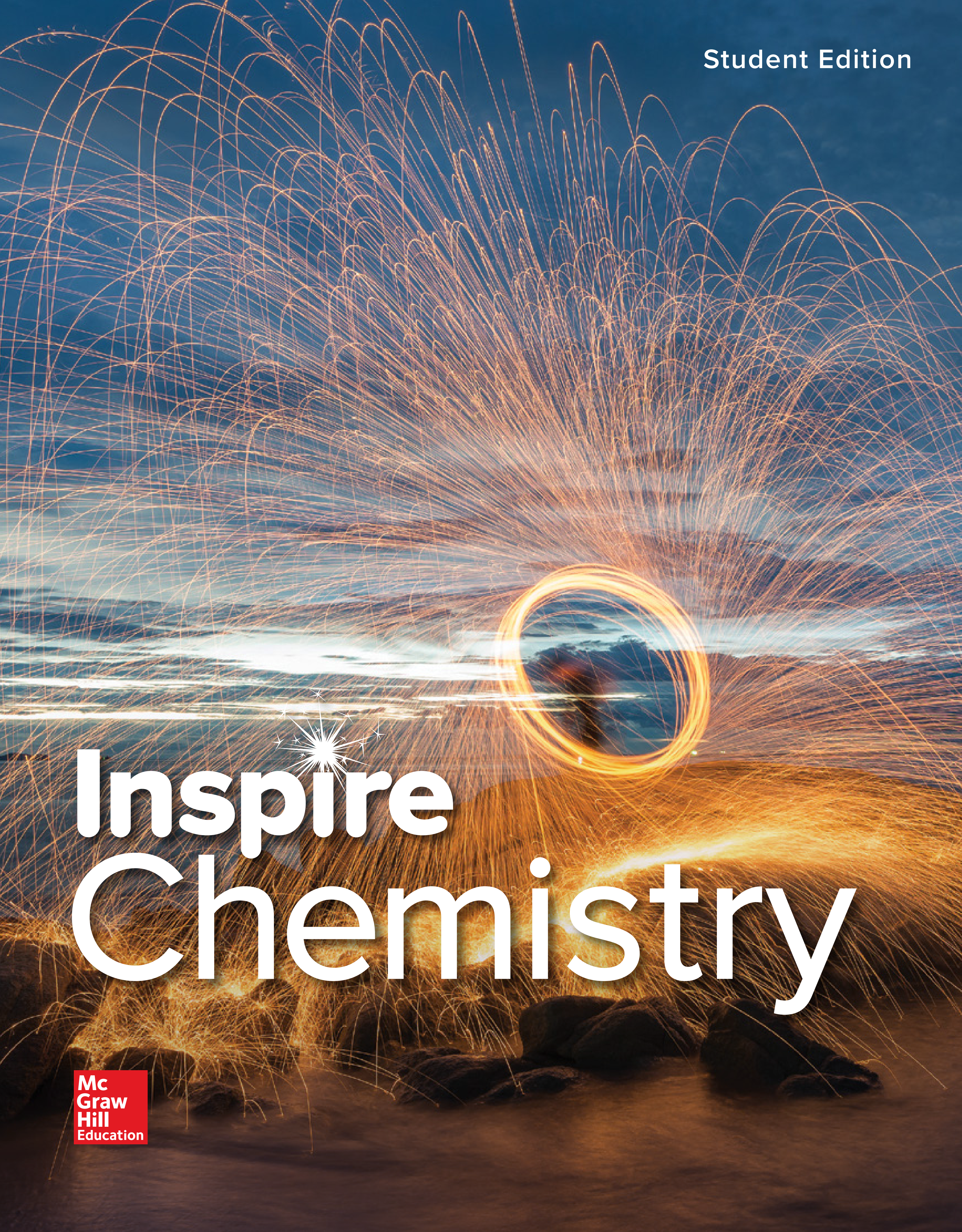

Student Edition

Inspire Chemistry

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Inspire Chemistry

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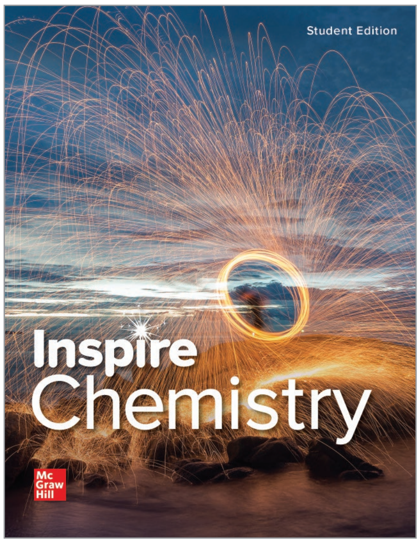

Phenomenon: Chemical Reactions

Thin strands of steel wool, which is made mostly of iron, react with oxygen to form iron(III) oxide, releasing energy as heat and light. The reaction, and therefore the glow, intensifies as the steel wool is spun in the air.

CAUTION: Burning steel wool gets extremely hot. Do not try this at home.

Fun Fact

Rusting is a slow reaction between iron and oxygen that also forms iron(III) oxide.

COVER: SupernanPhoto/iStock/Getty Images
BACK COVER: SupernanPhoto/iStock/Getty Images

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Welcome to

Inspire Chemistry

Explore Our Phenomenal World

The Inspire High School Series brings phenomena to the forefront of learning to engage and inspire students to investigate key science concepts through their three-dimensional learning experience.

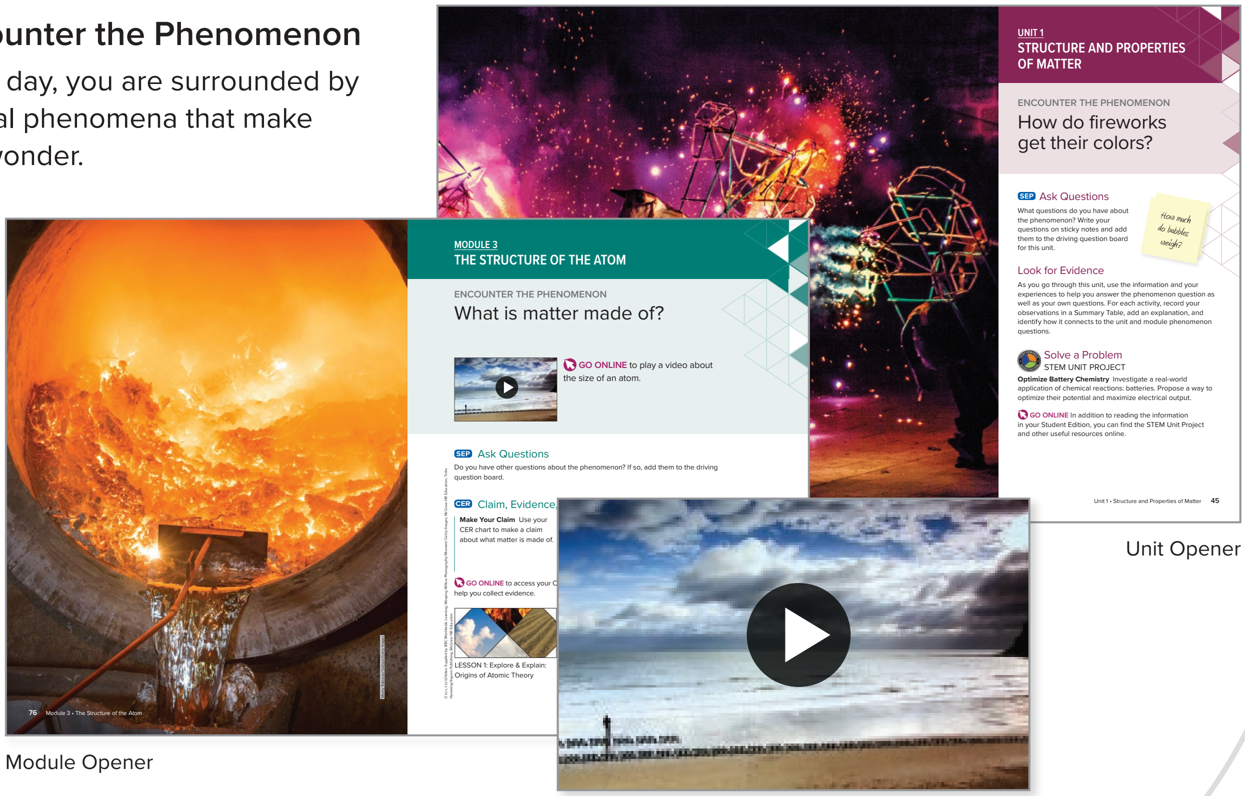
Start exploring now!

Inspire Curiosity • Inspire Investigation • Inspire Innovation

Owning Your Learning

1 Encounter the Phenomenon

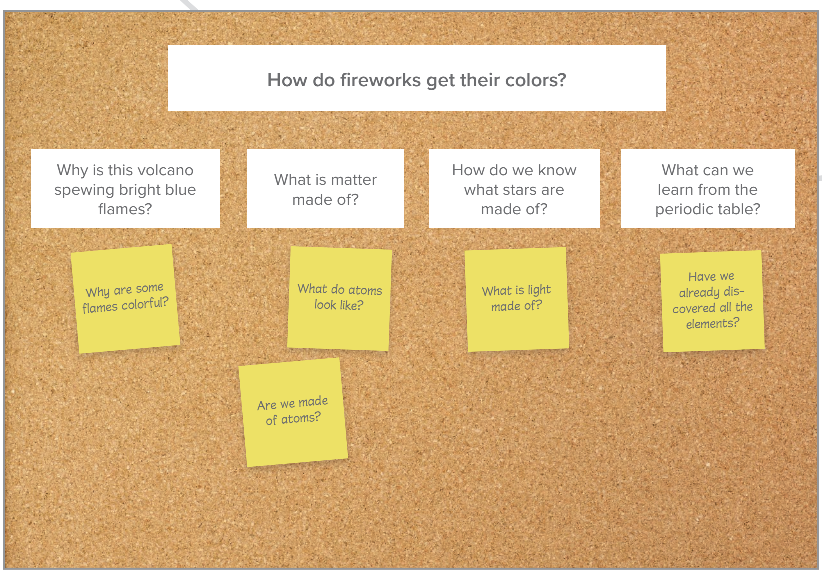
Every day, you are surrounded by natural phenomena that make you wonder.



Unit Opener

Module Opener

Phenomenon Video

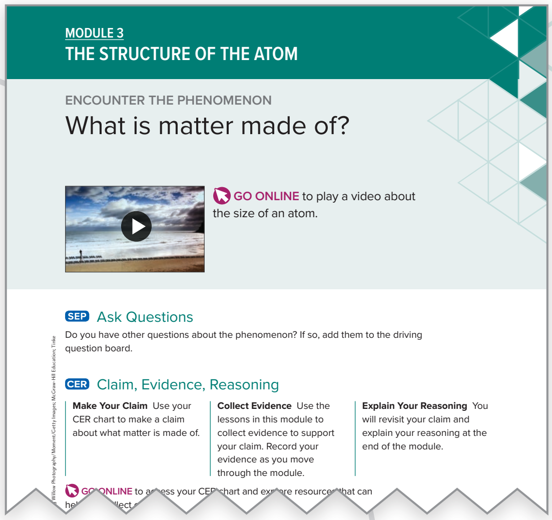


2 Ask Questions

At the beginning of each unit and module, make a list of the questions you have about the phenomenon. Share your questions with your classmates.

3 Claim, Evidence, Reasoning

As you investigate each phenomenon, you will write your claim, gather evidence by performing labs and completing reading assignments and Applying Practices, and explain your reasoning to answer the unit and module phenomena.


MODULE 3
THE STRUCTURE OF THE ATOM

ENCOUNTER THE PHENOMENON

What is matter made of?

GO ONLINE to play a video about the size of an atom.

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 matter is 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.

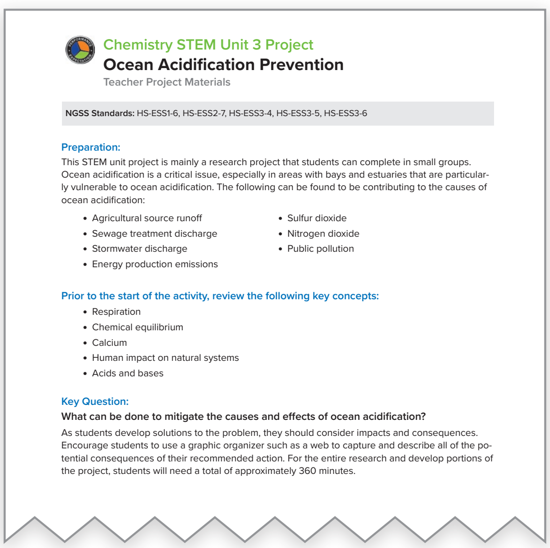
SUMMARY TABLE					
Activity Model	Observation Evidence	Explanation Reasoning	Connection to Phenom	Questions Answered	New Questions
Applying Practices: Is light a wave or a particle?	Light has both wavelike and particlelike properties.	Light is made up of bundles of energy called photons.	Unit: Firework colors are a result of the photons emitted by different elements. Module: Astronomers use light to determine the composition of the stars.	What is light made of?	Why do elements emit photons?

4 Summarize Your Work

When you collect evidence, you can record your data in a summary table and use the data to collaborate with others to answer the questions you had.

5 Apply Your Evidence and Reasoning

At the end of the unit, modules, and lessons, you can use all of the data you collected to help complete your STEM Unit Project.


Chemistry STEM Unit 3 Project
Ocean Acidification Prevention
Teacher Project Materials

NGSS Standards: HS-ESS1-6, HS-ESS2-7, HS-ESS3-4, HS-ESS3-5, HS-ESS3-6

Preparation:
This STEM unit project is mainly a research project that students can complete in small groups. Ocean acidification is a critical issue, especially in areas with bays and estuaries that are particularly vulnerable to ocean acidification. The following can be found to be contributing to the causes of ocean acidification:

- Agricultural source runoff
- Sewage treatment discharge
- Stormwater discharge
- Energy production emissions
- Sulfur dioxide
- Nitrogen dioxide
- Public pollution

Prior to the start of the activity, review the following key concepts:

- Respiration
- Chemical equilibrium
- Calcium
- Human impact on natural systems
- Acids and bases

Key Question:
What can be done to mitigate the causes and effects of ocean acidification?
As students develop solutions to the problem, they should consider impacts and consequences. Encourage students to use a graphic organizer such as a web to capture and describe all of the potential consequences of their recommended action. For the entire research and develop portions of the project, students will need a total of approximately 360 minutes.

ADVISORS AND CONSULTANTS

Teacher Advisory Board

The Teacher Advisory Board gave the editorial staff and design team feedback on the content and design of both the Student Edition and Teacher Edition. We thank these teachers for their hard work and creative suggestions.

Ann Cooper
Science Teacher
United Local Schools
Hanoverton, OH

David L. French
Chemistry Teacher
Milford High School
Milford, OH

Richard Glink
Chemistry/Physics Teacher
Indian Lake High School
Lewistown, OH

Susan Godez
Chemistry/Physics Teacher
Grandview Heights High School
Columbus, OH

Judith Johnston
Science Teacher, Department Chair
Wilmington High School
Wilmington, OH

Christine Lewis
Science Teacher
Martins Ferry High School
Martins Ferry, OH

Jennifer L. Most
Chemistry Teacher,
Science Department Chair
West Holmes High School
Millersburg, OH

Sandra Petrie-Forgey
National Board Certified
Science Teacher
Gallia Academy High School
Gallipolis, OH

Jason J. Zaros
Chemistry/Physics Teacher
Waterford High School
Waterford, OH

Content Consultants

Content consultants each reviewed selected chapters of ***Inspire Chemistry*** for content accuracy and clarity.

Alton J. Banks, PhD
Professor of Chemistry
North Carolina State University
Raleigh, NC

Howard Drossman, PhD
Professor of Chemistry and
Environmental Science
Colorado College
Colorado Springs, CO

Michael O. Hurst, Sr., PhD
Associate Professor of Chemistry
Georgia Southern University
Statesboro, GA

Kristen Kulinowski, PhD
Faculty Fellow, Department
of Chemistry
Rice University
Houston, TX

Maria Pacheco, PhD
Associate Professor of Chemistry
Buffalo State College
Buffalo, NY

Safety Consultant

The safety consultant reviewed labs and lab materials for safety and implementation.

Kenneth R. Roy, PhD
Director of Environmental Health and Safety
Glastonbury Public Schools
Glastonbury, CT

Contributing Writers

Additional science writers added feature content, teacher materials, assessment, and laboratory investigations.

Peter Carpico
Louisville, OH

Jennifer Gonya
Galena, OH

Cindy Klevickis
Elkton, VA

Jack Minot
Columbus, OH

Richard G. Smith
Ocean Isle Beach, NC

Stephen Whitt
Columbus, OH

Jenipher Willoughby
Forest, VA

Margaret K. Zorn
Yorktown, VA

KEY PARTNERS

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The PhET Interactive Simulations project at the University of Colorado Boulder provides teachers and students with interactive science and math simulations. Based on extensive education research, PhET sims engage students through an intuitive, game-like environment where students learn through exploration and discovery.

ABOUT THE AUTHORS

Thandi Buthelezi

Thandi is Associate Professor of Chemistry and Chemistry Department Chair at Wheaton College, Norton, MA. She earned her BA in Chemistry from Williams College, Williamstown, MA, and PhD in Experimental Physical Chemistry from the University of Florida, Gainesville, FL. Dr. Buthelezi has taught Chemistry at the undergraduate and graduate (master's) level. She was the co-founder and co-director of the Girls in Science Outreach Program at Western Kentucky University. She is a member of the American Chemical Society, the American Association for the Advancement of Science, and Sigma Xi.

Laurel Dingrando

Laurel is currently serving as the Secondary Science Coordinator for the Garland Independent School District. Mrs. Dingrando has a BS in Microbiology with a minor in Chemistry from Texas Tech University and an MAT in Science from the University of Texas at Dallas. She taught Chemistry for 35 years in the Garland Independent School District. She is a member of the American Chemical Society, National Science Teachers Association, Science Teachers Association of Texas, Texas Science Educators Leadership Association, and T3 (Teachers Teaching with Technology).

Nicholas Hainen

Nicholas taught chemistry and physics in the Worthington City Schools, Worthington, Ohio, for 31 years. Mr. Hainen holds BS and MA degrees in Science Education from The Ohio State University, majoring in chemistry and physics. His honors and awards include: American Chemical Society Outstanding Educator in Chemical Sciences; The Ohio State University Honor Roll of Outstanding High School Teachers; Ashland Oil Company Golden Apple Award; and Who's Who Among America's Teachers. Mr. Hainen is a member of the American Chemical Society and the ACS Division of Chemical Education.

Cheryl Wistrom

Cheryl is a Visiting Assistant Professor of Chemistry at Purdue University Northwest in Westville, Indiana. She has taught chemistry, biology, and science education courses at the college level since 1990 and is also a licensed pharmacist. She earned her BS degree in biochemistry at Northern Michigan University, a BS in pharmacy at Purdue University, and her MS and PhD in biological chemistry at the University of Michigan. Dr. Wistrom is a member of the Indiana Academy of Science, the National Science Teachers Association, and the American Society of Health-System Pharmacists.

Dinah Zike

Dinah is an international curriculum consultant and inventor who has developed educational products and three-dimensional, interactive graphic organizers for over 30 years. As president and founder of Dinah-Might Adventures, L.P., Dinah is the author of more than 100 award-winning educational publications, including *The Big Book of Science*. Dinah has a BS and an MS in educational curriculum and instruction from Texas A&M University. Dinah Zike's *Foldables* are an exclusive feature of McGraw-Hill textbooks.

UNIT 1

STRUCTURE AND PROPERTIES OF MATTER

ENCOUNTER THE PHENOMENON

How do fireworks get their colors?

UNIT 2

CHEMICAL BONDING AND REACTIONS

ENCOUNTER THE PHENOMENON

How can understanding chemical bonding and reactions support fire safety?

Rustin Gudim/ZUMA Press/Newscom

MODULE 6: IONIC COMPOUNDS AND METALS

MODULE 7: COVALENT BONDING

MODULE 8: CHEMICAL REACTIONS

MODULE 9: THE MOLE

MODULE 10: STOICHIOMETRY

Seaphotoart/Alamy Stock Photo

UNIT 3
MATTER, ENERGY, AND EQUILIBRIUM

ENCOUNTER THE PHENOMENON

How does ocean chemistry affect coral reefs?

UNIT 4

OXIDATION AND REDUCTION REACTIONS

ENCOUNTER THE PHENOMENON

Why do some metals tend to corrode?

MODULE 18: REDOX REACTIONS

MODULE 19: ELECTROCHEMISTRY

John A. Anderson/Shutterstock

UNIT 5

ORGANIC AND NUCLEAR CHEMISTRY

ENCOUNTER THE PHENOMENON

How has our use of carbon-based fuels affected our climate?

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MODULE 1
THE CENTRAL SCIENCE

ENCOUNTER THE PHENOMENON

What do plants and buildings have in common?

GO ONLINE to play a video about the discovery of potassium.

SEP Ask Questions

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

CER Claim, Evidence, Reasoning

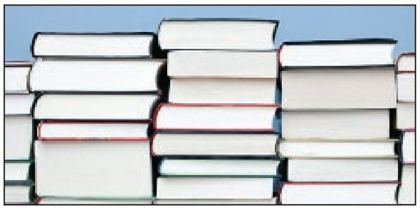
Make Your Claim Use your CER chart to make a claim about what plants and buildings have in common.

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: The Central Science



LESSON 1: Explore & Explain: Matter and Its Characteristics

Smithsonian

Additional Resources

LESSON 1

WHAT IS CHEMISTRY?

FOCUS QUESTION

How can chemistry help you understand the world?

Why study chemistry?

Take a moment to observe your surroundings and **Figure 1**. Where did all the "stuff" come from? All the stuff in the universe, including everything in the photos, is made from building blocks formed in stars. Scientists call these building blocks and the "stuff" made from these building blocks *matter.*

Figure 1 Everything in the universe, including particles in space and things around you, is composed of matter.

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?

As you begin your study of **chemistry**—the study of matter and the changes that it undergoes—you are probably asking yourself, "Why is chemistry important to me?" Whether you read this because of ink on a printed page or phosphors in a monitor or screen, you are relying on chemistry. Through chemistry, people have made the fibers in your clothing and those in the body panels of a car. If you cook food in a metal pan on a gas stove or in a glass bowl in a microwave oven, you make use of chemistry.

Chemistry: The Central Science

Because chemistry is the study of matter and the changes that it undergoes, a basic understanding of chemistry is central to all sciences—biology, physics, Earth science, ecology, and others. **Science** is the use of evidence to construct testable explanations and predictions of natural phenomena, as well as, the knowledge generated through this process. A testable explanation of a situation or phenomena is called a **hypothesis.** The nature, or essential characteristic, of science is scientific inquiry—the development of new explanations. Scientific inquiry is both a creative process and a process rooted in unbiased observations and investigation.

Because there are so many types of matter, there are many areas of study in the field of chemistry. Chemistry is traditionally broken down into branches that focus on specific areas, such as those listed in **Table 1**. Although chemistry is divided into specific areas of study, many of the areas overlap. For example, as you can see from **Table 1**, an organic chemist might study plastics, but an industrial chemist or a polymer chemist could also focus on plastics.

Table 1 Some Branches of Chemistry

Branch	Area of Emphasis	Examples of Emphasis
Organic chemistry	most carbon-containing chemicals	pharmaceuticals, plastics
Inorganic chemistry	in general, matter that does not contain carbon	minerals, metals and nonmetals, semiconductors
Physical chemistry	the behavior and changes of matter and the related energy changes	reaction rates, reaction mechanisms
Analytical chemistry	components and composition of substances	food nutrients, quality control
Biochemistry	matter and processes of living organisms	metabolism, fermentation
Environmental chemistry	matter and the environment	pollution, biochemical cycles
Industrial chemistry	chemical processes in industry	paints, coatings
Polymer chemistry	polymers and plastics	textiles, coatings, plastics
Theoretical chemistry	chemical interactions	many areas of emphasis
Thermochemistry	heat involved in chemical processes	heat of reaction

Figure 2 This car, which is powered by compressed air, and this tiny lens, which is about 0.4 mm wide, are examples of technologies that are made possible by the study of matter.

The Benefits of Chemistry

Chemists are an important part of the team of scientists that solve many of the problems or issues that we face today. Chemists are involved in resolving global issues, such as the ozone depletion problem. They are also involved in finding cures or vaccines for diseases, such as AIDS and influenza. Almost every situation that you can imagine involves a chemist, because everything in the universe is made of matter.

Figure 2 shows some of the advances in technology that are possible because of the study of matter. The car on the left is powered by compressed air. When the compressed air is allowed to expand, it pushes the pistons that move the car. Because the car is powered by compressed air, no pollutants are released. The photo on the right shows a type of lens that is made using nanofabrication techniques. Microscopic gears and even submarines have been developed through an understanding of matter.

Theory and Scientific Law

You might have heard of Einstein's theory of relativity or the atomic theory. In science, a **theory** is an explanation of a natural phenomenon based on many observations and investigations over time. A theory states a broad principle of nature that has been supported over time. All theories are still subject to new experimental data and can be modified. Also, theories often lead to new conclusions. A theory is considered valid if it can be used to make predictions that are proven true.

ACADEMIC VOCABULARY

validity

reasonableness; logical soundness

The inventor's claims seemed outrageous, but their validity was supported by careful testing.

Sometimes, many scientists come to the same conclusion about certain relationships in nature and they find no exceptions to these relationships. For example, you know that no matter how many times skydivers, like those shown in **Figure 3**, leap from a plane, they always return to Earth's surface. Sir Isaac Newton was so certain that an attractive force exists between all objects that he proposed his law of universal gravitation. Newton's law is a **scientific law**—a relationship in nature that is supported by many experiments. It is up to scientists to develop further hypotheses and experiments to explain why these relationships exist.

Figure 3 It does not matter how many times skydivers leap from a plane; Newton's law of universal gravitation applies every time.

Types of Scientific Investigations

Every day in the media—through TV, newspapers, magazines, or the Internet—the public is bombarded with the results of scientific investigations. Many deal with the environment, medicine, or health. As a consumer, you are asked to evaluate the results of scientific research and development. How do scientists use qualitative and quantitative data to solve different types of scientific problems?

Scientists conduct **pure research** to gain knowledge for the sake of knowledge itself. It is often driven by the curiosity of the researcher. For example, Harry Kroto, a chemist at the University of Sussex, had helped to discover long, linear-chain molecules of carbon in interstellar space. He was interested in investigating whether they might be formed in red giant stars. His investigations led to the discovery of buckminsterfullerenes, or buckyballs—a form of carbon composed of 60 carbon atoms arranged in the shape of a soccer ball.

Scientists conduct **applied research** to solve a specific problem. In the search for cancer therapies, scientists have developed ways to use fullerenes to destroy cancer cells through the process of photodynamic therapy (PDT). Scientists researched methods to modify fullerenes so that on exposure to light, the fullerene changes oxygen in a cell into highly reactive forms that damage cell components, which leads to cell death.

ACADEMIC VOCABULARY

qualitative

relating to or measuring the quality or characteristics of something rather than the quantity

The salesperson described the qualitative aspects of the car, like its color and its smooth ride—but never mentioned the quantitative aspects of gas mileage or cost.

Matter and Its Characteristics

Matter, the stuff of the universe, has many different forms. Everything around you, like the things in **Figure 4**, is matter. Some matter occurs naturally, such as ozone, and other substances are not natural, such as chlorofluorocarbons (CFCs). A **substance,** which is also known as a chemical, is matter that has a definite and uniform composition.

You might realize that everyday objects are composed of matter, but how do you define matter? Recall that matter is anything that has mass and takes up space. Also recall that **mass** is a measurement that reflects the amount of matter. You know that your textbook has mass and takes up space, but is air matter? You cannot see it and you cannot always feel it. However, when you inflate a balloon, it expands to make room for the air. The balloon gets heavier. Thus, air must be matter. Is everything matter? The thoughts and ideas that fill your head are not matter; neither are heat, light, radio waves, nor magnetic fields. What else can you name that is not matter?

Mass and weight

Have you ever used a bathroom scale to measure your weight? **Weight** is a measure not only of the amount of matter but also of the effect of Earth's gravitational pull on that matter. This force is not exactly the same everywhere on Earth and actually becomes less as you move away from Earth's surface at sea level. You might not notice a difference in your weight from one place to another, but subtle differences do exist.

Figure 4 Everything in this photo is matter and has mass and weight.

Compare and contrast *mass and weight.*

SCIENCE USAGE V. COMMON USAGE

weight

Science usage: the measure of the amount of matter in and the gravitational force exerted on an object

The weight of an object is the product of its mass and the local acceleration of gravity.

Common usage: the relative heaviness of an object

The puppy grew so quickly it doubled its weight in a matter of weeks.

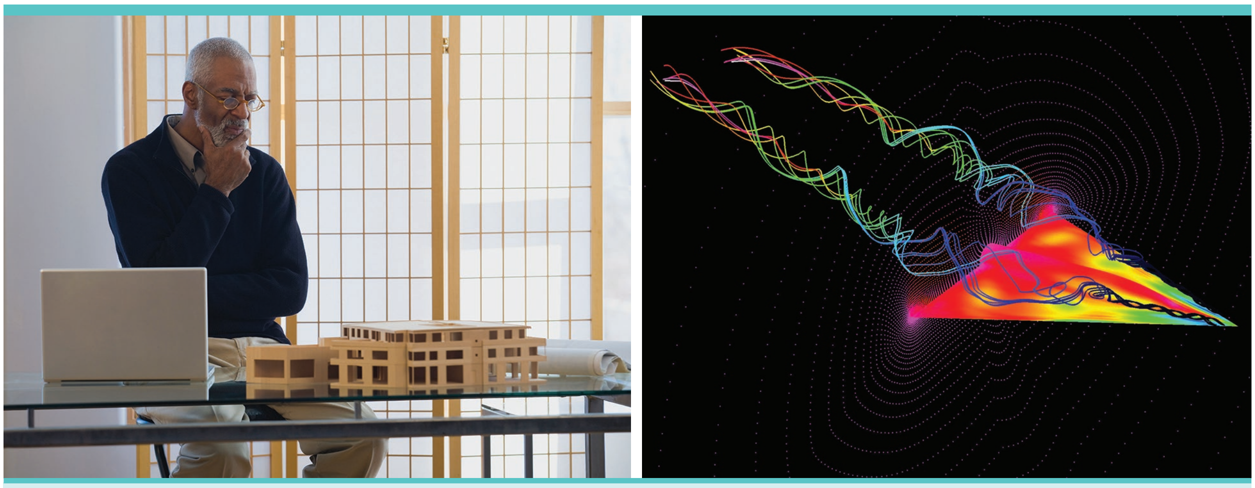
It might seem more convenient for scientists to simply use weight instead of mass. Why is it so important to think of matter in terms of mass? Scientists need to be able to compare the measurements that they make in different parts of the world. They could identify the gravitational force every time they weigh something, but that would not be practical or convenient. They use mass as a way to measure matter independently of gravitational force.

Structure and observable characteristics

What can you observe about the outside of your school building? You know that there is more to the building than what you can observe from the outside. Among other things, there are beams inside the walls that give the building structure, stability, and function. Consider another example. When you bend your arm at the elbow, you observe that your arm moves, but what you cannot see is that muscles under the skin contract and relax to move your arm.

Much of matter and its behavior is macroscopic; that is, you do not need a microscope to observe it. Chemistry seeks to explain the submicroscopic events that lead to macroscopic observations. One way this can be done is by making a model. A **model** is a visual, verbal, or mathematical explanation of experimental data. Scientists use many types of models to represent things that are hard to visualize, such as the structure and materials used in the construction of a building and the computer model of the airplane shown in **Figure 5**. Chemists also use several different types of models to represent matter, as you will soon learn.

Get It?
Identify two additional types of models that are used by scientists.



Office building model

Airplane model

Figure 5 Scientists use models to visualize complex ideas, such as the materials and structure used to build office buildings. They also use models to test a concept, such as a new airplane design, before it is mass produced.

Infer *why chemists use models to study atoms.*

(l)John Lund/Marc Romanelli/Blend Images/Alamy; (r)NASA/Science Source

You will learn in Module 2 that the tremendous variety of stuff around you can be broken down into more than a hundred types of matter called elements, and that elements are made up of particles called atoms. Atoms are so tiny that they cannot be seen even with optical microscopes. Thus, atoms are submicroscopic. They are so small that over a trillion atoms could fit onto the period at the end of this sentence. The structure, composition, and behavior of all matter can be explained on a submicroscopic level—or the atomic level. All that we observe about matter depends on atoms and the changes they undergo.

Check Your Progress

Summary

- Chemistry is the study of matter. There are several branches of chemistry, including organic, inorganic, physical, analytical, and biochemistry.
- Science is the use of evidence to construct testable explanations and predictions of natural phenomena, as well as, the knowledge generated through this process.
- Models are tools that scientists, including chemists, use. A hypothesis is a testable explanation of a situation or phenomena. A theory is a hypothesis that is supported by many experiments.

Demonstrate Understanding

1. **Explain** why the study of chemistry should be important to everyone.
2. **Define** *substance* and give two examples of things that are substances.
3. **Explain** why there are different branches of chemistry.
4. **Explain** why scientists use mass instead of weight for their measurements.
5. **Summarize** why it is important for chemists to study changes in the world at a submicroscopic level.
6. **Infer** why chemists use models to study submicroscopic matter.
7. **Identify** three models that scientists use, and explain why each model is useful.
8. **Evaluate** How would your mass and weight differ on the Moon? The gravitational force of the Moon is one-sixth the gravitational force on Earth.
9. **Evaluate** If you put a scale in an elevator and weigh yourself as you ascend and then descend, does the scale have the same reading in both instances? Explain your answer.
10. **Distinguish** Jacques Charles described the direct relationship between temperature and volume of all gases at constant pressure. Should this be called Charles's law or Charles's theory? Explain.

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LESSON 2

MEASUREMENT

FOCUS QUESTION

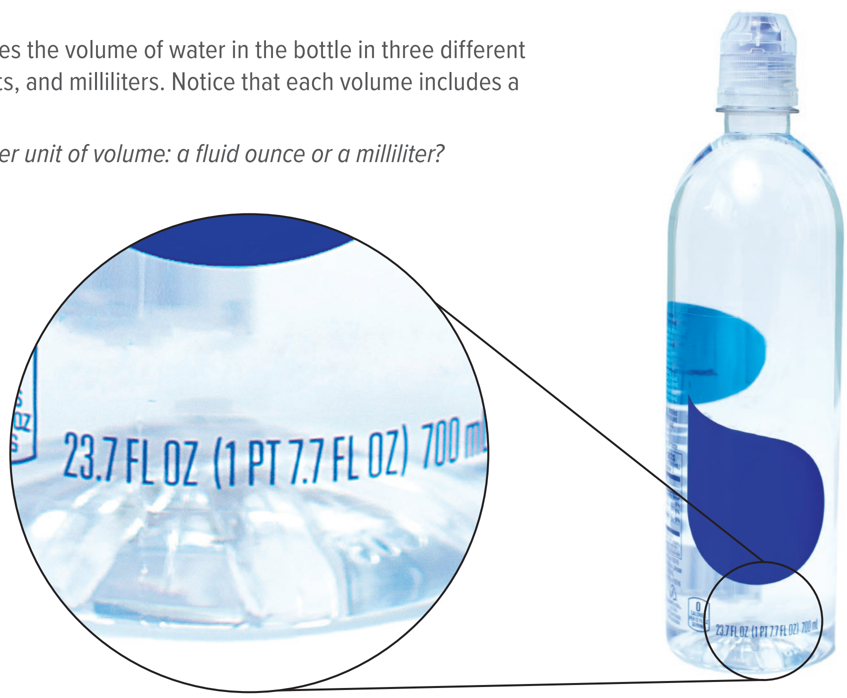
Why do scientists use standardized units?

Units

You use measurements almost every day. For example, reading the bottled water label in **Figure 6** helps you decide what size bottle to buy. Notice that the label uses a number and a unit, such as 700 mL, to give the volume. The label also gives the volume as 23.7 fluid ounces. Fluid ounces, pints, and milliliters are units used to measure volume.

Figure 6 The label gives the volume of water in the bottle in three different units: fluid ounces, pints, and milliliters. Notice that each volume includes a number and a unit.

Infer *Which is the larger unit of volume: a fluid ounce or a milliliter?*



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: Laboratory Techniques and Lab Safety
Plan and carry out an investigation to determine the quantity and scale of the mass of a substance and the volume of water.

Virtual Investigation: Density
Use models to identify patterns in the densities of several known and unknown solids and liquids.

