLUTION

Research one way in which scientists are trying to solve the problem of plastic in the environment. Create an infographic or other visual display about how this solution might work in your community.

A. & F. Michler/Photolibrary/Getty Images

MODULE 7
STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 THE COVALENT BOND

- Covalent bonds form when atoms share one or more pairs of electrons.
- Sharing one pair, two pairs, and three pairs of electrons forms single, double, and triple covalent bonds, respectively.
- Orbitals overlap directly in sigma bonds. Parallel orbitals overlap in pi bonds. A single covalent bond is a sigma bond, but multiple covalent bonds are made of both sigma and pi bonds.
- Bond dissociation energy is needed to break a covalent bond.

- covalent bond
- molecule
- Lewis structure
- sigma bond
- pi bond
- endothermic reaction
- exothermic reaction

Lesson 2 NAMING MOLECULES

- Names of covalent molecular compounds include prefixes for the number of each atom present. The final letter of the prefix is dropped if the element name begins with a vowel.
- Molecules that produce H^+ in solution are acids. Binary acids contain hydrogen and one other element. Oxyacids contain hydrogen and an oxyanion.

- oxyacid

Lesson 3 MOLECULAR STRUCTURES

- Different models can be used to represent molecules.
- Resonance occurs when more than one valid Lewis structure exists for the same molecule.
- Exceptions to the octet rule occur in some molecules.

- structural formula
- resonance
- coordinate covalent bond

Lesson 4 MOLECULAR SHAPES

- VSEPR model theory states that electron pairs repel each other and determine both the shape of and bond angles in a molecule.
- Hybridization explains the observed shapes of molecules by the presence of equivalent hybrid orbitals.

- VSEPR model
- hybridization

Lesson 5 ELECTRONEGATIVITY AND POLARITY

- The electronegativity difference determines the character of a bond between atoms.
- Polar bonds occur when electrons are not shared equally, forming a dipole.
- The spatial arrangement of polar bonds in a molecule determines the overall polarity of a molecule.
- Molecules attract each other by weak intermolecular forces. In a covalent network solid, each atom is covalently bonded to many other atoms.

- polar covalent bond

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

Why does water expand when it freezes?

CER Claim, Evidence, Reasoning

Explain your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

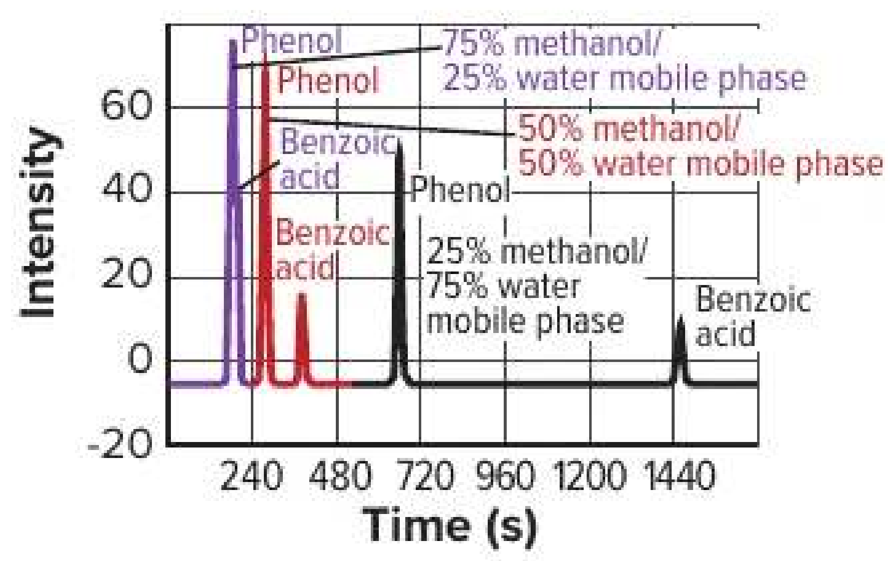
Based on Real Data*

SEP Data Analysis Lab

How does the polarity of the mobile phase affect chromatograms?

High-performance liquid chromatography, or HPLC, is used by analytical chemists to separate mixtures of solutes. During HPLC, components that are strongly attracted to the extracting solvent are retained longer by the moving phase and tend to appear early on a chromatograph. Several scientists performed HPLC using a methanol-water mixture as the extracting solvent to separate a phenol-benzoic acid mixture. Their results are shown in the graph. The peak areas on the chromatograph indicate the amount of each component present in the mixture.

Data and Observations



*Data obtained from: Joseph, Seema M. and Palasota, John A. 2001. The combined effects of pH and percent methanol on the HPLC separation of benzoic acid and phenol. Journal of Chemical Education 78:1381.

CER Analyze and Interpret Data

1. **Explain** the different retention times shown on the chromatograms.
2. **Infer** from the graph the component, phenol or benzoic acid, that is in excess. Explain your answer.
3. **Infer** which component of the mixture has more polar molecules.
4. **Determine** the most effective composition of the mobile phase (of those tested) for separating phenol from benzoic acid. Explain.

MODULE 8
CHEMICAL REACTIONS

MODULE 8

CHEMICAL REACTIONS

ENCOUNTER THE PHENOMENON

What happens to food when you cook it?

GO ONLINE to play a video of a chemical reaction that happens in your kitchen.

SEP Ask Questions

Do you have other questions about the phenomenon? If so, add them to the driving question board.

CER Claim, Evidence, Reasoning

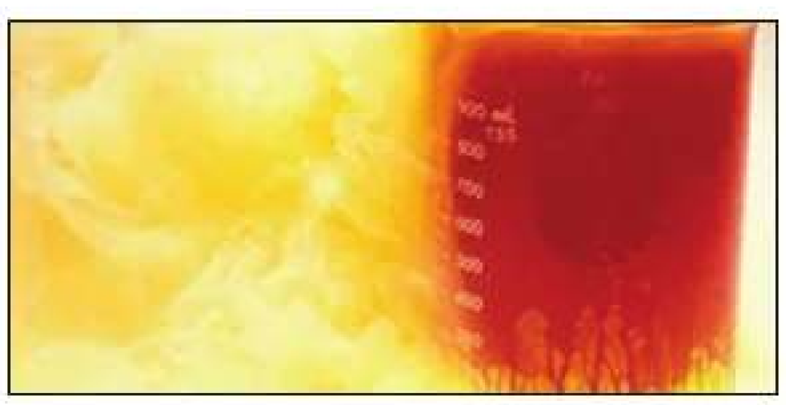
Make Your Claim Use your CER chart to make a claim about what happens to food when you cook it.

Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

Explain Your Reasoning You will revisit your claim and explain your reasoning at the end of the module.

GO ONLINE to access your CER chart and explore resources that can help you collect evidence.

LESSON 1: Explore & Explain: Evidence of Chemical Reactions



LESSON 1: Explore & Explain: Representing Chemical Reactions

Smithsonian

Additional Resources

(t)voshadhi/Creatas Video/Getty Images; (bl)Charles D. Winters/McGraw-Hill Education; (br)Charles D. Winters/Science Source

LESSON 1

REACTIONS AND EQUATIONS

FOCUS QUESTION

How are chemical reactions modeled?

Chemical Reactions

Do you know that the foods you eat, the fibers in your clothes, and the plastic in your CDs have something in common? Foods, fibers, and plastics are produced when the atoms in substances are rearranged to form different substances. Rearrangments of atoms are happening all the time. Atoms are rearranged during the cooking of food, inside the bodies of living things, in vehicle engines, in factories, and in the atmosphere.

The process by which the atoms of one or more substances are rearranged to form different substances is called a **chemical reaction.** A chemical reaction is another name for a chemical change, which you read about previously.

Chemical reactions affect every part of your life. They break down your food, producing the energy you need to live. Chemical reactions in the engines of cars and buses provide the energy to power the vehicles. They produce natural fibers, such as cotton and wool, in plants and animals. In factories, they produce synthetic fibers such as nylon, which is shown in **Figure 1.** Nylon is used in many familiar products, including carpeting, clothing, sports equipment, and tires.

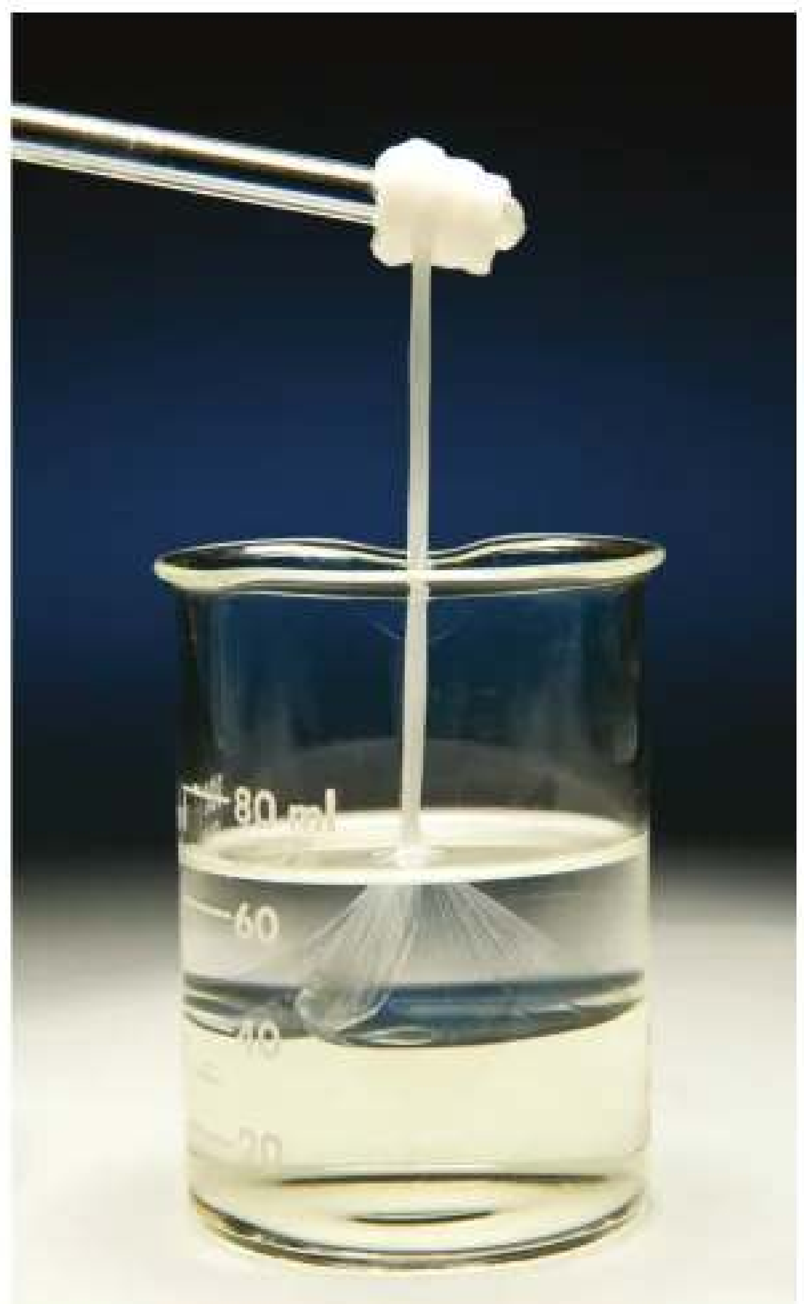


Figure 1 When adipoyl chloride in dichloromethane reacts with hexanediamine, nylon is formed.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Inquiry into Chemistry: Solve It: Mystery of the Moonlight Ride
Obtain information on a water decomposition reaction and describe its structure and function as it relates to fuel.

Virtual Investigation: Balancing Chemical Equations
Use mathematics and computational thinking to balance chemical equations.

Figure 2 Each of these photos illustrates evidence of a chemical reaction.

Evidence of a chemical reaction

Although some chemical reactions are hard to detect, many reactions provide physical evidence that they have occurred. A temperature change can indicate a chemical reaction. Many reactions, such as those that occur during the burning of wood, release energy in the form of heat and light. Other chemical reactions absorb heat.

In addition to a temperature change, color change can indicate that a chemical reaction has occurred. For example, you might have noticed that the color of some nails that are left outside changes from silver to orange-brown in a short time. The color change is evidence that a chemical reaction occurred between the iron in the nail and the oxygen in air. Odor, gas bubbles, and the formation of a solid are other indications of chemical change. Each of the photographs in **Figure 2** shows evidence of a chemical reaction.

Get It?
Cite Evidence Describe any evidence of a chemical reaction you note in the photo at the beginning of this module.

Representing Chemical Reactions

Chemists use statements called equations to represent chemical reactions. Equations show a reaction's **reactants,** which are the starting substances, and **products,** which are the substances formed during the reaction. Chemical equations do not express numerical equalities as mathematical equations do because during chemical reactions the reactants are used up as the products form. Instead, the equations used by chemists show the direction in which the reaction progresses. Therefore, an arrow rather than an equal sign is used to separate the reactants from the products. You read the arrow as react to produce or yield. The reactants are written to the left of the arrow, and the products are written to the right of the arrow. When there are two or more reactants, or when there are two or more products, a plus sign separates each reactant or each product. These elements of equation notation are shown below.

$$\text{Reactant 1} + \text{Reactant 2} \rightarrow \text{Product 1} + \text{Product 2}$$

Symbols for equations

You have already seen that the plus symbol is used to separate two or more reactants or products on each side of an equation, and that an arrow is used to separate the reactants from the products. When a reaction is reversible, a two-way arrow is used to show that the reaction can proceed in either direction.