Système Internationale d'Unités

For centuries, units of measurement were not exact. A person might measure distance by counting steps, or measure time using a sundial or an hourglass filled with sand. Such estimates worked for ordinary tasks. Because scientists need to report data that can be reproduced by other scientists, they need standard units of measurement. In 1960, an international committee of scientists met to update the existing metric system. The revised international unit system is called the Système Internationale d'Unités, which is abbreviated SI.

Table 2 SI Base Units

Quantity	Base Unit
Time	second (s)
Length	meter (m)
Mass	kilogram (kg)
Temperature	kelvin (K)
Amount of a substance	mole (mol)
Electric current	ampere (A)
Luminous intensity	candela (cd)

Base Units and SI Prefixes

There are seven base units in SI. A **base unit** is a defined unit in a system of measurement that is based on an object or event in the physical world. A base unit is independent of other units. **Table 2** lists the seven SI base units, the quantities they measure, and their abbreviations. Some familiar quantities that are expressed in base units are time, length, mass, and temperature.

To better describe the range of possible measurements, scientists add prefixes to the base units. This task is made easier because the metric system is a decimal system—a system based on units of 10. The prefixes in **Table 3** are based on factors of ten and can be used with all SI units. For example, the prefix *kilo-* means one thousand; therefore, 1 km equals 1000 m. Similarly, the prefix *milli-* means one-thousandth; therefore, 1 mm equals 0.001 m. Many mechanical pencils use lead that is 0.5 mm in diameter. How much of a meter is 0.5 mm?

Table 3 SI Prefixes

Prefix	Symbol	Numerical Value in Base Units	Power of 10 Equivalent
Giga	G	1,000,000,000	10^{9}
Mega	M	1,000,000	10^{6}
Kilo	k	1000	10^{3}
–	–	1	10^{0}
Deci	d	0.1	10^{-1}
Centi	c	0.01	10^{-2}
Milli	m	0.001	10^{-3}
Micro	μ	0.000001	10^{-6}
Nano	n	0.000000001	10^{-9}
Pico	p	0.000000000001	10^{-12}

Time

The SI base unit for time is the **second** (s). The physical standard used to define the second is the frequency of the radiation given off by a cesium-133 atom. Cesium-based clocks are used when highly accurate timekeeping is required. For everyday tasks, a second seems like a short amount of time. In chemistry, however, many chemical reactions take place within a fraction of a second.

Length

The SI base unit for length is the **meter** (m). A meter is the distance that light travels in a vacuum in 1/299,792,458 of a second. A vacuum exists where space contains no matter.

A meter is close in length to a yard and is useful for measuring the length and width of a small area, such as a room. For larger distances, such as between cities, you would use kilometers. Smaller lengths, such as the diameter of a pencil, are likely to be given in millimeters.

Mass

Recall that mass is a measure of the amount of matter an object contains. The SI base unit for mass is the **kilogram** (kg). Currently, a platinum and iridium cylinder kept in France defines the kilogram. The cylinder is stored in a vacuum under a triple bell jar to prevent the cylinder from oxidizing. As shown in **Figure 7**, scientists are working to redefine the kilogram using basic properties of nature.

A kilogram is equal to about 2.2 pounds. Because the masses measured in most laboratories are much smaller than a kilogram, scientists often measure quantities in grams (g) or milligrams (mg). For example, a laboratory experiment might ask you to add 35 mg of an unknown substance to 350 g of water. When working with mass values, it is helpful to remember that there are 1000 g in a kilogram. How many milligrams are in a gram?

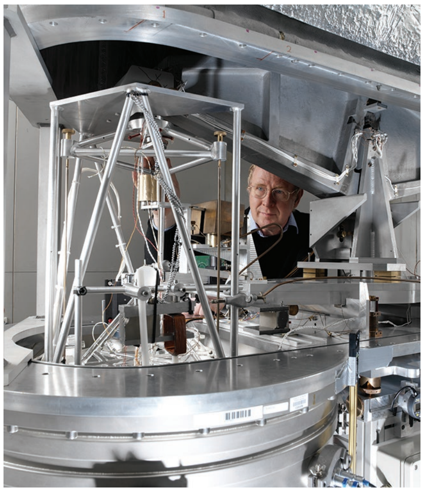


Figure 7 Scientists at the National Institute of Standards and Technology are experimenting with redefining the kilogram using an apparatus known as a watt balance. The watt balance uses electric current and a magnetic field to measure the force required to balance a one-kilogram mass against the force of gravity. Other scientists are counting the number of atoms in a one-kilogram mass to redefine the kilogram.

Andrew Brookes, National Physical Laboratory/Science Source

SCIENCE USAGE V. COMMON USAGE

meter

Science usage: the SI base unit of length

The metal rod was 1 m in length.

Common usage: a device used to measure

The time ran out on the parking meter.

Temperature

People often use qualitative descriptions, such as hot and cold, when describing the weather or the water in a swimming pool. Temperature, however, is a quantitative measurement of the average kinetic energy of the particles that make up an object. As the particle motion in an object increases, so does the temperature of the object.

Measuring temperature requires a thermometer or a temperature probe. A thermometer consists of a narrow tube that contains a liquid. The height of the liquid indicates the temperature. A change in temperature causes a change in the volume of the liquid, which results in a change in the height of the liquid in the tube. Electronic temperature probes make use of thermocouples. A thermocouple produces an electric current that can be calibrated to indicate temperature.

Several different temperature scales have been developed. Three temperature scales—Kelvin, Celsius, and Fahrenheit—are commonly used to describe how hot or cold an object is.

Fahrenheit In the United States, the Fahrenheit scale is used to measure temperature. German scientist Gabriel Daniel Fahrenheit devised the scale in 1724. On the Fahrenheit scale, water freezes at 32°F and boils at 212°F.

Celsius Another temperature scale, the Celsius scale, is used throughout much of the rest of the world. Anders Celsius, a Swedish astronomer, devised the Celsius scale. The scale is based on the freezing and boiling points of water. He defined the freezing point of water as 0 and the boiling point of water as 100. He then divided the distance between these two fixed points into 100 equal units, or degrees. To convert from degrees Celsius (°C) to degrees Fahrenheit (°F), you can use the following equation.

$$^\circ\text{F} = 1.8(^\circ\text{C}) + 32$$

Imagine a friend from Canada calls you and says that it is 35°C outside. What is the temperature in degrees Fahrenheit? To convert to degrees Fahrenheit, substitute 35°C into the above equation and solve.

$$1.8(35) + 32 = 95^\circ\text{F}$$

If it is 35°F outside, what is the temperature in degrees Celsius?

$$\frac{35^\circ\text{F} - 32}{1.8} = 1.7^\circ\text{C}$$

Get It?

Infer Which is warmer, 25°F or 25°C?

Kelvin The SI base unit for temperature is the **kelvin** (K). The Kelvin scale was devised by a Scottish physicist and mathematician, William Thomson, who was known as Lord Kelvin. Zero kelvin is a point where all particles are at their lowest possible energy state. On the Kelvin scale, water freezes at 273.15 K and boils at 373.15 K. Later, you will learn why scientists use the Kelvin scale to describe properties of a gas.

Figure 8 compares the Celsius and Kelvin scales. It is easy to convert between the Celsius scale and the Kelvin scale using the following equation.

Kelvin-Celsius Conversion Equation

$K = °C + 273$

K represents temperature in kelvins.
°C represents temperature in degrees Celsius.

Temperature in kelvins is equal to temperature in degrees Celsius plus 273.

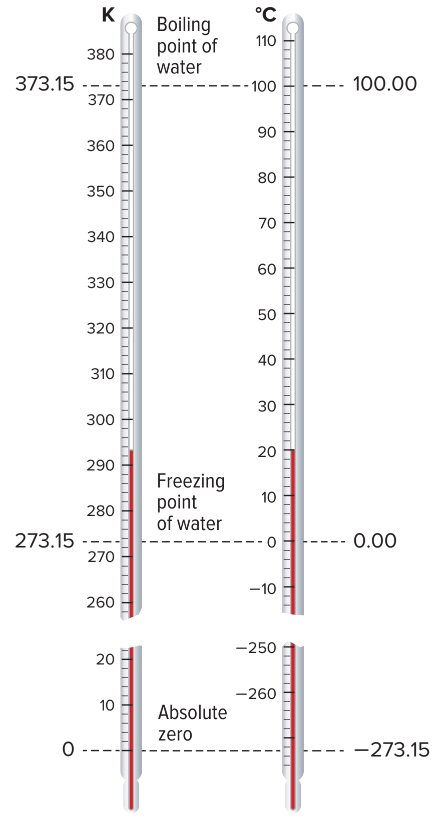


Figure 8 A change of 1 K on the Kelvin scale is equal in size to a change of 1°C on the Celsius scale. Notice also that the degree sign (°) is not used with the Kelvin scale.

As shown by the equation above, to convert temperatures reported in degrees Celsius to kelvins, you simply add 273. For example, consider the element mercury, which melts at −39°C. What is this temperature in kelvins?

$$-39°C + 273 = 234\ K$$

To convert from kelvins to degrees Celsius, just subtract 273.

For example, consider the element bromine, which melts at 266 K. What is this temperature in degrees Celsius?

$$266\ K - 273 = -7°C$$

You will use these conversions frequently throughout chemistry, especially when you study how gases behave. The gas laws you will learn are based on kelvin temperatures.

Derived Units

Not all quantities can be measured with SI base units. For example, the SI unit for speed is meters per second (m/s). Notice that meters per second includes two SI base units—the meter and the second. A unit that is defined by a combination of base units is called a **derived unit.** Two other quantities that are measured in derived units are volume (cm^3) and density (g/cm^3).

Volume

Volume is the space occupied by an object. The volume of an object with a cubic or rectangular shape can be determined by multiplying its length, width, and height dimensions. When each dimension is given in meters, the calculated volume has units of cubic meters (m^3). In fact, the derived SI unit for volume is the cubic meter.

It is easy to visualize a cubic meter; imagine a large cube whose sides are each 1 m in length. The volume of an irregularly shaped solid can be determined using the water displacement method.

The cubic meter is a large volume that is difficult to work with. For everyday use, a more useful unit of volume is the liter. A **liter** (L) is equal to one cubic decimeter (dm^3), that is, 1 L equals 1 dm^3. Liters are commonly used to measure the volume of water and beverage containers. One liter has about the same volume as one quart.

For smaller quantities of liquids in the laboratory, volume is often measured in cubic centimeters (cm^3) or milliliters (mL). A milliliter and a cubic centimeter are equal in size.

$$1 \text{ mL} = 1 \text{ cm}^3$$

Recall that the prefix *milli-* means one-thousandth. Therefore, one milliliter is equal to one-thousandth of a liter. In other words, there are 1000 mL in 1 L.

$$1 \text{ L} = 1000 \text{ mL}$$

Figure 9 shows the relationships among several different SI units of volume.

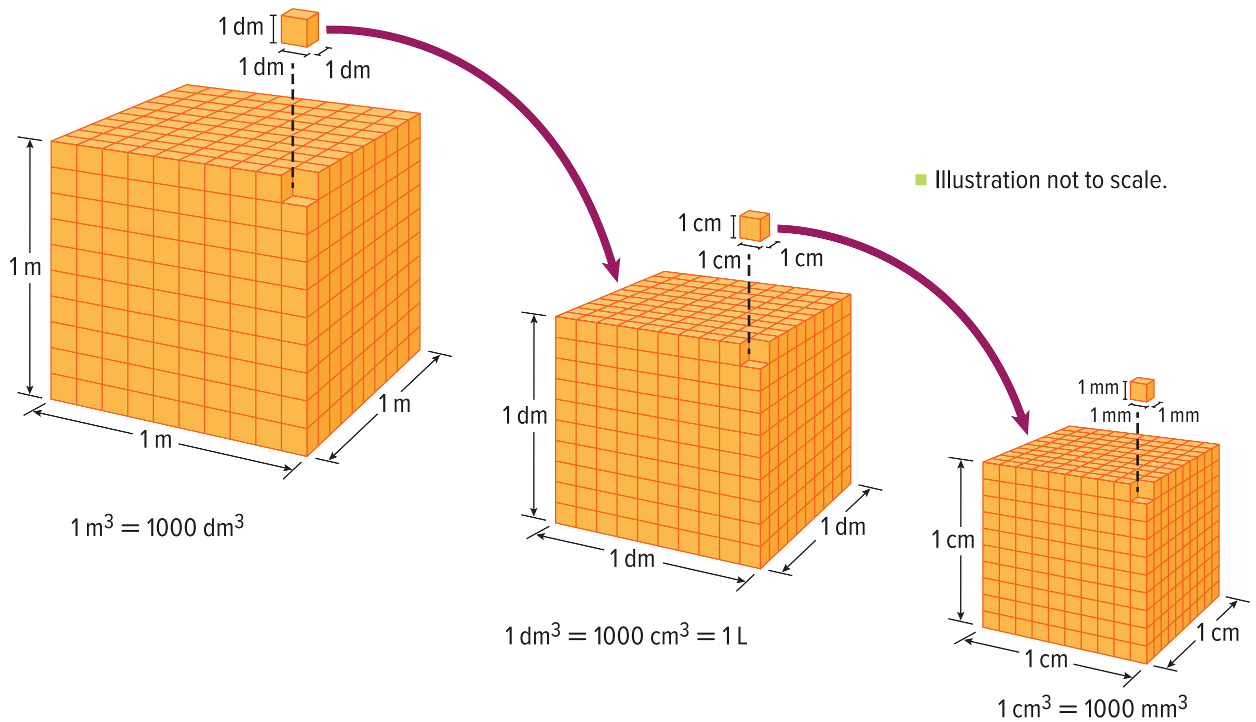


Figure 9 The three cubes show volume relationships between cubic meters (m^3), cubic decimeters (dm^3), cubic centimeters (cm^3), and cubic millimeters (mm^3). As you move from left to right, the volume of each cube gets $10 \times 10 \times 10$, or 1000 times, smaller.

Interpret *How many cubic centimeters (cm^3) are in 1 L?*

Get It?

Observe the photo at the beginning of this module. Explain why it is important for architects and builders to pay close attention to units when designing and constructing a new building.

Figure 10 The grape and the foam have the same mass, but the grape occupies less volume. The grape's density must be greater than the foam's.

Interpret *How would the masses compare if the volumes were equal?*

Density

You can explain why it is easier to lift a backpack filled with gym clothes than the same backpack filled with books in terms of density—the book-filled backpack contains more mass in the same volume. **Density** is a physical property of matter and is defined as the amount of mass per unit volume. Common units of density are grams per cubic centimeter (g/cm^3) for solids and grams per milliliter (g/mL) for liquids and gases. Another example is shown in **Figure 10**.

Get It?

State the quantities that must be known in order to calculate density.

The density of a substance usually cannot be measured directly. Rather, it is calculated using mass and volume measurements. You can calculate density using the following equation.

Density Equation

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

The density of an object or a sample of matter is equal to its mass divided by its volume.

Because density is a physical property of matter, it can sometimes be used to identify an unknown element. For example, imagine a piece of an unknown metallic element that has a volume of 5.0 cm^3 and a mass of 13.5 g. Substituting these values into the equation for density yields:

$$\text{density} = \frac{13.5 \text{ g}}{5.0 \text{ cm}^3} = 2.7 \text{ g/cm}^3$$

Real-World Chemistry

Liquid Density Measurement

HYDROMETERS A hydrometer is a device that measures the specific gravity (the ratio of the fluid's density to that of water) of a fluid. Fluids of different densities result in different readings. Hydrometers are often used at service stations to diagnose problems with an automobile's battery.

Now go to **Table R–7** in the Student Resources and find a density value that closely matches the calculated value of 2.7 g/cm^3. What is the identity of the unknown element?

Your textbook includes many Example Problems, each of which is solved using a three-step process. Read Example Problem 1 and follow the steps to calculate the mass of an object using density and volume.

THE PROBLEM

1. Read the problem carefully.
2. Be sure that you understand what is being asked.

ANALYZE THE PROBLEM

1. Read the problem again.
2. Identify what you are given, and list the known data. If needed, gather information from graphs, tables, or figures.
3. Identify and list the unknowns.
4. Plan the steps you will follow to find the answer.

SOLVE FOR THE UNKNOWN

1. Determine whether you need a sketch to solve the problem.
2. If the solution is mathematical, write the equation and isolate the unknown factor.
3. Substitute the known quantities into the equation.
4. Solve the equation.
5. Continue the solution process until you solve the problem.

EVALUATE THE ANSWER

1. Reread the problem. Is the answer reasonable?
2. Check your math. Are the units and the significant figures correct? (Refer to Lesson 3.)

EXAMPLE Problem 1

USING DENSITY AND VOLUME TO FIND MASS When a piece of aluminum is placed in a 25-mL graduated cylinder that contains 10.5 mL of water, the water level rises to 13.5 mL. What is the mass of the aluminum?

1 ANALYZE THE PROBLEM

The mass of aluminum is unknown. The known values include the initial and final volumes and the density of aluminum. The volume of the sample equals the volume of water displaced in the graduated cylinder. According to **Table R-7**, the density of aluminum is 2.7 g/mL. Use the density equation to solve for the mass of the aluminum sample.

Known	Unknown
density = 2.7 g/mL	mass = ? g
initial volume = 10.5 mL	
final volume = 13.5 mL	

2 SOLVE FOR THE UNKNOWN

volume of sample = final volume − initial volume	State the equation for volume.
volume of sample = 13.5 mL − 10.5 mL	Substitute final volume = 13.5 mL and initial volume = 10.5 mL.
volume of sample = 3.0 mL	
$\text{density} = \frac{\text{mass}}{\text{volume}}$	State the equation for density.
mass = volume × density	Solve the density equation for mass.
mass = 3.0 mL × 2.7 g/mL	Substitute volume = 3.0 mL and density = 2.7 g/mL.
$\text{mass} = 3.0\ \cancel{\text{mL}} \times 2.7\ \text{g}/\cancel{\text{mL}} = 8.1\ \text{g}$	Multiply, and cancel units.

3 EVALUATE THE ANSWER

Check your answer by using it to calculate the density of aluminum.

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{8.1\ \text{g}}{3.0\ \text{mL}} = 2.7\ \text{g/mL}$$

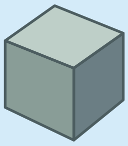
Because the calculated density for aluminum is correct, the mass value must also be correct.

PRACTICE Problems ADDITIONAL PRACTICE

11. Is the cube pictured at right made of pure aluminum? Explain your answer.

12. What is the volume of a sample that has a mass of 20 g and a density of 4 g/mL?

13. **CHALLENGE** A 147-g piece of metal has a density of 7.00 g/mL. A 50-mL graduated cylinder contains 20.0 mL of water. What is the final volume after the metal is added to the graduated cylinder?



Mass = 20 g
Volume = 5 cm^3

EARTH SCIENCE Connection As air at the equator is warmed, the particles in the air move farther apart and the air density decreases. At the poles, the air cools and its density increases as the particles move closer together. When a cooler, denser air mass sinks beneath a rising warm air mass, winds are produced. Weather patterns are created by moving air masses of different densities.

Scientific Notation

The Hope Diamond, which is shown in **Figure 11**, contains approximately 460,000,000,000,000,000,000,000 atoms of carbon. Each of these carbon atoms has a mass of 0.00000000000000000000002 g. If you were to use these numbers to calculate the mass of the Hope Diamond, you would find that the zeros would get in your way. Using a calculator offers no help, as it won't let you enter numbers this large or this small. Numbers such as these are best expressed in scientific notation. Scientists use this method to conveniently restate a number without changing its value.

Figure 11 At more than 45 carats, the Hope Diamond is the world's largest deep-blue diamond. Originally mined in India, the diamond's brilliant blue color is due to trace amounts of boron within the diamond. Diamonds are formed from a unique structure of carbon atoms, creating one of nature's hardest known substances. Note that a carat is a unit of measure used for gemstones (1 carat = 200 mg).

Scientific notation can be used to express any number as a number between 1 and 10 (known as the coefficient) multiplied by 10 raised to a power (known as the exponent). When written in scientific notation, the two numbers above appear as follows.

$$\text{carbon atoms in the Hope Diamond} = \overbrace{4.6}^{\text{Coefficient}} \times 10^{\overbrace{23}^{\text{Exponent}}}$$

$$\text{mass of one carbon atom} = \underbrace{2}_{} \times 10^{\underbrace{-23}_{}}\ \text{g}$$

Let's look at these two numbers more closely. In each case, the number 10 raised to an exponent replaced the zeros that preceded or followed the nonzero numbers. For numbers greater than 1, a positive exponent is used to indicate how many times the coefficient must be multiplied by 10 in order to obtain the original number. Similarly, for numbers less than 1, a negative exponent indicates how many times the coefficient must be divided by 10 in order to obtain the original number. Determining the exponent to use when writing a number in scientific notation is easy: simply count the number of places the decimal point must be moved to give a coefficient between 1 and 10. The number of places moved equals the value of the exponent. The exponent is positive when the decimal moves to the left and the exponent is negative when the decimal moves to the right.

$$460{,}000{,}000{,}000{,}000{,}000{,}000{,}000. \rightarrow 4.6 \times 10^{23}$$

Because the decimal point moves 23 places to the left, the exponent is 23.

$$0.00000000000000000000002 \rightarrow 2 \times 10^{-23}$$

Because the decimal point moves 23 places to the right, the exponent is −23.

EXAMPLE Problem 2

SCIENTIFIC NOTATION Write the following data in scientific notation.

a. The diameter of the Sun is 1,392,000 km.

b. The density of the Sun's lower atmosphere is 0.000000028 g/cm^3.

1 ANALYZE THE PROBLEM

You are given two values, one much larger than 1 and the other much smaller than 1. In both cases, the answers will have a coefficient between 1 and 10 multiplied by a power of 10.

2 SOLVE FOR THE UNKNOWN

Move the decimal point to give a coefficient between 1 and 10. Count the number of places the decimal point moves, and note the direction.

1,392,000.	Move the decimal point six places to the left.
0.000000028	Move the decimal point eight places to the right.
a. 1.392×10^6 km **b.** 2.8×10^{-8} g/cm^3	Write the coefficients, and multiply them by 10^n where n equals the number of places moved. When the decimal point moves to the left, n is positive; when the decimal point moves to the right, n is negative. Add units to the answers.

3 EVALUATE THE ANSWER

The answers are correctly written as a coefficient between 1 and 10 multiplied by a power of 10. Because the diameter of the Sun is a number greater than 1, its exponent is positive. Because the density of the Sun's lower atmosphere is a number less than 1, its exponent is negative.

PRACTICE Problems ADDITIONAL PRACTICE

14. Express each number in scientific notation.

a. 700 **c.** 4,500,000 **e.** 0.0054 **g.** 0.000000076

b. 38,000 **d.** 685,000,000,000 **f.** 0.00000687 **h.** 0.0000000008

15. CHALLENGE Express each quantity in regular notation along with its appropriate unit.

a. 3.60×10^5 s **c.** 5.060×10^3 km

b. 5.4×10^{-5} g/cm^3 **d.** 8.9×10^{10} Hz

Addition and subtraction

In order to add or subtract numbers written in scientific notation, the exponents must be the same. Suppose you need to add 7.35×10^2 m and 2.43×10^2 m. Because the exponents are the same, you can simply add the coefficients.

$$(7.35 \times 10^2 \text{ m}) + (2.43 \times 10^2 \text{ m}) = 9.78 \times 10^2 \text{ m}$$

How do you add numbers in scientific notation when they have different exponents? Consider the amounts of energy produced by renewable energy sources in the United States, such as the wind-powered turbines in **Figure 12.** In 2008, the energy production amounts from renewable sources were as follows.

Hydroelectric	2.643×10^{18} J*
Biomass	4.042×10^{18} J
Geothermal	3.89×10^{17} J
Wind	5.44×10^{17} J
Solar	7.8×10^{16} J

* J stands for joules, a unit of energy.

Figure 12 The uneven heating of Earth's surface causes wind, which powers these turbines and generates electricity.

To add these values, you must rewrite them with the same exponent. The two largest values have an exponent of 10^{18}, so convert the other numbers to values with this exponent. These other exponents must increase to become 10^{18}. As you learned earlier, each place the decimal shifts to the left increases the exponent by 1. Rewriting the values with exponents of 10^{18} and adding yields the following.

Hydroelectric	2.643×10^{18} J
Biomass	4.042×10^{18} J
Geothermal	0.389×10^{18} J
Wind	0.544×10^{18} J
Solar	0.078×10^{18} J
Total	7.696×10^{18} J

PRACTICE Problems ADDITIONAL PRACTICE

16. Solve each problem, and express the answer in scientific notation.

a. $(5 \times 10^{-5}) + (2 \times 10^{-5})$

b. $(7 \times 10^{8}) - (4 \times 10^{8})$

c. $(9 \times 10^{2}) - (7 \times 10^{2})$

d. $(4 \times 10^{-12}) + (1 \times 10^{-12})$

17. CHALLENGE Express each answer in scientific notation in the units indicated.

a. $(1.26 \times 10^{4}$ kg$) + (2.5 \times 10^{6}$ g$)$ in kg

b. $(7.06$ g$) + (1.2 \times 10^{-4}$ kg$)$ in kg

c. $(4.39 \times 10^{5}$ kg$) - (2.8 \times 10^{7}$ g$)$ in kg

d. $(5.36 \times 10^{-1}$ kg$) - (7.40 \times 10^{-2}$ kg$)$ in g

Multiplication and division

Multiplying and dividing numbers in scientific notation is a two-step process, but it does not require the exponents to be the same. For multiplication, multiply the coefficients and then add the exponents. For division, divide the coefficients, then subtract the exponent of the divisor from the exponent of the dividend.

Get It?

Restate the process used to multiply two numbers that are expressed in scientific notation.

To calculate the mass of the Hope Diamond, multiply the number of carbon atoms by the mass of a single carbon atom.

$$(4.6 \times 10^{23}\ \cancel{\text{atoms}})(2 \times 10^{-23}\ \text{g}/\cancel{\text{atom}}) = 9.2 \times 10^{0}\ \text{g} = 9.2\ \text{g}$$

Note that any number raised to a power of 0 is equal to 1; thus, 9.2×10^{0} g is equal to 9.2 g.

EXAMPLE Problem 3

MULTIPLYING AND DIVIDING NUMBERS IN SCIENTIFIC NOTATION Solve the following problems.

a. $(2 \times 10^{3}) \times (3 \times 10^{2})$

b. $(9 \times 10^{8}) \div (3 \times 10^{-4})$

1 ANALYZE THE PROBLEM

You are given numbers written in scientific notation to multiply and divide. For the multiplication problem, multiply the coefficients and add the exponents. For the division problem, divide the coefficients and subtract the exponent of the divisor from the exponent of the dividend.

$$\frac{9 \times 10^{8}}{3 \times 10^{-4}}$$

The exponent of the dividend is 8.

The exponent of the divisor is −4.

EXAMPLE Problem 3 (continued)

2 SOLVE FOR THE UNKNOWN

a. $(2 \times 10^3) \times (3 \times 10^2)$ — State the problem.

$2 \times 3 = 6$ — Multiply the coefficients.

$3 + 2 = 5$ — Add the exponents.

6×10^5 — Combine the parts.

b. $(9 \times 10^8) \div (3 \times 10^{-4})$ — State the problem.

$9 \div 3 = 3$ — Divide the coefficients.

$8 - (-4) = 8 + 4 = 12$ — Subtract the exponents.

3×10^{12} — Combine the parts.

3 EVALUATE THE ANSWER

To test the answers, write out the original data and carry out the arithmetic. For example, Problem **a** becomes $2000 \times 300 = 600{,}000$, which is the same as 6×10^5.

PRACTICE Problems — ADDITIONAL PRACTICE

18. Solve each problem, and express the answer in scientific notation.

a. $(4 \times 10^2) \times (1 \times 10^8)$

b. $(2 \times 10^{-4}) \times (3 \times 10^2)$

c. $(6 \times 10^2) \div (2 \times 10^1)$

d. $(8 \times 10^4) \div (4 \times 10^1)$

19. CHALLENGE Calculate the areas and densities. Report the answers in the correct units.

a. the area of a rectangle with sides measuring 3×10^1 cm and 3×10^{-2} cm

b. the area of a rectangle with sides measuring 1×10^3 cm and 5×10^{-1} cm

c. the density of a substance having a mass of 9×10^5 g and a volume of 3×10^{-1} cm^3

d. the density of a substance having a mass of 4×10^{-3} g and a volume of 2×10^{-2} cm^3

Dimensional Analysis

When planning a pizza party for a group of people, you might want to use dimensional analysis to figure out how many pizzas to order. **Dimensional analysis** is a systematic approach to problem solving that uses conversion factors to move, or convert, from one unit to another. A **conversion factor** is a ratio of equivalent values having different units. How many pizzas do you need to order if 32 people will attend a party, each person eats 3 slices of pizza, and each pizza has 8 slices? **Figure 13** shows how conversion factors are used to calculate the number of pizzas needed for the party.

Blend Images/Moxie Productions/Getty Images

Figure 13 Dimensional analysis can be used to calculate the number of pizzas that must be ordered for a party. How many pizzas will you need if 32 people eat 3 slices per person and there are 8 slices in each pizza?

$$(32\ \cancel{\text{people}})\left(\frac{3\ \cancel{\text{slices}}}{\cancel{\text{person}}}\right)\left(\frac{1\ \text{pizza}}{8\ \cancel{\text{slices}}}\right) = 12\ \text{pizzas}$$

Writing conversion factors

As you just read, conversion factors are ratios of equivalent values. Not surprisingly, these conversion factors are derived from equality relationships, such as 12 eggs = 1 dozen eggs, or 12 inches = 1 foot. Multiplying a quantity by a conversion factor changes the units of the quantity without changing its value.

Most conversion factors are written from relationships between units. For example, the prefixes in **Table 3** are the source of many conversion factors. From the relationship 1000 m = 1 km, the following conversion factors can be written.

$$\frac{1\text{ km}}{1000\text{ m}} \quad \text{and} \quad \frac{1000\text{ m}}{1\text{ km}}$$

A derived unit, such as a density of 2.5 g/mL, can also be used as a conversion factor. The value shows that 1 mL of the substance has a mass of 2.5 g. The following two conversion factors can be written.

$$\frac{2.5\text{ g}}{1\text{ mL}} \quad \text{and} \quad \frac{1\text{ mL}}{2.5\text{ g}}$$

Percentages can also be used as conversion factors. A percentage is a ratio; it relates the number of parts of one component to 100 total parts. For example, a fruit drink containing 10% sugar by mass contains 10 g of sugar in every 100 g of fruit drink. The conversion factors for the fruit drink are as follows.

$$\frac{10\text{ g sugar}}{100\text{ g fruit drink}} \quad \text{and} \quad \frac{100\text{ g fruit drink}}{10\text{ g sugar}}$$

PRACTICE Problems ADDITIONAL PRACTICE

20. Write two conversion factors for each of the following.

- **a.** a 16% (by mass) salt solution
- **b.** a density of 1.25 g/mL
- **c.** a speed of 25 m/s

21. CHALLENGE What conversion factors are needed to convert:

- **a.** nanometers to meters?
- **b.** density given in g/cm^3 to a value in kg/m^3?

Using conversion factors

A conversion factor used in dimensional analysis must accomplish two things: it must cancel one unit and introduce a new one. While working through a solution, all of the units except the desired unit must cancel. Suppose you want to know how many meters there are in 48 km. The relationship between kilometers and meters is 1 km = 1000 m. The conversion factors are as follows.

$$\frac{1\text{ km}}{1000\text{ m}} \quad \text{and} \quad \frac{1000\text{ m}}{1\text{ km}}$$

To convert km to m, use the conversion factor that causes the km unit to cancel.

$$48\ \cancel{\text{km}} \times \frac{1000\text{ m}}{1\ \cancel{\text{km}}} = 48{,}000\text{ m}$$

When converting a value with a large unit, such as km, to a value with a smaller unit, such as m, the numerical value increases. For example, 48 km (a value with a large unit) converts to 48,000 m (a larger numerical value with a smaller unit). **Figure 14** illustrates the connection between the numerical value and the size of the unit for a conversion factor.

$$\frac{1\text{ km}}{1000\text{ m}}$$

Figure 14 The two quantities shown above are equivalent; that is, 1 km = 1000 m. Note that a smaller numerical value (1) accompanies the larger unit (km), and a larger numerical value (1000) accompanies the smaller unit (m).

Now consider this question: How many eight-packs of water would you need if the 32 people attending your party each had two bottles of water? To solve the problem, identify the given quantities and the desired result. There are 32 people and each of them drinks two bottles of water. The desired result is the number of eight-packs. Using dimensional analysis yields the following.

$$32\ \cancel{\text{people}} \times \frac{2\ \cancel{\text{bottles}}}{\cancel{\text{person}}} \times \frac{1\text{ eight-pack}}{8\ \cancel{\text{bottles}}} = 8\text{ eight-packs}$$

EXAMPLE Problem 4

USING CONVERSION FACTORS In ancient Egypt, small distances were measured in Egyptian cubits. An Egyptian cubit was equal to 7 palms, and 1 palm was equal to 4 fingers. If 1 finger was equal to 18.75 mm, convert 6 Egyptian cubits to meters.

1 ANALYZE THE PROBLEM

A length of 6 Egyptian cubits needs to be converted to meters.

Known

length = 6 Egyptian cubits

7 palms = 1 cubit

1 palm = 4 fingers

1 finger = 18.75 mm

1 m = 1000 mm

Unknown

length = ? m

2 SOLVE FOR THE UNKNOWN

Use dimensional analysis to convert the units in the following order.

cubits → palms → fingers → millimeters → meters

$$6\text{ cubits} \times \frac{7\text{ palms}}{1\text{ cubit}} \times \frac{4\text{ fingers}}{1\text{ palm}} \times \frac{18.75\text{ mm}}{1\text{ finger}} \times \frac{1\text{ meter}}{1000\text{ mm}} = ?\text{ m}$$

Multiply by a series of conversion factors that cancels all the units except meter, the desired unit.

$$6\ \cancel{\text{cubits}} \times \frac{7\ \cancel{\text{palms}}}{1\ \cancel{\text{cubit}}} \times \frac{4\ \cancel{\text{fingers}}}{1\ \cancel{\text{palm}}} \times \frac{18.75\ \cancel{\text{mm}}}{1\ \cancel{\text{finger}}} \times \frac{1\text{ meter}}{1000\ \cancel{\text{mm}}} = 3.150\text{ m}$$

Multiply and divide the numbers as indicated, and cancel the units.

3 EVALUATE THE ANSWER

Each conversion factor is a correct restatement of the original relationship, and all units except for the desired unit, meters, cancel.

PRACTICE Problems

ADDITIONAL PRACTICE

Use Table 3 to solve each of the following.

22. **a.** Convert 360 s to ms.

b. Convert 4800 g to kg.

c. Convert 5600 dm to m.

d. Convert 72 g to mg.

e. Convert 2.45×10^2 ms to s.

f. Convert 5 μm to km.

g. Convert 6.800×10^3 cm to km.

h. Convert 2.5×10^1 kg to Mg.

23. **CHALLENGE** Write the conversion factors needed to determine the number of seconds in one year.

PRACTICE Problems | ADDITIONAL PRACTICE

24. The speedometer at right displays a car's speed in miles per hour. What is the car's speed in km/h? (1 km = 0.62 mile)

25. How many seconds are in 24 h?

26. **CHALLENGE** Vinegar is 5.00% acetic acid by mass and has a density of 1.02 g/mL. What mass of acetic acid, in grams, is present in 185 mL of vinegar?

Check Your Progress

Summary

- SI measurement units and prefixes allow scientists to report data easily.
- Density is a ratio of mass to volume and can be used to identify a substance.
- A number expressed in scientific notation is written as a coefficient between 1 and 10 multiplied by 10 raised to a power.
- Dimensional analysis uses conversion factors to solve problems.

Demonstrate Understanding

27. **Define** the SI units for length, mass, time, and temperature.
28. **Describe** how adding the prefix *mega-* to a unit affects the quantity being described.
29. **Compare** a base unit and a derived unit, and list the derived units used for density and volume.
30. **Calculate** Samples A, B, and C have masses of 80 g, 12 g, and 33 g, and volumes of 20 mL, 4 cm^3, and 11 mL, respectively. Which of the samples have the same density?
31. **Design** a concept map that shows the relationships among the following terms: volume, derived unit, mass, base unit, time, and length.
32. **Describe** how scientific notation makes it easier to work with very large or very small numbers.
33. **Write** a conversion factor relating cubic centimeters and milliliters.
34. **Explain** how dimensional analysis is used to solve problems.
35. **Apply Concepts** A classmate converts 68 km to meters and gets 0.068 m as the answer. Explain why this answer is incorrect, and identify the likely source of the error.
36. **Organize** Create a flowchart that outlines when to use dimensional analysis and when to use scientific notation.

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LESSON 3

UNCERTAINTY IN DATA

FOCUS QUESTION

Why are significant figures important?