As well, symbols are used to show the physical states of the reactants and products. Reactants and products can exist as solids, liquids, and gases. When they are dissolved in water, they are said to be aqueous. It is important to show the physical states of a reaction's reactants and products in an equation because the physical states provide clues about how the reaction occurs. Some basic symbols used in equations are shown in **Table 1**.

Table 1
Symbols Used in Equations

Symbol	Purpose
$+$	separates two or more reactants or products
$\rightarrow$	separates reactants from products
$\rightleftharpoons$	separates reactants from products and indicates a reversible reaction
(s)	identifies a solid state
(l)	identifies a liquid state
(g)	identifies a gaseous state
(aq)	identifies a water solution

Word equations

You can use statements called word equations to indicate the reactants and products of chemical reactions. The word equation below describes the reaction between aluminum (Al) and bromine (Br), which is shown in **Figure 3**. Aluminum is a solid, and bromine is a liquid. The brownish-red cloud in the photograph is excess bromine. The reaction's product, which is solid particles of aluminum bromide ($AlBr_3$), settles on the bottom of the beaker.

$$\text{Reactant 1} + \text{Reactant 2} \rightarrow \text{Product 1}$$

$$\text{aluminum(s)} + \text{bromine(l)} \rightarrow \text{aluminum bromide(s)}$$

The word equation reads, "Aluminum and bromine react to produce aluminum bromide."

Figure 3 Science, like all other disciplines, has a specialized language that allows specific information to be communicated in a uniform manner. This reaction between aluminum and bromine can be described by a word equation, a skeleton equation, or a balanced chemical equation.

Charles D. Winters/Science Source

Skeleton equations

Although word equations help to describe chemical reactions, they lack important information. A skeleton equation uses chemical formulas rather than words to identify the reactants and the products. For example, the skeleton equation for the reaction between aluminum and bromine uses the formulas for aluminum, bromine, and aluminum bromide in place of words.

$$Al(s) + Br_2(l) \rightarrow AlBr_3(s)$$

How would you write the skeleton equation that describes the reaction between carbon and sulfur to form carbon disulfide? Carbon and sulfur are solids. First, write the chemical formulas for each of the reactants to the left of the arrow. Then, separate the reactants with a plus sign. Add symbols to each formula to indicate the physical states of the compounds.

$$C(s) + S(s) \rightarrow$$

Finally, write the chemical formula for the product, liquid carbon disulfide, to the right of the arrow and indicate its physical state. The result is the skeleton equation for the reaction.

$$C(s) + S(s) \rightarrow CS_2(l)$$

This skeleton equation tells us that carbon in the solid state reacts with sulfur in the solid state to produce carbon disulfide in the liquid state. The skeleton equation does not, however, indicate the amounts of each compound involved in the reaction.

Get It?

Compare What information do skeleton equations provide that word equations lack?

PRACTICE Problems

ADDITIONAL PRACTICE

Write skeleton equations for the following word equations.

1. Hydrogen and bromine gases react to yield hydrogen bromide.

   hydrogen(g) + bromine(g) → hydrogen bromide(g)

2. When carbon monoxide and oxygen react, carbon dioxide forms.

   carbon monoxide(g) + oxygen(g) → carbon dioxide(g)

3. **CHALLENGE** Write the word equation and the skeleton equation for the following reaction: when heated, solid potassium chlorate yields solid potassium chloride and oxygen gas.

ACADEMIC VOCABULARY

formula

an expression using chemical symbols to represent a substance

The chemical formula for water is H_2O.

CCC CROSSCUTTING CONCEPTS

Energy and Matter The total amount of matter is conserved in a chemical reaction. Write a paragraph that argues that a skeleton equation does not provide enough information to show this to be true. Include evidence to support your argument.

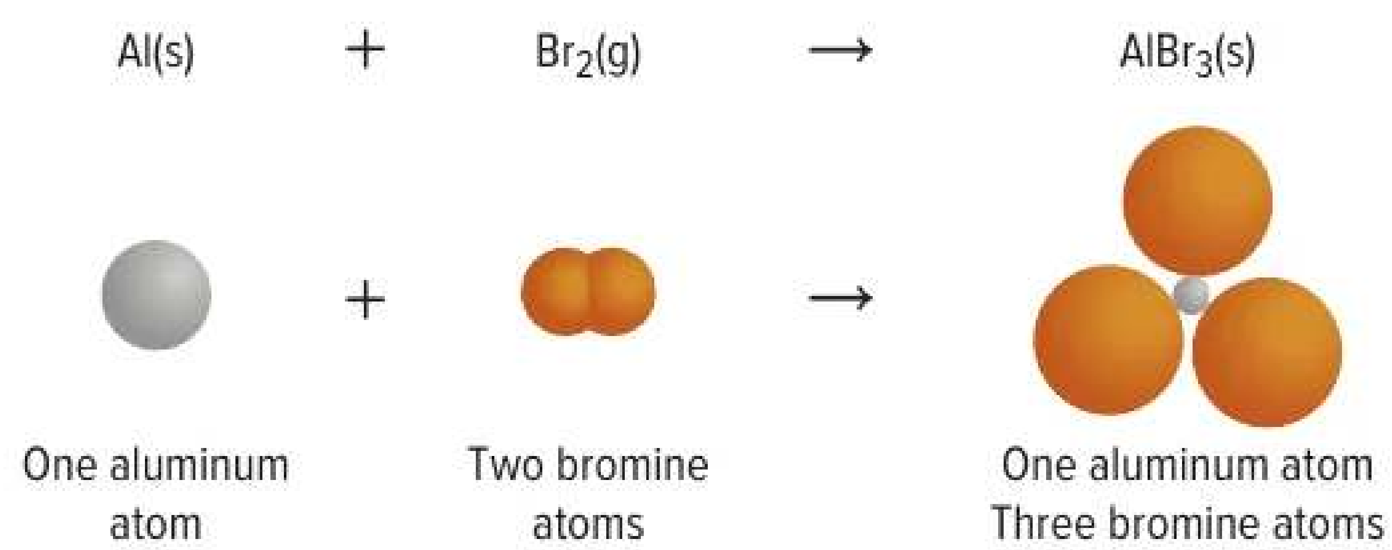


Figure 4 The information conveyed by skeleton equations is limited. In this case, the skeleton equation is correct, but it does not show the exact number of atoms that interact. Refer to **Table R-1** in the Student Resources for a key to atom color conventions.

Chemical equations

Like word equations, skeleton equations lack some information about reactions. Recall that the law of conservation of mass states that in a chemical change, matter is neither created nor destroyed. Chemical equations must show that atoms are conserved during a reaction. Skeleton equations lack that information.

Look at **Figure 4**. The skeleton equation for the reaction between aluminum and bromine shows that one aluminum atom and two bromine atoms react to produce a substance containing one aluminum atom and three bromine atoms. Was a bromine atom created in the reaction? Atoms are not created in chemical reactions, so more information is needed. To accurately represent a chemical reaction using an equation, the equation must show equal numbers of atoms of each reactant and each product on both sides of the reaction arrow. Such an equation is called a balanced chemical equation. A **chemical equation** is a statement that uses chemical formulas to show the identities and relative amounts of the substances involved in a chemical reaction.

Balancing Chemical Equations

The balanced equation for the reaction between aluminum and bromine is shown in **Figure 5**. To balance an equation, you must find the correct coefficients for the chemical formulas in the skeleton equation. A **coefficient** in a chemical equation is the number written in front of a reactant or product. Coefficients are usually whole numbers and are not usually written if the value is one. The coefficients in a balanced equation describe the lowest whole number ratio of the amounts of all of the reactants and products. **Table 2** shows the steps for balancing a chemical equation.

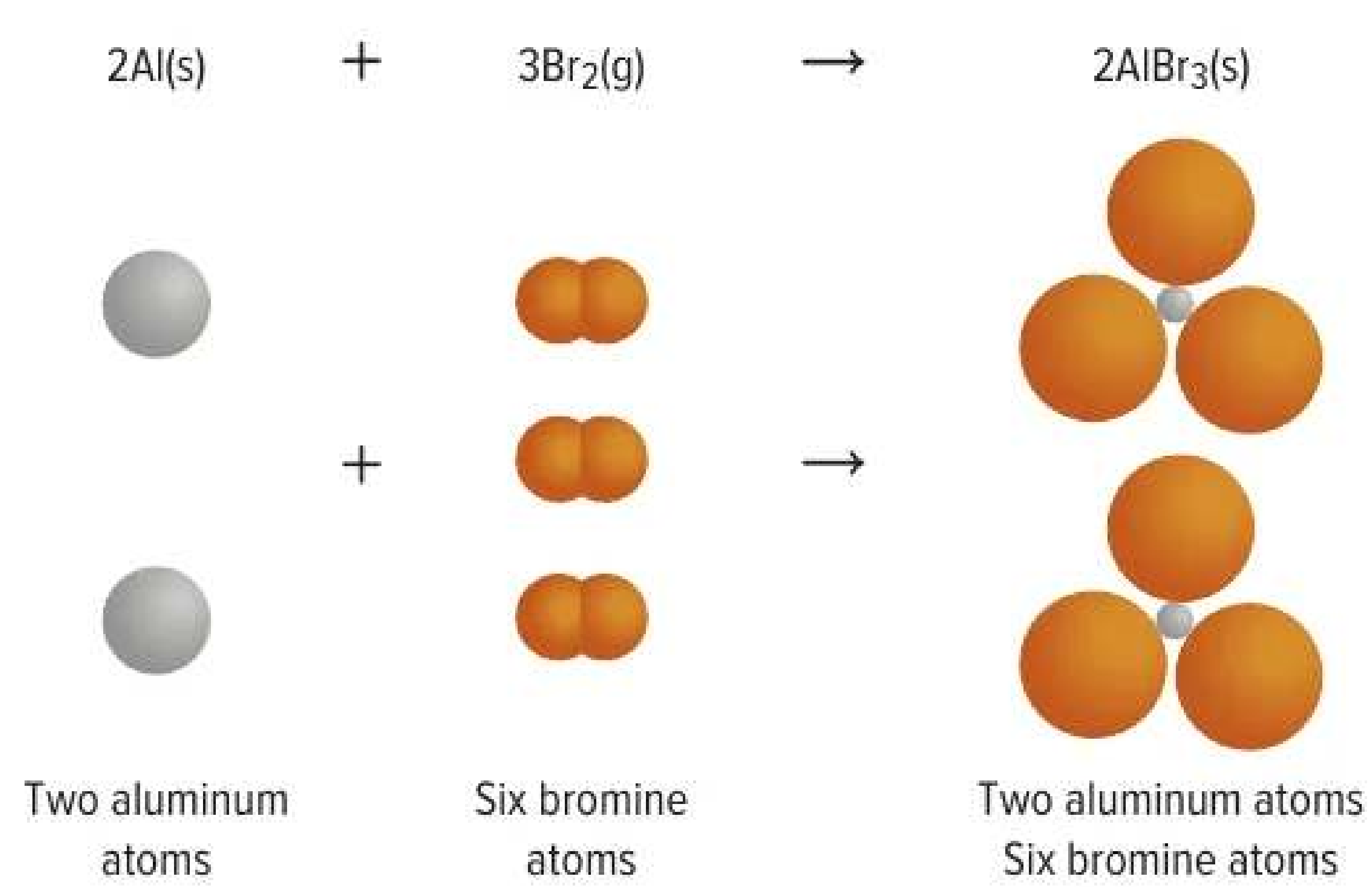
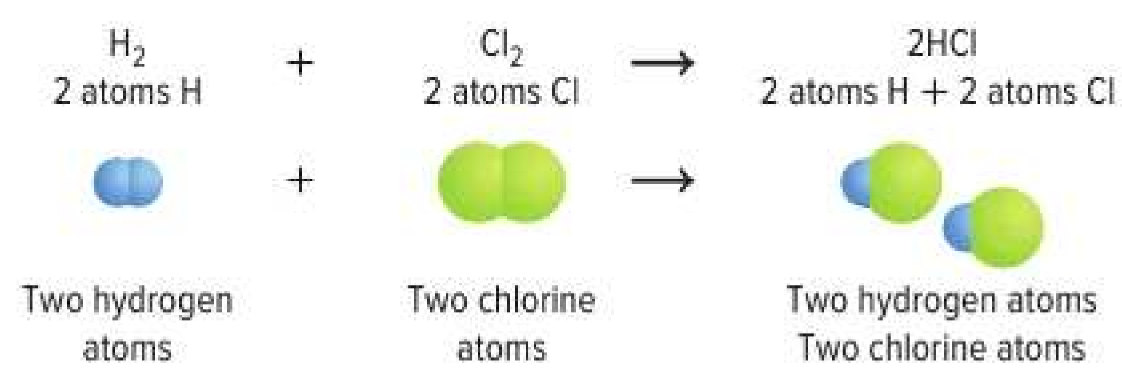


Figure 5 In a balanced chemical equation, the number of particles on the reactant side of the equation equals the number of particles on the product side of the equation. In this case, two aluminum atoms and six bromine atoms are needed on both sides of the equation.