Accuracy and Precision

To a scientist, the terms accuracy and precision have very different meanings. **Accuracy** refers to how close a measured value is to an accepted value. **Precision** refers to how close a series of measurements are to one another. The archery target in **Figure 15** illustrates the difference between accuracy and precision. For example, you measure in the lab the mass of an object three times. The arrows represent each measurement, and the center of the target is the accepted value.

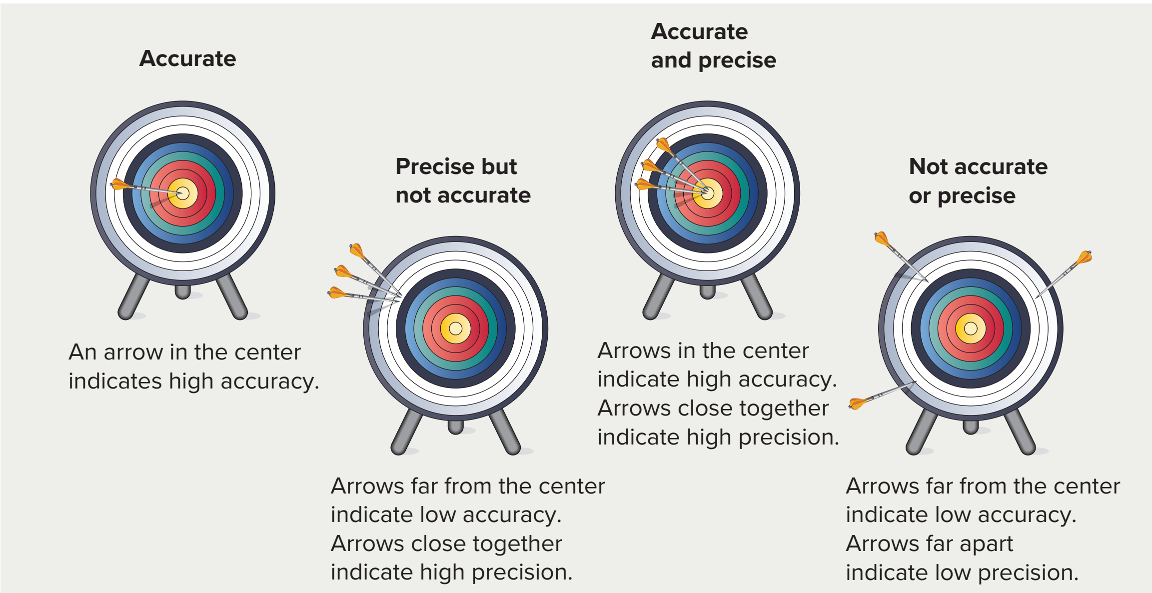

Figure 15 An archery target illustrates the difference between accuracy and precision. An accurate shot is located near the bull's-eye; precise shots are grouped closely together.

Apply *Why is it important to measure the same data more than once?*

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: Use Density to Date a Coin
Plan and carry out an investigation to predict the patterns in mass and volume of pre-1982 or post-1982 pennies.

Laboratory: Effective Use of a Bunsen Burner
Interpret data to predict the effect of beaker distance from a Bunsen burner for efficient heating.

Table 4 Student Density and Error Data (Unknown was sucrose; density = 1.59 g/cm^3)

	Student A		Student B		Student C	
	Density	Error (g/cm^3)	Density	Error (g/cm^3)	Density	Error (g/cm^3)
Trial 1	1.54 g/cm^3	−0.05	1.40 g/cm^3	−0.19	1.70 g/cm^3	+0.11
Trial 2	1.60 g/cm^3	+0.01	1.68 g/cm^3	+0.09	a 1.69 g/cm^3	+0.10
Trial 3	1.57 g/cm^3	−0.02	1.45 g/cm^3	−0.14	1.71 g/cm^3	+0.12
Average	b 1.57 g/cm^3		1.51 g/cm^3		1.70 g/cm^3	

a These trial values are the most precise.

b This average is the most accurate.

Consider the data in **Table 4**. Students were asked to find the density of an unknown white powder. Each student measured the volume and mass of three separate samples. They reported calculated densities for each trial and an average of the three calculations. The powder, sucrose (table sugar), has a density of 1.59 g/cm^3. Which student collected the most accurate data? Who collected the most precise data? Student A's measurements are the most accurate because they are closest to the accepted value of 1.59 g/cm^3. Student C's measurements are the most precise because they are the closest to one another.

Recall that precise measurements might not be accurate. Looking at just the average of the densities can be misleading. Based solely on the average, Student B appears to have collected fairly reliable data. However, on closer inspection, Student B's data are neither accurate nor precise. The data are not close to the accepted value, nor are they close to one another.

Error and percent error

The density values reported in **Table 4** are experimental values, which means they are values measured during an experiment. The known density of sucrose is an accepted value, which is a value that is considered true. To evaluate the accuracy of experimental data, you can compare how close the experimental value is to the accepted value. **Error** is defined as the difference between an experimental value and an accepted value. The errors for the experimental density values are also given in **Table 4**.

Error Equation

$$\text{error} = \text{experimental value} - \text{accepted value}$$

The error associated with an experimental value is the difference between the experimental value and the accepted value.

Scientists often want to know what percent of the accepted value an error represents. **Percent error** expresses error as a percentage of the accepted value.

Percent Error Equation

$$\text{percent error} = \frac{|\text{error}|}{\text{accepted value}} \times 100$$

The percent error of an experimental value equals the absolute value of its error divided by the accepted value, multiplied by 100.

Figure 16 This digital caliper is being used to check the size of a nut to one-hundredth of a millimeter (0.01 mm). Skill is required to correctly position the part in the caliper. Experienced machinists will obtain more precise and more accurate readings than inexperienced machinists.

Notice that the percent-error equation uses the absolute value of the error. This is because only the size of the error matters; it does not matter whether the experimental value is larger or smaller than the accepted value.

Get It?

Summarize why error is important.

Percent error is an important concept for the machinist who made the nut shown in **Figure 16**. The machinist must check the tolerances of the nut. Tolerances are a narrow range of allowable dimensions based on acceptable amounts of error. If the dimensions of the nut do not fall within the acceptable range—that is, the nut exceeds its tolerances—it will be retooled or possibly discarded.

EXAMPLE Problem 5

CALCULATING PERCENT ERROR Use Student A's density data in **Table 4** to calculate the percent error in each trial. Report your answers to two places after the decimal point.

1 ANALYZE THE PROBLEM

You are given the errors for a set of density calculations. To calculate percent error, you need to know the accepted value for density, the errors, and the equation for percent error.

Known

accepted value for density = 1.59 g/cm^3

errors: −0.05 g/cm^3; 0.01 g/cm^3; −0.02 g/cm^3

Unknown

percent errors = ?

2 SOLVE FOR THE UNKNOWN

$$\text{percent error} = \frac{|\text{error}|}{\text{accepted value}} \times 100$$

State the percent error equation.

$$\text{percent error} = \frac{|-0.05\ \cancel{g/cm^3}|}{1.59\ \cancel{g/cm^3}} \times 100 = 3.14\%$$

Substitute error = −0.05 g/cm^3, and solve.

$$\text{percent error} = \frac{|0.01\ \cancel{g/cm^3}|}{1.59\ \cancel{g/cm^3}} \times 100 = 0.63\%$$

Substitute error = 0.01 g/cm^3, and solve.

$$\text{percent error} = \frac{|-0.02\ \cancel{g/cm^3}|}{1.59\ \cancel{g/cm^3}} \times 100 = 1.26\%$$

Substitute error = −0.02 g/cm^3, and solve.

EXAMPLE Problem 5 (continued)

3 EVALUATE THE ANSWER

The percent error is greatest for Trial 1, which had the largest error, and smallest for Trial 2, which was closest to the accepted value.

PRACTICE Problems ADDITIONAL PRACTICE

Answer the following questions using data from Table 4.

37. Calculate the percent errors for Student B's trials.

38. Calculate the percent errors for Student C's trials.

39. **CHALLENGE** Based on your calculations in questions 37 and 38, which student's trial was the most accurate? The least accurate?

Significant Figures

Often, precision is limited by the tools available. For example, a digital clock that displays the time as 12:47 or 12:48 can record the time only to the nearest minute. With a stopwatch, however, you might record time to the nearest hundredth second. As scientists have developed better measuring devices, they have been able to make more precise measurements. Of course, for measurements to be both accurate and precise, the measuring devices must be in good working order. Additionally, accurate and precise measurements rely on the skill of the person using the instrument; the user must be trained and use proper techniques.

The precision of a measurement is indicated by the number of digits reported. A value of 3.52 g is more precise than a value of 3.5 g. The reported digits are called significant figures. **Significant figures** include all known digits plus one estimated digit.

Consider the rod in **Figure 17**. The end of the rod falls between 5.2 cm and 5.3 cm. The 5 and 2 are known digits corresponding to marks on the ruler. To these known digits, an estimated digit is added. This last digit estimates the rod's location between the second and third millimeter marks. Because it is an estimate, one person might report the measurement as 5.22 cm and another as 5.23 cm. Either way, the measurement has three significant figures—two known and one estimated.

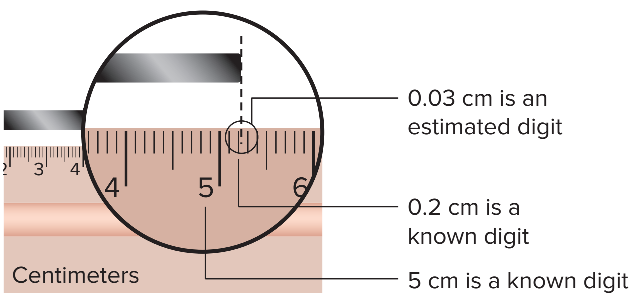

Figure 17 The markings on the ruler represent known digits. The reported measurement includes the known digits plus the estimated digit. The measurement is 5.23 cm.

Infer *What is the estimated digit if the length of an object being measured falls exactly on the 5-cm mark?*

Remember that measurements reported with a lot of significant figures might be precise but not accurate. For example, some chemistry labs have balances that report mass to the nearest hundredth of a gram. If you and each of your classmates measured the same copper cylinder on the same balance, you would probably have a group of very precise measurements. But what if the balance had been previously damaged by an object that was too large for it? Your precise measurements would not be very accurate.

PROBLEM-SOLVING STRATEGY

Recognizing Significant Figures

Learning these five rules for recognizing significant figures will help you when solving problems. Examples of each rule are shown below. Note that each of the highlighted examples has three significant figures.

Rule 1.	Nonzero numbers are always significant.	**72.3** g has three.
Rule 2.	All final zeros to the right of the decimal are significant.	**6.20** g has three.
Rule 3.	Any zero between significant figures is significant.	**60.5** g has three.
Rule 4.	Placeholder zeroes are not significant. To remove placeholder zeros, rewrite the number in scientific notation.	0.0**253** g and **432**0 g (each has three)
Rule 5.	Counting numbers and defined constants have an infinite number of significant figures.	6 molecules 60 s = 1 min

EXAMPLE Problem 6

SIGNIFICANT FIGURES Determine the number of significant figures in the following masses.

a. 0.00040230 g

b. 405,000 kg

1 ANALYZE THE PROBLEM

You are given two measured mass values. Apply the appropriate rules to determine the number of significant figures in each value.

2 SOLVE FOR THE UNKNOWN

Count all nonzero numbers, zeros between nonzero numbers, and final zeros to the right of the decimal place. (Rules 1, 2, and 3)

Ignore zeros that act as placeholders. (Rule 4)

a. 0.000**40230** g has five significant figures.

b. **405**,000 kg has three significant figures.

3 EVALUATE THE ANSWER

One way to verify your answers is to write the values in scientific notation: 4.0230×10^{-4} g and 4.05×10^{5} kg. Without the placeholder zeros, it is clear that 0.00040230 g has five significant figures and that 405,000 kg has three significant figures.

PRACTICE Problems ADDITIONAL PRACTICE

Determine the number of significant figures in each measurement.

40. a. 508.0 L **c.** 1.0200×10^5 kg

b. 820,400.0 L **d.** 807,000 kg

41. a. 0.049450 s **c.** 3.1587×10^{-4} g

b. 0.000482 mL **d.** 0.0084 mL

42. CHALLENGE Write the numbers 10, 100, and 1000 in scientific notation with two, three, and four significant figures, respectively.

Rounding Numbers

Calculators perform flawless arithmetic, but they are not aware of the number of significant figures that should be reported in the answer. For example, a density calculation should not have more significant figures than the original data with the fewest significant figures. To report a value correctly, you often need to round. Consider an object with a mass of 22.44 g and volume of 14.2 cm^3. When you calculate the object's density using a calculator, the displayed answer is 1.5802817 g/cm^3, as shown in **Figure 18**. Because the measured mass had four significant figures and the measured volume had three, it is not correct to report the calculated density value with eight significant figures. Instead, the density must be rounded to three significant figures, or 1.58 g/cm^3.

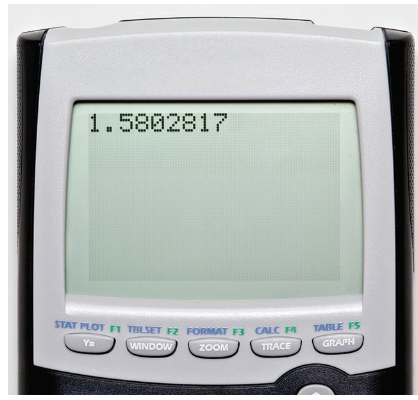

Figure 18 You need to apply the rules of significant figures and rounding to report a calculated value correctly.

Consider the value 3.515014. How would you round this number to five significant figures? To three significant figures? In each case, you need to look at the digit that follows the desired last significant figure. To round to five digits, first identify the fifth significant figure, in this case 0, and then look at the number to its right, in this case 1.

Last significant figure

3.515014

Number to right of last significant figure

Do not change the last significant figure if the digit to its right is less than five. Because a 1 is to the right, the number rounds to 3.5150. If the number had been 5 or greater, you would have rounded up. To round to three digits, identify the third significant figure, in this case 1, and then look at the number to its right, in this case 5.

Last significant figure

3.515014

Number to right of last significant figure

If the digits to the right of the last significant figure are a 5 followed by 0, then look at the last significant figure. If it is odd, round it up; if it is even, do not round up. Because the last significant digit is odd (1), the number rounds up to 3.52.

PROBLEM-SOLVING STRATEGY

Rounding Numbers

Learn these four rules for rounding, and use them when solving problems. Examples of each rule are shown below. Note that each example has three significant figures.

2.532 → 2.53	**Rule 1.**	If the digit to the right of the last significant figure is less than 5, do not change the last significant figure.
2.536 → 2.54	**Rule 2.**	If the digit to the right of the last significant figure is greater than 5, round up the last significant figure.
2.5351 → 2.54	**Rule 3.**	If the digits to the right of the last significant figure are a 5 followed by a nonzero digit, round up the last significant figure.
2.5350 → 2.54 2.5250 → 2.52	**Rule 4.**	If the digits to the right of the last significant figure are a 5 followed by 0 or no other number at all, look at the last significant figure. If it is odd, round it up; if it is even, do not round up.

PRACTICE Problems

ADDITIONAL PRACTICE

43. Round each number to four significant figures.

a. 84,791 kg **c.** 256.75 cm

b. 38.5432 g **d.** 4.9356 m

44. **CHALLENGE** Round each number to four significant figures, and write the answer in scientific notation.

a. 0.00054818 g **c.** 308,659,000 mm

b. 136,758 kg **d.** 2.0145 mL

Addition and subtraction

When you add or subtract measurements, the answer must have the same number of digits to the right of the decimal as the original value having the fewest number of digits to the right of the decimal. For example, the measurements 1.24 mL, 12.4 mL, and 124 mL have two, one, and zero digits to the right of the decimal, respectively. When adding or subtracting, arrange the values so that the decimal points align. Identify the value with the fewest places after the decimal point, and round the answer to that number of places.

Get It?

Determine how many places after the decimal point should the sum of 22.7 m, 5.16 m, 16.287 m, and 124.31 m have?

Multiplication and division

When you multiply or divide numbers, your answer must have the same number of significant figures as the measurement with the fewest significant figures.

EXAMPLE Problem 7

ROUNDING NUMBERS WHEN ADDING A student measured the length of his lab partners' shoes. If the lengths are 28.0 cm, 23.538 cm, and 25.68 cm, what is the total length of the shoes?

1 ANALYZE THE PROBLEM

The three measurements need to be aligned on their decimal points and added. The measurement with the fewest digits after the decimal point is 28.0 cm, with one digit. Thus, the answer must be rounded to only one digit after the decimal point.

2 SOLVE FOR THE UNKNOWN

$$\begin{array}{r} 28.0 \text{ cm} \\ 23.538 \text{ cm} \\ +\ 25.68 \text{ cm} \\ \hline 77.218 \text{ cm} \end{array}$$

Align the measurements and add the values.

The answer is **77.2 cm.**

Round to one place after the decimal; Rule 1 applies.

3 EVALUATE THE ANSWER

The answer, 77.2 cm, has the same precision as the least-precise measurement, 28.0 cm.

PRACTICE Problems

ADDITIONAL PRACTICE

45. Add and subtract as indicated. Round off when necessary.

a. 43.2 cm + 51.0 cm + 48.7 cm

b. 258.3 kg + 257.11 kg + 253 kg

46. **CHALLENGE** Add and subtract as indicated. Round off when necessary.

a. $(4.32 \times 10^3 \text{ cm}) - (1.6 \times 10^4 \text{ mm})$

b. $(2.12 \times 10^7 \text{ mm}) + (1.8 \times 10^3 \text{ cm})$

EXAMPLE Problem 8

ROUNDING NUMBERS WHEN MULTIPLYING Calculate the volume of a book with the following dimensions: length = 28.3 cm, width = 22.2 cm, height = 3.65 cm.

1 ANALYZE THE PROBLEM

Volume is calculated by multiplying length, width, and height. Because all of the measurements have three significant figures, the answer also will.

Known
length = 28.3 cm
width = 22.2 cm
height = 3.65 cm

Unknown
Volume = ? cm^3

2 SOLVE FOR THE UNKNOWN

Calculate the volume, and apply the rules of significant figures and rounding.

Volume = length × width × height

State the formula for the volume of a rectangle.

Volume = 28.3 cm × 22.2 cm × 3.65 cm = 2293.149 cm^3

Substitute values and solve.

Volume = **2290 cm^3**

Round the answer to three significant figures.

EXAMPLE Problem 8 (continued)

3 EVALUATE THE ANSWER

To check if your answer is reasonable, round each measurement to one significant figure and recalculate the volume. Volume = 30 cm × 20 cm × 4 cm = 2400 cm^3. Because this value is close to your calculated value of 2290 cm^3, it is reasonable to conclude the answer is correct.

PRACTICE Problems — ADDITIONAL PRACTICE

Perform the following calculations. Round the answers.

47. **a.** 24 m × 3.26 m **c.** 1.23 m × 2.0 m
b. 120 m × 0.10 m **d.** 53.0 m × 1.53 m

48. **a.** 4.84 m ÷ 2.4 s **c.** 102.4 m ÷ 51.2 s
b. 60.2 m ÷ 20.1 s **d.** 168 m ÷ 58 s

49. **CHALLENGE** $(1.32 \times 10^3\ g) \div (2.5 \times 10^2\ cm^3)$

Check Your Progress

Summary

- An accurate measurement is close to the accepted value. A set of precise measurements shows little variation. The measurement device determines the degree of precision possible.
- Error is the difference between the measured value and the accepted value. Percent error gives the percent deviation from the accepted value.
- The number of significant figures reflects the precision of reported data.
- Calculations are often rounded to the correct number of significant figures.

Demonstrate Understanding

50. **State** how a measured value is reported in terms of known and estimated digits.
51. **Define** *accuracy and precision*.
52. **Identify** the number of significant figures in each of these measurements of an object's length: 76.48 cm, 76.47 cm, and 76.59 cm.
53. **Apply** The object in Question 52 has an actual length of 76.49 cm. Are the measurements in Question 52 accurate? Are they precise?
54. **Calculate** the error and percent error for each measurement in Question 52.
55. **Apply** Write an expression for the quantity 506,000 cm in which it is clear that all the zeros are significant.
56. **Analyze Data** Students collected mass data for a group of coins. The mass of a single coin is 5.00 g. Determine the accuracy and precision of the measurements.

Number of coins	5	10	20	30	50
Mass (g)	23.2	54.5	105.9	154.5	246.2

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LESSON 4
REPRESENTING DATA

FOCUS QUESTION

How can displaying data help you interpret it?

Graphing

A goal of many experiments is to discover whether a pattern exists in a certain situation. When data are listed as shown in **Table 5**, a pattern might not be obvious. However, using data to create a graph can help to reveal a pattern if one exists. A **graph** is a visual display of data.

Circle graphs

The circle graph shown in **Figure 19** is based on the percentage data given in **Table 5**. A circle graph is sometimes called a pie chart because it is divided into wedges that look like a pie. A circle graph is useful for showing parts of a fixed whole. The parts are usually labeled as percents with the whole circle representing 100%.

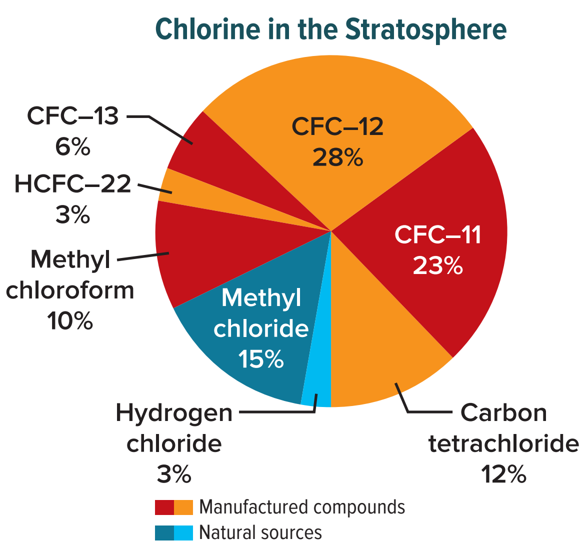

Table 5 Sources of Chlorine in the Stratosphere

Source	Percent
Hydrogen chloride (HCl)	3
Methyl chloride (CH_3Cl)	15
Carbon tetrachloride (CCl_4)	12
Methyl chloroform ($C_2H_3Cl_3$)	10
CFC-11	23
CFC-12	28
CFC-13	6
HCFC-22	3

Figure 19 Although the percentage data presented in the table and the circle graph are basically the same, the circle graph makes it much easier to analyze.

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: Making a Graph

Plan and carry out an investigation to determine the patterns in information graphs provide beyond the initial data.

Review the News

Obtain information from a current news story that uses graphs to convey information. Evaluate your source and communicate your findings to the class.

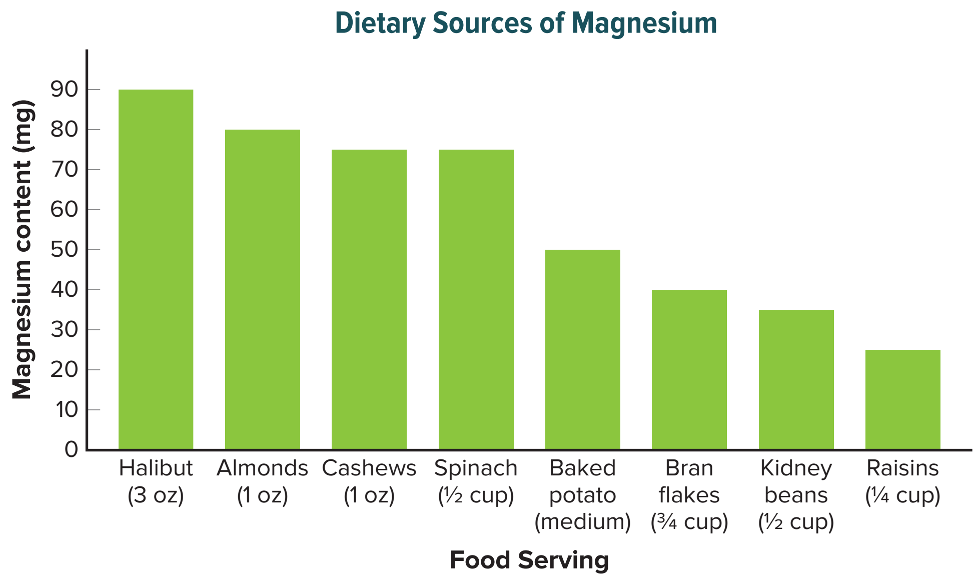

Figure 20 A bar graph is an effective way to present and compare data. This graph shows various dietary sources of the element magnesium. Magnesium plays an important role in the health of your muscles, nerves, and bones.

Bar graphs

A bar graph is often used to show how a quantity varies across categories. Examples of categories include time, location, and temperature. The quantity being measured appears on the vertical axis (y-axis). The independent variable appears on the horizontal axis (x-axis). The relative heights of the bars show how the quantity varies. A bar graph can be used to compare population figures for a single country by decade or the populations of multiple countries at the same point in time. In **Figure 20**, the quantity being measured is magnesium, and the category being varied is food servings. When examining the graph, you can quickly see how the magnesium content varies for these food servings.

Get It?

Interpret Which two food servings provide equal amounts of magnesium?

Line graphs

In chemistry, most graphs that you create and interpret will be line graphs. The points on a line graph represent the intersection of data for two variables—the independent variable and the dependent variable.

STEM CAREER Connection

Graphic Designer

Do you enjoy art as much as you enjoy science? Graphic designers are good at making complicated scientific data more accessible. They can help scientists communicate their work to broad audiences in visually appealing ways. Sometimes designers work by hand, but specialized design software is also commonly used. In general, graphic designers have a bachelor's degree in graphic design or a related field.

Independent and dependent variables The **independent variable** is the variable that a scientist deliberately changes during an experiment. The independent variable is plotted on the *x*-axis. The variable whose value depends on, or changes in response to, the independent variable is called the **dependent variable.** It is plotted on the *y*-axis. In **Figure 21a**, the independent variable is volume and the dependent variable is mass. What are the values for the independent variable and the dependent variable at Point B? **Figure 21b** is a graph of temperature versus elevation. Because the data points do not fit perfectly, the line cannot pass exactly through all of the points. The line must be drawn so that about as many points fall above the line as fall below it. This line is called a best-fit line.

Relationships between variables If the best-fit line for a set of data is straight, there is a linear relationship between the variables and the variables are said to be directly related. The relationship between the variables can be described further by analyzing the steepness, or slope, of the line.

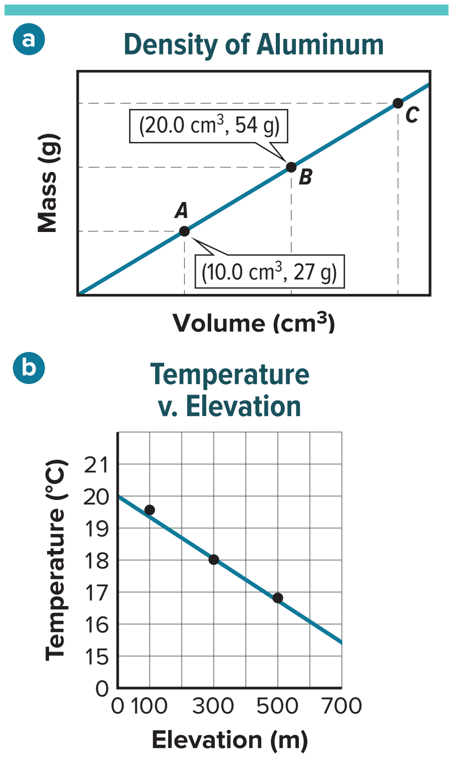

Figure 21 Both of these line graphs show linear relationships. The slope of each line is defined as the ratio of rise over run.

Get It?

Identify the graph that shows a direct relationship.

If the best-fit line rises to the right, then the slope of the line is positive. A positive slope indicates that the dependent variable increases as the independent variable increases. If the best-fit line sinks to the right, then the slope of the line is negative. A negative slope indicates that the dependent variable decreases as the independent variable increases. In either case, the slope of the line is constant. You can use two pairs of data points to calculate the slope of the line. The slope is the rise (Δy) divided by the run (Δx).

Slope Equation

$$\text{slope} = \frac{\text{rise}}{\text{sun}} = \frac{\Delta y}{\Delta x} = \frac{y_2 - y_1}{x_2 - x_1}$$

y_2, y_1, x_2, and x_1 are values from data points (x_1, y_1) and (x_2, y_2).

The slope of a line is equal to the change in *y* divided by the change in *x*.

When the mass of a material is plotted against its volume, the slope of the line represents the material's density. An example of this is shown in **Figure 21a**. To calculate the slope of the line, substitute the *x* and *y* values for Points A and B in the slope equation and solve.

$$\begin{aligned}\text{slope} &= \frac{54\text{ g} - 27\text{ g}}{20.0\text{ cm}^3 - 10.0\text{ cm}^3}\\ &= \frac{27\text{ g}}{10.0\text{ cm}^3}\\ &= 2.7\text{ g/cm}^3\end{aligned}$$

Thus, the slope of the line, and the density, is 2.7 g/cm^3. When the best-fit line is curved, the relationship between the variables is nonlinear. In chemistry, you will study nonlinear relationships called inverse relationships. Refer to the Math Handbook for more discussion of graphs.

Interpreting Graphs

You should use an organized approach when analyzing graphs. First, note the independent and dependent variables. Next, decide if the relationship between the variables is linear or nonlinear. If the relationship is linear, is the slope positive or negative?

Interpolation and extrapolation

When points on a line graph are connected, the data are considered to be continuous. Continuous data allow you to read the value from any point that falls between the recorded data points. This process is called interpolation. For example, from **Figure 21b**, what is the temperature at an elevation of 350 m? To interpolate this value, first locate 350 m on the *x*-axis; it is located halfway between 300 m and 400 m. Project upward until you hit the plotted line, and then project that point horizontally to the left until you reach the *y*-axis. The temperature at 350 m is approximately 17.8°C.

You can also extend a line beyond the plotted points in order to estimate values for the variables. This process is called extrapolation. It is important to be very careful with extrapolation, however, as it can easily lead to errors and result in very inaccurate predictions.

Get It?

Explain why extrapolation might be less reliable than interpolation.

Interpreting ozone data

The value of using graphs to visualize data is illustrated by **Figure 22**. These important ozone measurements were taken at the Halley Research Station in Antarctica. The graph shows how ozone levels vary from August to April. The independent and dependent variables are the month and the total ozone, respectively.

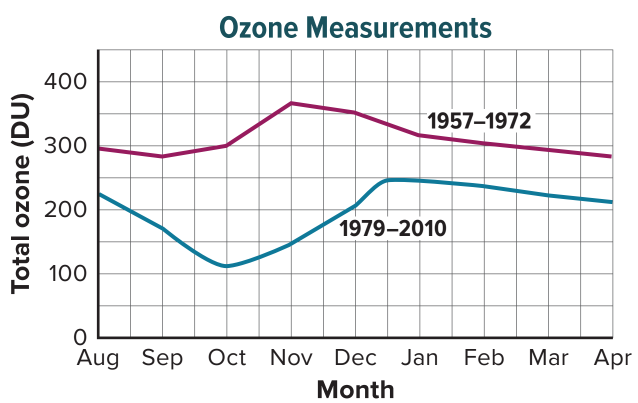

Figure 22 The two lines in this graph represent average ozone levels for two time periods, 1957–1972 and 1979–2010. The graph shows clearly that ozone levels in recent years have been lower overall than in 1957–1972. The ozone hole is generally considered to be the area where total ozone is less than 220 Dobson Units (DU).

Each line on the graph represents a different period of time. The red line represents average ozone levels from 1957 to 1972, during which time ozone levels varied from about 285 DU (Dobson units) to 360 DU. The blue line shows the ozone levels from 1979 to 2010. At no point during this period were the ozone levels as high as they were at corresponding times during 1957–1972.

The graph makes the ozone hole clearly evident—the dip in the blue line indicates the presence of the ozone hole. Having data from two time periods on the same graph allows scientists to compare recent data with data from a time before the ozone hole existed. Graphs similar to **Figure 22** helped scientists identify a significant trend in ozone levels and verify the depletion in ozone levels over time.

Get It?

Interpret By how much did the total ozone vary during the 9-month period shown for 1979–2010?

Check Your Progress

Summary

- Circle graphs show parts of a whole. Bar graphs show how a factor varies with time, location, or temperature.
- Independent (*x*-axis) variables and dependent (*y*-axis) variables can be related in a linear or a nonlinear manner. The slope of a straight line is defined as rise/run, or $\Delta y/\Delta x$.
- Because line graph data are considered continuous, you can interpolate between data points or extrapolate beyond them.

Demonstrate Understanding

57. **Explain** why graphing can be an important tool for analyzing data.
58. **Infer** What type of data must be plotted on a graph for the slope of the line to represent density?
59. **Relate** If a linear graph has a negative slope, what can you say about the dependent variable?
60. **Summarize** What data are best displayed on a circle graph? On a bar graph?
61. **Construct** a circle graph for the composition of air: 78.08% N_2, 20.95% O_2, 0.93% Ar, and 0.04% CO_2 and other gases.
62. **Infer** from **Figure 22** how long the ozone hole lasts.
63. **Apply** Graph mass versus volume for the data given in the table. What is the slope of the line?

Volume (cm^3)	7.5	12	15	22
Mass (g)	24.1	38.5	48.0	70.1

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SCIENCE & SOCIETY

A Deadly Chemical Attraction

Over the last few decades, heroin addiction has claimed the lives of thousands of people in the United States. The heroin-related overdose death rate has quadrupled since 2010. In recent years, however, the mixing of heroin with a drug called fentanyl has sent the overdose death rate skyrocketing. In 2016, for example, more than 20,000 of the total estimated 64,000 U.S. drug overdose deaths were attributed to fentanyl and related synthetic opioids.

The chemical formula of fentanyl is $C_{22}H_{28}N_2O$, but parts of the molecule can be modified to form slightly different drugs.

Fentanyl and Its Analogs

Opioids are chemicals that interact with receptors on nerve cells to reduce pain. They can also cause pleasurable feelings. Because of the effects that opioids have on the body, they can be highly addictive. Fentanyl is an opiate that is used to relieve post-surgical pain, as well as the pain of some cancers. It is also powerfully euphoric, which is why it is illegally mixed with heroin. However, just a few grains of fentanyl are fatal to most people—thus the uptick in overdoses.

It is illegal in the U.S. to sell or use fentanyl, unless a person has a prescription. In response, chemists in illegal underground labs—many located in China and other countries, beyond the reach of U.S. law enforcement—have slightly altered the formula of the fentanyl molecule, creating substances known as analogs that are weaker or stronger than fentanyl. Chemists do this by adding atoms to the molecule or by moving groups of atoms within the molecule.

For example, chemists add two atoms of carbon and two atoms of hydrogen to the fentanyl molecule to create the analog carfentanil, which is 100 times as strong as fentanyl. Illegal overseas labs also sell the chemicals needed to make fentanyl and its analogs, including carfentanil, to underground chemists working with drug dealers in the U.S.

Analogs have effects similar to fentanyl, but they are not illegal because their chemical structures are different. Lawmakers are constantly playing catch-up; they outlaw one substance, only to find another, similar substance being sold on the streets a short time later.

COMMUNICATE SCIENTIFIC INFORMATION

Using digital or print resources, find more information about fentanyl and its analogs. Create a short video that warns your peers about the deadly substances that are often mixed with heroin. Include information about the differing chemical formulas of the analogs.

MODULE 1
STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 WHAT IS CHEMISTRY?

- Chemistry is the study of matter. There are several branches of chemistry, including organic, inorganic, physical, analytical, and biochemistry.
- Science is the use of evidence to construct testable explanations and predictions of natural phenomena, as well as, the knowledge generated through this process.
- Models are tools that scientists, including chemists, use. A hypothesis is a testable explanation of a situation or phenomena. A theory is a hypothesis that is supported by many experiments.

- chemistry
- science
- hypothesis
- theory
- scientific law
- pure research
- applied research
- substance
- mass
- weight
- model

Lesson 2 MEASUREMENT

- SI measurement units and prefixes allow scientists to report data.
- Density is a ratio of mass to volume and can be used to identify a substance.
- A number expressed in scientific notation is written as a coefficient between 1 and 10 multiplied by 10 raised to a power.
- Dimensional analysis uses conversion factors to solve problems.

- base unit
- second
- meter
- kilogram
- kelvin
- derived unit
- liter
- density
- scientific notation
- dimensional analysis
- conversion factor

Lesson 3 UNCERTAINTY IN DATA

- An accurate measurement is close to the accepted value. A set of precise measurements shows little variation. The measurement device determines the degree of precision possible.
- Error is the difference between the measured value and the accepted value. Percent error gives the percent deviation from the accepted value.
- The number of significant figures reflects the precision of reported data.
- Calculations are often rounded to the correct number of significant figures.

- accuracy
- precision
- error
- percent error
- significant figures

Lesson 4 REPRESENTING DATA

- Circle graphs show parts of a whole. Bar graphs show how a factor varies with time, location, or temperature.
- Independent (x-axis) variables and dependent (y-axis) variables can be related in a linear or a nonlinear manner. The slope of a straight line is defined as rise/run, or $\Delta y/\Delta x$.

$$\text{slope} = \frac{y_2 - y_1}{x_2 - x_1} = \frac{\Delta y}{\Delta x}$$

- Because line graph data are considered continuous, you can interpolate between data points or extrapolate beyond them.

- graph
- independent variable
- dependent variable

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

What do plants and buildings have in common?

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.

GO FURTHER

SEP Data Analysis Lab

How can mass and volume data for an unknown sample be used to identify it?

A student collected from a stream bed several samples that looked like gold. She measured the mass of each sample and used water displacement to determine each sample's volume. Her data are given in the table.

Mass and Volume Data for Unknown Samples

Sample	Mass	Initial Volume (water only)	Final Volume (water + sample)
1	50.25 g	50.1 mL	60.3 mL
2	63.56 g	49.8 mL	62.5 mL
3	57.65 g	50.2 mL	61.5 mL
4	55.35 g	45.6 mL	56.7 mL
5	74.92 g	50.3 mL	65.3 mL
6	67.78 g	47.5 mL	60.8 mL

CER Analyze and Interpret Data

For a given sample, the difference in the volume measurements made with the graduated cylinder yields the volume of the sample. Thus, for each sample, the mass and volume are known, and the density can be calculated. Note that density is a property of matter that can often be used to identify an unknown sample.

1. **Calculate** the volume and density for each sample and the average density of the six samples. Be sure to use significant figure rules.
2. **Claim** The student hopes the samples are gold, which has a density of 19.3 g/cm^3. A local geologist suggested the samples might be pyrite, which is a mineral with a density of 5.01 g/cm^3. What is the identity of the unknown sample?
3. **Calculate** the error and percent error of each sample. Use the density values given in Question 2 as the accepted values.
4. **Evidence, Reasoning** Was the data collected by the student accurate? Explain your answer.

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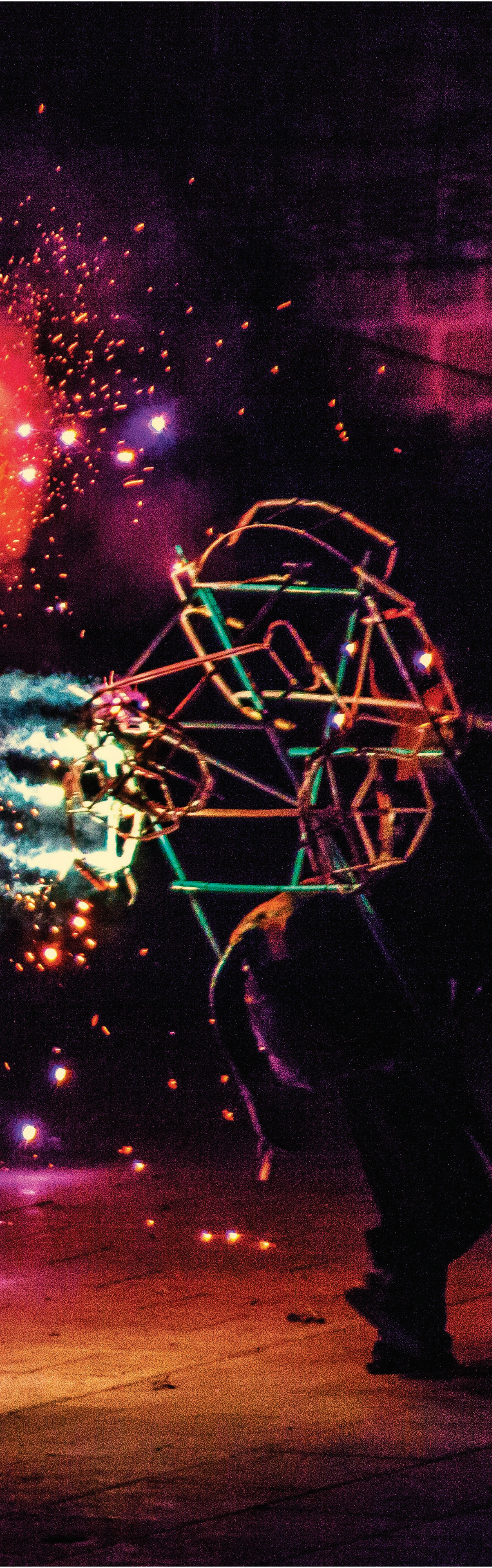


UNIT 1

STRUCTURE AND PROPERTIES OF MATTER

ENCOUNTER THE PHENOMENON

How do fireworks get their colors?

SEP Ask Questions

What questions do you have about the phenomenon? Write your questions on sticky notes and add them to the driving question board for this unit.

How much do bubbles weigh?

Look for Evidence

As you go through this unit, use the information and your experiences to help you answer the phenomenon question as well as your own questions. For each activity, record your observations in a Summary Table, add an explanation, and identify how it connects to the unit and module phenomenon questions.

Solve a Problem

PERFORMANCE EXPECTATION

STEM UNIT PROJECT

Construction Projects How do the properties of matter affect the materials chosen for a construction project? Investigate the ways in which materials can be engineered to meet our needs.

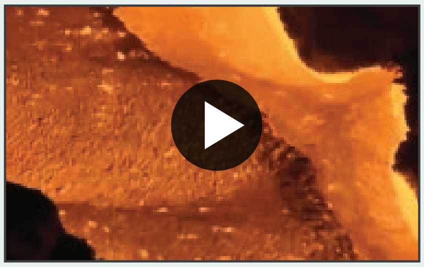
GO ONLINE In addition to reading the information in your Student Edition, you can find the STEM Unit Project and other useful resources online.

Sopon Chienwittayakun/Moment/Getty Images

MODULE 2
MATTER—PROPERTIES AND CHANGES

ENCOUNTER THE PHENOMENON

Why is this volcano spewing bright blue flames?



GO ONLINE to play a video about the mystery of the Kawah Ijen volcano.

SEP Ask Questions

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

CER Claim, Evidence, Reasoning

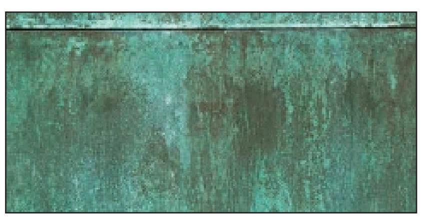
Make Your Claim Use your CER chart to make a claim about why this volcano is spewing bright blue flames.

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: Physical Properties



LESSON 2: Explore & Explain: Chemical Changes

Smithsonian

Additional Resources

(t)Video Supplied by BBC Worldwide Learning; (bl)Design Pics/Kelly Redinger; (br)Pablo Scapinachis/Shutterstock.com

LESSON 1
PROPERTIES OF MATTER

FOCUS QUESTION

What are the differences between physical and chemical properties?

Substances

As you know, matter is anything that has mass and takes up space. Everything around us is matter, including things that we cannot see, such as air and microbes. For example, table salt is a simple type of matter that you are probably familiar with. Table salt has a unique and unchanging chemical composition. Its chemical name is sodium chloride. It is always 100% sodium chloride, and its composition does not change from one sample to another. Salt harvested from the sea or extracted from a mine, as shown in **Figure 1**, always has the same composition and properties.

Salt from the sea

Salt from a mine

Figure 1 Harvested from the sea or extracted from a mine, salt always has the same composition.

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 Density of Wood
Plan and carry out an investigation to determine the patterns in mass and volume of different types of wood.

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

Recall that matter with a uniform and unchanging composition is called a substance, also known as a pure substance. Table salt is a pure substance. Another example of a pure substance is pure water. Water is always composed of hydrogen and oxygen. Seawater and tap water, on the other hand, are not pure substances because samples taken from different locations will often have different compositions. That is, the samples will contain different amounts of water, minerals, and other dissolved substances. Substances are important; much of your chemistry course will be focused on the composition of substances and how they interact with one another.

States of Matter

Imagine you are sitting on a sunny bench, breathing heavily and drinking water after playing a game of soccer. You are interacting with four different forms of matter—the bench is a solid, the water is a liquid, the air you breathe is a gas, and the Sun is a plasma. Although scientists recognize other forms, matter that exists naturally on Earth can be classified as one of these physical forms, which are called **states of matter.** Each state of matter can be distinguished by its properties.

Get It?

Name four states of matter.

Solids A **solid** is a form of matter that has its own definite shape and volume. Wood, iron, paper, and sugar are all examples of solids. The particles of matter in a solid are tightly packed. When heated, a solid expands, but only slightly. Because its shape is definite, a solid might not conform to the shape of the container in which it is placed. If you place a rock into a container, the rock will not take the shape of the container, as shown in **Figure 2**. The tight packing of particles in a solid makes it incompressible; that is, it cannot be pressed into a smaller volume. It is important to understand that a solid is not defined by its rigidity or hardness. For instance, although concrete is rigid and wax is soft, they are both solids.

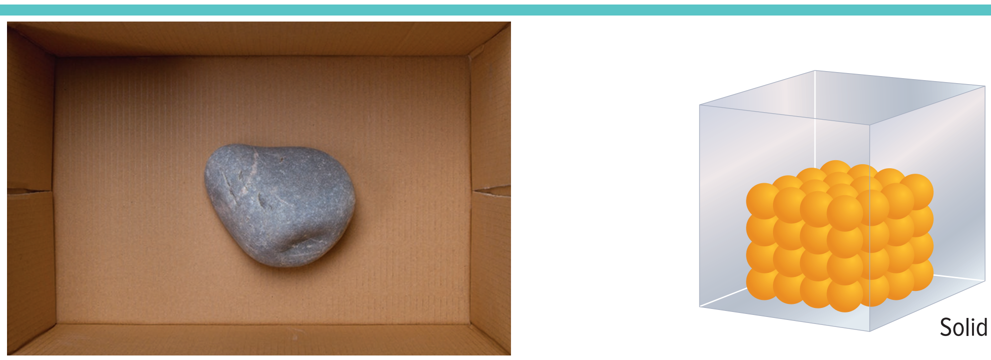


Figure 2 A solid has a definite shape and does not take the shape of its container. Particles in a solid are tightly packed.

Liquids A **liquid** is a form of matter that flows, has constant volume, and takes the shape of its container. Common examples of liquids include water, blood, and mercury. The particles in a liquid are not rigidly held in place and are less closely packed than the particles in a solid. Liquid particles are able to move past each other. This property allows a liquid to flow and take the shape of its container, as shown in **Figure 3**, although it might not completely fill the container. A liquid's volume is constant: regardless of the size and shape of the container in which the liquid is held, the volume of the liquid remains the same. Because of the way the particles of a liquid are packed, liquids are virtually incompressible. Like solids, however, liquids tend to expand when they are heated.

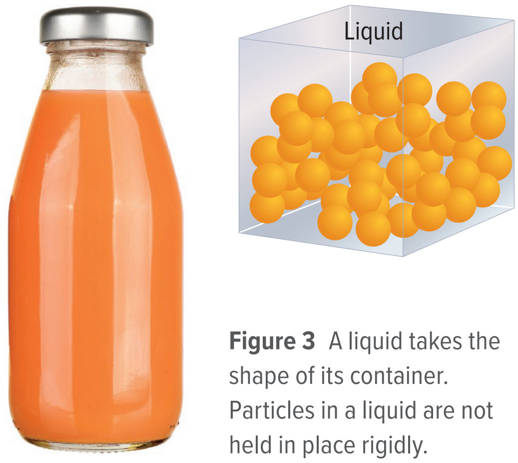


Figure 3 A liquid takes the shape of its container. Particles in a liquid are not held in place rigidly.

Get It?

Compare the properties of solids and liquids in terms of their particle arrangements.

Gases A **gas** is a form of matter that not only flows to conform to the shape of its container but also fills the entire volume of its container, as shown in **Figure 4**. If you flow gas into a container and close the container, the gas will expand to fill the container. Compared to solids and liquids, the particles of gases are far apart. Because of the significant amount of space between particles, gases are easily compressed.

You are probably familiar with the word *vapor* as it relates to the word *gas*. However, the words *gas* and *vapor*, while similar, do not mean the same thing, and should not be used interchangeably. The word *gas* refers to a substance that is naturally in the gaseous state at room temperature. The word **vapor** refers to the gaseous state of a substance that is a solid or a liquid at room temperature. For example, steam is a vapor because water exists as a liquid at room temperature.

Get It?

Differentiate between gas and vapor.

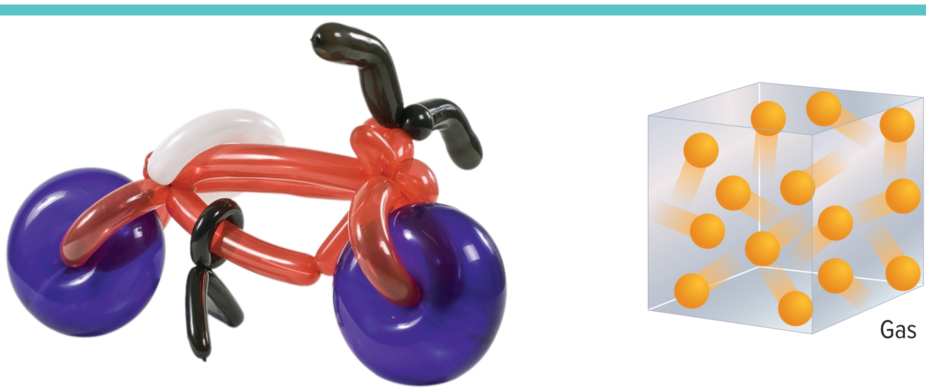


Figure 4 Gases take the shape and volume of their containers. Particles in a gas are very far apart.

Figure 5 Plasmas are ionized gases.

Plasmas The fluorescent lights in a classroom or the bright, colorful lights in the sign in **Figure 5** involve matter as a plasma. A plasma is a form of matter that results when the particles of a gas become ionized and are broken apart into smaller charged particles. Ionization requires energy, such as that provided by an electric current in a neon or fluorescent bulb or the high temperatures of lightning or a star. Although plasmas can behave like gases in many ways, ionization causes plasmas to exhibit a strong response to electric and magnetic fields.

Physical Properties of Matter

You are probably used to identifying objects by their properties—their characteristics and behavior. For example, you can easily identify a pencil in your backpack because you recognize its shape, color, weight, or some other property. These characteristics are all physical properties of the pencil. A **physical property** is a characteristic of matter that can be observed or measured without changing the sample's composition. Physical properties also describe pure substances. Because substances have uniform and unchanging compositions, they also have consistent and unchanging physical properties. Density, color, odor, hardness, melting point, and boiling point are common physical properties that scientists record as identifying characteristics of a substance.

Get It?
Define *physical property* and provide examples.

Real-World Chemistry
Physical Properties

MINERALS Scientists use physical properties such as color and hardness to identify minerals. For instance, malachite is always green and relatively soft. Malachite was used as a pigment in paint and is now mainly used to make jewelry.

Table 1 Physical Properties of Common Substances

Substance	Color	State at 25°C	Melting Point (°C)	Boiling Point (°C)	Density (g/cm^3) at 25°C
Oxygen	colorless	gas	–219	–183	0.0013
Mercury	silver	liquid	–39	357	13.5
Water	colorless	liquid	0	100	0.997
Sucrose	white	solid	186	decomposes	1.58
Sodium chloride	white	solid	801	1465	2.17

Extensive and intensive properties **Table 1** lists several common substances and their physical properties. Physical properties can be further described as being one of two types. **Extensive properties** are dependent on the amount of substance present. For example, mass is an extensive property. Length and volume are also extensive properties. Density, on the other hand, is an example of an intensive property of matter. **Intensive properties** are independent of the amount of substance present. For example, the density of a substance (at constant temperature and pressure) is the same no matter how much substance is present.

A substance can often be identified by its intensive properties. In some cases, a single intensive property is unique enough for identification. For instance, most of the spices shown in **Figure 6** can be identified by their scent.

Chemical Properties of Matter

Some properties of a substance are not obvious unless the substance has changed composition as a result of its contact with other substances or the application of thermal or electric energy. The ability or inability of a substance to combine with or change into one or more other substances is called a **chemical property.**

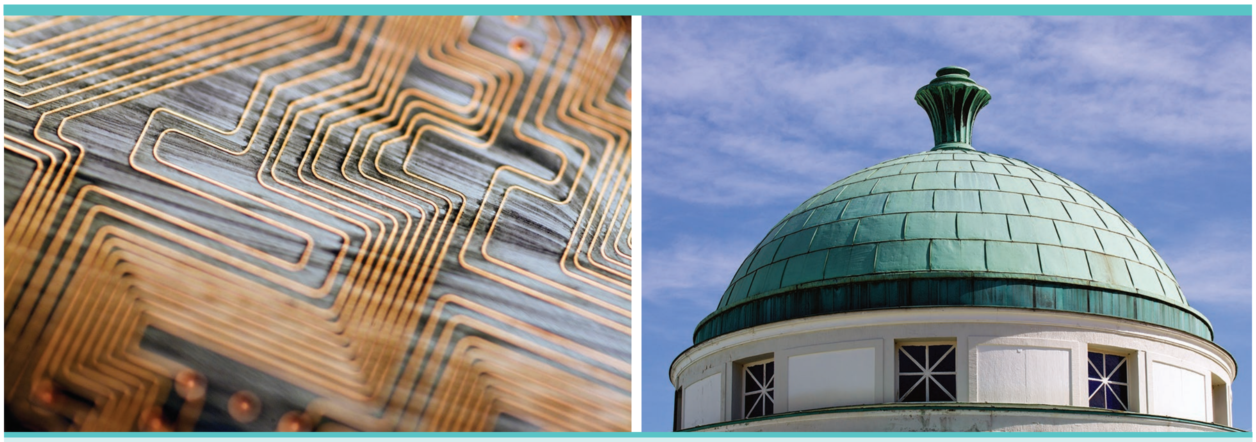
Iron forming rust when combined with the oxygen in air is an example of a chemical property of iron. Similarly, the inability of a substance to change into another substance is also a chemical property. For example, when iron is placed in nitrogen gas at room temperature, no chemical change occurs.

Get It?

Compare physical and chemical properties.

Figure 6 Many spices can be identified by their scent, which is an intensive property.

Infer *Name an extensive property of one of the spices pictured in this figure.*



Copper wires

Copper roof

Figure 7 One of the physical properties of copper is that it can be shaped into different forms, such as the wires on circuit boards. The fact that copper turns from reddish to green when reacting with substances in the air is a chemical property.

Observing Properties of Matter

Every substance has its own unique set of physical and chemical properties. **Figure 7** shows physical and chemical properties of copper. Copper can be shaped into different forms, which is a physical property. When copper is in contact with air for a long time, it reacts with the substances in the air and turns green. This is a chemical property. **Table 2** lists several physical and chemical properties of copper.

Properties and states of matter

The properties of copper listed in **Table 2** might vary depending on the conditions under which they are observed. Because the particular form, or state, of a substance is a physical property, changing the state introduces or adds another physical property to its characteristics. It is important to state the specific conditions, such as temperature and pressure, under which observations are made because both physical and chemical properties depend on these conditions. Resources that provide tables of physical and chemical properties of substances, such as the *CRC Handbook of Chemistry and Physics,* generally include the physical properties of substances in all of the states in which they can exist.

Table 2 **Properties of Copper**

Physical Properties	Chemical Properties
• reddish brown, shiny • easily shaped into sheets (malleable) and drawn into wires (ductile) • a good conductor of heat and electricity • density = 8.96 g/cm^3 • melting point = 1085°C • boiling point = 2562°C	• forms green copper carbonate compound when in contact with moist air • reacts with nitric acid and sulfuric acid, forming new substances • one type of compound forms a deep-blue solution when in contact with ammonia

Figure 8 Because the density of ice is lower than the density of water, icebergs float on the ocean.

Consider the properties of water, for example. You might think of water as a liquid (physical property) which is not particularly chemically reactive (chemical property). You might also find that water has a density of 1.00 g/cm^3 (physical property). These properties, however, apply only to water at standard temperature and pressure. At temperatures greater than 100°C, water is a gas (physical property) with a density of about 0.0006 g/cm^3 (physical property) that reacts rapidly with many different substances (chemical property). Below 0°C, water is a solid (physical property) with a density of about 0.92 g/cm^3 (physical property). The lower density of ice accounts for the fact that icebergs float on the ocean, as shown in **Figure 8**. Clearly, the properties of water are dramatically different under different conditions.

Check Your Progress

Summary

- Four states of matter are solid, liquid, gas, and plasma.
- Physical properties can be observed without altering a substance's composition.
- Chemical properties describe a substance's ability to combine with or change into one or more new substances.
- External conditions can affect both physical and chemical properties.

Demonstrate Understanding

1. **Create** a table that describes four states of matter in terms of their properties.
2. **Describe** the characteristics that identify a sample of matter as a substance.
3. **Classify** each of the following as a physical or a chemical property.
   a. Iron and oxygen form rust.
   b. Iron is more dense than aluminum.
   c. Magnesium burns brightly when ignited.
   d. Oil and water do not mix.
   e. Mercury melts at −39°C.
4. **Organize** Create a chart that compares physical and chemical properties. Give two examples for each type of property.

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LESSON 2
CHANGES IN MATTER

FOCUS QUESTION

How does a substance change when it burns?

Physical Changes

Imagine crumpling aluminum foil from a smooth, flat, mirrorlike sheet to a round, compact ball. The substance looks very different, but its composition is unchanged—it is still aluminum. A change such as this, which alters a substance without changing its composition, is a **physical change.**

Phase change

As with other physical properties, the state of matter depends on the temperature and pressure of the surroundings. As temperature and pressure change, most substances undergo a change from one state (or phase) to another. A **phase change** is a transition of matter from one state to another. **Figure 9** shows two common phase changes.

Condensation | Solidification

Figure 9 Condensation can occur when a gas is in contact with a cool surface, causing droplets to form. Solidification occurs when a liquid cools. Water dripping from the roof forms icicles as it cools.

(l)Andrei Kuzmik/Shutterstock.com, (r)Stephen Bedtelyon/Shutterstock.com

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: Properties of Water
Obtain, evaluate, and communicate information to determine the structure and function of water in Earth's surface processes.

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

EARTH SCIENCE Connection **The water cycle** Phase changes associated with water make up the water cycle, which allows life to exist on Earth. At atmospheric pressure and at temperatures of 0°C and below, water is in its solid state, which is known as ice. As heat is added to the ice, it melts and becomes liquid water. This change of state is a physical change because even though ice and water have different appearances, they have the same composition. If the temperature of the water increases to 100°C, the water begins to boil, and liquid water is converted to steam. Melting and formation of a gas are both physical changes and phase changes. Terms such as *boil, freeze, condense, vaporize,* or *melt* in chemistry generally refer to a phase change in matter.

The temperature and pressure at which a substance undergoes a phase change are important physical properties. These properties are called the melting and boiling points of the substance. Look again at **Table 1** to see this information for several common substances. Like density, the melting and boiling points are intensive physical properties. Tables of intensive properties, such as those given at the end of this textbook or in the *CRC Handbook of Chemistry and Physics,* are useful tools in identifying unknown substances from experimental data.

Chemical Changes

A process that involves one or more substances changing into new substances is called a **chemical change,** commonly referred to as a chemical reaction. The new substances formed in the reaction have different compositions and different properties from the substances present before the reaction occurred. For example, the formation of rust when iron reacts with oxygen in moist air is a chemical change. Rust, shown in **Figure 10**, is a chemical combination of iron and oxygen.

In chemical reactions, the starting substances are called reactants, and the new substances that are formed are called products. Terms such as *decompose, explode, rust, oxidize, corrode, tarnish, ferment, burn,* or *rot* generally refer to chemical reactions.

Get It?
Define *chemical change.*

Figure 10 When iron rusts, new substances are formed due to chemical change.
Identify *the reactants and products in the formation of rust.*

Figure 11 When food rots, new substances are formed due to chemical change.

Evidence of a chemical reaction

Take another look at **Figure 10.** As shown, rust is a brownish-orange powdery substance. It looks very different from the elements iron and oxygen that compose it. Rust is not attracted to a magnet, whereas iron is. The observation that the product (rust) has different properties than the reactants (iron and oxygen) is evidence that a chemical reaction has taken place. A chemical reaction always produces a change in properties.

Spoiled food, such as rotten fruit and bread, is another example of chemical reactions. The properties of spoiled food differ from the properties of fresh food. How the food tastes, how it smells, and its digestibility might all change as food rots. Examples of food that have undergone chemical reactions are shown in **Figure 11**.

Conservation of Mass

It was only in the late eighteenth century that scientists began to use quantitative tools to study and monitor chemical changes. The analytical balance, which was capable of measuring small changes in mass, was developed at that time. By carefully measuring mass before and after many chemical reactions, it was observed that, although chemical changes occurred, the total mass involved in the reaction remained constant. Assuming this was true for all reactions, chemists summarized this observation in a scientific law.

The **law of conservation of mass** states that mass is neither created nor destroyed during a chemical reaction—it is conserved. In other words, the mass of the reactants equals the mass of the products. The equation form of the law of conservation of mass is as follows:

The Law of Conservation of Mass

$$\text{mass}_{\text{reactants}} = \text{mass}_{\text{products}}$$

Mass is conserved in a chemical reaction; products have the same mass as reactants.

EXAMPLE Problem 1

CONSERVATION OF MASS In an experiment, 10.00 g of red mercury(II) oxide powder is placed in an open flask and heated until it is converted to liquid mercury and oxygen gas. The liquid mercury has a mass of 9.26 g. What is the mass of oxygen formed in the reaction?

1 ANALYZE THE PROBLEM

You are given the mass of a reactant and the mass of one of the products in a chemical reaction. According to the law of conservation of mass, the total mass of the products must equal the total mass of the reactants.

Known

$m_{\text{mercury(II) oxide}} = 10.00$ g

$m_{\text{mercury}} = 9.26$ g

Unknown

$m_{\text{oxygen}} = ?$ g

2 SOLVE FOR THE UNKNOWN

$\text{Mass}_{\text{reactants}} = \text{Mass}_{\text{products}}$ State the law of conservation of mass.

$m_{\text{mercury(II) oxide}} = m_{\text{mercury}} + m_{\text{oxygen}}$

$m_{\text{oxygen}} = m_{\text{mercury(II) oxide}} - m_{\text{mercury}}$ Solve for m_{oxygen}

$m_{\text{oxygen}} = 10.00 \text{ g} - 9.26 \text{ g}$ Substitute $m_{\text{mercury(II) oxide}} = 10.00$ g and $m_{\text{mercury}} = 9.26$ g.

$m_{\text{oxygen}} = 0.74 \text{ g}$

3 EVALUATE THE ANSWER

The sum of the masses of the two products equals the mass of the reactant, verifying that mass has been conserved. The answer is correctly expressed to the hundredths place, making the number of significant digits correct.

PRACTICE Problems

ADDITIONAL PRACTICE

5. Use the data in the table to answer the following questions.

Aluminum and Liquid Bromine Reaction

Substance	Before Reaction	After Reaction
Aluminum	10.3 g	0.0 g
Liquid bromine	100.0 g	8.2 g
Compound	0.0 g	

How many grams of bromine reacted? How many grams of compound were formed?

6. From a laboratory process designed to separate water into hydrogen and oxygen gas, a student collected 10.0 g of hydrogen and 79.4 g of oxygen. How much water was originally involved in the process?

7. A student carefully placed 15.6 g of sodium in a reactor supplied with an excess quantity of chlorine gas. When the reaction was complete, the student obtained 39.7 g of sodium chloride. Calculate how many grams of chlorine gas reacted. How many grams of sodium reacted?

8. A 10.0-g sample of magnesium reacts with oxygen to form 16.6 g of magnesium oxide. How many grams of oxygen reacted?

9. **CHALLENGE** 106.5 g of HCl(g) react with an unknown amount of NH_3(g) to produce 156.3 g of NH_4Cl(s). How many grams of NH_3(g) reacted? Is the law of conservation of mass observed in the reaction? Justify your answer.

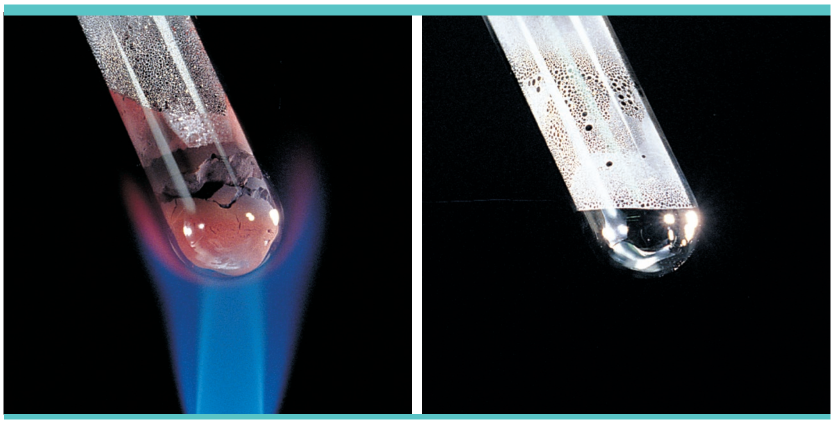

Figure 12 When mercury(II) oxide is heated, the powdery red solid reacts to form silvery liquid mercury and colorless oxygen gas. The sum of the masses of mercury and oxygen produced during the reaction equals the mass of mercury(II) oxide.

French scientist Antoine Lavoisier (1743–1794) was one of the first to use an analytical balance to monitor chemical reactions. He studied the thermal decomposition of mercury(II) oxide, shown in **Figure 12.** The color change and production of a gas are indicators of a chemical reaction. In a closed container, the oxygen gas cannot escape and the mass before and after the reaction can be measured. The masses are the same. The law of conservation of mass is one of the most fundamental concepts in the study of chemistry and chemical reactions.

Check Your Progress

Summary

- A physical change alters the physical properties of a substance without changing its composition.
- A chemical change, also known as a chemical reaction, involves a change in a substance's composition.
- In a chemical reaction, reactants form products.
- The law of conservation of mass states that mass is neither created nor destroyed during a chemical reaction; it is conserved.

Demonstrate Understanding

10. **Classify** each change as physical or chemical.
    a. crushing an aluminum can
    b. recycling used aluminum cans to make new ones
    c. aluminum combining with oxygen to form aluminum oxide
11. **Describe** the results of a physical change and list three examples of physical change.
12. **Describe** the results of a chemical change. List four indicators of chemical change.
13. **Calculate** Solve each of the following.
    a. If 22.99 g of sodium and 35.45 g of chlorine fully react, how much sodium chloride forms?
    b. A 12.2-g sample of *X* reacts with a sample of *Y* to form 78.9 g of *XY*. What mass of *Y* reacted?
14. **Evaluate** A friend tells you, "Because composition does not change during a physical change, the appearance of a substance does not change." Is your friend correct? Explain.

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LESSON 3

ELEMENTS AND COMPOUNDS

FOCUS QUESTION

What are elements and compounds?

Elements

Earlier in this module, you considered the diversity of your surroundings in terms of matter. Although matter can take many different forms, all matter can be broken down into a relatively small number of basic building blocks called elements. An **element** is a pure substance that cannot be separated into simpler substances by physical or chemical means. On Earth, over 90 elements occur naturally. Copper, oxygen, and gold are examples of naturally-occurring elements. Elements are found in different physical states in normal conditions, as shown in **Figure 13**. There are also several elements that do not exist naturally but have been developed by scientists.

Each element has a unique chemical name and chemical symbol. The chemical symbol consists of one, two, or three letters; the first letter is always capitalized, and the remaining letter(s) are always lowercase. The names and symbols of the elements are universally accepted by scientists world-wide in order to make the communication of chemical information possible.

Copper pot—solid	Mercury switch—liquid	Helium balloon—gas

Figure 13 In normal conditions, elements exist in different states.

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: Identify the Products of a Chemical Reaction
Obtain, evaluate, and communicate information to determine the cause and effect of a chemical reaction.

Inquiry into Chemistry: Is it gold or fool's gold?
Construct an explanation for the importance of examining the structure of an unknown material.

The naturally-occurring elements are not equally abundant. For example, hydrogen is estimated to make up approximately 75% of the mass of the universe. Oxygen and silicon together comprise almost 75% of the mass of Earth's crust, while oxygen, carbon, and hydrogen account for more than 90% of the human body. Francium, on the other hand, is one of the least-abundant naturally-occurring elements. There is probably less than 20 g of francium dispersed throughout Earth's crust.

A first look at the periodic table As many new elements were being discovered in the early nineteenth century, chemists began to observe and study patterns of similarities in the chemical and physical properties of particular sets of elements. In 1869, Russian chemist Dmitri Mendeleev (1834–1907) devised a chart, shown in **Figure 14**, which organized all of the elements that were known at the time. His classification was based on the similarities and masses of the elements. Mendeleev's table was the first version of what has been further developed into the periodic table of the elements. The **periodic table** organizes the elements into a grid of horizontal rows called periods and vertical columns called groups or families. Elements in the same group have similar chemical and physical properties. The table is called periodic because the pattern of similar properties repeats from period to period.

Tabelle I.

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	—	—	—

der chemischen Elemente.

Figure 14 Mendeleev was one of the first scientists to organize elements in a periodic manner, as shown in this chart, and to observe periodic patterns in the properties of the elements.

SCIENCE USAGE V. COMMON USAGE

element

Science usage: a pure substance that cannot be separated into simpler substances by ordinary chemical means

Lead is one of the heaviest elements.

Common usage: the state or sphere that is natural or suited to any person or thing

In snow, huskies are in their element.

Compounds

Many pure substances can be classified as compounds. A **compound** is made up of two or more different elements that are combined chemically in a fixed ratio. Most matter in the universe exists in the form of compounds. Today, there are more than 50 million known compounds, and new compounds continue to be developed and discovered at the rate of about 100,000 per year. There appears to be no limit to the number of compounds that can be made or that will be discovered. Considering this virtually limitless potential, several organizations have assumed the task of collecting data and indexing the known chemical compounds. The information is stored in databases.

Get It?
Define *element* and *compound*.

Separating compounds into components As you have read earlier in this module, elements can never be separated into simpler substances. However, compounds can be broken down into simpler substances by chemical means. In general, compounds that occur naturally are more stable than the individual component elements. Separating a compound into its elements often requires external energy, such as heat or electricity. **Figure 15** shows the setup used to produce the chemical change of water into its component elements—hydrogen and oxygen—through a process called electrolysis. During electrolysis, one end of a long platinum electrode is exposed to the water in a tube and the other end is attached to a power source. An electric current splits water into hydrogen gas in the compartment on the right and oxygen gas in the compartment on the left. Because water is composed of two parts hydrogen and one part oxygen, there is twice as much hydrogen gas as there is oxygen gas.

Get It?
Explain the process of electrolysis.

The chemical symbols of the periodic table make it easy to write the formulas for chemical compounds. For example, table salt, which is called sodium chloride, is composed of one part sodium (Na) and one part chlorine (Cl), and its chemical formula is NaCl. Water is composed of two parts hydrogen (H) and one part oxygen (O), and its chemical formula is H_2O. The subscript 2 indicates that two hydrogen atoms combine with one oxygen atom to form water.

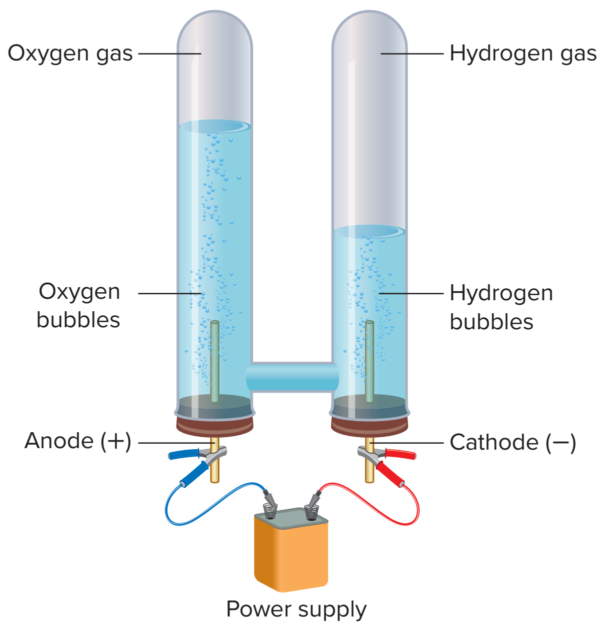


Figure 15 An electric current breaks down water into its components, oxygen and hydrogen.