Conservation of mass A fundamental concept of chemistry is the law of conservation of mass. All chemical reactions obey the law that matter, including atoms, is neither created nor destroyed. Therefore, it is important that chemical equations are balanced to show that reactions obey this law. The fact that atoms are conserved is used to describe and balance chemical equations, as shown in **Table 2**.

Table 2 Steps for Balancing Equations

Step	Process	Example
1	*Write the skeleton equation for the reaction*. Make sure that the chemical formulas correctly represent the substances. An arrow separates the reactants from the products, and a plus sign separates multiple reactants and products. Show the physical states of all reactants and products.	
2	*Count the atoms of the elements in the reactants*. If a reaction involves identical polyatomic ions in the reactants and products, count each polyatomic ion as a single element. This reaction does not involve any polyatomic ions. Two atoms of hydrogen and two atoms of chlorine are reacting.	H_2 + Cl_2 → 2 atoms H 2 atoms Cl
3	*Count the atoms of the elements in the products*. One atom of hydrogen and one atom of chlorine are produced.	HCl 1 atom H + 1 atom Cl
4	*Change the coefficients to make the number of atoms of each element equal on both sides of the equation, showing that atoms are conserved*. Never change a subscript in a chemical formula to balance an equation because doing so changes the identity of the substance.	
5	*Write the coefficients in their lowest possible ratio*. The coefficients should be the smallest possible whole numbers. The ratio 1 hydrogen to 1 chlorine to 2 hydrogen chloride (1:1:2) is the lowest-possible ratio because the coefficients cannot be reduced further and still remain whole numbers.	$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ 1:1:2 1 H_2 to 1 Cl_2 to 2 HCl
6	*Check your work*. Make sure that the chemical formulas are written correctly. Then, check that the number of atoms of each element is equal on both sides of the equation.	H_2 + Cl_2 → 2HCl 2 atoms H 2 atoms Cl 2 atoms H + 2 atoms Cl There are two hydrogen atoms and two chlorine atoms on both sides of the equation.

EXAMPLE Problem 1

WRITING A BALANCED CHEMICAL EQUATION Write the balanced chemical equation for the reaction in which aqueous sodium hydroxide and aqueous calcium bromide react to produce solid calcium hydroxide and aqueous sodium bromide.

1 ANALYZE THE PROBLEM

You are given the reactants and products in a chemical reaction. Start with a skeleton equation, and use the steps given in **Table 2** for balancing chemical equations.

2 SOLVE FOR THE UNKNOWN

Write the skeleton equation for the chemical reaction. Be sure to put the reactants on the left side of the arrow and the products on the right. Separate the substances with plus signs, and indicate their physical states.

$$NaOH(aq) + CaBr_2(aq) \rightarrow Ca(OH)_2(s) + NaBr(aq)$$

1 Na, 1 O, 1 H, 1 Ca, 2 Br — Count the atoms of each element in the reactants.

1 Na, 2 O, 2 H, 1 Ca, 1 Br — Count the atoms of each element in the products.

$2NaOH + CaBr_2 \rightarrow Ca(OH)_2 + NaBr$ — Insert the coefficient 2 in front of NaOH to balance the hydroxide ions.

$2NaOH + CaBr_2 \rightarrow Ca(OH)_2 + 2NaBr$ — Insert the coefficient 2 in front of NaBr to balance the Na and Br atoms.

The ratio of the coefficients is 2:1:1:2. — Write the coefficients in their lowest-possible ratio.

Reactants: 2 Na, 2 OH, 1 Ca, 2 Br
Products: 2 Na, 2 OH, 1 Ca, 2 Br — Check to make sure that the number of atoms of each element is equal on both sides of the equation.

3 EVALUATE THE ANSWER

The chemical formulas for all substances are written correctly.

The number of atoms of each element is equal on both sides of the equation. The coefficients are written in the lowest possible ratio. The balanced chemical equation for the reaction is

$$2NaOH(aq) + CaBr_2(aq) \rightarrow Ca(OH)_2(s) + 2NaBr(aq)$$

Real-World Chemistry
Calcium Hydroxide

REEF AQUARIUMS An aqueous solution of calcium hydroxide is used in reef aquariums to provide calcium for animals such as snails and corals. Calcium hydroxide reacts with the carbon dioxide in the water to produce calcium and bicarbonate ions. Reef animals use the calcium to grow shells and strong skeletal systems.

PRACTICE Problems

ADDITIONAL PRACTICE

Write chemical equations for each of the following reactions.

4. In water, iron(III) chloride reacts with sodium hydroxide, producing solid iron(III) hydroxide and sodium chloride.

5. Liquid carbon disulfide reacts with oxygen gas, producing carbon dioxide gas and sulfur dioxide gas.

6. CHALLENGE A piece of zinc metal is added to a solution of dihydrogen sulfate. This reaction produces a gas and a solution of zinc sulfate.

Balancing Chemical Equations

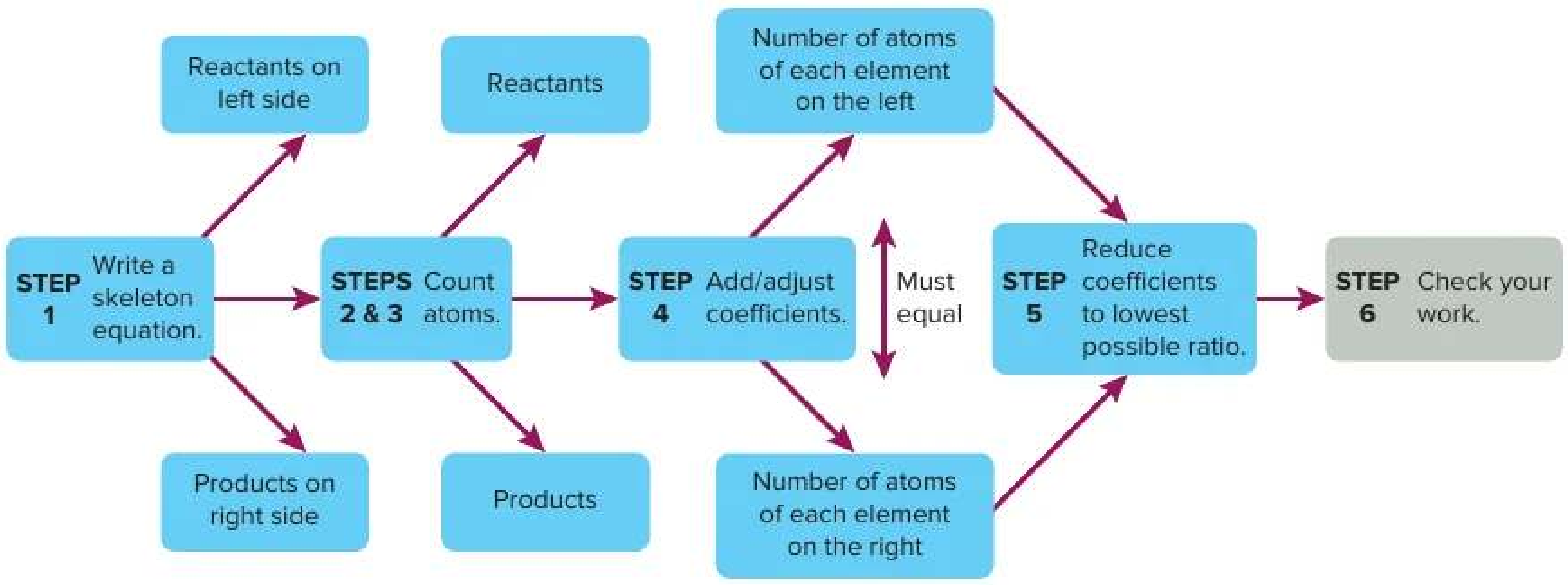


Figure 6 Use this flowchart to help you master the skill of balancing equations.

Figure 6 summarizes the steps for balancing equations. Most chemical equations can be balanced by the process you learned in this lesson.

Check Your Progress

Summary

- Some physical changes are evidence that indicate a chemical reaction has occurred.
- Word equations and skeleton equations provide important information about a chemical reaction.
- A chemical equation gives the identities and relative amounts of the reactants and products that are involved in a chemical reaction.
- Balancing an equation involves adjusting the coefficients until the number of atoms of each element is equal on both sides of the equation.

Demonstrate Understanding

7. **Explain** why it is important that a chemical equation be balanced in terms of the fact that atoms are conserved.
8. **List** three types of physical evidence that indicate a chemical reaction has occurred.
9. **Compare and contrast** a skeleton equation and a chemical equation.
10. **Explain** why it is important to reduce coefficients in a balanced equation to the lowest-possible whole-number ratio.
11. **Analyze** When balancing a chemical equation, can you adjust the subscript in a formula? Explain.
12. **Assess** Is the following equation balanced? If not, correct the coefficients to balance the equation.
    $2K_2CrO_4(aq) + Pb(NO_3)_2(aq) \rightarrow 2KNO_3(aq) + PbCrO_4(s)$
13. **Evaluate** Aqueous phosphoric acid and aqueous calcium hydroxide react to form solid calcium phosphate and water. Write a balanced chemical equation for this reaction.

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LESSON 2

CLASSIFYING CHEMICAL REACTIONS

FOCUS QUESTION

What are the different types of chemical reactions?

Types of Chemical Reactions

The fact that atoms are conserved, together with knowledge of the chemical properties of the elements involved, can be used to describe and predict chemical reactions. Chemists classify chemical reactions into several categories. Knowing the categories of chemical reactions can help you remember and understand them. It can also help you recognize patterns and predict the products of many reactions.

Chemists distinguish among four reaction types: synthesis, combustion, decomposition, and replacement reactions. By analyzing and comparing the reactants and products of a variety of chemical reactions, you will notice patterns that will help you to classify them. Note, however, that some reactions fit into more than one of these types.

Synthesis Reactions

In **Figure 7**, two sodium atoms react with a molecule of chlorine to produce sodium chloride. This reaction is a **synthesis reaction**—a chemical reaction in which two or more substances (A and B) react to produce a single product (AB).

$$A + B \rightarrow AB$$

When two elements react, the reaction is always a synthesis reaction.

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

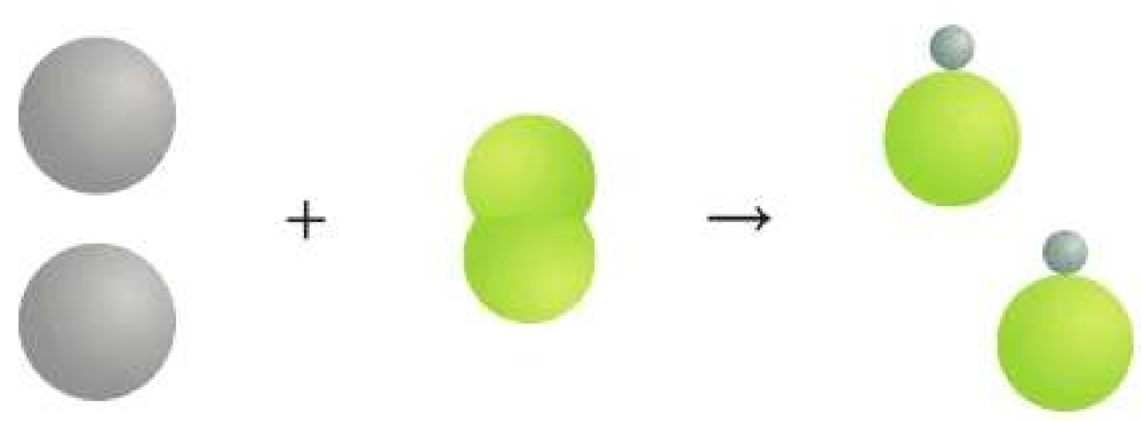


Figure 7 In this synthesis reaction, two elements, sodium and chlorine, react to produce one compound, sodium chloride.

3D THINKING DCI Disciplinary Core Ideas CCC Crosscutting Concepts SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

ChemLAB: Develop an Activity Series

Obtain and evaluate information on the effects of metal reactivity to predict chemical reactions.

Laboratory: Single-Replacement Reactions

Obtain and evaluate information on the effects of a single replacement reaction to understand the conservation of atoms.

Two compounds can also combine to form one compound. For example, the reaction between calcium oxide (CaO) and water (H_2O) to form calcium hydroxide ($Ca(OH)_2$) is a synthesis reaction.

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

Another type of synthesis reaction involves a reaction between a compound and an element, as happens when sulfur dioxide gas (SO_2) reacts with oxygen gas (O_2) to form sulfur trioxide (SO_3).

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

Combustion Reactions

The synthesis reaction between sulfur dioxide and oxygen can also be classified as a **combustion reaction.** In a combustion reaction, such as the one shown in **Figure 8**, oxygen combines with a substance and releases energy in the form of heat and light. Oxygen can combine in this way with many different substances, making combustion reactions common.

Figure 8 The light produced by a sparkler is the result of a combustion reaction between oxygen and different metals.

A combustion reaction occurs between hydrogen and oxygen when hydrogen is heated, as illustrated in **Figure 9**. Water is formed during the reaction, and a large amount of energy is released. Another important combustion reaction occurs when coal is burned to produce energy. Coal is called a fossil fuel because it contains the remains of plants that lived long ago. It is composed primarily of the element carbon. Coal-burning power plants generate electric power in many parts of the United States. The primary reaction that occurs in these plants is between carbon and oxygen.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

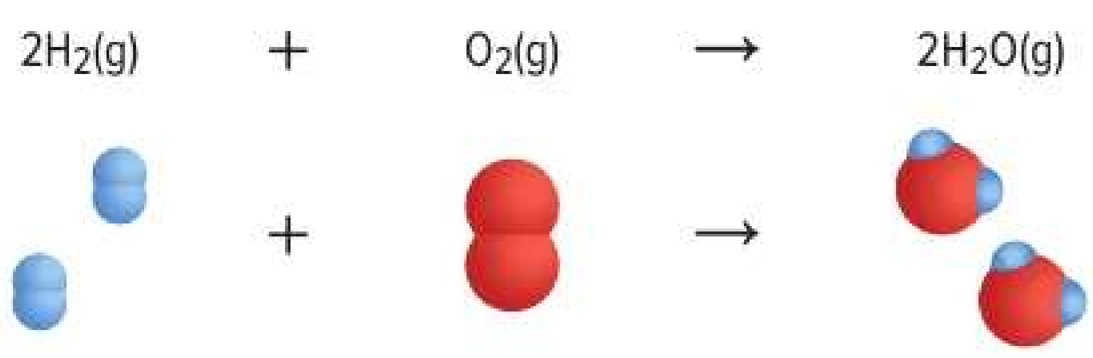


Figure 9 During a combustion reaction between oxygen and hydrogen, water is formed.

WORD ORIGIN

combustion

comes from the Latin word *comburere*, meaning *to burn*

CCC CROSSCUTTING CONCEPTS

Patterns Study the different types of chemical reactions in this lesson. Then create a poster that identifies the different patterns observed for the reaction types. What evidence supports the patterns you used in your poster?

Note that the combustion reactions just mentioned are also synthesis reactions. However, not all combustion reactions are synthesis reactions. For example, the reaction involving methane gas (CH_4) and oxygen illustrates a combustion reaction in which one substance replaces another in the formation of products.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

Methane, which belongs to a group of substances called hydrocarbons, is the major component of natural gas. A fireplace that uses natural gas as fuel is shown in **Figure 10.** All hydrocarbons contain carbon and hydrogen and burn in oxygen to yield carbon dioxide and water.

Figure 10 The combustion of natural gas in this fireplace provides warmth and light.

PRACTICE Problems ADDITIONAL PRACTICE

Write chemical equations for the following reactions. Classify each reaction into as many categories as possible.

14. The solids aluminum and sulfur react to produce aluminum sulfide.

15. Water and dinitrogen pentoxide gas react to produce aqueous hydrogen nitrate.

16. The gases nitrogen dioxide and oxygen react to produce dinitrogen pentoxide gas.

17. CHALLENGE Sulfuric acid (H_2SO_4) and sodium hydroxide solutions react to produce aqueous sodium sulfate and water.

Decomposition Reactions

Some chemical reactions are essentially the opposite of synthesis reactions. These reactions are classified as **decomposition reactions.** A decomposition reaction is one in which a single compound breaks down into two or more elements or new compounds. In generic terms, decomposition reactions can be represented as follows.

$$AB \rightarrow A + B$$

Decomposition reactions often require an energy source, such as heat, light, or electricity, to occur. For example, ammonium nitrate breaks down into dinitrogen monoxide and water when the reactant is heated to a high temperature.

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(g)$$

Notice that this decomposition reaction involves one reactant compound breaking down into two product compounds. Other decomposition reactions involve one reactant compound breaking down into more than two products. The products of a decomposition reaction may be elements, compounds, or one or more of each.

Figure 11 The decomposition of sodium azide, which produces a gas, is the chemical reaction that inflates air bags.

The outcome of another decomposition reaction is shown in **Figure 11**. Automobile safety air bags inflate rapidly as sodium azide pellets decompose. A device that can provide an electric signal to start the reaction is packaged inside air bags along with the sodium azide pellets. When the device is activated, sodium azide decomposes, producing nitrogen gas that quickly inflates the air bag.

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

Notice that the decomposition of sodium azide produces sodium metal in addition to the harmless gas nitrogen. Sodium metal is highly reactive and caustic and therefore poses a safety concern when the air bag deploys. Designers overcame this problem by adding iron(III) oxide. The sodium reacts with the iron(III) oxide to produce sodium oxide, which then reacts with carbon dioxide and water vapour in the air to produce the much safer substance sodium hydrogen carbonate, also known as baking soda. These reactions all take place very quickly.

PRACTICE Problems ADDITIONAL PRACTICE

Write chemical equations for the following decomposition reactions.

18. Aluminum oxide(s) decomposes when electricity passes through it.

19. Nickel(II) hydroxide(s) decomposes to produce nickel(II) oxide(s) and water.

20. CHALLENGE Heating sodium hydrogen carbonate(s) produces sodium carbonate(aq) and water. Carbon dioxide gas is also produced.

Replacement Reactions

In contrast to synthesis, combustion, and decomposition reactions, many chemical reactions are replacement reactions and involve the replacement of an element in a compound. These replacement reactions are also known as displacement reactions. There are two types of replacement reactions: single-replacement reactions and double-replacement reactions.

Single-replacement reactions

The reaction between lithium and water is shown in **Figure 12**. The following chemical equation shows that a lithium atom replaces one of the hydrogen atoms in a water molecule.

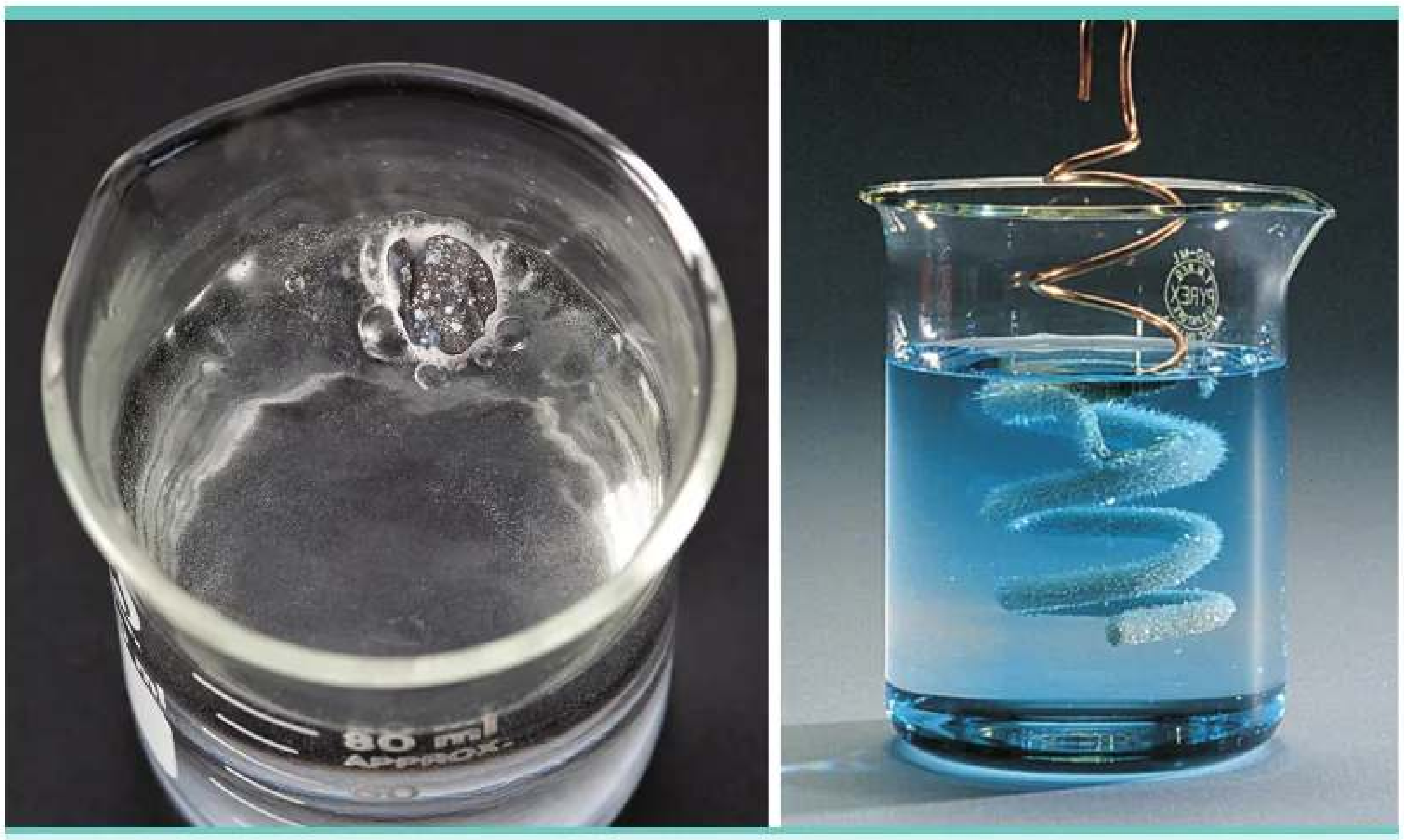
$$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$$

A reaction in which the atoms of one element replace the atoms of another element in a compound is called a **single-replacement reaction.** The following generic equation can be used to represent single-replacement reactions such as the reaction of lithium with water to form lithium hydroxide and hydrogen.

$$A + BX \rightarrow AX + B$$

Get It?

Analyze In the reaction between lithium and water, which element replaces hydrogen in water?



Lithium + Water

Copper + Silver Nitrate

Figure 12 In a single-replacement reaction, the atoms of one element replace the atoms of another element in a compound.

Metal replaces hydrogen or another metal The reaction between lithium and water is one type of single-replacement reaction, in which a metal replaces a hydrogen atom in a water molecule. Another type of single-replacement reaction occurs when one metal replaces another metal in a compound dissolved in water. **Figure 12** shows a single-replacement reaction occurring when copper wire is placed in aqueous silver nitrate. The crystals that are accumulating on the copper bar are the silver atoms that the copper atoms replaced.

$$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(s) + Cu(NO_3)_2(aq)$$

A metal will not always replace another metal in a compound dissolved in water because metals differ in their reactivities. Reactivity is the ability to react with another substance. An activity series of some metals is shown in **Figure 13**. This series orders metals by reactivity with other metals. Single-replacement reactions are used to determine a metal's position on the list. The most active metals are at the top of the list. The least active metals are at the bottom. Similarly, the reactivity of each halogen has been determined and listed, as shown in **Figure 13**.

Figure 13 An activity series is a useful tool for determining the result of a single-replacement reaction.

You can use the activity series to predict whether or not certain reactions will occur. A specific metal can replace any metal listed below it that is in a compound. It cannot replace any metal listed above it. For example, copper atoms replace silver atoms in a solution of silver nitrate. However, if you place a silver wire in aqueous copper(II) nitrate, the silver atoms will not replace the copper. Silver is listed below copper in the activity series, so no reaction occurs. The letters NR (no reaction) are commonly used to indicate that a reaction will not occur.

$$Ag(s) + Cu(NO_3)_2(aq) \rightarrow NR$$

Nonmetal replaces nonmetal A third type of single-replacement reaction involves the replacement of a nonmetal in a compound by another nonmetal. Halogens are frequently involved in these types of reactions. Like metals, halogens exhibit different activity levels in single-replacement reactions. The reactivities of halogens, determined by single-replacement reactions, are also shown in **Figure 13**.

The most active halogen is fluorine, and the least active is iodine. A more reactive halogen replaces a less reactive halogen that is part of a compound dissolved in water. A halogen cannot replace any halogen listed above it. For example, fluorine replaces bromine in water containing dissolved sodium bromide, as shown in the following chemical equation. However, bromine does not replace fluorine in water containing dissolved sodium fluoride.

$$F_2(g) + 2NaBr(aq) \rightarrow 2NaF(aq) + Br_2(l)$$

$$Br_2(g) + 2NaF(aq) \rightarrow NR$$

Get It?

Explain how a single-replacement reaction works.

EXAMPLE Problem 2

SINGLE-REPLACEMENT REACTIONS Predict the products that will result when these reactants combine, and write a balanced chemical equation for each reaction.

a. $Fe(s) + CuSO_4(aq) \rightarrow$

b. $Br_2(l) + MgCl_2(aq) \rightarrow$

c. $Mg(s) + AlCl_3(aq) \rightarrow$

1 ANALYZE THE PROBLEM

You are given three sets of reactants. Using **Figure 13,** you must first determine if each reaction occurs. Then, if a reaction is predicted, you can determine the product(s) of the reaction. With this information you can write a skeleton equation for the reaction. Finally, you can use the steps for balancing chemical equations to write the complete balanced chemical equation.