Determine *What is the ratio between the amount of hydrogen and the amount of oxygen released during electrolysis?*

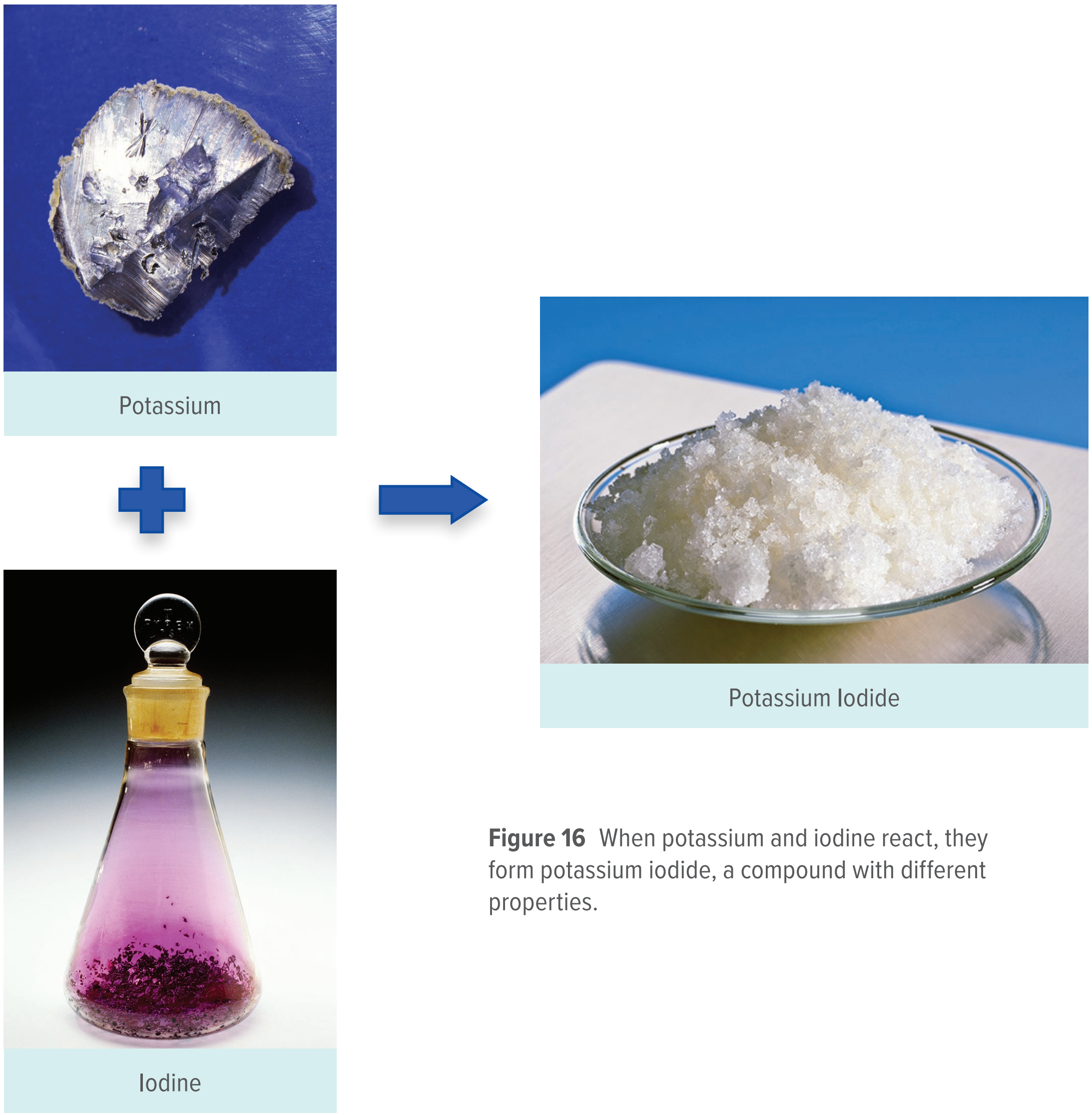


Figure 16 When potassium and iodine react, they form potassium iodide, a compound with different properties.

(t)Andrew Lambert Photography/Science Source, (b)Charles D. Winters/Science Source, (r)Martyn F. Chillmaid/Science Source

Properties of compounds The properties of a compound are different from those of its component elements. The example of water in **Figure 15** illustrates this fact. Water is a stable compound that is liquid at room temperature. When water is broken down, its components, hydrogen and oxygen, are dramatically different than the liquid they form when combined. Oxygen and hydrogen are colorless, odorless gases that undergo vigorous chemical reactions with many elements. This difference in properties is a result of a chemical reaction between the elements.

Figure 16 shows the component elements—potassium and iodine—of the compound called potassium iodide. Note how different the properties of potassium iodide are from its component elements. Potassium is a light silvery metal that reacts with water. Iodine is a black solid that changes into a purple gas at room temperature. Potassium iodide is a white salt.

Get It?
Summarize how the properties of a compound and the properties of its component elements compare.

Law of Definite Proportions

An important characteristic of compounds is that the elements comprising them always combine in definite proportions by mass. This observation is so fundamental that it is summarized as the law of definite proportions. The **law of definite proportions** states that a compound is always composed of the same elements in the same proportion by mass, no matter how large or small the sample. The mass of the compound is equal to the sum of the masses of the elements that make up the compound.

The relative amounts of the elements in a compound can be expressed as percent by mass. The **percent by mass** is the ratio of the mass of each element to the total mass of the compound expressed as a percentage.

Percent by Mass

$$\text{percent by mass (\%)} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100$$

Percent by mass is obtained by dividing the mass of the element by the mass of the compound and then by multiplying this ratio by 100 to express it as a percentage.

Get It?

State the law of definite proportions.

For example, consider the compound granulated sugar (sucrose). This compound is composed of three elements—carbon, hydrogen, and oxygen. The analysis of 20.00 g of sucrose from a bag of granulated sugar is given in **Table 3**. Note that the sum of the individual masses of the elements found in the sugar equals 20.00 g, which is the amount of the granulated sugar sample that was analyzed. This demonstrates the law of conservation of mass as applied to compounds: the mass of a compound is equal to the sum of the masses of the elements that make up the compound.

Table 3 Sucrose Analysis

	20.00 g of Granulated Sugar		500.0 g of Sugarcane	
Element	Analysis by Mass (g)	Percent by Mass (%)	Analysis by Mass (g)	Percent by Mass (%)
Carbon	8.44	$\frac{8.44 \text{ g C}}{20.00 \text{ g sucrose}} \times 100 = 42.20\%$	211.0	$\frac{211.0 \text{ g C}}{500.0 \text{ g sucrose}} \times 100 = 42.20\%$
Hydrogen	1.30	$\frac{1.30 \text{ g H}}{20.00 \text{ g sucrose}} \times 100 = 6.50\%$	32.5	$\frac{32.50 \text{ g H}}{500.0 \text{ g sucrose}} \times 100 = 6.500\%$
Oxygen	10.26	$\frac{10.26 \text{ g O}}{20.00 \text{ g sucrose}} \times 100 = 51.30\%$	256.5	$\frac{256.5 \text{ g O}}{500.0 \text{ g sucrose}} \times 100 = 51.30\%$
Total	20.00	100%	500.0	100%

Suppose you analyzed 500.0 g of sucrose from a sample of sugarcane. The analysis is shown in **Table 3**. The percent-by-mass values for the sugarcane are equal to the values obtained for the granulated sugar. According to the law of definite proportions, samples of a compound from any source must have the same mass proportions. Conversely, compounds with different mass proportions must be different compounds. Thus, you can conclude that samples of sucrose will always be composed of 42.20% carbon, 6.50% hydrogen, and 51.30% oxygen, no matter their sources.

PRACTICE Problems ADDITIONAL PRACTICE

15. A 78.0-g sample of an unknown compound contains 12.4 g of hydrogen. What is the percent by mass of hydrogen in the compound?

16. 1.0 g of hydrogen reacts completely with 19.0 g of fluorine. What is the percent by mass of hydrogen in the compound that is formed?

17. If 3.5 g of element X reacts with 10.5 g of element Y to form the compound XY, what is the percent by mass of element X in the compound? The percent by mass of element Y?

18. Two unknown compounds are tested. Compound I contains 15.0 g of hydrogen and 120.0 g of oxygen. Compound II contains 2.0 g of hydrogen and 32.0 g of oxygen. Are the compounds the same? Explain your answer.

19. **CHALLENGE** All you know about two unknown compounds is that they have the same percent by mass of carbon. With only this information, can you be sure the two compounds are the same? Explain.

Law of Multiple Proportions

Compounds composed of different elements are obviously different compounds. However, different compounds can also be composed of the same elements. This happens when those different compounds have different mass compositions. The **law of multiple proportions** states that when different compounds are formed by a combination of the same elements, different masses of one element combine with the same fixed mass of the other element in a ratio of small whole numbers. Ratios compare the relative amounts of any items or substances. The comparison can be expressed using numbers separated by a colon or as a fraction. With regard to the law of multiple proportions, ratios express the relationship of elements in a compound.

Water and hydrogen peroxide The two distinct compounds water (H_2O) and hydrogen peroxide (H_2O_2) illustrate the law of multiple proportions. Each compound contains the same elements (hydrogen and oxygen). Water is composed of two parts hydrogen and one part oxygen. Hydrogen peroxide is composed of two parts hydrogen and two parts oxygen. Hydrogen peroxide differs from water in that it has twice as much oxygen. When you compare the mass of oxygen in hydrogen peroxide to the mass of oxygen in water, you get the ratio 2:1.

Get It?

State the law of multiple proportions in your own words.

Table 4 **Analysis Data of Two Copper Compounds**

Compound	% Cu	% Cl	Mass Cu (g) in 100.0 g of Compound	Mass Cl (g) in 100.0 g of Compound	Mass Ratio $\left(\frac{\text{mass Cu}}{\text{mass Cl}}\right)$
I	64.20	35.80	64.20	35.80	$\frac{1.793 \text{ g Cu}}{1 \text{ g Cl}}$
II	47.27	52.73	47.27	52.73	$\frac{0.8964 \text{ g Cu}}{1 \text{ g Cl}}$

Compounds made of copper and chlorine In another example, copper (Cu) reacts with chlorine (Cl) under different sets of conditions to form two different compounds. **Table 4** provides an analysis of their compositions. How can you tell that the compounds are different?

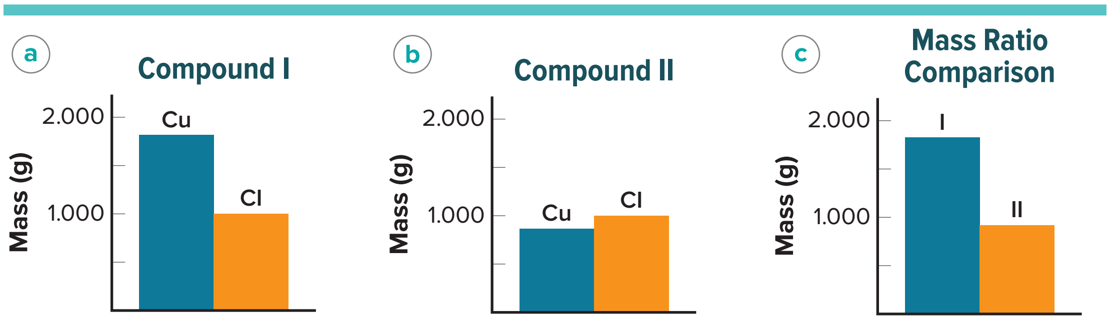
The two copper compounds must be different because they have different percents by mass. Compound I is composed of 64.20% copper, but Compound II contains only 47.27% copper. Compound I contains 35.80% chlorine, but Compound II contains 52.73% chlorine.

Using **Figure 17** and **Table 4**, compare the ratio of the mass of copper to the mass of chlorine for each compound. Notice that the mass ratio of copper to chlorine in Compound I (1.793) is exactly 2 times the mass ratio of copper to chlorine in Compound II (0.8964).

$$\frac{\text{mass ratio of Compound I}}{\text{mass ratio of Compound II}} = \frac{1.793 \text{ g Cu/g Cl}}{0.8964 \text{ g Cu/g Cl}} = 2.000$$

Get It?

Explain why the ratio of relative masses of copper in both compounds is 2:1.

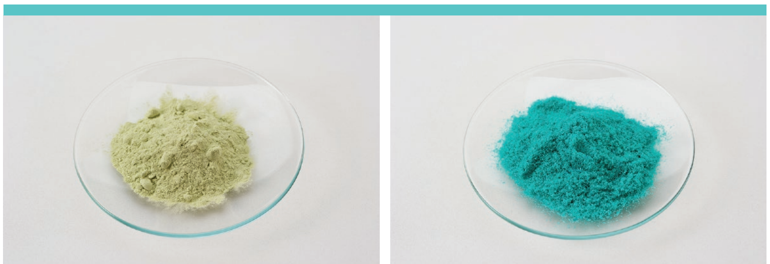


Bar graph **a** compares the relative masses of copper and chlorine in Compound I.

Bar graph **b** compares the relative masses of copper and chlorine in Compound II.

Bar graph **c** shows a comparison between the relative masses of copper in both compounds. The ratio is 2:1.

Figure 17 Copper and chlorine can form different compounds.



Compound I—copper(I) chloride Compound II—copper(II) chloride

Figure 18 Different compounds are formed when different relative masses of each element are combined. Although they are both made of copper and chlorine, Compound I has a greenish color, whereas Compound II has a bluish color.

Figure 18 shows the two compounds formed by the combination of copper and chlorine and presented in **Table 4** and **Figure 17**. These compounds are called copper(I) chloride and copper(II) chloride. As the law of multiple proportions states, the different masses of copper that combine with a fixed mass of chlorine in the two different copper compounds can be expressed as a small whole-number ratio. In this case, the ratio is 2:1. Considering that there is a finite number of elements that exist today and an exponentially greater number of compounds that are composed of these elements under various conditions, it becomes clear how important the law of multiple proportions is in chemistry.

Check Your Progress

Summary

- Elements cannot be broken down into simpler substances.
- Elements are organized in the periodic table of the elements.
- Compounds are chemical combinations of two or more elements, and their properties differ from the properties of their component elements.
- The law of definite proportions states that a compound is always composed of the same elements in the same proportions.
- The law of multiple proportions states that if elements form more than one compound, those compounds will have compositions that are whole-number multiples of each other.

Demonstrate Understanding

20. **Compare and contrast** elements and compounds.
21. **Describe** the basic organizational feature of the periodic table of the elements.
22. **Explain** how the law of definite proportions applies to compounds.
23. **State** the type of compounds that are compared in the law of multiple proportions.
24. **Complete** the table. Then analyze the data to determine if Compounds I and II are the same compound. If the compounds are different, use the law of multiple proportions to show the relationship between them.

ANALYSIS DATA OF TWO IRON COMPOUNDS

Compound	Total Mass (g)	Mass Fe (g)	Mass O (g)	Mass Percent Fe	Mass Percent O
I	75.00	52.46	22.54		
II	56.00	43.53	12.47		

25. **Calculate** the mass percent of each element in water.
26. **Graph** Create a graph that illustrates the law of multiple proportions.

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Matt Meadows/McGraw-Hill Education

LESSON 4

MIXTURES OF MATTER

FOCUS QUESTION

Would the substance still burn blue if it were mixed with another substance?

Mixtures

You have already read that a pure substance has a uniform and unchanging composition. What happens when two or more substances are combined? A **mixture** is a combination of two or more pure substances in which each pure substance retains its individual chemical properties. The composition of mixtures is variable, and the number of mixtures that can be created by combining substances is infinite. Although much of the focus of chemistry is the behavior of substances, it is important to remember that most everyday matter occurs as mixtures. Substances tend to mix naturally; it is difficult to keep any substance pure.

Examine the mixtures in **Figure 19**. When oil, seasonings, and vinegar are mixed, you can still distinguish all of the components. If the mixture remains undisturbed long enough, the oil will form a layer on top of the vinegar. On the other hand, you cannot distinguish between the components of the mercury-silver mixture. You could, however, separate them by heating the mixture. The mercury will evaporate before the silver does, and you will obtain two separate substances: mercury vapor and solid silver. The mercury and silver physically mixed to form the mixture but did not chemically react with each other. They could be separated by the physical method of boiling. The vapor can be collected and condensed into liquid mercury.

Figure 19 There are different types of mixtures. The components of some mixtures, like this salad dressing, are visible. It is not possible to see the different components of some mixtures, such as this mercury-silver filling.

(t)McGraw-Hill Education, (b) Szasz-Fabian Jozsef/Shutterstock.com

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.

Small-Scale Lab: Separation of Aspirin

Carry out an investigation to determine the scale, proportion, and quantity of the components of a common medication.

Identify Crosscutting Concepts

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

Types of mixtures

The combinations of pure substances shown in **Figure 19** are both mixtures, despite their obvious visual differences. Mixtures can be defined in different ways and are classified as either heterogeneous or homogeneous.

A **heterogeneous mixture** is a mixture that does not blend smoothly throughout and in which the individual substances remain distinct. The salad dressing mixture is an example of a heterogeneous mixture. Its composition is not uniform—the substances have not blended smoothly and remain distinct. We can therefore say that the existence of two or more distinct areas indicates a heterogeneous mixture.

A **homogeneous mixture** is a mixture that has constant composition throughout; it always has a single phase. If you cut two pieces out of a silver-mercury amalgam, their compositions will be the same. They will contain the same relative amounts of silver and mercury, no matter the size of each piece.

Homogeneous mixtures are also referred to as **solutions.** You are probably most familiar with solutions in a liquid form, such as tea and lemonade, but solutions can be solids, liquids, or gases. They can be a mixture of a solid and a gas, a solid and a liquid, a gas and a liquid, and so on. **Table 5** lists the various types of solution systems and examples.

The solid-solid solution known as steel is called an alloy. An alloy is a homogeneous mixture of metals, or a mixture of a metal and a nonmetal in which the metal substance is the major component. For instance, steel is a mixture of iron and carbon. Adding carbon atoms increases the hardness of the metal.

Manufacturers combine the properties of various metals in an alloy to achieve greater strength and durability in their products. Jewelry is often made of alloys such as bronze, sterling silver, pewter, and 14-karat gold.

Get It?

List six examples of solutions that you encounter as you go through your daily activities. Indicate the system to which each belongs.

Table 5 Types of Solution Systems

System	Example
Gas-gas	Air in a scuba tank is primarily a mixture of nitrogen, oxygen, and argon gases.
Gas-liquid	Oxygen and carbon dioxide are dissolved in seawater.
Liquid-gas	Moist air exhaled by the scuba diver contains water droplets.
Liquid-liquid	When it is raining, fresh water mixes with seawater.
Solid-liquid	Solid salts are dissolved in seawater.
Solid-solid	The air tank is made of an alloy—a mixture of two metals.

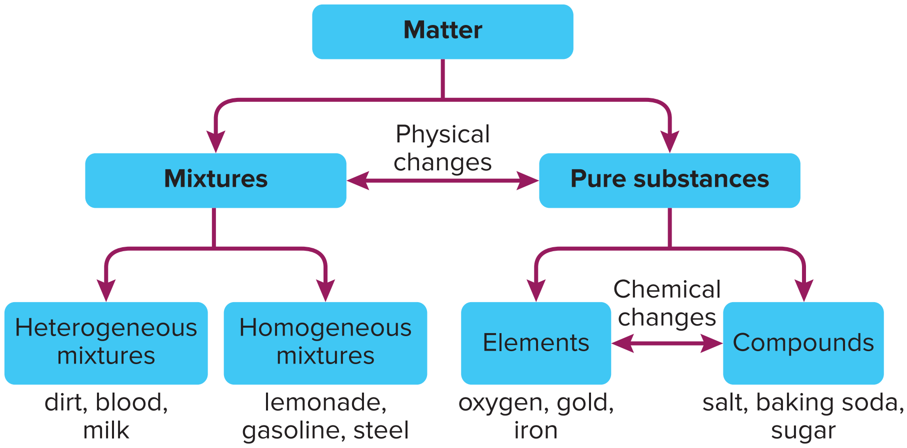


Figure 20 Matter can be classified into different categories that have defined properties.

Examine *How are mixtures and substances related? Elements and compounds?*

Recall what you have learned about the organization of matter. Matter is classified as pure substances and mixtures. An element is a pure substance that cannot be separated into simpler substances, whereas a compound is a chemical combination of two or more elements and can be separated into its components. A mixture can be homogeneous or heterogeneous. **Figure 20** summarizes the classification of matter and how these components are related to one another.

Get It?

Summarize the different types of matter and how they are related to each other.

Separating Mixtures

Most matter exists naturally in the form of mixtures. To gain a thorough understanding of matter, it is important to be able to separate mixtures into their component substances. Because the substances in a mixture are physically combined, the processes used to separate a mixture are physical processes that are based on differences in the physical properties of the substances. For instance, a mixture of iron and sand can be separated into its components with a magnet because a magnet will attract iron but not sand. Numerous techniques have been developed that take advantage of different physical properties in order to separate various mixtures.

STEM CAREER Connection

High School Chemistry Teacher

Do you enjoy helping others? Do you like chemistry? Are you able to effectively explain yourself when talking to others? You may wish to consider a career teaching high school chemistry. High school teachers prepare students for life after high school. They prepare students to be valuable members of society by teaching lessons and skills needed to be successful in post-secondary education or to enter the workforce.

WORD ORIGIN

mixture

from the Latin word *misceo*, meaning *to mix* or *blend*

Figure 21 As the mixture passes through the filter, the solids remain in the filter, while the filtrate (the remaining liquid) is collected in the beaker.

Filtration Heterogeneous mixtures composed of solids and liquids are easily separated by filtration. **Filtration** is a technique that uses a porous barrier to separate a solid from a liquid. As **Figure 21** shows, the mixture is poured through a piece of filter paper that has been folded into a cone shape. The liquid passes through, leaving the solids trapped in the filter paper.

Get It?

Predict the result of trying to use filtration to separate the components of salad dressing. Imagine using salad dressing that is a mixture of oil, vinegar, and spices and a paper filter for the procedure.

Distillation Most homogeneous mixtures can be separated by distillation. **Distillation** is a physical separation technique that is based on differences in the boiling points of the substances involved. In distillation, a mixture is heated until the substance with the lowest boiling point boils to a vapor that can then be condensed into a liquid and collected. When precisely controlled, distillation can separate substances that have boiling points differing by only a few degrees.

Sublimation Mixtures can also be separated by **sublimation,** which is the process during which a solid changes to vapor without melting, i.e., without going through the liquid phase. Sublimation can be used to separate two solids present in a mixture when one of the solids sublimates but not the other.

Chromatography **Chromatography** is a technique that separates the components of a mixture dissolved in either a gas or a liquid (called the mobile phase) based on the ability of each component to travel or to be drawn across the surface of a fixed substrate (called the stationary phase).

For example, chromatography paper is a stationary phase with a solid substrate. During paper chromatography, the separation occurs because the various components of the mixture in the liquid mobile phase spread through the paper at different rates. Components with the strongest attraction for the paper travel slower.

Get It?

Compare and contrast the mobile phase with the stationary phase in chromatography.

Martyn F. Chillmaid/Science Source

Crystallization Making rock candy from a sugar solution is an example of separating a mixture by crystallization. **Crystallization** is a separation technique that results in the formation of pure solid particles of a substance from a solution containing the dissolved substance. When the solution contains as much dissolved substance as it can possibly hold, the addition of even a tiny amount more often causes the dissolved substance to come out of solution and collect as crystals on any available surface.

In the rock candy example, as water evaporates from the sugar-water solution, the solution becomes more concentrated. This is equivalent to adding more of the dissolved substance to the solution. As more water evaporates, the sugar forms a solid crystal on the string, as shown in **Figure 22**. Crystallization produces highly pure solids.

Get It?

Classify which techniques for separating mixtures depend on phase changes and identify the change used by each.

Figure 22 As the water evaporates from the water-sugar solution, the sugar crystals form on the string.

Check Your Progress

Summary

- A mixture is a physical blend of two or more pure substances in any proportion.
- Solutions are homogeneous mixtures.
- Mixtures can be separated by physical means. Common separation techniques include filtration, distillation, sublimation, chromatography, and crystallization.

Demonstrate Understanding

27. **Classify** each of the following as either a heterogeneous or a homogeneous mixture.

    a. tap water b. air c. raisin muffin

28. **Compare** mixtures and substances.
29. **Describe** the separation technique that could be used to separate each of the following mixtures.

    a. two colorless liquids

    b. a nondissolving solid mixed with a liquid

    c. red and blue marbles of the same size and mass

30. **Design** a concept map that summarizes the relationships among the categories of matter, elements, mixtures, compounds, pure substances, and homogeneous and heterogeneous mixtures.

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

SCIENCE & SOCIETY

In Rare Form

What do smartphones, hybrid vehicles, and lasers have in common? Components used in all three are manufactured with rare earth elements (REE). Without REEs, many products could not function. Currently, the United States obtains REEs by trading with China, where most REE mines are located.

Rare earth elements are difficult to mine without negative effects on the environment.

High demand, low supply?

Rare earth elements (REEs) include scandium (SC-21), yttrium (Y-39), and 15 others. *Rare* is a misnomer; many (but not all) REEs are actually quite common. They are found in Earth's crust all over the world. However, it is not easy to find a high concentration of these elements in one place. Without a high concentration, mining is not cost-effective. In addition, it can be difficult to separate the elements from ore and from one another when they occur in the same area, and mining REEs in a way that does not harm the environment is complicated and expensive.

REEs are in demand because of their desirable qualities, which vary from element to element. These include magnetic, electrical, mechanical, chemical, and elastic properties. REEs are extremely important in the production of high-tech products, such as magnets, radar, and sonar, as well as everyday products, such as batteries and televisions. Many renewable energy technologies, from wind turbines to solar cells, also rely on REEs.

China mines the vast majority of the total world production of REEs. As China becomes more developed, it may need more REEs for its own use, and it might start exporting smaller amounts of REEs to other countries. Soon, worldwide demand for REEs may exceed the supply that China provides. However, other countries—including Australia, Russia, India, and Brazil—are mining REEs, although China's output is much greater than all the rest put together.

The United States also has REE deposits and is working to open its own mines. Until the U.S. can open its own mines, what can the country do to ensure its supply of REEs? Recycling REEs from e-waste is difficult and expensive, so it is not as common as it could be. New processes, however, have made the process easier and cheaper, and REE recycling will most likely be on the rise in the future.

SEP DCI CCC

ASK QUESTIONS TO CLARIFY

After reading this feature, brainstorm three questions about the topic of REEs. Use print or online resources to research answers to your questions. Share what you have learned with your class.

MODULE 2
STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 PROPERTIES OF MATTER

- Four states of matter are solid, liquid, gas, and plasma.
- Physical properties can be observed without altering a substance's composition.
- Chemical properties describe a substance's ability to combine with or change into one or more new substances.
- External conditions can affect both physical and chemical properties.

- states of matter
- solid
- liquid
- gas
- vapor
- physical property
- extensive property
- intensive property
- chemical property

Lesson 2 CHANGES IN MATTER

- A physical change alters the physical properties of a substance without changing its composition.
- A chemical change, also known as a chemical reaction, involves a change in a substance's composition.
- In a chemical reaction, reactants form products.
- The law of conservation of mass states that mass is neither created nor destroyed during a chemical reaction; it is conserved.

$$\text{mass}_{\text{reactants}} = \text{mass}_{\text{products}}$$

- physical change
- phase change
- chemical change
- law of conservation of mass

Lesson 3 ELEMENTS AND COMPOUNDS

- Elements cannot be broken down into simpler substances.
- Elements are organized in the periodic table of the elements.
- Compounds are chemical combinations of two or more elements, and their properties differ from the properties of their component elements.
- The law of definite proportions states that a compound is always composed of the same elements in the same proportions.

$$\text{percent by mass} = \frac{\text{mass of the element}}{\text{mass of the compound}} \times 100$$

- The law of multiple proportions states that if elements form more than one compound, those compounds will have compositions that are whole-number multiples of each other.

- element
- periodic table
- compound
- law of definite proportions
- percent by mass
- law of multiple proportions

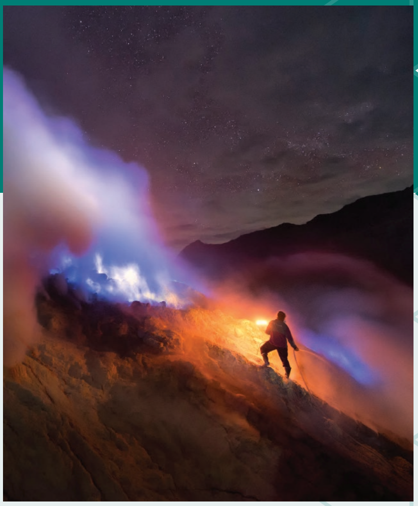
Lesson 4 MIXTURES OF MATTER

- A mixture is a physical blend of two or more pure substances in any proportion.
- Solutions are homogeneous mixtures.
- Mixtures can be separated by physical means. Common separation techniques include filtration, distillation, sublimation, chromatography, and crystallization.

- mixture
- heterogeneous mixture
- homogeneous mixture
- solution
- filtration
- distillation
- sublimation
- chromatography
- crystallization

THREE-DIMENSIONAL THINKING

Module Wrap-Up



REVISIT THE PHENOMENON

Why is this volcano spewing bright blue flames?

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 is compressed gas released?

Tanks of compressed gases are common in chemistry laboratories. For example, nitrogen is often flowed over a reaction in progress to displace other gases that might interfere with the experiment. Given what you know about gases, explain how the release of compressed nitrogen is controlled.

CER Analyze and Interpret Data

The particles of gases are far apart, and gases tend to fill their containers—even if the container is a laboratory room. Tanks of compressed gas come from the supplier capped to prevent the gas from escaping. In the lab, a chemist or technician attaches a regulator to the tank and secures the tank to a stable fixture.

1. **Claim** Explain why the flow of a compressed gas must be controlled for practical and safe use.
2. **Evidence, Reasoning** Predict what would happen if the valve on a full tank of compressed gas were suddenly opened all the way or if the tank were accidentally punctured. Explain your prediction.


Monty Rakusen/Cultura/Getty Images

MODULE 3

THE STRUCTURE OF THE ATOM

ENCOUNTER THE PHENOMENON

What is matter made of?

GO ONLINE to play a video about the size of an atom.

SEP Ask Questions

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

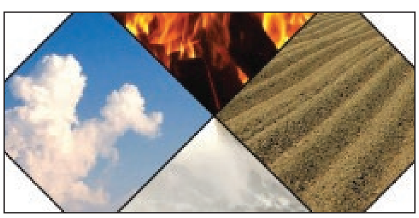
CER Claim, Evidence, Reasoning

Make Your Claim Use your CER chart to make a claim about what matter is made of.

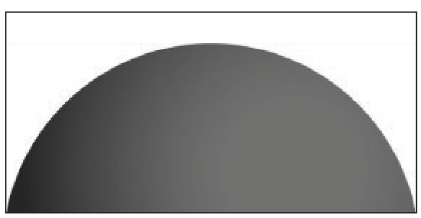
Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

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: Origins of Atomic Theory



LESSON 2: Explore & Explain: The Modern Model of the Atom

Smithsonian

Additional Resources

(l to r, t to b)Video Supplied by BBC Worldwide Learning; Weeping Willow Photography/Moment/Getty Images; McGraw-Hill Education; Tinke Hamming/Ingram Publishing; McGraw-Hill Education

LESSON 1

EARLY IDEAS ABOUT MATTER

FOCUS QUESTION

How has our understanding of matter changed over time?

The Roots of Atomic Theory

Science as we know it today did not exist several thousand years ago. No one knew what a controlled experiment was, and there were few tools for scientific exploration. In this setting, the power of the mind and intellectual thought were considered the primary avenues to the truth. Curiosity sparked the interest of scholarly thinkers known as philosophers who considered the many mysteries of life. As they speculated about the nature of matter, many of the philosophers formulated explanations based on their own life experiences.

Many of these philosphers concluded that matter was composed of things such as earth, water, air, and fire, as shown in **Figure 1**. It was also commonly accepted that matter could be endlessly divided into smaller and smaller pieces. While these early ideas were creative, there was no method available to test their validity, and so these ideas could not be shown to be correct.

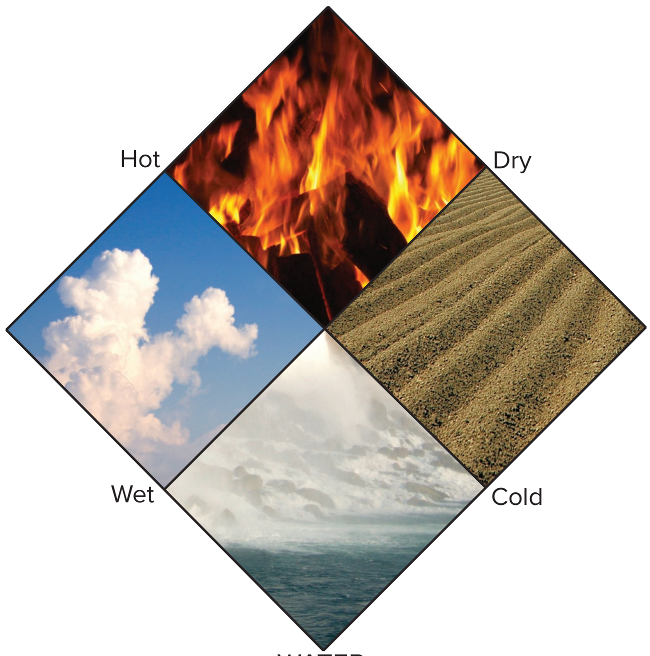


Figure 1 Many Greek philosophers thought that matter was composed of four elements: earth, air, water, and fire. They also associated properties with each element. The pairing of opposite properties such as hot and cold, and wet and dry, mirrored the symmetry they observed in nature. These early ideas were non-scientific.

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 chemical processes, their rates, and whether or not energy is conserved. Evaluate your source and communicate your findings to the class.

Table 1 Ancient Greek Ideas About Matter

Philosopher	Ideas
Democritus (460–370 B.C.)	• Matter is composed of atoms, which move through empty space. • Atoms are solid, homogeneous, indestructible, and indivisible. • Different kinds of atoms have different sizes and shapes. • Size, shape, and movement of atoms determine the properties of matter.
Aristotle (384–322 B.C.)	• Empty space cannot exist. • Matter is made of earth, fire, and water.

Democritus

The Greek philosopher Democritus (460–370 B.C.) was the first person to propose the idea that matter was not infinitely divisible. He believed matter was made up of tiny particles called atomos, from which the English word atom is derived. Democritus believed that *atoms* could not be created, destroyed, or further divided. Democritus and a summary of his ideas are shown in **Table 1**.

While a number of Democritus's ideas do not agree with modern atomic theory, his belief in the existence of atoms was amazingly ahead of his time. However his ideas were met with criticism from other philosophers who asked, "What holds the atoms together?" Democritus could not answer this question.

Aristotle

Other criticisms came from Aristotle (384–322 B.C.), one of the most influential Greek philosophers. He rejected the notion of atoms because it did not agree with his own ideas about nature. One of Aristotle's major criticisms concerned the idea that atoms moved through empty space. He did not believe that empty space could exist. His ideas are also presented in **Table 1**. Because Aristotle was one of the most influential philosophers of his time, Democritus's atomic theory was eventually rejected.

In fairness to Democritus, it was impossible for him or anyone else of his time to determine what held atoms together. More than two thousand years would pass before scientists would know the answer. However, it is important to realize that Democritus's ideas were just that—ideas, not science. Without the ability to conduct controlled experiments, Democritus could not identify patterns, test the validity of his ideas, and collect evidence to support his hypotheses.

Unfortunately for the advancement of science, Aristotle was able to gain wide acceptance for his ideas on nature—ideas that denied the existence of atoms. Incredibly, the influence of Aristotle was so great and the development of science so primitive that his denial of the existence of atoms went largely unchallenged for two thousand years!

STEM CAREER Connection

Chemistry Professor

If you like spending time with people and have a passion for chemistry and research, professors spend their time instructing students at a wide variety of colleges, universities, and community colleges. In addition, they conduct research and publish scholarly papers and books. You will need a master's degree or Ph.D. depending on the institution's requirements.

CCC CROSSCUTTING CONCEPTS

Patterns Examine the contributions of Democritus, Aristotle, and Dalton. In groups of two, make a poster that compares and contrasts any methods that were used or patterns that were identified that provided evidence in explaining the nature of matter and the atom.

Table 2 Dalton's Atomic Theory

Scientist	Ideas
Dalton (1766–1844)	• Matter is composed of extremely small particles called atoms. • Atoms are indivisible and indestructible. • Atoms of a given element are identical in size, mass, and chemical properties. • Atoms of a specific element are different from those of another element. • Different atoms combine in simple whole-number ratios to form compounds. • In a chemical reaction, atoms are separated, combined, or rearranged.