2 SOLVE FOR THE UNKNOWN

a. Iron is listed above copper in the activity series. Therefore, the first reaction will occur because iron is more reactive than copper. In this case, iron will replace copper. The skeleton equation for this reaction is

$$Fe(s) + CuSO_4(aq) \rightarrow FeS0_4(aq) + Cu(s)$$

This equation is balanced.

b. In the second reaction, chlorine is more reactive than bromine because bromine is listed below chlorine in the activity series. Therefore, the reaction will not occur. The skeleton equation for this situation is

$$Br(l) + MgCl_2(aq) \rightarrow NR$$

No balancing is required.

c. Magnesium is listed above aluminum in the activity series. Therefore, the third reaction will occur because magnesium is more reactive than aluminum. In this case, magnesium will replace aluminum. The skeleton equation for this reaction is

$$Mg(s) + AlCl_3(aq) \rightarrow Al(s) + MgCl_2(aq)$$

This equation is not balanced. The balanced equation is

$$3Mg(s) + 2AlCl_3(aq) \rightarrow 2Al(s) + 3MgCl_2(aq)$$

3 EVALUATE THE ANSWER

The activity series shown in **Figure 13** supports the reaction predictions. The chemical equations balance because the number of atoms of each substance is equal on both sides of the equation.

Real-World Chemistry

Single-Replacement Reactions

ZINC PLATING Tools made of steel are often covered with a layer of zinc to prevent corrosion. Zinc is more reactive than the lead in steel. During zinc plating, the zinc replaces some of the surface lead, coating the steel.

PRACTICE Problems — ADDITIONAL PRACTICE

Predict whether the following single-replacement reactions will occur. If a reaction occurs, write a balanced equation for the reaction.

21. $K(s) + ZnCl_2(aq) \rightarrow$

22. $Cl_2(g) + HF(aq) \rightarrow$

23. $Fe(s) + Na_3PO_4(aq) \rightarrow$

24. CHALLENGE $Al(s) + Pb(NO_3)_2(aq) \rightarrow$

$$AX + BY \rightarrow AY + BX$$

$$Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$$

Figure 14 The color-coding in the generic equation for a double-replacement reaction and in the equation for the reaction between calcium hydroxide and hydrochloric acid shows the anions changing places.

Double-replacement reactions

The final type of replacement reaction, which involves an exchange of ions between two compounds, is called a **double-replacement reaction.**

In the generic equation in **Figure 14**, A and B represent positively charged ions (cations), and X and Y represent negatively charged ions (anions). Notice that the anions have switched places and are now bonded to the other cations in the reaction. In other words, X replaces Y and Y replaces X—a double replacement. More simply, the positive and negative ions of two compounds switch places.

The reaction between calcium hydroxide and hydrochloric acid is a double-replacement reaction.

$$Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$$

The ionic components of the reaction are Ca^{2+}, OH^-, H^+, and Cl^-. Knowing this, you can now see the two replacements of the reaction. The anions (OH^- and Cl^-) have changed places and are now bonded to the other cations (Ca^{2+} and H^+), as shown in **Figure 14**.

The reaction between sodium hydroxide and copper(II) chloride in solution is also a double-replacement reaction.

$$2NaOH(aq) + CuCl_2(aq) \rightarrow 2NaCl(aq) + Cu(OH)_2(s)$$

In this case, the anions (OH^- and Cl^-) changed places and bonded to the other cations (Na^+ and Ca^{2+}). **Figure 15** shows that the result of this reaction is a solid product, copper(II) hydroxide. A solid produced during a chemical reaction in a solution is called a **precipitate.**

Figure 15 When aqueous sodium hydroxide is added to a solution of copper(II) chloride, the anions (OH^- and Cl^-) change places. The resulting products are sodium chloride, which remains in solution, and copper(II) hydroxide, the blue solid in the test tube.

Get It?

Compare How do single-replacement reactions and double-replacement reactions differ?

Table 3 **Guidelines for Writing Double-Replacement Reactions**

Step	Example
1. Write the components of the reactants in a skeleton equation.	$Al(NO_3)_3 + H_2SO_4$
2. Identify the cations and the anions in each compound.	$Al(NO_3)_3$ has Al^{3+} and NO_3^- H_2SO_4 has H^+ and SO_4^{2-}
3. Pair up each cation with the anion from the other compound.	Al^{3+} pairs with SO_4^{2-} H^+ pairs with NO_3^-
4. Write the formulas for the products using the pairs from Step 3.	$Al_2(SO_4)_3$ HNO_3
5. Write the complete equation for the double-replacement reaction.	$Al(NO_3)_3 + H_2SO_4 \rightarrow Al_2(SO_4)_3 + HNO_3$
6. Balance the equation.	$2Al(NO_3)_3 + 3H_2SO_4 \rightarrow$ $Al_2(SO_4)_3 + 6HNO_3$

Products of double-replacement reactions

One of the key characteristics of double-replacement reactions is the type of product that is formed when the reaction takes place. All double-replacement reactions produce either water, a precipitate, or a gas. Refer back to the two double-replacement reactions previously discussed in this section. The reaction between calcium hydroxide and hydrochloric acid produces water. A precipitate is produced in the reaction between sodium hydroxide and copper(II) chloride. An example of a double-replacement reaction that forms a gas is that of potassium cyanide and hydrobromic acid.

$$KCN(aq) + HBr(aq) \rightarrow KBr(aq) + HCN(g)$$

The basic steps to write double-replacement reactions are given in **Table 3**.

Get It?

Describe what happens to the anions in a double-replacement reaction.

PRACTICE Problems

ADDITIONAL PRACTICE

Write the balanced chemical equations for the following double-replacement reactions.

25. The two substances at right react to produce solid silver iodide and aqueous lithium nitrate.

26. Aqueous barium chloride and aqueous potassium carbonate react to produce solid barium carbonate and aqueous potassium chloride.

27. Aqueous sodium oxalate and aqueous lead(II) nitrate react to produce solid lead(II) oxalate and aqueous sodium nitrate.

28. **CHALLENGE** Acetic acid (CH_3COOH) and potassium hydroxide react to produce potassium acetate and water.

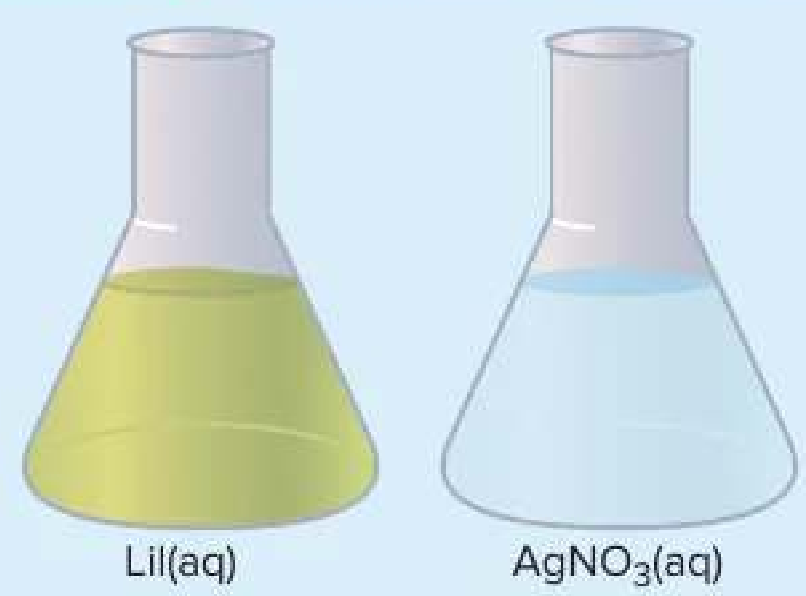


Table 4 Predicting Products of Chemical Reactions

Type of Reaction	Reactants	Probable Products	Generic Equation
Synthesis	• two or more substances	• one compound	$A + B \rightarrow AB$
Combustion	• a metal and oxygen • a nonmetal and oxygen • a compound and oxygen	• the oxide of the metal • the oxide of the nonmetal • two or more oxides	$A + O_2 \rightarrow AO$
Decomposition	• one compound	• two or more elements and/or compounds	$AB \rightarrow A + B$
Single-replacement	• a metal and a compound • a nonmetal and a compound	• a new compound and the replaced metal • a new compound and the replaced non-metal	$A + BX \rightarrow AX + B$
Double-replacement	• two compounds	• two different compounds, one of which is a solid, water, or a gas	$AX + BY \rightarrow AY + BX$

Table 4 summarizes the types of chemical reactions. Use this table to identify reactions and predict their products. First, write the chemical equation. Second, determine what is happening in the reaction. How many reactants are there? How many products are there? What happened to the elements and compounds in the reaction? Third, use your analysis to classify the reaction. Finally, compare the reaction to the generic equations in the table to check your answer.

Check Your Progress

Summary

- Classifying chemical reactions makes them easier to under-stand, remember, and recognize.
- Activity series of metals and halogens can be used to predict if single-replacement reactions will occur.

Demonstrate Understanding

29. Describe the four types of chemical reactions and their characteristics.
30. **Explain** how an activity series of metals is organized.
31. **Compare and contrast** single-replacement reactions and double-replacement reactions.
32. **Describe** the result of a double-replacement reaction.
33. Predict Use the fact that atoms are conserved, and your knowledge of the properties of the elements involved, to describe and predict the reaction most likely to occur when barium reacts with fluorine. Write the chemical equation for the reaction.
34. **Interpret Data** Could the following reaction occur? Explain your answer.

$$3Ni + 2AuBr_3 \rightarrow 3NiBr_2 + 2Au$$

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LESSON 3

REACTIONS IN AQUEOUS SOLUTIONS

FOCUS QUESTION

What is unique about reactions that take place in water?

Aqueous Solutions

You read previously that a solution is a homogeneous mixture. Many of the reactions discussed in the previous section involve substances dissolved in water. When a substance dissolves in water, a solution forms. An **aqueous solution** contains one or more substances called **solutes** dissolved in the water. In this case, water is the **solvent**—the most plentiful substance in the solution.

Molecular compounds in solution

Although water is always the solvent in aqueous solutions, there are many possible solutes. Some solutes, such as sucrose (table sugar) and ethanol (grain alcohol), are molecular compounds that exist as molecules in aqueous solutions. Other solutes are molecular compounds that form ions when they dissolve in water. For example, the molecular compound hydrogen chloride forms hydrogen ions and chloride ions when it dissolves in water, as shown in **Figure 16**.

An equation can be used to show this ionization process.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Compounds such as hydrogen chloride that produce hydrogen ions in aqueous solution are acids. In fact, an aqueous solution of hydrogen chloride is usually referred to as hydrochloric acid.

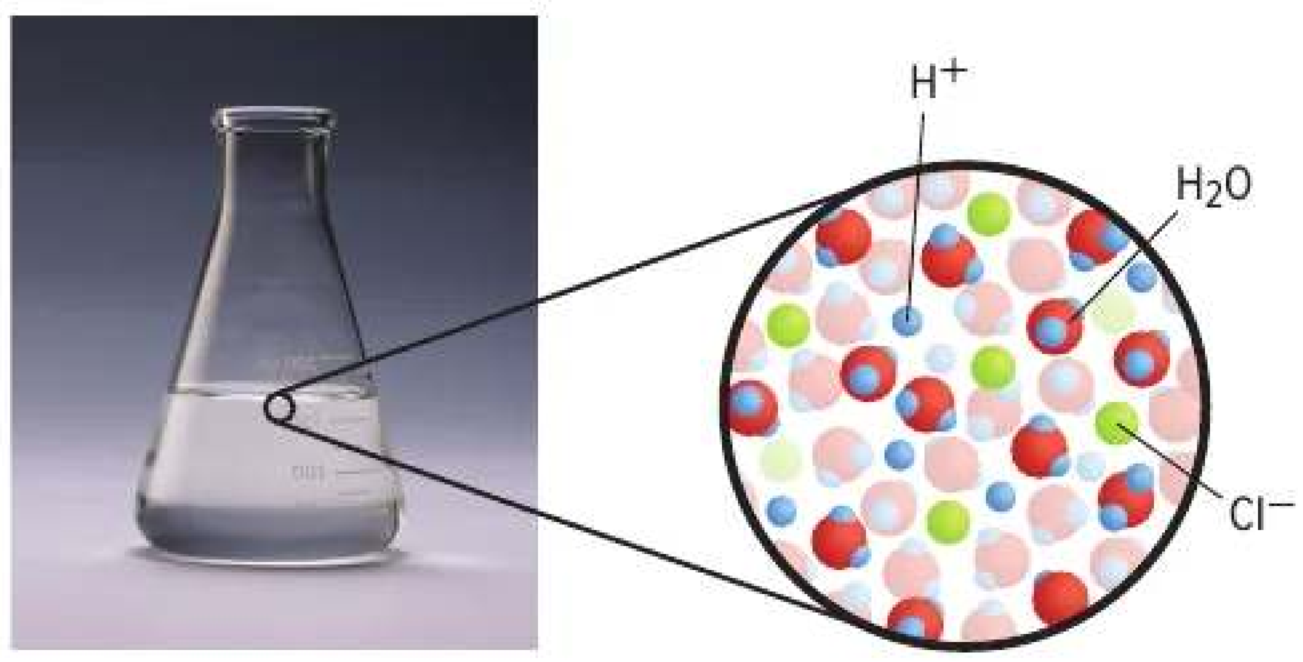

Figure 16 In water, hydrogen chloride (HCl) breaks apart into hydrogen ions (H^+) and chloride ions (Cl^-).

3D THINKING | DCI Disciplinary Core Ideas | CCC Crosscutting Concepts | SEP Science & Engineering Practices

COLLECT EVIDENCE

Use your Science Journal to record the evidence you collect as you complete the readings and activities in this lesson.

INVESTIGATE

GO ONLINE to find these activities and more resources.

Design Your Own Lab: How thick is the coating on a galvanized nail?
Plan and carry out an investigation to discover the structure and function of the coating on a galvanized nail by using your knowledge of the chemical properties involved.

Ionic compounds in solution

In addition to molecular compounds, ionic compounds might be solutes in aqueous solutions. Recall that ionic compounds consist of positive ions and negative ions held together by ionic bonds. When ionic compounds dissolve in water, their ions can separate—a process called dissociation. For example, an aqueous solution of the ionic compound sodium chloride contains Na^+ and Cl^- ions because when sodium chloride is added to water, the ions in the compound dissociate and become dispersed throughout the resulting solution.

Types of Reactions in Aqueous Solutions

When two aqueous solutions that contain ions as solutes are combined, the ions might react with one another. These reactions are always double-replacement reactions. The solvent molecules, which are all water molecules, do not usually react. Three types of products can form from the double-replacement reaction: a precipitate, water, or a gas.

Reactions that form precipitates

Some reactions that occur in aqueous solutions produce precipitates. For example, recall from Lesson 2 that when aqueous solutions of sodium hydroxide and copper(II) chloride are mixed, a double-replacement reaction occurs in which the precipitate copper(II) hydroxide forms. This reaction is shown in **Figure 17.**

$$2NaOH(aq) + CuCl_2(aq) \rightarrow 2NaCl(aq) + Cu(OH)_2(s)$$

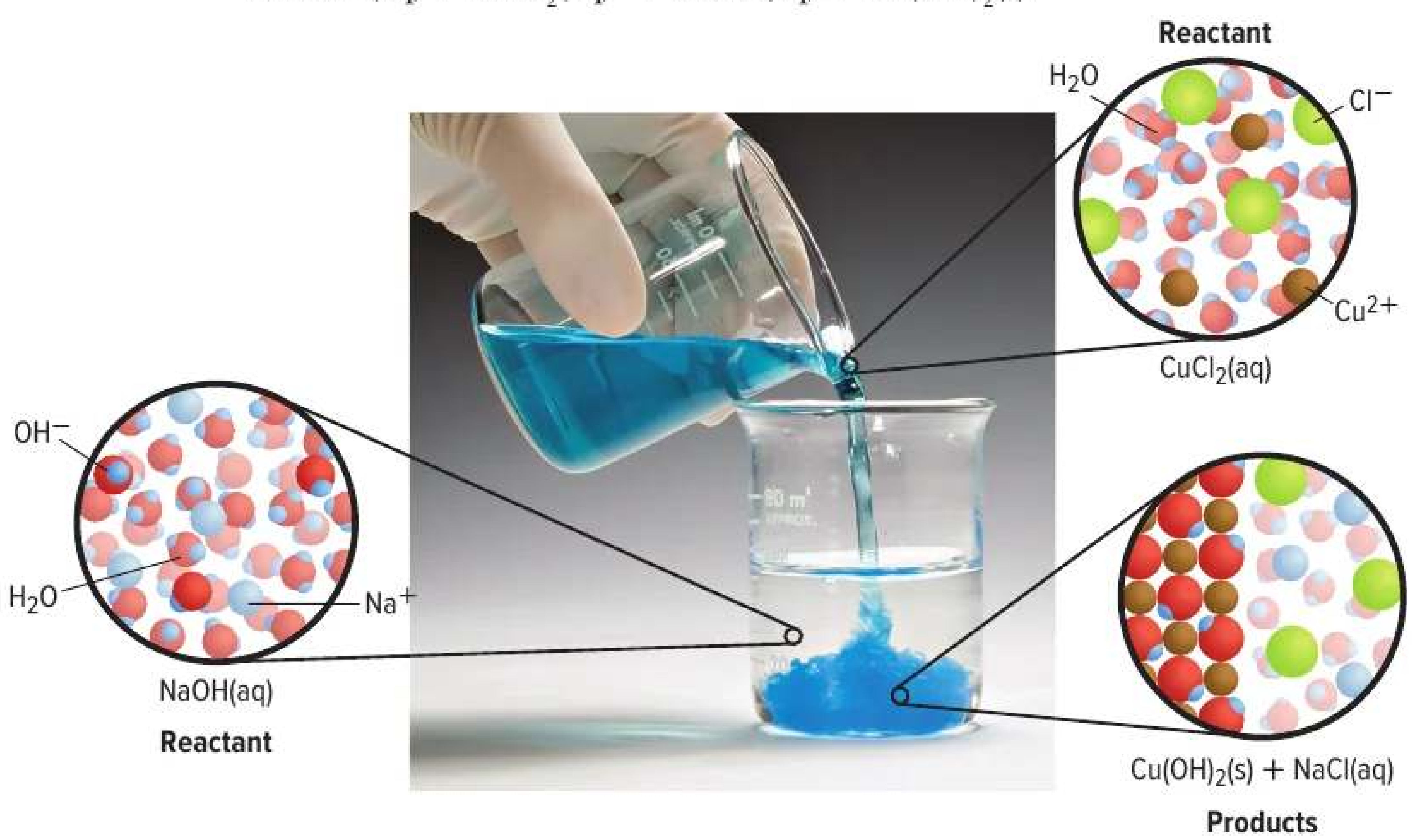

Figure 17 Like the aqueous solution of HCl in **Figure 16**, sodium hydroxide (NaOH) in an aqueous solution dissociates into sodium (Na^+) and hydroxide (OH^-) ions. Copper(II) chloride ($CuCl_2$) also dissociates into Cu^{2+} and Cl^- ions.

Interpret *What is the identity of the blue solid that is forming in the beaker?*

Note that the chemical equation does not show some details of this reaction. Sodium hydroxide and copper(II) chloride are ionic compounds. Therefore, in aqueous solutions they exist as Na^+, OH^-, Cu^{2+}, and Cl^- ions. When their solutions are combined, Cu^{2+} ions in one solution and OH^- ions in the other solution react to form the precipitate copper(II) hydroxide, $Cu(OH)_2(s)$. The Na^+ and Cl^- ions remain dissolved in the newly formed solution.

Ionic equations To show the details of reactions that involve ions in aqueous solutions, chemists use ionic equations. Ionic equations differ from chemical equations in that substances that are ions in solution are written as ions in the equation. Look again at the reaction between aqueous solutions of sodium hydroxide and copper(II) chloride. To write the ionic equation for this reaction, you must show the reactants, NaOH(aq) and $CuCl_2(aq)$, and the product, NaCl(aq), as ions.

$$2Na^+(aq) + 2OH^-(aq) + Cu^{2+}(aq) + 2Cl^-(aq) \rightarrow 2Na^+(aq) + 2Cl^-(aq) + Cu(OH)_2(s)$$

An ionic equation that shows all of the particles in a solution as they exist is called a **complete ionic equation.**

Note that the sodium ions and the chloride ions are both reactants and products. Because they are both reactants and products, they do not participate in the reaction. Ions that do not participate in a chemical reaction are called **spectator ions** and are not usually shown in ionic equations. Spectator ions are like spectators at a baseball game. The ions are present for the reaction but they do not affect the outcome of the reaction, just as spectators at a baseball game are present to watch the game but do not directly affect its outcome.

Net ionic equations are ionic equations that include only the particles that participate in the reaction. Net ionic equations are written from complete ionic equations by removing all spectator ions. For example, a net ionic equation is what remains after the sodium and chloride ions are crossed out of this complete ionic equation.

$$\cancel{2Na^+(aq)} + 2OH^-(aq) + Cu^{2+}(aq) + \cancel{2Cl^-(aq)} \rightarrow \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)} + Cu(OH)_2(s)$$

Only the hydroxide and copper ions are left in the net ionic equation shown below.

$$2OH^-(aq) + Cu^{2+}(aq) \rightarrow Cu(OH)_2(s)$$

Get It?

Compare How are complete ionic equations and net ionic equations different from chemical equations?

SCIENCE USAGE V. COMMON USAGE

compound

Science usage: a chemical combination of two or more different elements

Salt is a compound comprised of the elements sodium and chlorine.

Common usage: a word that consists of two or more words

Two compound words are basketball *and* textbook.

EXAMPLE Problem 3

REACTIONS THAT FORM A PRECIPITATE Write the chemical, complete ionic, and net ionic equations for the reaction between aqueous solutions of barium nitrate and sodium carbonate that forms the precipitate barium carbonate.

1 ANALYZE THE PROBLEM

You are given the word equation for the reaction between barium nitrate and sodium carbonate. You must determine the chemical formulas and relative amounts of all reactants and products to write the balanced chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation, you can write the net ionic equation. The net ionic equation will include fewer substances than the other equations.

2 SOLVE FOR THE UNKNOWN

Write the correct chemical formulas and physical states for all substances involved in the reaction.

$Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(s) + NaNO_3(aq)$

$Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(s) + 2NaNO_3(aq)$ — Balance the skeleton equation.

$Ba^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + 2Na^+(aq) + 2NO_3^-(aq)$ — Show the ions of the reactants and the products.

$Ba^{2+}(aq) + \cancel{2NO_3^-(aq)} + \cancel{2Na^+(aq)} + CO_3^{2-}(aq) \rightarrow BaCO_3(s) + \cancel{2Na^+(aq)} + \cancel{2NO_3^-(aq)}$ — Cross out the spectator ions from the complete ionic equation.

$Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$ — Write the net ionic equation.

3 EVALUATE THE ANSWER

The net ionic equation includes fewer substances than the other equations because it shows only the reacting particles. The particles composing the solid precipitate that is the result of the reaction are no longer ions.

PRACTICE Problems — ADDITIONAL PRACTICE

Write chemical, complete ionic, and net ionic equations for each of the following reactions that might produce a precipitate. Use *NR* to indicate that no reaction occurs.

35. Aqueous solutions of potassium iodide and silver nitrate are mixed, forming the precipitate silver iodide.

36. Aqueous solutions of ammonium phosphate and sodium sulfate are mixed. No precipitate forms and no gas is produced.

37. Aqueous solutions of aluminum chloride and sodium hydroxide are mixed, forming the precipitate aluminum hydroxide.

38. Aqueous solutions of lithium sulfate and calcium nitrate are mixed, forming the precipitate calcium sulfate.

39. **CHALLENGE** When aqueous solutions of sodium carbonate and manganese(V) chloride are mixed, a precipitate forms. The precipitate is a compound containing manganese.

Reactions that form water

Another type of double-replacement reaction that occurs in an aqueous solution produces water molecules. For example, when you mix hydrobromic acid (HBr) with a sodium hydroxide solution (NaOH), as shown in **Figure 18**, a double-replacement reaction occurs and water is formed. The chemical equation is shown below.

$$HBr(aq) + NaOH(aq) \rightarrow H_2O(l) + NaBr(aq)$$

In this case, the reactants and the product sodium bromide exist as ions in an aqueous solution. The complete ionic equation for this reaction shows these ions.

$$H^+(aq) + Br^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Na^+(aq) + Br^-(aq)$$

Examine the complete ionic equation. The reacting solute ions are the hydrogen ions and hydroxide ions because the sodium ions and bromine ions are spectator ions. If you cross out the spectator ions, you are left with the ions that take part in the reaction.

$$H^+(aq) + \cancel{Br^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \rightarrow H_2O(l) + \cancel{Na^+(aq)} + \cancel{Br^-(aq)}$$

This equation is the net ionic equation for the reaction.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Get It?

Analyze In the reaction between hydrobromic acid and sodium hydroxide, why are the sodium ions and bromine ions called spectator ions?