John Dalton

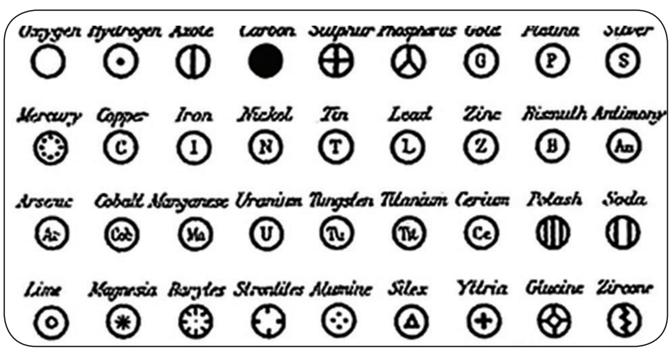
Although the concept of the atom was revived in the eighteenth century, it took another hundred years before significant progress was made. The work done in the nineteenth century by John Dalton (1766–1844), a schoolteacher in England, marks the beginning of the development of modern atomic theory. Dalton revived and revised Democritus's ideas based on the results of scientific research he conducted. In many ways, Democritus's and Dalton's ideas are similar.

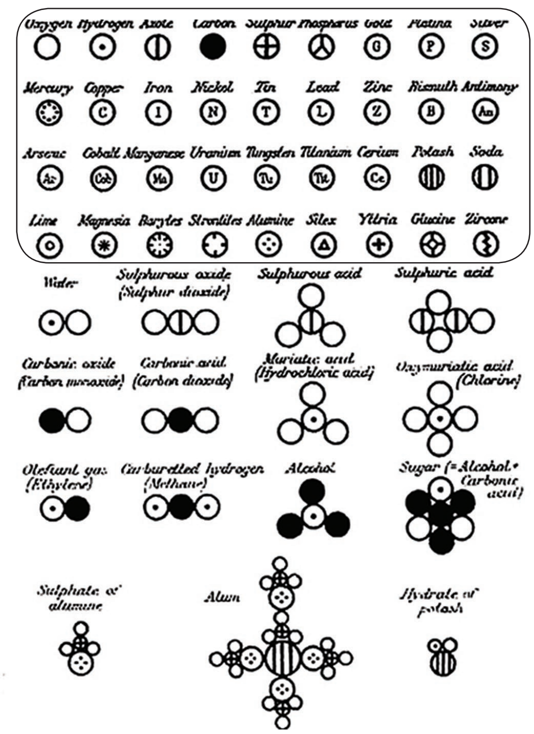
Thanks to advancements in science since Democritus's day, Dalton was able to perform experiments that allowed him to identify patterns that refined previous understandings and developed new explanations that were supported by his research. He studied numerous chemical reactions, making careful observations and measurements along the way. He was able to determine the mass ratios of the elements involved in those reactions. The mass ratio describes how many atoms of each element make up a compound. The results of his research are known as **Dalton's atomic theory,** which he proposed in 1803. The main points of his theory are summarized in **Table 2**. Dalton published his ideas in a book, an extract of which is shown in **Figure 2**.

Get It?

Explain how Dalton built upon the ideas that Democritus had about matter.

Figure 2 In his book, *A New System of Chemical Philosophy*, John Dalton presented his symbols for the elements known at that time and their possible combinations.





Science Source

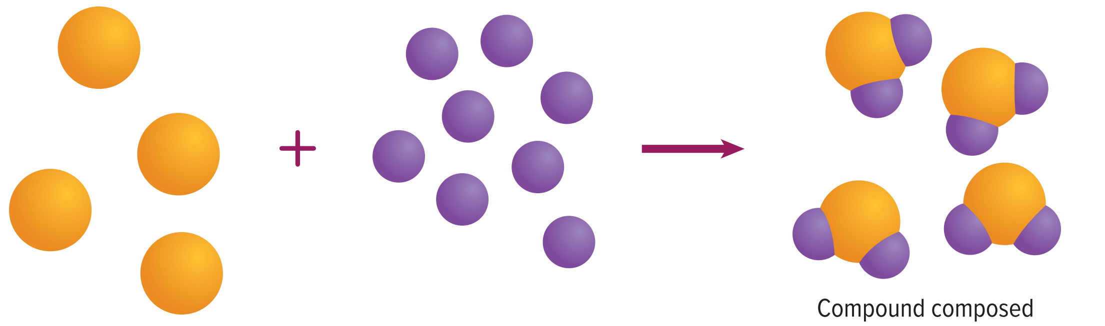


Figure 3 When atoms of two or more elements combine to form a compound, the number of atoms of each element is conserved. Thus, the mass is conserved as well.

Conservation of mass

The law of conservation of mass states that mass is conserved in any process. Dalton's atomic theory explains that the conservation of mass in chemical reactions is the result of the rearrangement of atoms during the reaction. They are not created or destroyed. The formation of a compound by combining elements is shown in **Figure 3.** Notice that the mass ratio for the compound consists of 1 atom of element A and 2 atoms of element B, a 1:2 ratio. The total number of atoms of each element is the same before and after the reaction.

Dalton's convincing experimental evidence and clear explanation of the composition of compounds, and conservation of mass, led to the general acceptance of his atomic theory. However, not all of Dalton's theory was accurate. As our undertanding of atoms evolved, new information was learned, leading to the revision and improvement of Dalton's atomic theory by later scientists.

Check Your Progress

Summary

- Democritus was the first person to propose the existence of atoms.
- According to Democritus, atoms are solid, homogeneous, and indivisible.
- Aristotle did not believe in the existence of atoms.
- Dalton revised the ideas of Democritus based on the results of scientific research.

Demonstrate Understanding

1. **Summarize** how matter was described by many early Greek philosophers.
2. **Define** *atom* using your own words.
3. **Summarize** Dalton's atomic theory.
4. **Explain** the relationship between Dalton's atomic theory and the conservation of mass in chemical reactions.
5. **Design** a concept map that compares and contrasts how Democritus, Aristotle, and Dalton viewed matter.

LESSON 2
DEFINING THE ATOM

FOCUS QUESTION

What does an atom look like?

The Atom

Many experiments since Dalton's time have proven that atoms do exist. So what exactly is the definition of an atom? To answer this question, consider a gold ring. Suppose you decide to grind the ring down into a pile of gold dust. Each fragment of gold dust still retains all of the properties of gold. If it were possible—which it is not without special equipment—you could continue to divide the gold dust particles into still smaller particles. Eventually, you would encounter a particle that could not be divided any further and still retain the properties of gold. This smallest particle of matter that retains the properties of the element is called an **atom.**

To get an idea of its size, consider the population of the world, which was about 6.5×10^9 in 2006. By comparison, a typical solid-copper penny contains 2.9×10^{22} atoms, almost five trillion times the world population! The diameter of a single copper atom is 1.28×10^{-10} m. Placing 6.5×10^9 copper atoms side by side would result in a line of copper atoms less than 1 m long. **Figure 4** illustrates another way to visualize the size of an atom. Imagine that you increase the size of an atom to be as big as an orange. To keep the proportions between the real sizes of the atom and the orange, you would have to increase the size of the orange and make it as big as Earth. This illustrates how small atoms are. It was because of this that scientists could not see the atom they were attempting to model.

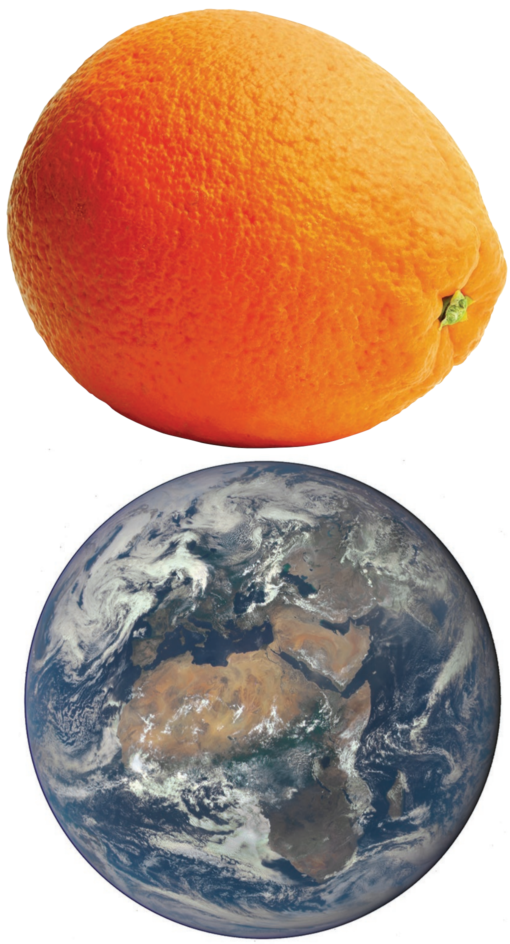


Figure 4 Imagine that you could increase the size of an atom to make it as big as an orange. At this new scale, an orange would be as big as Earth.

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—The Early Atom
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.

Revisit the Encounter the Phenomenon Question
What information from this lesson can help you answer the module question?

You might think that because atoms are so small, there would be no way to see them. However, an instrument called the scanning tunneling microscope (STM) allows individual atoms to be seen. An STM works as follows: a fine point is moved above a sample and the interaction of the point with the superficial atoms is recorded electronically. **Figure 5** illustrates how individual atoms look when observed with an STM. Scientists are now able to move individual atoms around to form shapes, patterns, and even simple machines. This capability has led to the exciting new field of nanotechnology. The promise of nanotechnology is molecular manufacturing—the atom-by-atom building of machines the size of molecules.

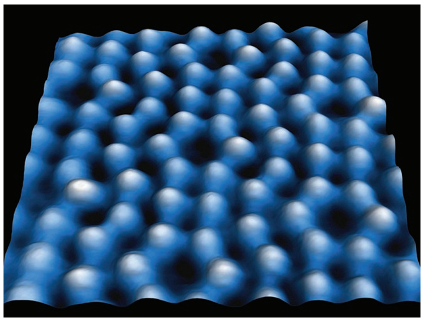


Figure 5 This image, recorded with an STM, shows the individual atoms of a fatty acid on a graphite surface. The false colors were added later on to improve the contrast between each atom.

The Electron

Once scientists were convinced of the existence of atoms, a new set of questions emerged. What is an atom like? Is the composition of an atom uniform throughout, or is it composed of still-smaller particles? Although many scientists researched the atom in the 1800s, it was not until almost 1900 that some of these questions were answered. The development of new technologies allowed scientists to develop models that were used to simulate systems to determine interactions within and between systems at different scales.

The cathode-ray tube

As scientists tried to unravel the atom, they began to make connections between matter and electric charge. Scientist learned, through experimentation, using models to simulate systems, that the structure and interactions of matter are determined by electrical forces. With the help of the newly invented vacuum pump, they passed electricity through glass tubes from which most of the air had been removed. Such tubes are called cathode-ray tubes and were invented and investigated by scientists from 1869–1875.

A typical cathode-ray tube used by researchers for studying the relationship between mass and charge is illustrated in **Figure 6**. Note that metal electrodes are located at opposite ends of the tube. The electrode connected to the negative terminal of the battery is called the cathode, and the electrode connected to the positive terminal is called the anode.

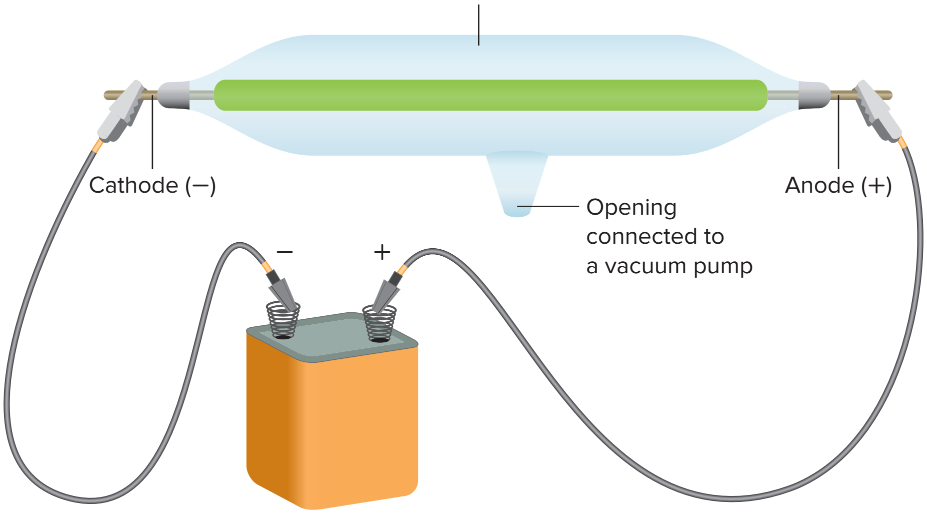
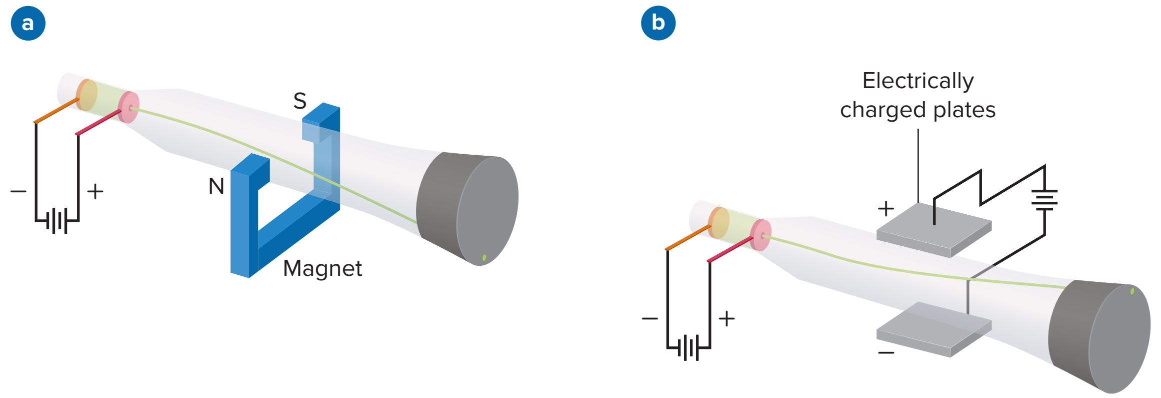


Figure 6 A cathode-ray tube is a tube with an anode at one end and a cathode at the other end. When a voltage is applied, electricity travels from the cathode to the anode.



Because the cathode ray is deflected in a magnetic field, the particles in the ray must be charged.

Because the cathode ray is deflected toward the positively charged plate by an electric field, the particles in the ray must have a negative charge.

Figure 7 A tiny hole located at the center of the anode produces a thin beam of electrons. A phosphor coating allows the position of the beam to be determined as it strikes the end of the tube.

Sir William Crookes While working in a darkened laboratory, English physicist Sir William Crookes noticed a flash of light within one of the cathode-ray tubes. A green flash was produced by some form of radiation striking a zinc-sulfide coating that had been applied to the end of the tube. Further work showed that there was a ray (radiation) going through the tube. This ray, originating from the cathode and traveling to the anode, was called a **cathode ray.** The accidental discovery of the cathode ray led to the invention of television, which is nothing more than a cathode-ray tube. Scientists continued their research using cathode-ray tubes, and they were fairly convinced by the end of the 1800s that cathode rays were a stream of charged particles and that these particles carried a negative charge. The exact value of the negative charge was not yet known.

Because changing the metal that makes up the electrodes or varying the gas (at very low pressure) in the cathode-ray tube did not affect the cathode ray produced, researchers concluded that the ray's negative particles were found in all forms of matter. These negatively charged particles that are part of all forms of matter are now known as **electrons.** Some of the experiments used to determine the properties of the cathode ray are shown in **Figure 7**.

Mass and charge of the electron

In spite of the progress made from all of the cathode-ray tube experiments, no one succeeded in determining the mass of a single cathode-ray particle. Unable to measure the particle's mass directly, English physicist J. J. Thomson (1856–1940) began a series of cathode-ray tube experiments at Cambridge University in the late 1890s to determine the ratio of its charge to its mass.

Charge-to-mass ratio By carefully measuring the effects of both magnetic and electric fields on a cathode ray, Thomson was able to determine the charge-to-mass ratio of the charged particle. He then compared that ratio to other known ratios.

Thomson concluded that the mass of the charged particle was much less than that of a hydrogen atom, the lightest known atom. The conclusion was shocking because it meant there were particles smaller than the atom. In other words, Dalton had been incorrect—atoms were divisible into smaller subatomic particles. Because Dalton's atomic theory had become so widely accepted and Thomson's conclusion was so revolutionary, many other scientists found it hard to accept this new discovery. But Thomson was correct. He had identified the first subatomic particle—the electron. He received a Nobel Prize in 1906 for this discovery.

Get It?

Summarize how Crookes and Thomson contributed to our understanding of the nature of the electron.

The oil-drop experiment and the charge of an electron The next significant development came in the early 1910s, when the American physicist Robert Millikan (1868–1953) determined the charge of an electron using the oil-drop apparatus shown in **Figure 8**. In this apparatus, oil is sprayed into the chamber above the two parallel charged plates. The top plate has a small hole through which the oil drops. X-rays knock out electrons from the air particles between the plates and the electrons stick to the droplets, giving them a negative charge. By varying the intensity of the electric field, Millikan could control the rate of a droplet's fall. He determined that the magnitude of the charge on each drop increased in discrete amounts and determined that the smallest common denominator was 1.602×10^{-19} coulombs. He identified this number as the charge of the electron. This charge was later equated to a single unit of negative charge noted 1–; in other words, a single electron carries a charge of 1–. So good was Millikan's experimental setup and technique that the charge he measured almost one hundred years ago is within 1% of the currently accepted value.

Mass of an electron Knowing the electron's charge and using the known charge-to-mass ratio, Millikan calculated the mass of an electron. The equation below shows how small the mass of an electron is.

$$\text{Mass of an electron} = 9.1 \times 10^{-28}\ \text{g} = \frac{1}{1840}\ \text{the mass of a hydrogen atom}$$

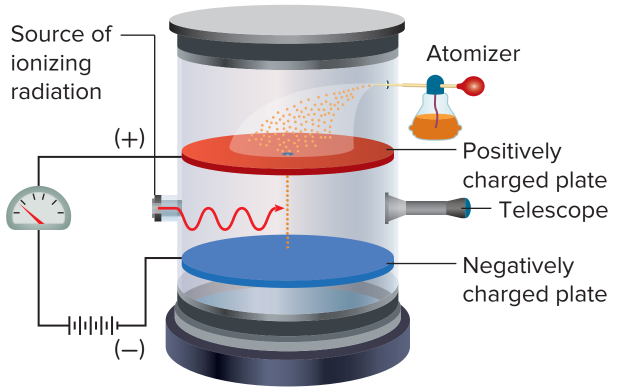


Figure 8 The motion of oil droplets within Millikan's apparatus depends on the charge of droplets on the electric field. Millikan observed the droplets with the telescope. He could make droplets fall more slowly, rise, or pause has he varied the strength of the electric field. From his observations, he calculated the charge on each droplet.

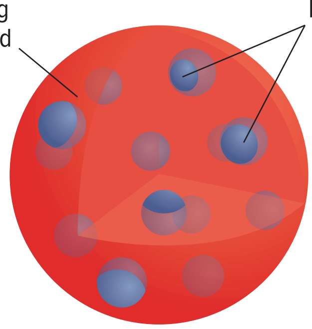


Figure 9 J.J. Thomson's plum pudding model of the atom states that the atom is a uniform, positively charged sphere containing electrons.

The plum pudding model

The existence of the electron and the knowledge of some of its properties raised some interesting new questions about the nature of atoms. It was known that matter is neutral—it has no electric charge. You know that matter is neutral from everyday experience: you do not receive an electric shock (except under certain conditions) when you touch an object. If electrons are part of all matter and they possess a negative charge, how can all matter be neutral? Also, if the mass of an electron is so small, what accounts for the rest of the mass in a typical atom?

In an attempt to answer these questions, J. J. Thomson proposed a model of the atom that became known as the plum pudding model. As you can see in **Figure 9**, Thomson's model consisted of a spherically shaped atom composed of a uniformly distributed positive charge in which the individual negatively charged electrons resided. However, as you are about to read, Thomson's plum pudding model of the atom did not last for very long.

The Nucleus

In 1911, Ernest Rutherford (1871–1937) began to study how positively charged alpha particles (radioactive particles you will read more about later in this module) interacted with solid matter. With a small group of scientists, Rutherford conducted an experiment to see if alpha particles would be deflected as they passed through a thin gold foil. If they were deflected, by how much would they deflect?

Rutherford's experiment

In the experiment, a narrow beam of alpha particles was aimed at a thin sheet of gold foil. A zinc-sulfide coated screen surrounding the gold foil produced a flash of light when struck by an alpha particle. By noting where the flashes occurred, the scientists could determine if the atoms in the gold foil deflected the alpha particles. As with any good experimental design, the procedure was repeated over and over again to ensure that all observations were taken and that all results observed were repeated in each trial. As the number of times an experiment is repeated increases, and the same results are obtained, the results are easier for scientists to accept.

Rutherford was aware of Thomson's plum pudding model of the atom. He expected the paths of the massive and fast-moving alpha particles to be only slightly altered by a collision with an electron. And because the positive charge within the gold atoms was thought to be uniformly distributed, he thought it would not alter the paths of the alpha particles, either. **Figure 10** shows the results Rutherford expected from the experiment.

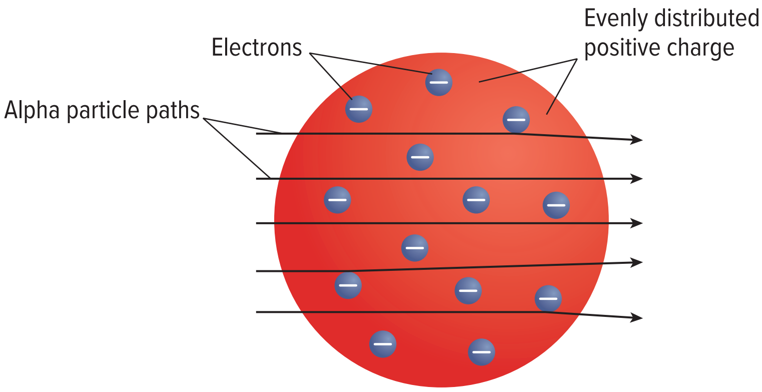


Figure 10 Based on Thomson's model, Rutherford expected the alpha particles to pass through the gold atoms.

The actual results observed by Rutherford and his colleagues are shown in **Figure 11**. Some alpha particles were deflected at various angles. Several particles were deflected straight back toward the source. Rutherford likened the results to firing a large artillery shell at a sheet of paper and the shell coming back at the cannon.

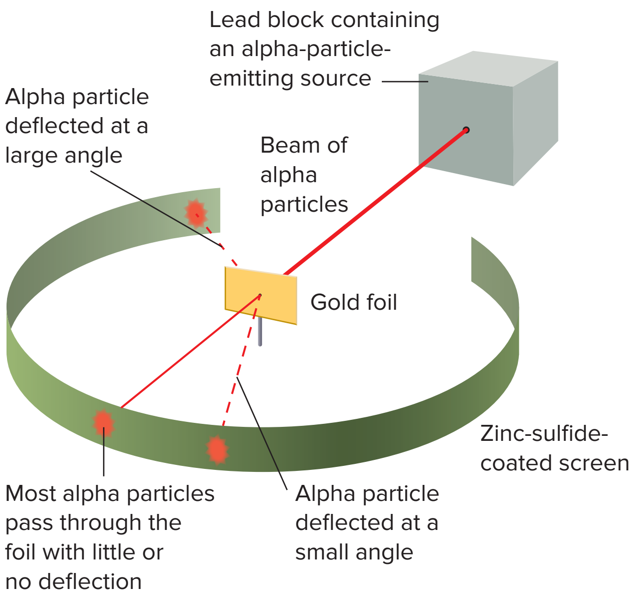


Figure 11 During Rutherford's experiment, a beam of alpha particles bombarded a thin gold foil. Most of the alpha particles went through the gold foil. However, a few of them bounced back, some at large angles.

Rutherford's model of the atom Rutherford concluded that the plum pudding model was incorrect because it could not explain the results of the gold foil experiment. Considering the properties of the alpha particles and the electrons, and the frequency of the deflections, he calculated that an atom consisted mostly of empty space through which the electrons move. He also concluded that almost all of the atom's positive charge and almost all of its mass were contained in a tiny, dense region in the center of the atom, which he called the **nucleus.**

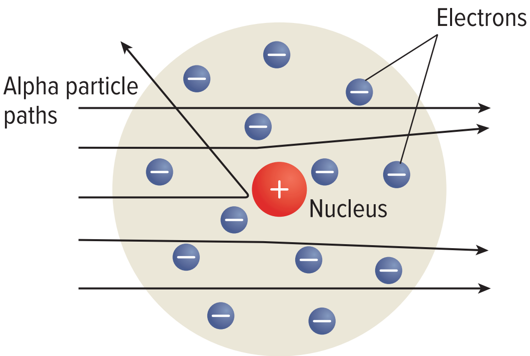


Figure 12 In Rutherford's nuclear model, the atom is composed of a dense, positively charged nucleus that is surrounded by negative electrons. Alpha particles passing far from the nucleus are only slightly deflected, if at all. Those passing closer to the nucleus are deflected more. Alpha particles directly approaching the nucleus are deflected at large angles.

The negatively charged electrons are held within the atom by their attraction to the positively charged nucleus. Rutherford's nuclear atomic model is shown in **Figure 12**.

Because the nucleus occupies such a small space and contains most of an atom's mass, it is incredibly dense. If a nucleus were the size of the dot in an exclamation point, its mass would be amost as much as 70 automobiles! The volume of space through which the electrons move is huge compared to the volume of the nucleus. If an atom had a diameter of two football fields, the nucleus would be the size of a nickel.

Get It?

Describe how the photograph at the beginning of the module is related to Rutherford's ideas about what atoms are made of.

The repulsive force produced between the positive nucleus and the positive alpha particles causes the deflections. **Figure 12** illustrates how Rutherford's nuclear atomic model explained the results of the gold foil experiment. The nuclear model also explains the neutral nature of matter: the positive charge of the nucleus balances the negative charge of the electrons. However, the model still could not account for all of the atom's mass.

The proton and the neutron

By 1920, Rutherford concluded that the nucleus contained positively charged particles called protons. A **proton** is a subatomic particle carrying a charge equal to but opposite that of an electron. In 1932, Rutherford's coworker, James Chadwick (1891–1974), showed that the nucleus also contained another subatomic neutral particle, called the neutron. A **neutron** is a subatomic particle that has a mass nearly equal to that of a proton, but carries no electric charge.

Completing the model of the atom All atoms have a charged substructure consisting of a small, dense nucleus, which is made of protons and neutrons, surrounded by one or more electrons. **Figure 13,** on the next page, summarizes the history and development of modern atomic theory.

SCIENCE USAGE V. COMMON USAGE

neutral

Science usage: to have no electric charge

Neutrons have a charge of zero. They are neutral particles.

Common usage: not engaged in either side

Switzerland remained neutral during World War II.

CCC CROSSCUTTING CONCEPTS

System and System Models Review the system models used by Crookes, Thomson, Millikan, Rutherford, and Chadwick that led to the development of the completed model of the atom. Develop a time line describing their model simulations, instruments, and contributions to atomic theory.

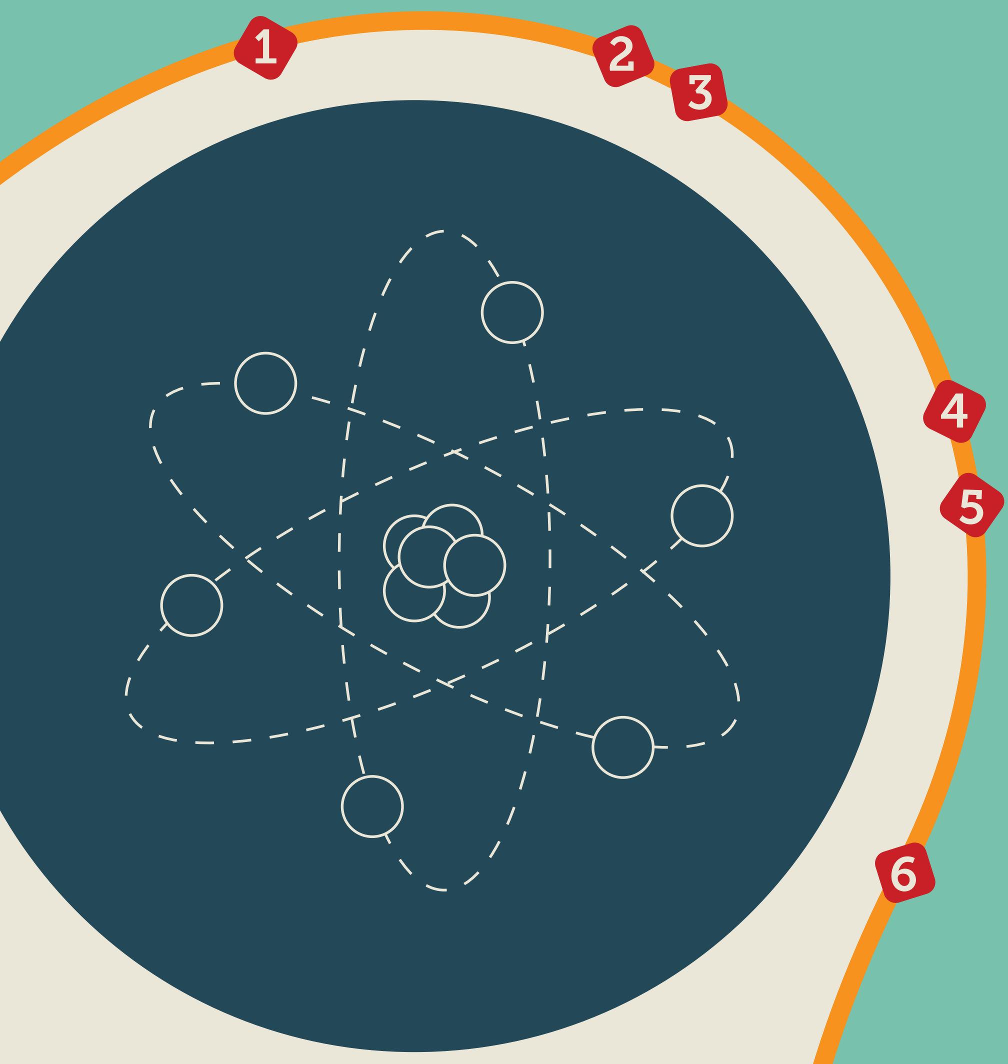


Figure 13

Development of Modern Atomic Theory

Current understanding of the properties and behavior of atoms and subatomic particles is based on the work of scientists worldwide during the past two centuries.

1 **1897** Using cathode-ray tubes, J.J. Thomson identifies the electron and determines the ratio of the mass of an electron to its electric charge.

2 **1911** With the gold foil experiment, Ernest Rutherford determines properties of the nucleus, including charge, relative size, and density.

3 **1913** Niels Bohr publishes a theory of atomic structure relating the electron arrangement in atoms to atomic chemical properties.

4 **1932** James Chadwick proves the existence of neutrons.

5 **1938** Lise Meitner, Otto Hahn, and Fritz Straussman split uranium atoms in a process they called fission.

6 **1954** CERN, the world's largest nuclear physics research center, located in Switzerland, is founded to study particle physics.

7 **1968** Scientists provide the first experimental evidence for subatomic particles known as quarks.

8 **2010–2013** The Large Hadron Collider at CERN smashes beams of protons at the highest energy ever, allowing scientists to find a particle consistent with the predicted properties of the Higgs boson.

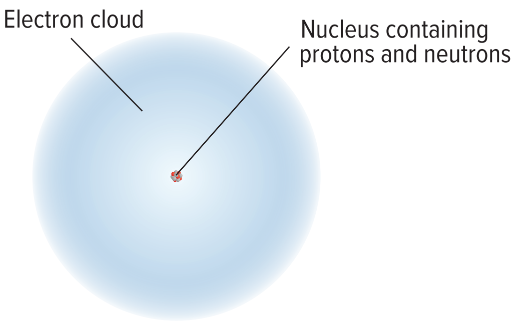


Figure 14 Atoms are composed of a nucleus containing protons and neutrons, and surrounded by a cloud of electrons.

Table 3 Properties of Subatomic Particles

Particle	Symbol	Relative Electric Charge	Relative Mass	Actual Mass (g)
Electron	e^-	1−	$\frac{1}{1840}$	9.11×10^{-28}
Proton	p	1+	1	1.673×10^{-24}
Neutron	n	0	1	1.675×10^{-24}

Most of an atom consists of fast-moving electrons traveling through empty space around the nucleus. The electrons are held within the atom by their attraction to the positively charged nucleus. The nucleus, which is composed of neutral neutrons (hydrogen's single-proton nucleus is an exception) and positively charged protons, contains more than 99.97% of an atom's mass. It occupies only about one ten-thousandth of the volume of the atom. Because an atom is electrically neutral, the number of protons in the nucleus equals the number of electrons around the nucleus. The typical atom is shown in **Figure 14**, and the properties of the subatomic particles are summarized in **Table 3.**

Scientists have determined that protons and neutrons have their own structures. They are composed of subatomic particles called quarks. These particles will not be covered in this textbook because scientists do not yet understand if or how they affect chemical behavior. Chemical behavior can be explained by considering only an atom's electrons.

Check Your Progress

Summary

- An atom is the smallest particle of an element that maintains the properties of that element.
- Electrons have a 1− charge, protons have a 1+ charge, and neutrons have no charge.
- An atom consists mostly of empty space surrounding the nucleus.

Demonstrate Understanding

6. **Describe** the structure of a typical atom. Identify the location and charge of each subatomic particle.
7. **Compare and contrast** Thomson's plum pudding atomic model with Rutherford's nuclear atomic model.
8. **Describe** the research, including instrumentation, which led to the conclusion that electrons are negatively charged particles found in all matter.
9. **Compare** the relative charge and mass of each subatomic particle.
10. **Calculate** What is the difference expressed in kilograms between the mass of a proton and the mass of an electron?

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LESSON 3
HOW ATOMS DIFFER

FOCUS QUESTION

Are all atoms identical?

Atomic Number

There are more than 110 different elements on the periodic table of the elements. What makes an atom of one element different from an atom of another element? After Rutherford's gold foil experiment, the English scientist Henry Moseley (1887–1915) discovered that an element contains a unique positive charge in its nucleus. Thus, the number of protons in an atom identifies it as an atom of a particular element. The number of protons in an atom is referred to as the **atomic number.** Hydrogen's square on the periodic table is shown in **Figure 15**. The number 1 to the left of the symbol is the atomic number. Moving across the periodic table to the right, you will next come to helium (He). It has two protons in its nucleus; thus, it has an atomic number of 2. The next row begins with lithium (Li), atomic number 3, and so on. The periodic table is organized left-to-right and top-to-bottom by increasing atomic number. The pattern of repeating order is called periodicity, which accounts for the name given to this table.

Elements are listed in their neutral state on the periodic table. Therefore, once you know the atomic number of an element, you know the number of protons and electrons it contains. Lithium, with an atomic number of 3, has three protons and three electrons. Thus, the periodicity of the elements in the periodic table provides a clue as to how the structure of each element differs.

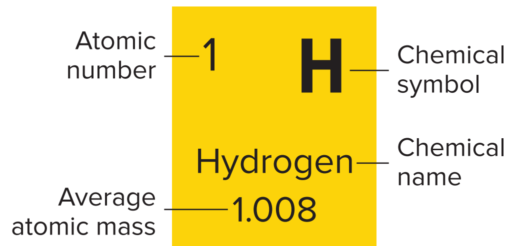


Figure 15 In the periodic table, each element is represented by its chemical name, atomic number, chemical symbol, and average atomic mass.

Atomic number

$$\begin{aligned}\text{atomic number} &= \text{number of protons}\\ &= \text{number of electrons}\end{aligned}$$

The atomic number of an atom equals its number of protons and its number of 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.

Virtual Investigation: Decoding the Periodic Table
Analyze and interpret data to identify the patterns in the structure and properties of matter.

Quick Investigation: Model Isotopes
Plan and carry out an investigation to create a model of stable forms of matter.

EXAMPLE Problem 1

ATOMIC NUMBER Complete the following table.

Composition of Several Elements

	Element	Atomic Number	Protons	Electrons
a.	Pb	82		
b.			8	
c.				30

1 ANALYZE THE PROBLEM

Apply the relationship among atomic number, number of protons, and number of electrons to complete most of the table. Then, use the periodic table to identify the element.

Known

a. element = Pb, atomic number = 82

b. number of protons = 8

c. number of electrons = 30

Unknown

a. number of protons (N_p), number of electrons (N_e) = ?

b. element, atomic number (Z), N_e = ?

c. element, Z, N_p = ?

2 SOLVE FOR THE UNKNOWN

a. number of protons = atomic number — Apply the atomic-number relationship.

$N_p = 82$ — Substitute atomic number = 82.

number of electrons = number of protons

$N_e = 82$

The number of protons and the number of electrons is 82.

b. atomic number = number of protons — Apply the atomic-number relationship.