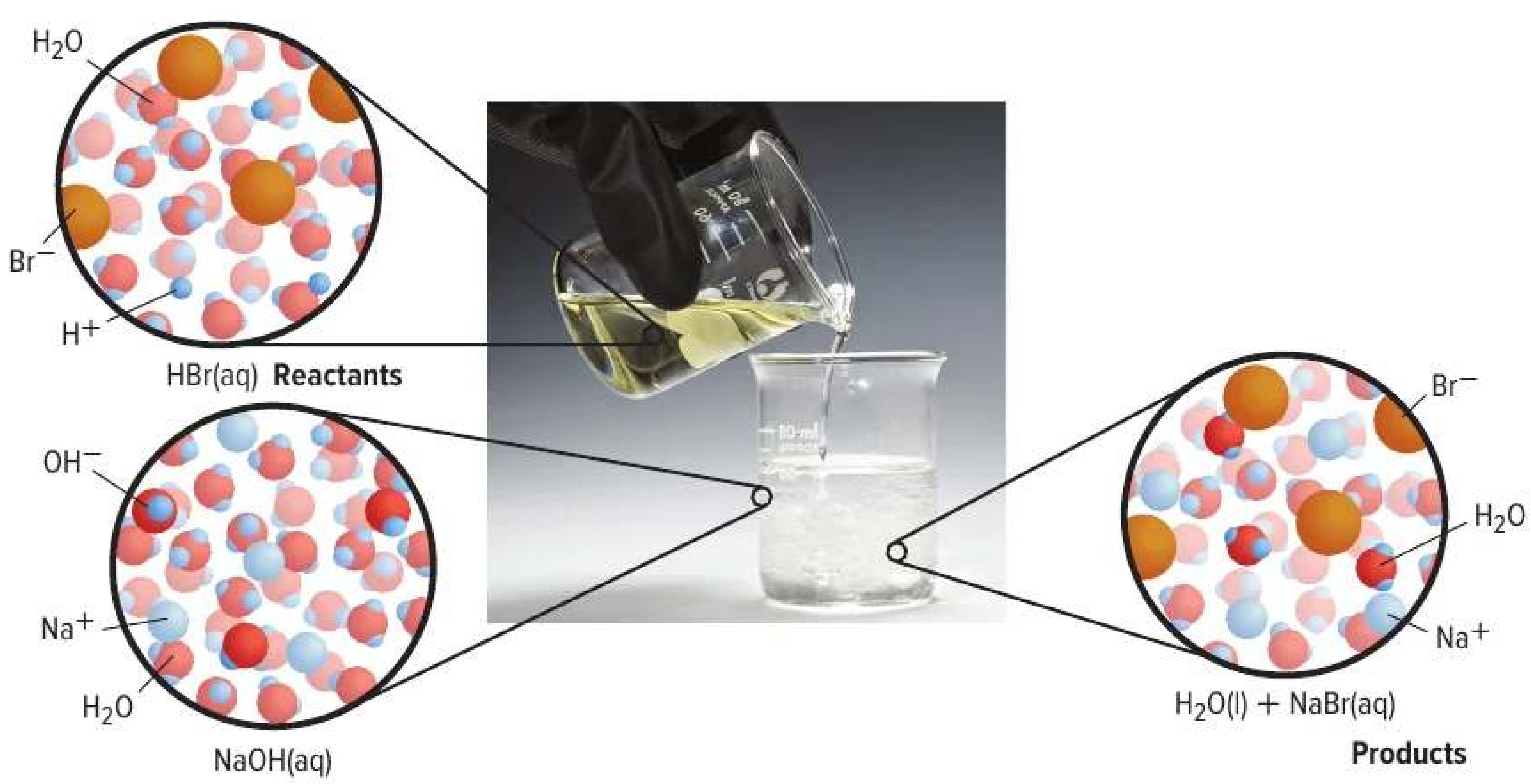

Figure 18 In water, hydrogen bromide (HBr) ionizes into H^+ and Br^- ions. Sodium hydroxide (NaOH) also dissociates into Na^+ and OH^- ions. The hydrogen ions and hydroxide ions react to form water.

Determine *Which ions are the anions in this reaction? The cations?*

After you have written the chemical, complete ionic, and net ionic equations for several double-replacement reactions that produce water, you will notice that the net ionic equation is the same for all of them. It shows hydrogen ions and hydroxide ions reacting to form water. If after writing the chemical and complete ionic equations for these types of reactions you end up with a different net ionic equation, go back and check your work.

EXAMPLE Problem 4

REACTIONS THAT FORM WATER Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous lithium hydroxide. This reaction produces water and aqueous lithium chloride.

1 ANALYZE THE PROBLEM

You are given the word equation for the reaction that occurs between hydrochloric acid and aqueous lithium hydroxide to produce water and aqueous lithium chloride. You must determine the chemical formulas for and relative amounts of all reactants and products to write the balanced chemical equation. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions from the complete ionic equation, you can write the net ionic equation.

2 SOLVE FOR THE UNKNOWN

Write the skeleton equation for the reaction and balance it.

$HCl(aq) + LiOH(aq) \rightarrow H_2O(l) + LiCl(aq)$

$H^+(aq) + Cl^-(aq) + Li^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Li^+(aq) + Cl^-(aq)$

Show the ions of the reactants and the products.

$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Li^+(aq)} + OH^-(aq) \rightarrow H_2O(l) + \cancel{Li^+(aq)} + \cancel{Cl^-(aq)}$

Cross out the spectator ions from the complete ionic equation.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Write the net ionic equation.

3 EVALUATE THE ANSWER

The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produces water. The particles that compose the product water are no longer ions.

PRACTICE Problems

ADDITIONAL PRACTICE

Write chemical, complete ionic, and net ionic equations for the reactions between the following substances, which produce water.

40. Mixing sulfuric acid (H_2SO_4) and aqueous potassium hydroxide produces water and aqueous potassium sulfate.

41. Mixing hydrochloric acid (HCl) and aqueous calcium hydroxide produces water and aqueous calcium chloride.

42. Mixing nitric acid (HNO_3) and aqueous ammonium hydroxide produces water and aqueous ammonium nitrate.

43. Mixing hydrosulfuric acid (H_2S) and aqueous calcium hydroxide produces water and aqueous calcium sulfide.

44. **CHALLENGE** When benzoic acid (C_6H_5COOH) and magnesium hydroxide are mixed, water and magnesium benzoate are produced.

Reactions that form gases

A third type of double-replacement reaction that occurs in aqueous solutions results in the formation of a gas. Some gases commonly produced in these reactions are carbon dioxide, hydrogen cyanide, and hydrogen sulfide. A gas-producing reaction occurs when you mix hydroiodic acid (HI) with an aqueous solution of lithium sulfide. Bubbles of hydrogen sulfide gas form in the container during the reaction. Lithium iodide is also produced.

$$2HI(aq) + Li_2S(aq) \rightarrow H_2S(g) + 2LiI(aq)$$

The reactants hydroiodic acid and lithium sulfide exist as ions in aqueous solution. Therefore, you can write an ionic equation for this reaction. The complete ionic equation includes all of the substances in the solution.

$$2H^+(aq) + 2I^-(aq) + 2Li^+(aq) + S^{2-}(aq) \rightarrow H_2S(g) + 2Li^+(aq) + 2I^-(aq)$$

Note that there are many spectator ions in the equation. When the spectator ions are crossed out, only the substances involved in the reaction remain in the equation. In this case, only hydrogen ions, sulfide ions, and hydrogen sulfide remain.

$$2H^+(aq) + \cancel{2I^-(aq)} + \cancel{2Li^+(aq)} + S^{2-}(aq) \rightarrow H_2S(g) + \cancel{2Li^+(aq)} + \cancel{2I^-(aq)}$$

This is the net ionic equation.

$$2H^+(aq) + S^{2-}(aq) \rightarrow H_2S(g)$$

Get It?

Explain why lithium ions are not shown in the above equation.

A reaction that produces carbon dioxide gas occurs in your kitchen when you mix vinegar and baking soda. Vinegar is an aqueous solution of acetic acid and water. Baking soda essentially consists of sodium hydrogen carbonate. Rapid bubbling occurs when vinegar and baking soda are combined. The bubbles are carbon dioxide gas escaping from the solution. You can see this reaction occurring in **Figure 19**.

Figure 19 When vinegar and baking soda (sodium hydrogen carbonate, $NaHCO_3$) combine, the result is a vigorous bubbling that releases carbon dioxide (CO_2).

A reaction similar to the one between vinegar and baking soda occurs when you combine any acidic solution and sodium hydrogen carbonate. In all cases, two reactions must occur almost simultaneously in the solution to produce the carbon dioxide gas. One of these is a double-replacement reaction and the other is a decomposition reaction.

For example, when you dissolve sodium hydrogen carbonate in hydrochloric acid, a gas-producing double-replacement reaction occurs. The hydrogen in the hydrochloric acid and the sodium in the sodium hydrogen carbonate replace each other.

$$HCl(aq) + NaHCO_3(aq) \rightarrow H_2CO_3(aq) + NaCl(aq)$$

Sodium chloride is an ionic compound, and its ions remain separate in the aqueous solution. However, as the carbonic acid (H_2CO_3) forms, it decomposes immediately into water and carbon dioxide.

$$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$

EXAMPLE Problem 5

REACTIONS THAT FORM GASES Write the chemical, complete ionic, and net ionic equations for the reaction between hydrochloric acid and aqueous sodium sulfide, which produces hydrogen sulfide gas.

1 ANALYZE THE PROBLEM

You are given the word equation for the reaction between hydrochloric acid (HCl) and sodium sulfide (Na_2S). You must write the skeleton equation and balance it. To write the complete ionic equation, you need to show the ionic states of the reactants and products. By crossing out the spectator ions in the complete ionic equation, you can write the net ionic equation.

2 SOLVE FOR THE UNKNOWN

Write the correct skeleton equation for the reaction.

$HCl(aq) + Na_2S(aq) \rightarrow H_2S(g) + NaCl(aq)$

$2HCl(aq) + Na_2S(aq) \rightarrow H_2S(g) + 2NaCl(aq)$ — Balance the skeleton equation.

$2H^+(aq) + 2Cl^-(aq) + 2Na^+(aq) + S^{2-}(aq) \rightarrow H_2S(g) + 2Na^+(aq) + 2Cl^-(aq)$ — Show the ions of the reactants and the products.

$2H^+(aq) + \cancel{2Cl^-(aq)} + \cancel{2Na^+(aq)} + S^{2-}(aq) \rightarrow H_2S(g) + \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)}$ — Cross out the spectator ions from the complete ionic equation.

$2H^+(aq) + S^{2-}(aq) \rightarrow H_2S(g)$ — Write the net ionic equation in its smallest whole-number ratio.

3 EVALUATE THE ANSWER

The net ionic equation includes fewer substances than the other equations because it shows only those particles involved in the reaction that produce hydrogen sulfide. The particles that compose the product are no longer ions.

PRACTICE Problems

ADDITIONAL PRACTICE

Write chemical, complete ionic, and net ionic equations for these reactions.

45. Perchloric acid ($HClO_4$) reacts with aqueous potassium carbonate, forming carbon dioxide gas and water.

46. Sulfuric acid (H_2S0_4) reacts with aqueous sodium cyanide, forming hydrogen cyanide gas and aqueous sodium sulfate.

47. Hydrobromic acid (HBr) reacts with aqueous ammonium carbonate, forming carbon dioxide gas and water.

48. Nitric acid (HNO_3) reacts with aqueous potassium rubidium sulfide, forming hydrogen sulfide gas.

49. **CHALLENGE** Aqueous potassium iodide reacts with lead nitrate in solution. forming solid lead iodide.

Double-replacement reaction

$$\begin{array}{ccccccc} AX & + & BY & \longrightarrow & AY & + & BX \\ HCl(aq) & + & NaHCO_3(aq) & \longrightarrow & H_2CO_3(aq) & + & NaCl(aq) \end{array}$$

$$\begin{array}{ccccc} AB & \longrightarrow & A & + & B \\ H_2CO_3(aq) & \longrightarrow & H_2O(l) & + & CO_2(g) \end{array}$$

Decomposition reaction

Figure 20 When HCl is combined with $NaHCO_3$, a double-replacement reaction takes place, followed immediately by a decomposition reaction.

Overall equations

Recall that when you combine an acidic solution, such as hydrochloric acid, and sodium hydrogen carbonate, two reactions occur—a double-replacement reaction and a decomposition reaction. These reactions are shown in **Figure 20.** The two reactions can be combined and represented by one chemical equation in a process similar to adding mathematical equations. An equation that combines two reactions is called an overall equation. To write an overall equation, the reactants in the two reactions are written on the reactant side of the combined equation, and the products of the two reactions are written on the product side. Then, any substances that are on both sides of the equation are crossed out.

Reaction 1 $HCl(aq) + NaHCO_3(aq) \rightarrow H_2CO_3(aq) + NaCl(aq)$

Reaction 2 $H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$

Combined equation $HCl(aq) + NaHCO_3(aq) + \cancel{H_2CO_3(aq)} \rightarrow \cancel{H_2CO_3(aq)} + NaCl(aq) + H_2O(l) + CO_2(g)$

Overall equation $HCl(aq) + NaHCO_3(aq) \rightarrow H_2O(l) + CO_2(g) + NaCl(aq)$

In this case, the reactants in the overall equation exist as ions in aqueous solution. Therefore, a complete ionic equation can be written for the reaction.

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g) + Na^+(aq) + Cl^-(aq)$$

Note that the sodium and chloride ions are the spectator ions. When you cross them out, only the substances that take part in the reaction remain.

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g) + \cancel{Na^+(aq)} + \cancel{Cl^-(aq)}$$

The net ionic equation shows that the reaction produces water and carbon dioxide gas.

$$H^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g)$$

Get It?

Describe What is an overall equation?

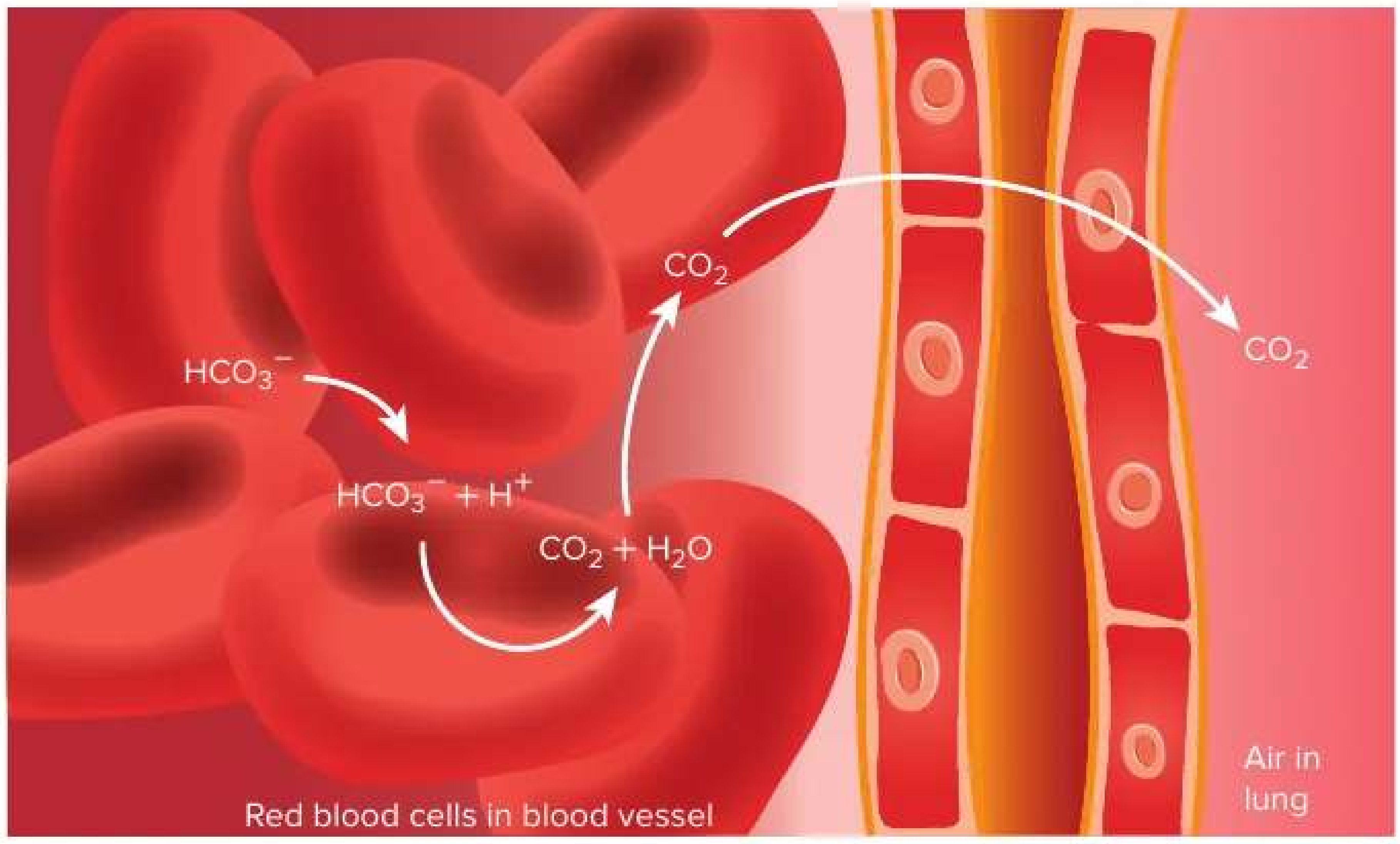

Figure 21 After a hydrogen carbonate ion (HCO_3^-) enters a red blood cell, it reacts with a hydrogen ion (H^+) to form water and carbon dioxide (CO_2). The CO_2 is exhaled from the lungs during respiration.

LIFE SCIENCE Connection The reaction between hydrogen ions and hydrogen carbonate ions to produce water and carbon dioxide is an important one in your body. This reaction is occurring in the blood vessels of your lungs as you read these words. As shown in **Figure 21**, the carbon dioxide gas produced in your cells is transported in your blood as hydrogen carbonate ions (HCO_3^-). In the blood vessels of your lungs, the HCO_3^- ions combine with H^+ ions to produce water and CO_2, which you exhale.

As you have read, the reaction between an acid and sodium hydrogen carbonate also occurs in products that are made with baking soda, which contains sodium hydrogen carbonate. Sodium hydrogen carbonate makes baked goods rise. This is partly because the sodium hydrogen carbonate reacts with an acid in the batter, such as lemon juice, to produce carbon dioxide. The gas results in bubbles that give the baked item a fluffy texture. In addition, at high temperatures, sodium hydrogen carbonate decomposes to form products that include carbon dioxide. The chemical and physical properties of sodium hydrogen carbonate mean it has many other uses. It is used as an antacid and in deodorants to absorb moisture and odors. Baking soda can be added to toothpaste to whiten teeth and freshen breath. As a paste, sodium bicarbonate can be used in cleaning and scrubbing. It is even used as a fire-suppression agent in some fire extinguishers.

Get It?

Infer Why is a mechanism needed to remove carbon dioxide gas from your cells?

STEM CAREER Connection

Hair Stylist

Most hair stylists are savvy with scissors and up to date on the latest hair trends, but did you know that they also use chemical reactions in their career? Hair stylists work with many processes, like those that permanently curl or relax hair, that involve chemical reactions. They also need to know how to handle hazardous chemicals safely.

Table 5 Reactions that Take Place in Aqueous Solutions

Type of Reaction	Description
Reactions that form precipitates	Dissolved substances are mixed. When they react, a solid is produced, visible as a white or colored cloudiness in the reaction mixture.
Reactions that form water	Evidence of the reaction may not be observable because water is colorless, odorless, and already makes up most of the solution.
Reactions that form gases	Gases such as carbon dioxide, hydrogen cyanide, and hydrogen sulfide are produced. Bubbles are produced as the reaction proceeds.

Table 5 lists the types of reactions that occur in aqueous solutions. The descriptions summarize how physical evidence such as the formation of a precipitate or the production of bubbles can help you classify a chemical reaction.

Check Your Progress

Summary

- In aqueous solutions, the solvent is always water. There are many possible solutes.
- Many molecular compounds form ions when they dissolve in water. When some ionic compounds dissolve in water, their ions separate.
- When two aqueous solutions that contain ions as solutes are combined, the ions might react with one another. The solvent molecules do not usually react.
- Reactions that occur in aqueous solutions are double-replacement reactions.

Demonstrate Understanding

50. **List** three common types of products produced by reactions that occur in aqueous solutions.
51. **Describe** solvents and solutes in aqueous solution.
52. **Distinguish** between a complete ionic equation and a net ionic equation.
53. **Write** complete ionic and net ionic equations for the reaction between sulfuric acid (H_2SO_4) and calcium carbonate ($CaCO_3$).
    $H_2SO_4(aq) + CaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + CaSO_4(aq)$
54. **Analyze** Complete and balance the following equation.
    $$CO_2(g) + HCl(aq) \rightarrow$$
55. **Predict** What type of product would the following reaction be most likely to produce? Explain your reasoning.
    $$Ba(OH)_2(aq) + 2HCl(aq) \rightarrow$$
56. **Formulate Equations** A reaction occurs when nitric acid (HNO_3) is mixed with an aqueous solution of potassium hydrogen carbonate. Aqueous potassium nitrate is produced. Write the chemical and net ionic equations for the reaction.

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SCIENCE & SOCIETY

How One Woman Led the FDA to Save Lives

When a drug company could not provide data to support its claims of safety for a new drug, one woman's persistence shaped the future of the Food and Drug Administration.

Strict warnings about the potential for birth defects are displayed on thalidomide's packaging.

Drug Safety

Drug metabolism is the series of chemical reactions that breaks down a drug into compounds that are used in the body. Before a new drug is approved for human use, scientists perform research to ensure that no harmful side effects are produced as a result of these reactions. Pharmacologists – scientists who study the reactions of drugs in the body – work with the Food and Drug Administration (FDA) to evaluate data and make decisions regarding the safety and effectiveness of a proposed medication.

A Historical Example: Thalidomide

In 1960, the FDA received an application for the approval of thalidomide, a drug used to treat a variety of symptoms from nausea to sleeplessness. In Europe and other parts of the world, many doctors gave the drug to pregnant women as a treatment for morning sickness.

The application fell to pharmacologist Dr. Frances Oldham Kelsey. When she reviewed clinical studies of the drug, the lack of research to support the drug's safety concerned her. She and her colleagues were further alarmed when insistence for more data from the drug company did not result in additional evidence.

While the FDA held off approving the application, reports of devastating birth defects in babies born to mothers who had taken thalidomide surfaced around the world. By November 1961, German officials took the drug off the market, and other countries soon followed suit. By early 1962, the distributor withdrew its application for FDA approval.

Because of her persistence in the pursuit of evidence, Kelsey impacted countless lives. In 1962, President John F. Kennedy honored Dr. Kelsey with the Distinguished Federal Civilian Service award. In part because of her work, the United States Congress passed the Kefauver-Harris bill, forcing major regulatory reforms on the pharmaceutical industry.

COMMUNICATE SCIENTIFIC INFORMATION

Research the metabolism of atorvastatin, a commonly prescribed medication used to lower cholesterol. Make a presentation that explains these metabolic reactions. Include any potential harmful effects or products of the reactions.

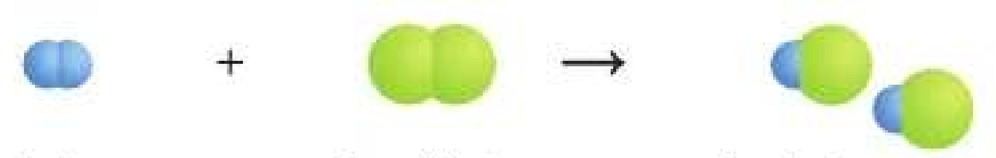
MODULE 8
STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 REACTIONS AND EQUATIONS

- Some physical changes are evidence that indicate a chemical reaction has occurred.
- Word equations and skeleton equations provide important information about a chemical reaction.
- A chemical equation gives the identities and relative amounts of the reactants and products that are involved in a chemical reaction.
- Balancing an equation involves adjusting the coefficients until the number of atoms of each element is equal on both sides of the equation.

H_2	+	Cl_2	$\longrightarrow$	$2HCl$
2 atoms H		2 atoms Cl		2 atoms H + 2 atoms Cl



Two hydrogen atoms	Two chlorine atoms	Two hydrogen atoms Two chlorine atoms

- chemical reaction
- reactant
- product
- chemical equation
- coefficient

Lesson 2 CLASSIFYING CHEMICAL REACTIONS

- Classifying chemical reactions makes them easier to understand, remember, and recognize.
- Activity series of metals and halogens can be used to predict if single-replacement reactions will occur.

- synthesis reaction
- combustion reaction
- decomposition reaction
- single-replacement reaction
- double-replacement reaction
- precipitate

Lesson 3 REACTIONS IN AQUEOUS SOLUTIONS

- In aqueous solutions, the solvent is always water. There are many possible solutes.
- Many molecular compounds form ions when they dissolve in water. When some ionic compounds dissolve in water, their ions separate.
- When two aqueous solutions that contain ions as solutes are combined, the ions might react with one another. The solvent molecules do not usually react.
- Reactions that occur in aqueous solutions are double-replacement reactions.

- aqueous solution
- solute
- solvent
- complete ionic equation
- spectator ion
- net ionic equation

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

What happens to food when you cook it?

CER Claim, Evidence, Reasoning

Explain your Reasoning Revisit the claim you made when you encountered the phenomenon. Summarize the evidence you gathered from your investigations and research and finalize your Summary Table. Does your evidence support your claim? If not, revise your claim. Explain why your evidence supports your claim.

STEM UNIT PROJECT

Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

SEP Data Analysis Lab

How can you explain the reactivities of halogens?

The location of all the halogens in group 17 in the periodic table tells you that halogens have common characteristics. Indeed, halogens are all nonmetals and have seven electrons in their outermost orbitals. However, each halogen also has its own characteristics, such as the ability to react with other substances. Examine the data table. It includes data about the atomic radii, ionization energies, and electronegativities of the halogens.

Data and Observations

Properties of Halogens

Halogen	Atomic Radius (pm)	Ionization Energy (kJ/mol)	Electro-negativity
Fluorine	72	1681	3.98
Chlorine	100	1251	3.16
Bromine	114	1140	2.96
Iodine	133	1008	2.66
Astatine	140	920	2.2

CER Analyze and Interpret Data

1. **Make graphs** Use the information in the data table to make three line graphs.
2. **Claim** Describe any periodic trends that you identify in the data.
3. **Evidence, Reasoning** Relate any periodic trends that you identify among the halogens to the activity series of the halogens.
4. **Claim, Evidence, Reasoning** Predict the location of the element astatine in the activity series of halogens. Explain.

Credits

Student Notes

Student Notes

Student Notes

Student Notes

Student Notes

Student Notes

Student Notes