$Z = 8$ — Substitute number of protons = 8.

number of electrons = number of protons

$N_e = 8$

The atomic number and the number of electrons is 8. The **element** is **oxygen (O).** — Consult the periodic table to identify the element.

c. number of protons = number of electrons — Apply the atomic-number relationship.

$N_p = 30$ — Substitute number of electrons = 30.

atomic number = number of protons

$Z = 30$

The atomic number and the number of protons is 30. The **element** is **zinc (Zn).** — Consult the periodic table to identify the element.

3 EVALUATE THE ANSWER

The answers agree with atomic numbers and element symbols given in the periodic table.

PRACTICE Problems

ADDITIONAL PRACTICE

11. How many protons and electrons are in each atom?

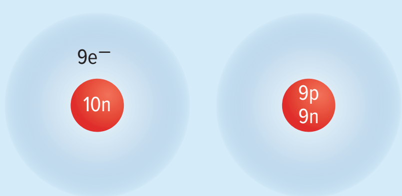
a. radon

b. magnesium

12. An atom of an element contains 66 electrons. Which element is it?

13. An atom of an element contains 14 protons. Which element is it?

14. **CHALLENGE** Do the atoms shown in the figure to the right have the same atomic number?



Isotopes and Mass Number

All atoms of an element have the same number of protons and electrons, but the number of neutrons might differ. For example, there are three types of potassium atoms that occur naturally and all three types contain 19 protons and 19 electrons. However, one type of potassium atom contains 20 neutrons, another 21 neutrons, and still another 22 neutrons. Atoms with the same number of protons but different numbers of neutrons are called **isotopes.**

Mass of isotopes While isotopes that contain more neutrons have a greater mass, all isotopes of an atom have the same chemical behavior. The **mass number** identifies each isotope of an element and is the sum of the atomic number (or number of protons) and neutrons in the nucleus.

For example, copper has two isotopes. The isotope with 29 protons and 34 neutrons has a mass number of 63 ($29 + 34 = 63$), and is called copper-63 (also written ^{63}Cu or Cu-63). The isotope with 29 protons and 36 neutrons is called copper-65. Chemists often write out isotopes using a notation shown in **Figure 16.**

Natural abundance of isotopes In nature, most elements are found as mixtures of isotopes. Usually, the relative abundance of each isotope is constant. Different sources containing atoms of potassium would have the same percent composition of potassium isotopes. The three potassium isotopes are summarized in **Figure 17**.

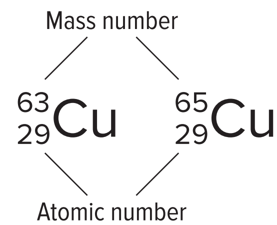


Figure 16 Cu is the chemical symbol for copper. Copper, which was used to make this Chinese gong, is composed of 69.2% copper-63 and 30.8% copper-65.

Mass number

$$\text{mass number} = \text{atomic number} + \text{number of neutrons}$$

The mass number of an atom is the sum of its atomic number and its number of neutrons.

	Potassium-39	Potassium-40	Potassium-41
Protons	19	19	19
Neutrons	20	21	22
Electrons	19	19	19

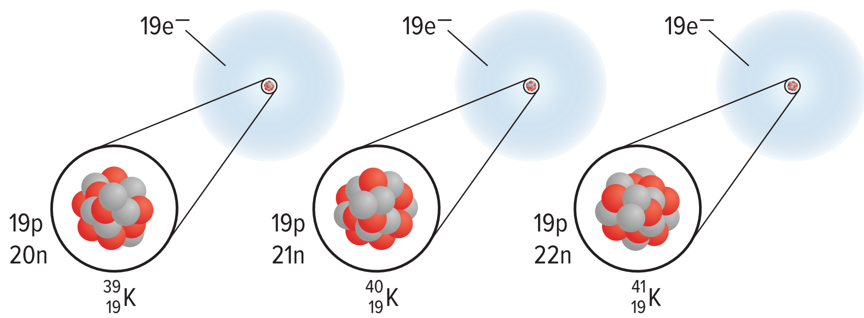


Figure 17 Potassium has three naturally occurring isotopes, potassium-39, potassium-40, and potassium 41. The table lists the number of protons, neutrons, and electrons in each potassium isotope.

EXAMPLE Problem 2

USE ATOMIC NUMBER AND MASS NUMBER A chemistry laboratory has analyzed the composition of isotopes of several elements. The composition data is given in the table below. Determine the number of protons, electrons, and neutrons in the isotope of neon. Name the isotope and give its symbol.

Isotope Composition Data

	Element	Atomic Number	Mass Number
a.	Neon	10	22
b.	Calcium	20	46
c.	Oxygen	8	17
d.	Iron	26	57
e.	Zinc	30	64
f.	Mercury	80	204

1 ANALYZE THE PROBLEM

You are given some data for neon in the table. The symbol for neon can be found on the periodic table. From the atomic number, the number of protons and electrons in the isotope are known. The number of neutrons in the isotope can be found by subtracting the atomic number from the mass number.

Known

element: neon

atomic number = 10

mass number = 22

Unknown

number of protons (N_p), electrons (N_e), and neutrons (N_n) = ?

name of isotope = ?

symbol for isotope = ?

2 SOLVE FOR THE UNKNOWN

number of protons = atomic number = 10

number of electrons = atomic number = 10

Apply the atomic number relationship.

number of neutrons = mass number − atomic number

Use the atomic number and the mass number to calculate the number of neutrons.

$N_n = 22 - 10 = 12$

Substitute mass number = 22 and atomic number = 10

The name of the isotope is neon-22.

Use the element name and mass number to write the isotope's name.

The symbol for the isotope is $^{22}_{10}Ne$.

Use the chemical symbol, mass number, and atomic number to write out the isotope in symbolic notation form.

3 EVALUATE THE ANSWER

The relationships among number of electrons, protons, and neutrons have been applied correctly. The isotope's name and symbol are in the correct format. Refer to the Elements Handbook to learn more about neon.

PRACTICE Problems

ADDITIONAL PRACTICE

15. Determine the number of protons, electrons, and neutrons for isotopes **b.—f.** in the table above. Name each isotope, and write its symbol.

16. **CHALLENGE** An atom has a mass number of 55. Its number of neutrons is the sum of its atomic number and five. How many protons, neutrons, and electrons does this atom have? What is the identity of this atom?

Mass of Atoms

Recall from **Table 3** that the masses of both protons and neutrons are approximately 1.67×10^{-24} g. While this is a small mass, the mass of an electron is even smaller—only about 1/1840 that of a proton or a neutron.

Table 4 Masses of Subatomic Particles

Particle	Mass (amu)
Electron	0.000549
Proton	1.007276
Neutron	1.008665

Atomic mass unit Because these extremely small masses expressed in scientific notation are difficult to work with, chemists have developed a method of measuring the mass of an atom relative to the mass of a specific atomic standard. That standard is the carbon-12 atom. Scientists assigned the carbon-12 atom a mass of exactly 12 atomic mass units. Thus, one **atomic mass unit (amu)** is defined as one-twelfth the mass of a carbon-12 atom. Although a mass of 1 amu is nearly equal to the mass of a single proton or a single neutron, it is important to realize that the values are slightly different. **Table 4** gives the masses of the subatomic particles in terms of amu.

Atomic mass Because an atom's mass depends mainly on the number of protons and neutrons it contains, and because protons and neutrons have masses close to 1 amu, you might expect the atomic mass of an element to always be nearly a whole number. However, this is often not the case. The explanation involves how atomic mass is defined. The **atomic mass** of an element is the weighted average mass of the isotopes of that element. Because isotopes have different masses, the weighted average is not a whole number. The calculation of the atomic mass of chlorine is illustrated in **Figure 18**.

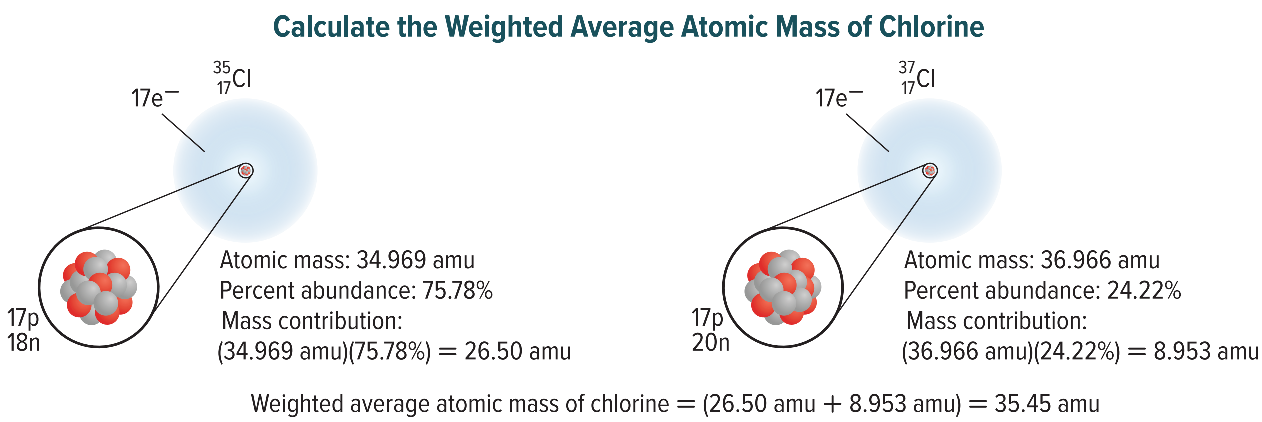


Figure 18 To calculate the weighted average atomic mass of chlorine, you first need to calculate the mass contribution of each isotope.

ACADEMIC VOCABULARY

specific

characterized by a precise formulation or accurate restriction; Some diseases have specific symptoms.

CCC CROSSCUTTING CONCEPTS

Patterns Identify the pattern found in the periodic table as you move horizontally across the table from potassium to iron by plotting the atomic number (*x*-axis) vs. the atomic mass (*y*-axis). Describe the trend in your graph, including an explanation of how the pattern of electrons, neutrons, and protons accounts for this trend.

Chlorine exists naturally as a mixture of about 76% chlorine-35 and 24% chlorine-37. It has an atomic mass of 35.453 amu. Because atomic mass is a weighted average, the chlorine-35 atoms, which exist in greater abundance than the chlorine-37 atoms, have a greater effect in determining the atomic mass. The atomic mass of chlorine is calculated by multiplying each isotope's percent abundance by its atomic mass and then adding the products. The process is similar to calculating an average grade. You can calculate the atomic mass of any element if you know the number of naturally occurring isotopes, their masses, and their percent abundances.

Get It?

Infer how atomic mass changes from element to element as you move horizontally across the periodic table.

Isotope abundances Analyzing an element's mass can indicate the most abundant isotope for that element. For example, fluorine (F) has an atomic mass that is extremely close to 19 amu. If fluorine had several fairly abundant isotopes, its atomic mass would not likely be so close to a whole number. Thus, you might conclude that all naturally occurring fluorine is probably in the form of fluorine-19 $\left(^{19}_{9}\text{F}\right)$. Indeed, 100% of naturally occurring fluorine is in the form of fluorine-19. While this type of reasoning generally works well, it is not foolproof. Consider bromine (Br). It has an atomic mass of 79.904 amu. With a mass so close to 80 amu, it seems likely that the most common bromine isotope would be bromine-80. However, bromine's two isotopes are bromine-79 (78.918 amu, 50.69%) and bromine-81 (80.917 amu, 49.31%). There is no bromine-80 isotope.

Figure 19 Bromine is extracted from sea water and salt lakes. The Dead Sea area of Israel is one of the major bromine production sites in the world. Applications of bromine include microbe and algae control in swimming pools. It is also used in medicines, oils, paints, pesticides, and flame retardants.

Oleg Lopatkin/iStock/Getty Images

Figure 19 shows one of the major production sites of bromine, located in the Dead Sea area.

EXAMPLE Problem 3

CALCULATE ATOMIC MASS Given the data in the table, calculate the atomic mass of unknown Element X. Then, identify the unknown element, which is used medically to treat some mental disorders.

1 ANALYZE THE PROBLEM

Calculate the atomic mass and use the periodic table to confirm.

Known

^{6}X: mass = 6.015 amu

abundance = 7.59% = 0.0759

^{7}X: mass = 7.016 amu

abundance = 92.41% = 0.9241

Unknown

atomic mass of X = ? amu

element X = ?

Isotope Abundance for Element X

Isotope	Mass (amu)	Percent Abundance
^{6}X	6.015	7.59%
^{7}X	7.016	92.41%

EXAMPLE Problem 3 (Continued)

2 SOLVE FOR THE UNKNOWN

^{6}X: mass contribution = (mass)(percent abundance)	Calculate ^{6}X's contribution.
mass contribution = (6.015 amu)(0.0759) = 0.4565 amu	Substitute mass = 6.015 amu and abundance = 0.0759.
^{7}X: mass contribution = (mass)(percent abundance)	Calculate ^{7}X's contribution.
mass contribution = (7.016 amu)(0.9241) = 6.483 amu	Substitute mass = 7.016 amu and abundance = 0.9241.
atomic mass of X = (0.4565 amu + 6.483 amu) = 6.939 amu	Total the mass contributions to find the atomic mass.
The element with a mass nearest 6.939 amu is lithium (Li).	Identify the element using the periodic table.

3 EVALUATE THE ANSWER

The result agrees with the atomic mass given in the periodic table. The isotope masses have four significant figures, so the atomic mass needs to have four significant figures.

PRACTICE Problems

ADDITIONAL PRACTICE

17. Boron (B) has two naturally occurring isotopes: boron-10 (abundance = 19.8%, mass = 10.013 amu) and boron-11 (abundance = 80.2%, mass = 11.009 amu). Calculate the atomic mass of boron.

18. CHALLENGE Nitrogen has two naturally occurring isotopes, N-14 and N-15. Its atomic mass is 14.007. Which isotope is more abundant? Explain your answer.

Check Your Progress

Summary

- The atomic number of an atom is given by its number of protons. The mass number of an atom is the sum of its neutrons and protons.
- Atoms of the same element with different numbers of neutrons are called isotopes.
- The atomic mass of an element is a weighted average of the masses of all of its naturally occurring isotopes.

Demonstrate Understanding

19. **Recall** Which subatomic particle identifies an atom as belonging to a particular element?
20. **Describe** the difference between atomic number and mass number.
21. **Explain** how isotopes are related to why atomic masses are not whole numbers.
22. **Calculate** Copper has two isotopes Cu-63 (abundance = 69.2%, mass = 62.930 amu) and Cu-65 (abundance = 30.8%, mass = 64.928 amu). Calculate the atomic mass of copper.
23. **Explain** how patterns in the periodic table can provide evidence as to how the structures of elements differ.

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LESSON 4
UNSTABLE NUCLEI AND RADIOACTIVE DECAY

FOCUS QUESTION

How can atoms change?

Radioactivity

Nuclear reactions In the late 1890s, scientists noticed that some substances spontaneously emitted radiation in a process they named **radioactivity.** The rays and particles emitted by the radioactive material were called **radiation.** Scientists discovered that radioactive atoms undergo changes that can alter their identities. A reaction that involves a change in an atom's nucleus is called a **nuclear reaction.** Prior to this discovery, it was not known that a reaction could result in the formation of a new element. Radioactive atoms emit radiation because their nuclei are unstable. Unstable systems, whether they are atoms or people doing handstands, as shown in **Figure 20**, gain stability by losing energy.

Radioactive decay Unstable nuclei lose energy by emitting radiation in a spontaneous process called **radioactive decay.** Unstable atoms undergo radioactive decay until they form stable atoms, often of a different element. An atom can lose energy and reach a stable state when emitting radiation.

Figure 20 Being in a handstand position is an unstable state. Like unstable atoms, people doing handstands eventually return to a more stable state – standing on their feet – by losing potential energy.

TinnaPong/Shutterstock.com

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: Half-Life of Barium-137m

Construct an explanation to determine the stability and change of an atom experiencing spontaneous radioactive decay.

CCC **Identify Crosscutting Concepts**

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

Types of Radiation

Scientists began researching radioactivity in the late 1800s. By directing radiation from a radioactive source between two electrically charged plates, scientists were able to identify three different types of radiation based on their electric charge. As shown in **Figure 21**, radiation was deflected toward the negative plate, the positive plate, or not at all.

Alpha radiation The radiation that was deflected toward the negatively charged plate was named **alpha radiation.** It is made up of alpha particles. An **alpha particle** contains two protons and two neutrons, and thus has a 2+ charge, which explains why alpha particles are attracted to the negatively charged plate as shown in **Figure 21**. An alpha particle is equivalent to a helium-4 nucleus and is represented by ${}^{4}_{2}He$ or α. The alpha decay of radioactive radium-226 into radon-222 is shown below.

$$\underset{\text{radium-226}}{{}^{226}_{88}Ra} \rightarrow \underset{\text{radon-222}}{{}^{222}_{86}Rn} + \underset{\text{alpha particle}}{\alpha}$$

Note that a new element, radon (Rn), is created as a result of the alpha decay of the unstable radium-226 nucleus. The type of equation shown above is known as a **nuclear equation.** It shows the atomic numbers and mass numbers of the particles involved. Note that in a nuclear equation, although elements are not conserved, the total number of protons plus neutrons is conserved.

Beta radiation The radiation that was deflected toward the positively charged plate was named **beta radiation.** This radiation consists of fast-moving beta particles. Each **beta particle** is an electron with a 1– charge. The negative charge of the beta particle explains why it is attracted to the positively charged plate shown in **Figure 21**. Beta particles are represented by the symbol β or e^- . The beta decay of carbon-14 into nitrogen-14 is shown below. The beta decay of unstable carbon-14 results in the formation of the new element, nitrogen (N).

$$\underset{\text{carbon-14}}{{}^{14}_{6}C} \rightarrow \underset{\text{nitrogen-14}}{{}^{14}_{7}N} + \underset{\text{beta particle}}{\beta}$$

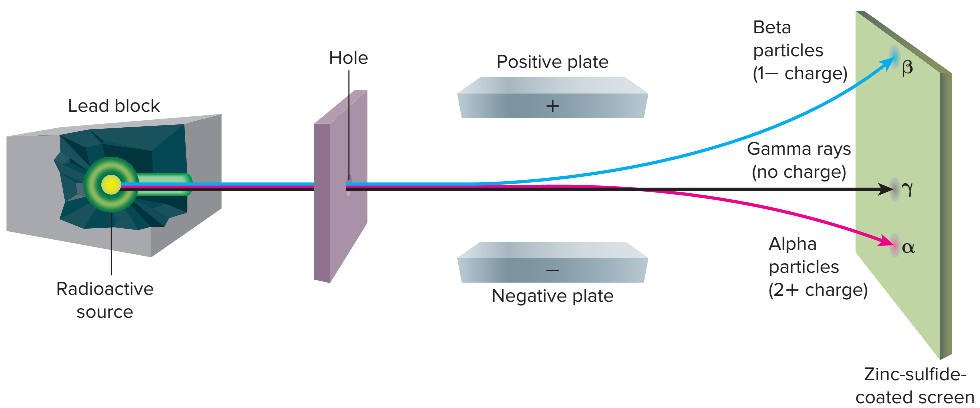


Figure 21 An electric field will deflect radiation in different directions, depending on the electric charge of the radiation.

Get It?
Compare and contrast alpha decay and beta decay.

Gamma radiation The third common type of radiation is called gamma radiation, or gamma rays. A **gamma ray** is high-energy radiation that has no mass and is denoted by the symbol γ. Gamma rays are neutral, and so are not deflected by electric or magnetic fields. They often accompany alpha and beta radiation, and account for most of the energy lost during radioactive decay. For example, gamma rays accompany the decay of uranium-238.

Table 5 Characteristics of Radiation

	Alpha	Beta	Gamma
Symbol	$^{4}_{2}He$ or α	e^- or β	γ
Mass (amu)	4	$\frac{1}{1840}$	0
Mass (kg)	6.65×10^{-27}	9.11×10^{-31}	0
Charge	2+	1–	0

$$\underset{\text{uranium-238}}{^{238}_{92}U} \rightarrow \underset{\text{thorium-234}}{^{234}_{90}Th} + \underset{\text{alpha particle}}{\alpha} + \underset{\text{gamma rays}}{2\gamma}$$

Because gamma rays are massless, the emission of gamma rays by themselves cannot result in the formation of a new element. **Table 5** summarizes the characteristics of alpha, beta, and gamma radiation.

Nuclear stability Much of science deals with understanding how things change and how they remain stable. An atom's stability is governed by its ratio of neutrons to protons. Atoms that contain either too many or too few neutrons are unstable and lose energy as they decay to form a stable nucleus. They emit alpha and beta particles, which affect the neutron-to-proton ratio of the newly created nucleus. Eventually, radioactive atoms undergo enough radioactive decay to form stable, nonradioactive atoms.

Check Your Progress

Summary

- The radioactive decay of unstable nuclei involves the release of energy.
- There are three types of radiation: alpha (charge of 2+), beta (charge of 1–) and gamma (no charge).
- The neutron-to-proton ratio of an atom's nucleus determines its stability.

Demonstrate Understanding

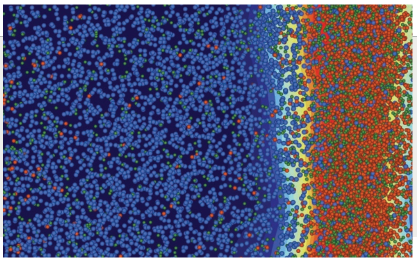
24. **Describe** the difference between radioactivity and radioactive decay.
25. **State** what quantities are conserved and which are not conserved in a nuclear reaction.
26. **Explain** why beta particles are deflected towards a positive plate, alpha particles are deflected towards a negative plate, and gamma rays are not deflected.
27. **Calculate** How much more mass does an alpha particle have compared to an electron?
28. **Create** a table showing how each type of radiation affects the atomic number and mass number of an atom.

LEARNSMART®

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

Mapping the Mysteries of Materials

Visualizing and mapping the arrangement and bonding of atoms in a material provides scientists with a vast amount of information about its structure, properties, and potential uses.



Scientists have mapped the atoms in this copper-cobalt-manganese catalyst. Cobalt atoms are blue, copper atoms are orange, and manganese atoms are green.

Atoms and Properties

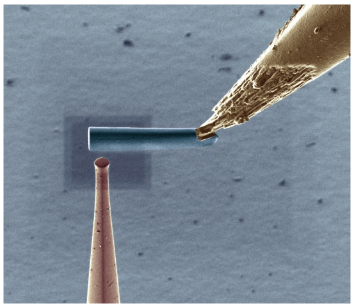
Everything on Earth is composed of atoms. The differences among atoms and how they are combined result in the great variety of materials and substances on Earth. The properties of any material, such as flexibility and strength, are determined by the types of atoms it contains and how they are combined.

How Atoms Are Mapped

Recent advances have enabled mapping the precise positions of billions of individual atoms in a material. Scanning transmission electron microscopy scans samples with a beam of electrons and gathers information about the interactions of the electrons at many different angles to produce two-dimensional images. The two-dimensional images are then combined using advanced computer programs, resulting in a detailed three-dimensional atomic "map".

Real-World Applications

The applications of this advanced technology are virtually limitless. Using this technique, scientists can detect defects at specific locations in materials—places where the "wrong" atom was inserted, or an atom was missing. These atomic-level defects can cause large scale problems that could lead to catastrophic damage. For example, if the material used in a bridge had defects at the atomic level that reduced its strength, it would compromise the safety of the bridge.



A tiny sample of aluminum (the blue rod) is being prepared for analysis to map its atomic structure.

MAKE AND DEFEND A CLAIM

With a small group, write a short informational statement to explain to individuals who do not have a background in chemistry how scientists map materials at the atomic level.

MODULE 3
STUDY GUIDE

GO ONLINE to study with your Science Notebook.

Lesson 1 EARLY IDEAS ABOUT MATTER

- Democritus was the first person to propose the existence of atoms.
- According to Democritus, atoms are solid, homogeneous, and indivisible.
- Aristotle did not believe in the existence of atoms.
- Dalton revised the ideas of Democritus based on the results of scientific research.

- Dalton's atomic theory

Lesson 2 DEFINING THE ATOM

- An atom is the smallest particle of an element that maintains the properties of that element.
- Electrons have a 1− charge, protons have a 1+ charge, and neutrons have no charge.
- An atom consists mostly of empty space surrounding the nucleus.

- atom
- cathode ray
- electron
- nucleus
- proton
- neutron

Lesson 3 HOW ATOMS DIFFER

- The atomic number of an atom is given by its number of protons. The mass number of an atom is the sum of its neutrons and protons.

$$\text{atomic number} = \text{number of protons} = \text{number of electrons}$$

$$\text{mass number} = \text{atomic number} + \text{number of neutrons}$$

- Atoms of the same element with different numbers of neutrons are called isotopes.
- The atomic mass of an element is a weighted average of the masses of all of its naturally occurring isotopes.

- atomic number
- isotope
- mass number
- atomic mass unit (amu)
- atomic mass

Lesson 4 UNSTABLE NUCLEI AND RADIOACTIVE DECAY

- The radioactive decay of unstable nuclei involves the release of energy.
- There are three types of radiation: alpha (charge of 2+), beta (charge of 1−) and gamma (no charge).
- The neutron-to-proton ratio of an atom's nucleus determines its stability.

- radioactivity
- radiation
- nuclear reaction
- radioactive decay
- alpha radiation
- alpha particle
- nuclear equation
- beta radiation
- beta particle
- gamma ray

THREE-DIMENSIONAL THINKING

Module Wrap-Up

REVISIT THE PHENOMENON

What is matter 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

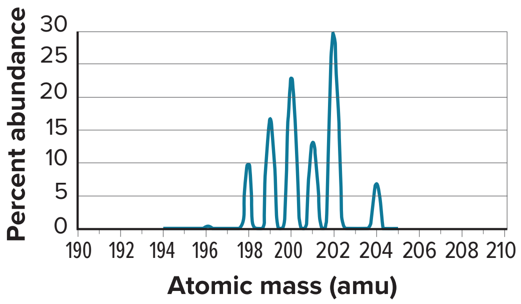
Now that you've completed the module, revisit your STEM unit project. You will summarize your evidence and apply it to the project.

GO FURTHER

SEP Data Analysis Lab

How can you identify an element from mass spectrometer data?

A mass spectrometer separates atoms and molecules according to their masses. A substance is heated in a vacuum and then ionized. The ions are accelerated through a magnetic field, which separates ions of different masses.



CER Analyze and Interpret Data

1. **Claim, Evidence, Reasoning** Without performing any calculations, predict the approximate atomic mass of element X. Explain the reasoning behind your prediction.
2. **Evidence, Reasoning** Calculate the weighted average mass of the isotopes of element X. Then, use the periodic table to identify the element.

Monty Rakusen/Cultura/Getty Images

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.

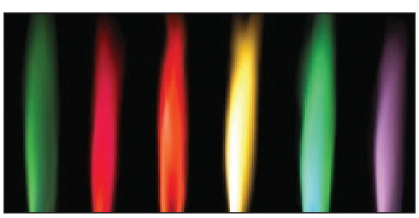
CER Claim, Evidence, Reasoning

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

(t)Video Supplied by BBC Worldwide Learning; (bl)SPL/Science Source; (br)NASA, ESA, and M. Livio and the Hubble 20th Anniversary Team (STScI)

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.

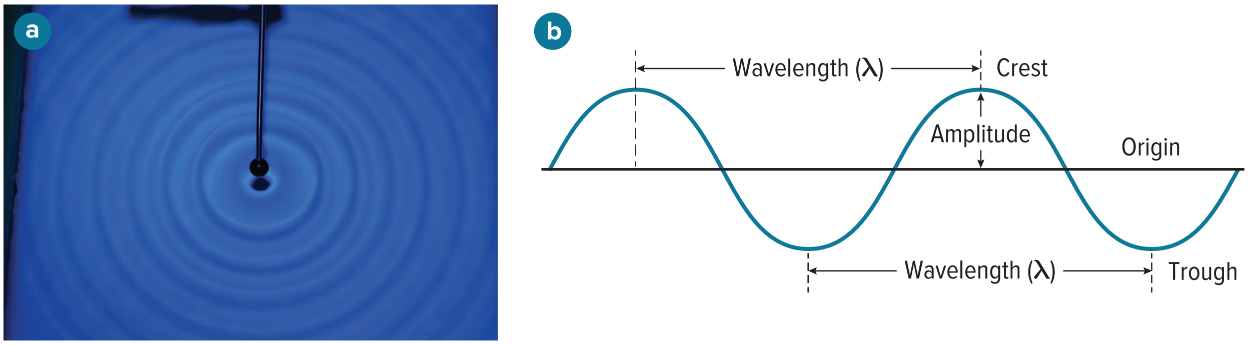
Characteristics of waves

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 ($1 \text{ nm} = 1 \times 10^{-9} \text{ m}$).

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: $652 \text{ Hz} = 652 \text{ waves/second} = 652\text{/s} = 652 \text{ s}^{-1}$.

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.



Matt Meadows/McGraw-Hill Education

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.*

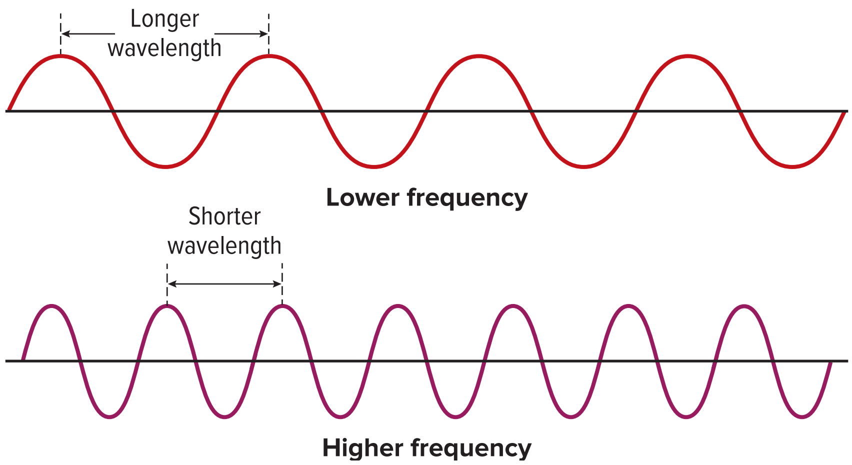


Figure 3 These waves illustrate the relationship between wavelength and frequency. As frequency increases, wavelength decreases.

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.

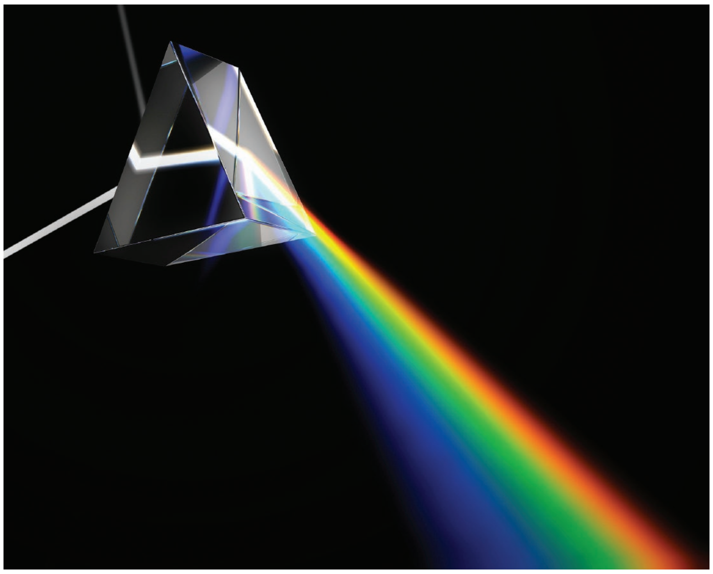
The speed of light in a vacuum is equal to the product of the wavelength and the frequency.

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.



Andrzej Wojcicki/Science Photo Library/Alamy

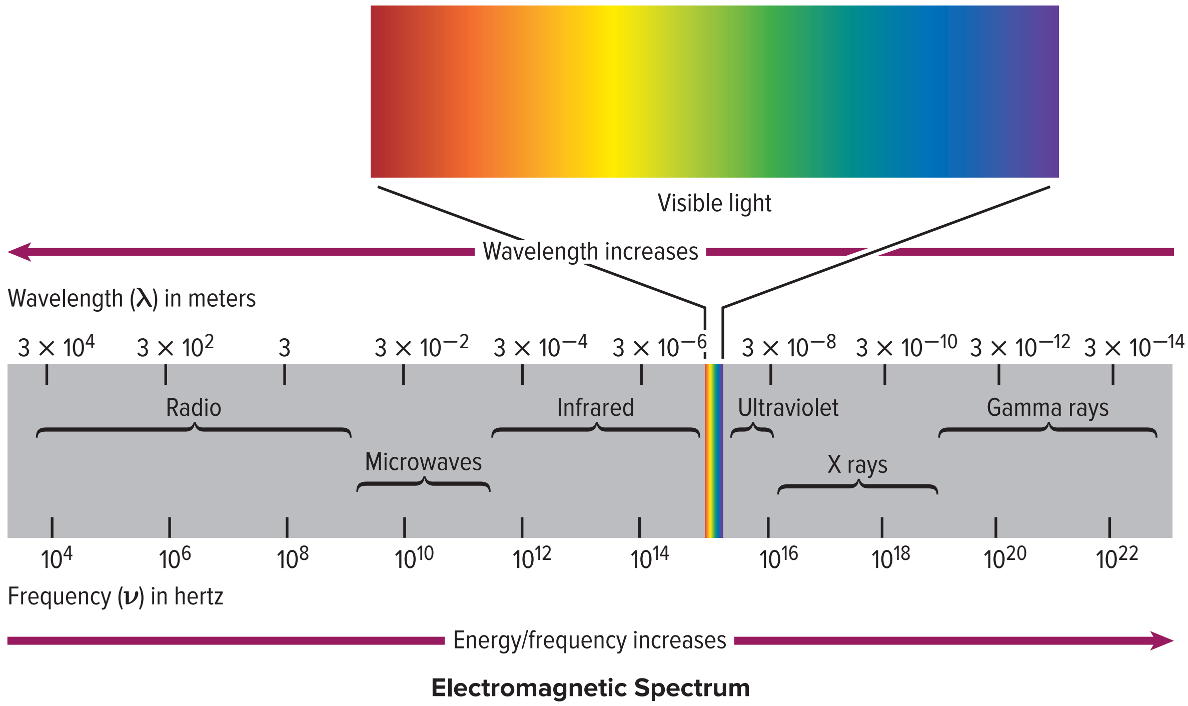


Figure 5 The electromagnetic spectrum covers a wide range of frequencies. The visible-light section of the spectrum is very narrow. As frequency and energy increase, wavelength decreases.

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

$\nu = 3.44 \times 10^9$Hz

$c = 3.00 \times 10^8$ m/s

Unknown

λ = ? m

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.

$\lambda = \dfrac{3.00 \times 10^8\ \text{m/s}}{3.44 \times 10^9\ \text{Hz}}$ — 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}.

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

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

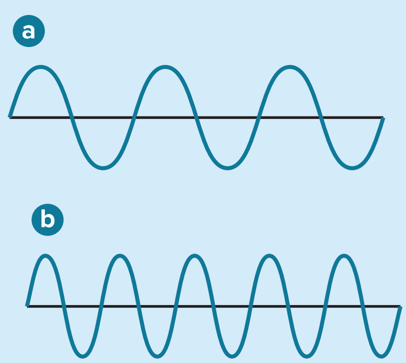
3 EVALUATE THE ANSWER

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

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

$E_{quantum}$ 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.

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.

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.

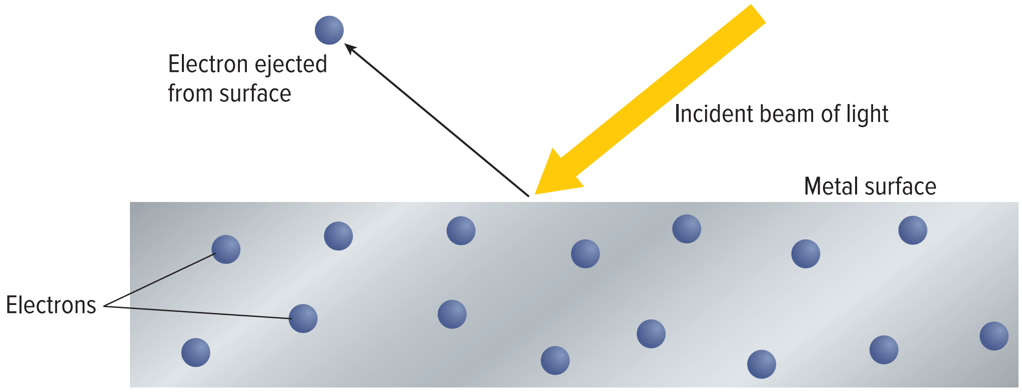


Figure 7 The photoelectric effect occurs when light of a certain frequency strikes a metal surface and ejects electrons. When the intensity of the light increases, the number of electrons ejected increases. When the frequency (energy) of the light increases, the energy of the ejected electrons increases.

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 $7.230 \times 10^{14}\ s^{-1}$?

1 ANALYZE THE PROBLEM

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

2 SOLVE FOR THE UNKNOWN

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

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

$E_{\text{photon}} = 4.791 \times 10^{-19}\ J$ — 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. $6.32 \times 10^{20}\ s^{-1}$ **b.** 9.50×10^{13} Hz **c.** $1.05 \times 10^{16}\ s^{-1}$
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.

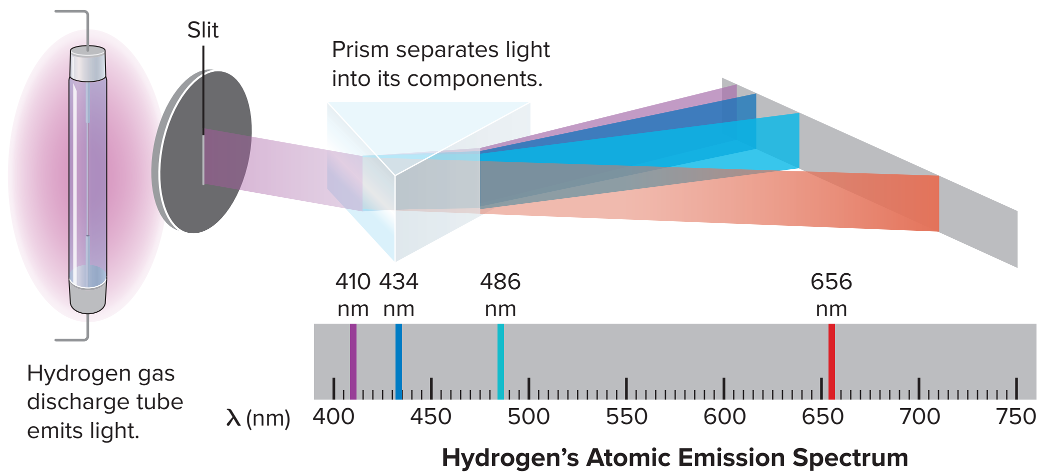


Figure 8 The purple light emitted by hydrogen can be separated into its different components using a prism. Hydrogen has an atomic emission spectrum that comprises four lines of different wavelengths.
Determine *Which line has the highest energy?*

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.*

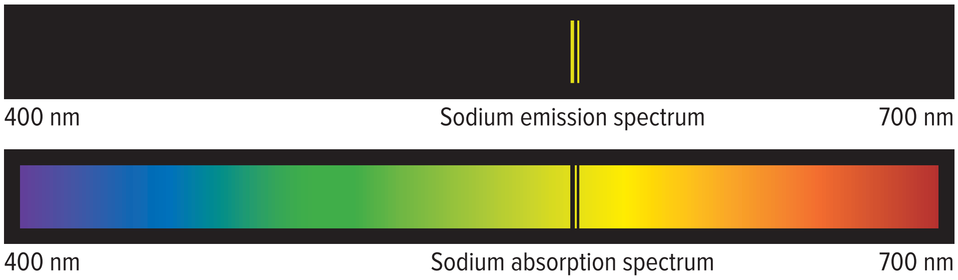


Figure 9 When excited sodium atoms return to a less excited state, they emit light at certain frequencies, producing an emission spectrum. When a continuous spectrum of light passes through sodium gas, atoms in the gas absorb light at those same frequencies, producing an absorption spectrum with dark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

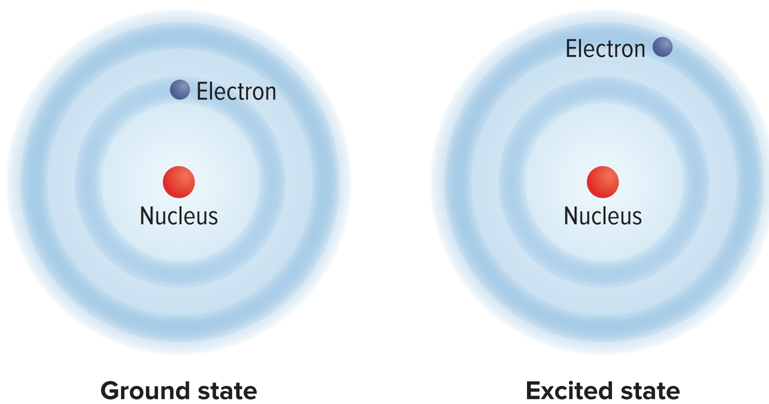


Figure 10 The figure shows an atom that has one electron. Note that the illustration is not to scale. In its ground state, the electron is associated with the lowest energy level. When the atom is in an excited state, the electron is associated with a higher energy level.

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.**

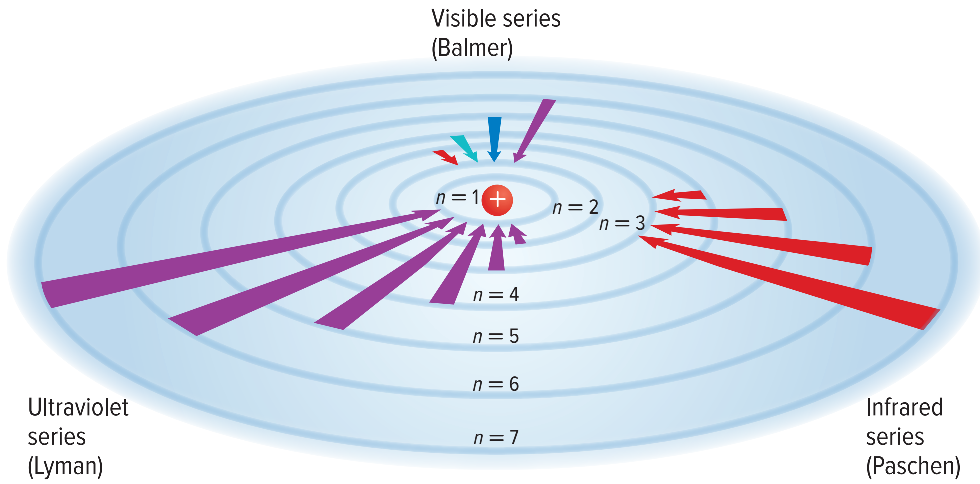


Figure 11 When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon is emitted. The ultraviolet (Lyman), visible (Balmer), and infrared (Paschen) series correspond to electrons dropping to $n = 1$, $n = 2$, and $n = 3$, respectively.

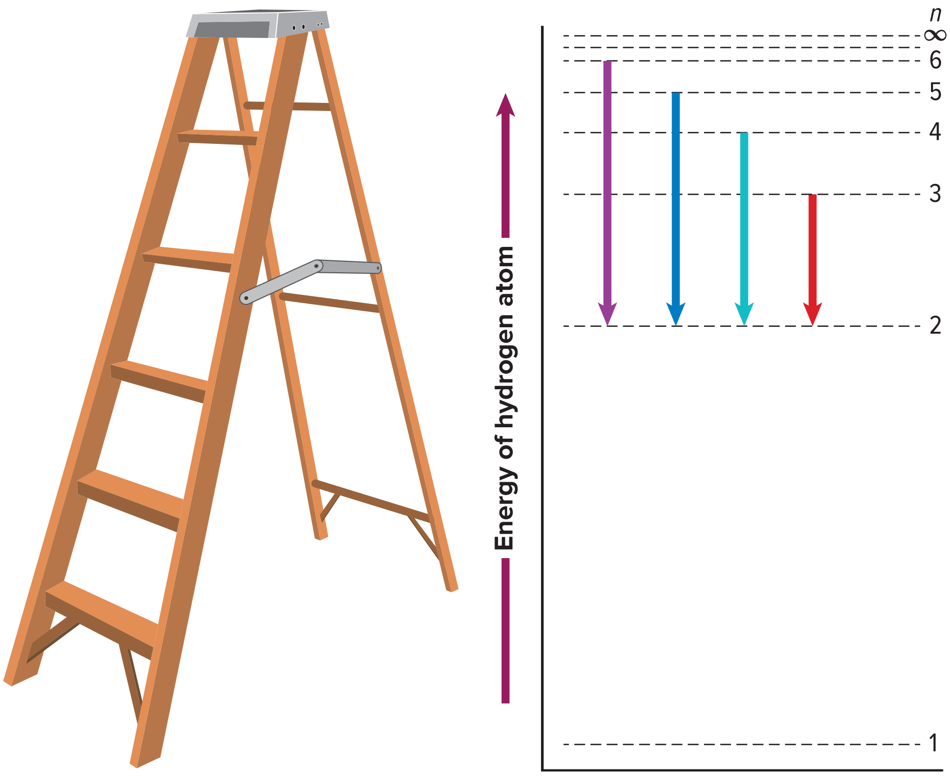


Figure 12 Only certain energy levels are allowed. The energy levels are similar to the rungs of a ladder. The four visible lines correspond to electrons dropping from a higher n to the orbit $n = 2$. As n increases, the hydrogen atom's energy levels are closer to each other.

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.

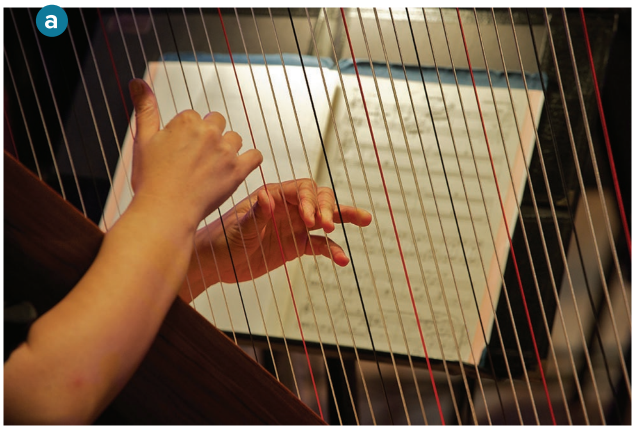
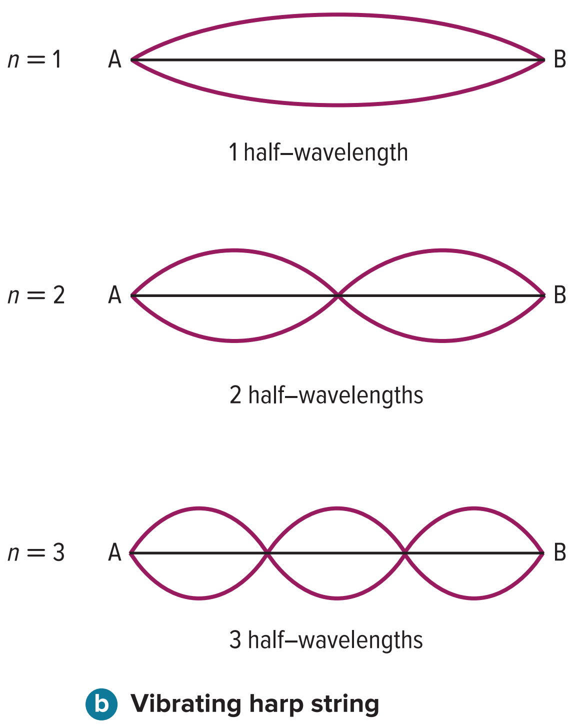
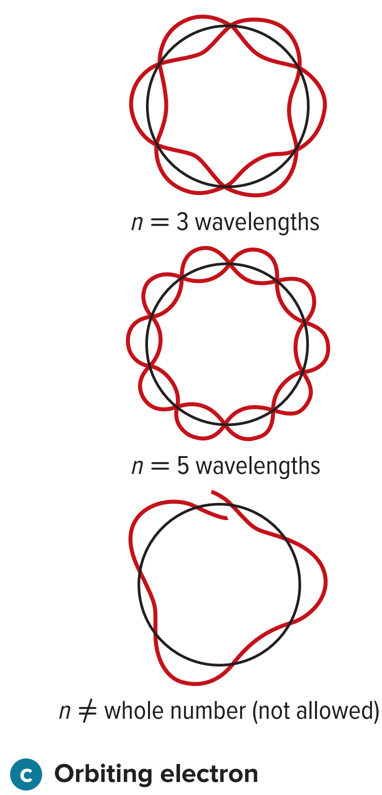


Figure 13 **a.** The string on the harp vibrates between two fixed endpoints. **b.** The vibrations of a string between the two fixed endpoints labeled A and B are limited to multiples of half-wavelengths. **c.** Electrons on circular orbits can only have whole numbers of wavelengths.





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.

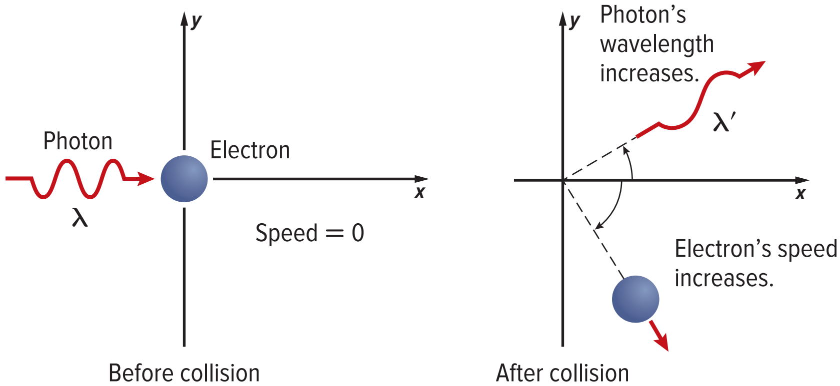


Figure 14 When a photon interacts with an electron at rest, both the velocity and the position of the electron are modified. This illustrates the Heisenberg uncertainty principle. It is impossible to know at the same time the position and the velocity of a particle.

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.

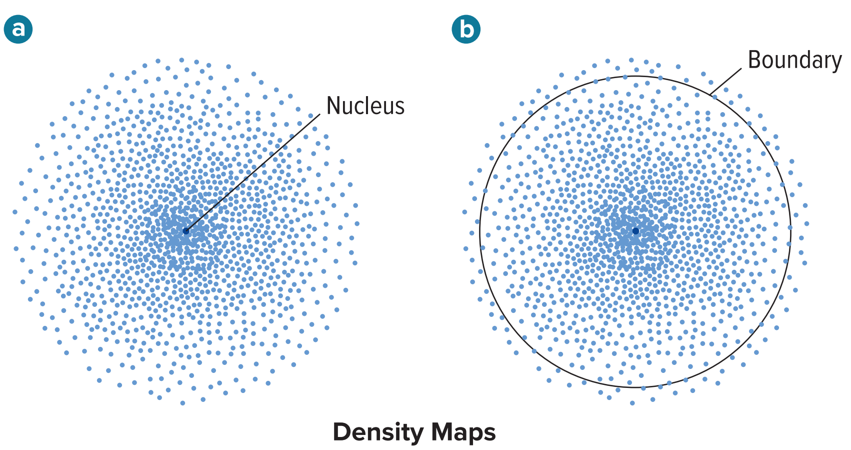


Figure 15 The density map represents the probability of finding an electron at a given position around the nucleus. **a.** The higher density of points near the nucleus shows that the electron is more likely to be found close to the nucleus. **b.** At any given time, there is a 90% probability of finding 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.

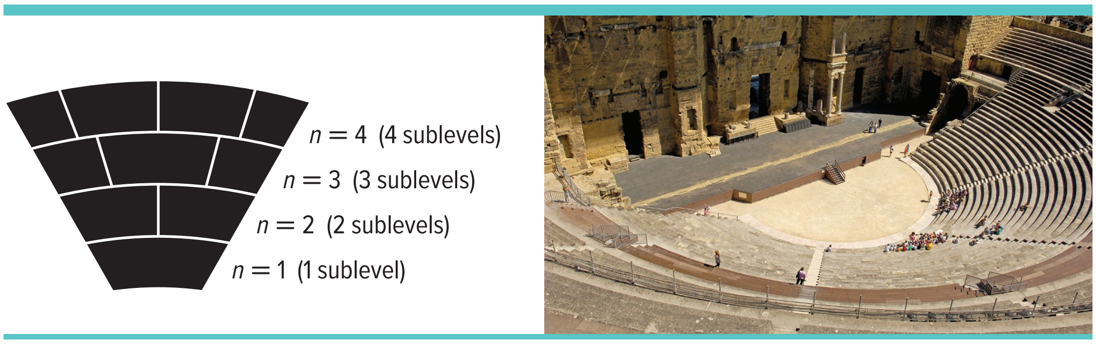


Figure 16 Energy levels can be thought of as rows of seats in a theater. The rows that are higher up and farther from the stage contain more seats. Similarly, energy levels related to orbitals farther from the nucleus contain more sublevels.

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.

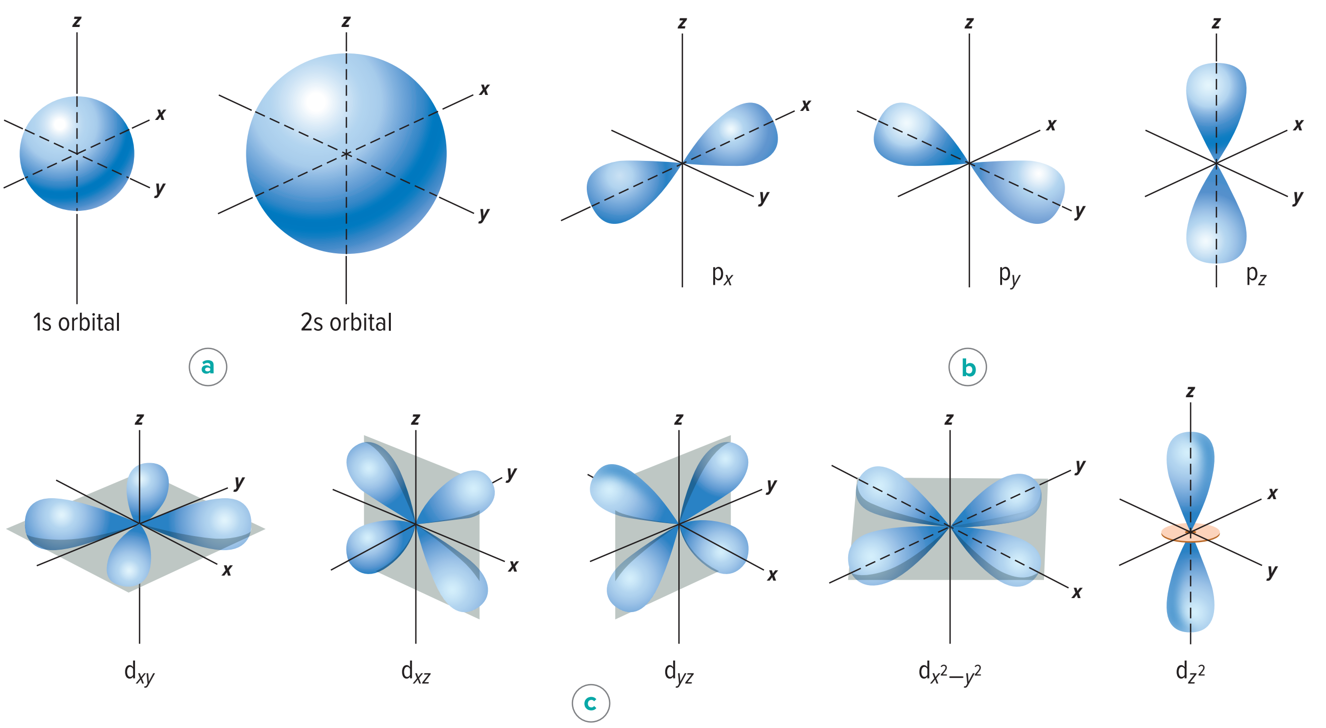


Figure 17 The shapes of atomic orbitals describe the probable distribution of electrons in energy sublevels.
a. All s orbitals are spherical, and their size increases with increasing principal quantum number.
b. The three p orbitals are dumbbell-shaped and are oriented along the three perpendicular *x*, *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.

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.

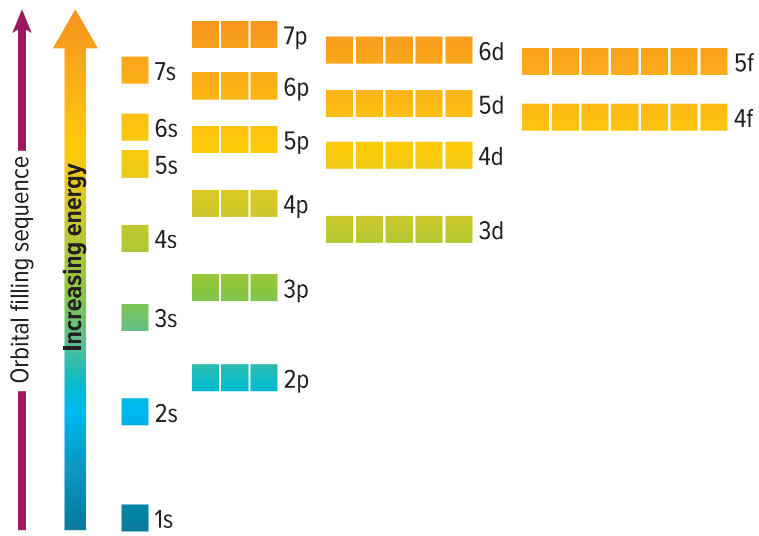


Figure 18 The aufbau diagram shows the energy of each sublevel relative to the energy of other sublevels.

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.

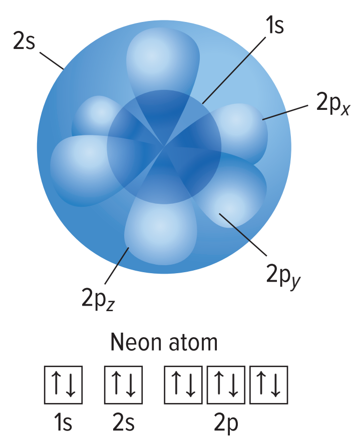
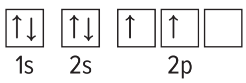


Figure 19 The 1s, 2s, and 2p orbitals of a neon atom overlap.

Determine *how many electrons a neon atom has.*

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.



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.

$$1s^22s^22p^63s^1$$

↑↓	↑↓	↑↓ ↑↓ ↑↓	↑
1s	2s	2p	3s

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, $1s^22s^22p^6$.

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	$1s^22s^22p^63s^1$	$[Ne]3s^1$
Magnesium	12	$1s^22s^22p^63s^2$	$[Ne]3s^2$
Aluminum	13	$1s^22s^22p^63s^23p^1$	$[Ne]3s^23p^1$
Silicon	14	$1s^22s^22p^63s^23p^2$	$[Ne]3s^23p^2$
Phosphorus	15	$1s^22s^22p^63s^23p^3$	$[Ne]3s^23p^3$
Sulfur	16	$1s^22s^22p^63s^23p^4$	$[Ne]3s^23p^4$
Chlorine	17	$1s^22s^22p^63s^23p^5$	$[Ne]3s^23p^5$
Argon	18	$1s^22s^22p^63s^23p^6$	$[Ne]3s^23p^6$ 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

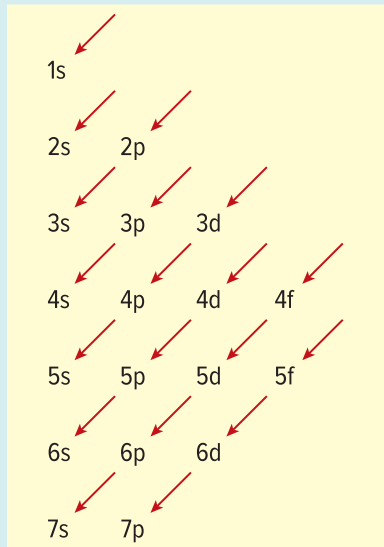
You can use the aufbau diagram to write correct ground-state electron configurations for all elements up to and including vanadium, atomic number 23. However, if you were to proceed in this manner, your configurations for chromium, [Ar]$4s^2 3d^4$, and copper, [Ar]$4s^2 3d^9$, would be incorrect. The 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) **c.** antimony (Sb) **e.** terbium (Tb)

b. strontium (Sr) **d.** rhenium (Re) **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, $[Ne]3s^23p^4$. 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	$1s^22s^1$	Li·
Beryllium	4	$1s^22s^2$	·Be·
Boron	5	$1s^22s^22p^1$	·Ḃ·
Carbon	6	$1s^22s^22p^2$	·Ċ· (with one dot below)
Nitrogen	7	$1s^22s^22p^3$	·N̈· (with one dot below)
Oxygen	8	$1s^22s^22p^4$	:Ö· (with one dot below)
Fluorine	9	$1s^22s^22p^5$	:F̈· (with two dots below)
Neon	10	$1s^22s^22p^6$	:N̈e: (with two dots below)

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.

$[Kr]5s^24d^{10}5p^2$ 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.

·Ṡn· (with one dot below)

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.

•X•

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.

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.

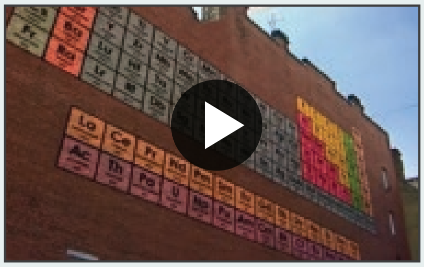
Will & Deni McIntyre/Science Source

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.

Video Supplied by BBC Worldwide Learning

SEP Ask Questions

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

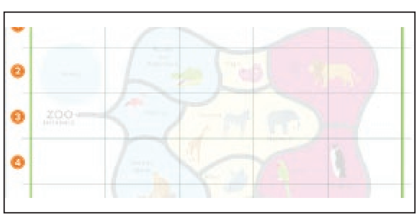
CER Claim, Evidence, Reasoning

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

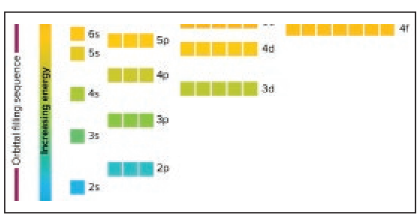
Collect Evidence Use the lessons in this module to collect evidence to support your claim. Record your evidence as you move through the module.

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.

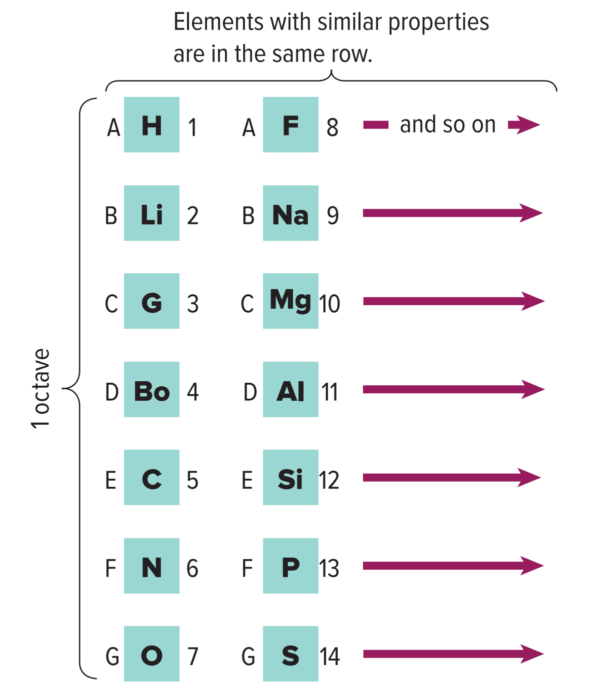


Figure 1 John Newlands noticed that the properties of elements repeated every eighth element in the same way musical notes repeat every eighth note and form octaves.

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.

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.

			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?	—
Typische Elemente			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**.

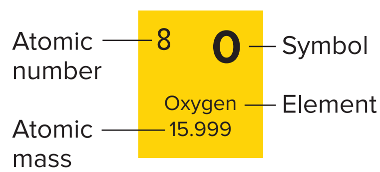


Figure 3 A typical box from the periodic table contains the element's name, its chemical symbol, its atomic number, and its atomic mass.

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**.

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.

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.

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?

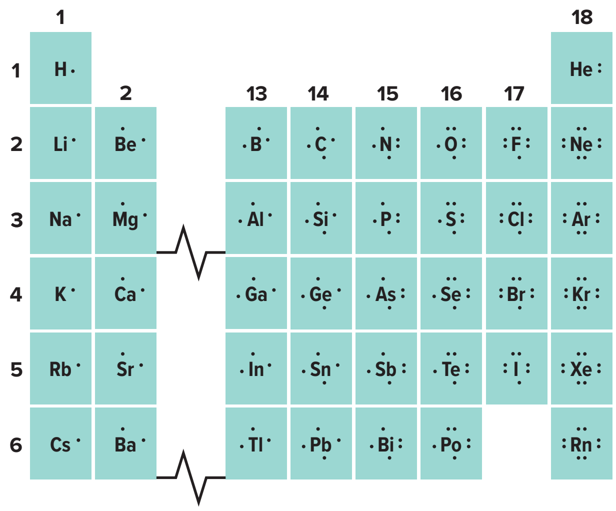

Figure 8 The figure shows the electron-dot structure of most representative elements.

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 $[Ar]4s^2 3d^{10} 4p^1$. 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.

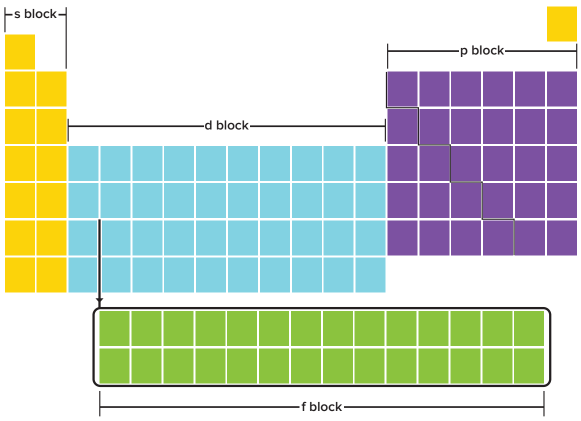

Figure 9 The periodic table is divided into four blocks—s, p, d, and f.

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	$[He]2s^22p^6$
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.

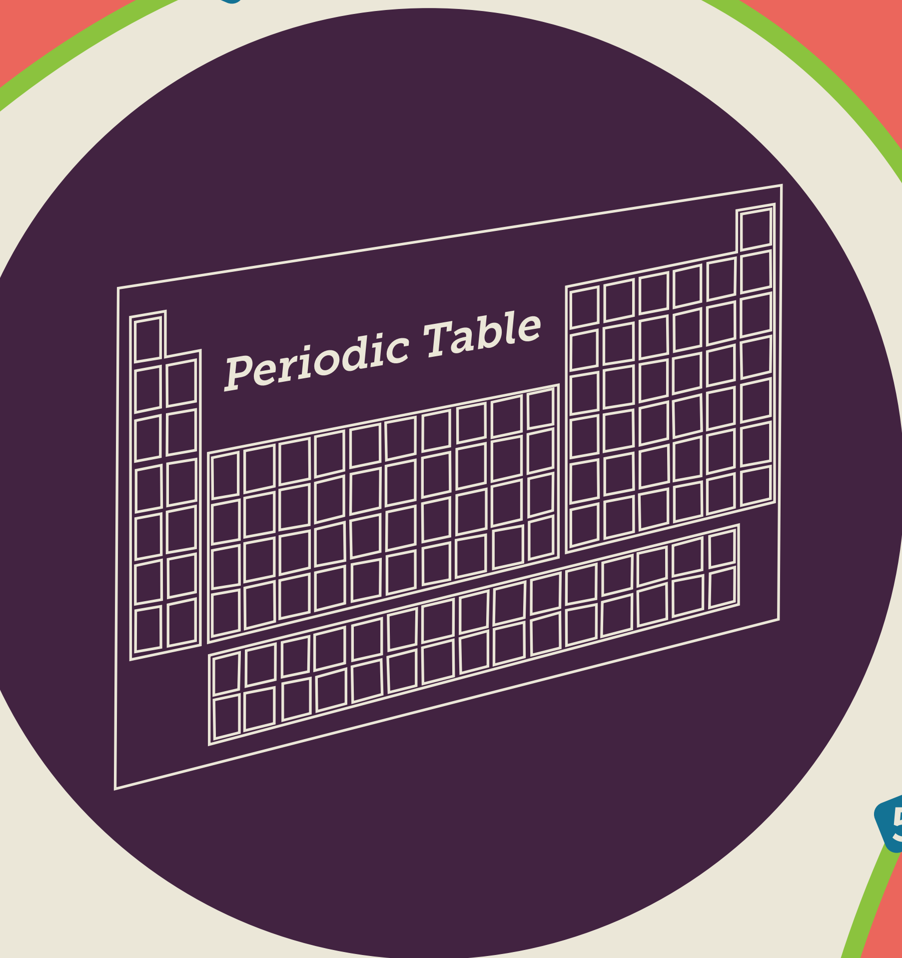

Figure 10

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

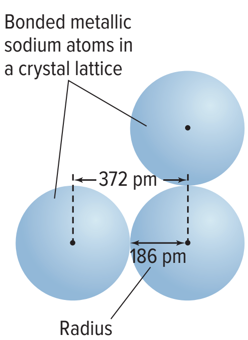
How can you use the periodic table to predict an element's properties?

Atomic Radius

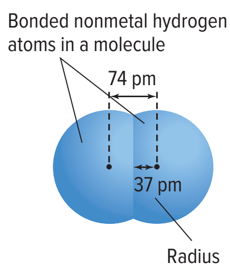
Many properties of the elements tend to change in a predictable way, known as a trend, as you move across a period or down a group. Atomic size is one such periodic trend. The sizes of atoms are influenced by electron configuration.

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**.



The radius of a metal atom is one-half the distance between two adjacent atoms in the crystal.



The radius of a nonmetal atom is often determined from a molecule of two identical atoms.

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

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: 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.

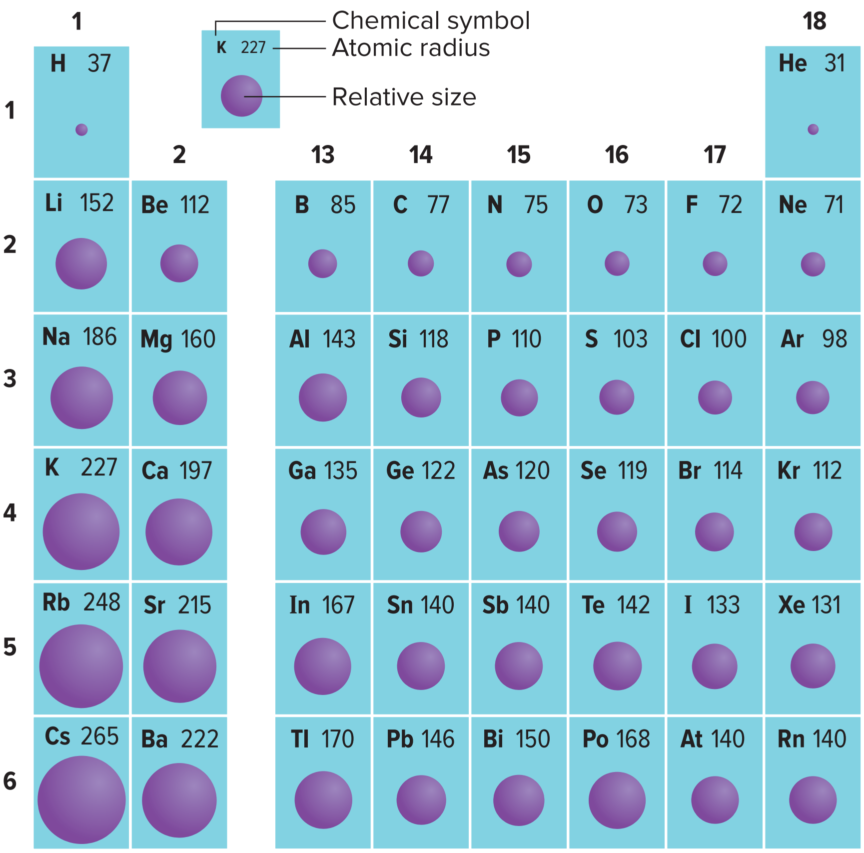


Figure 12 The atomic radii of the representative elements, given in picometers (10^{-12} m), vary as you move from left to right within a period and down a group.

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?

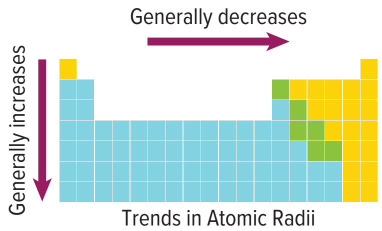


Figure 13 Atomic radii generally decrease from left to right in a period and generally increase as you move down a group.

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.

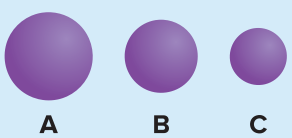
PRACTICE Problems

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.

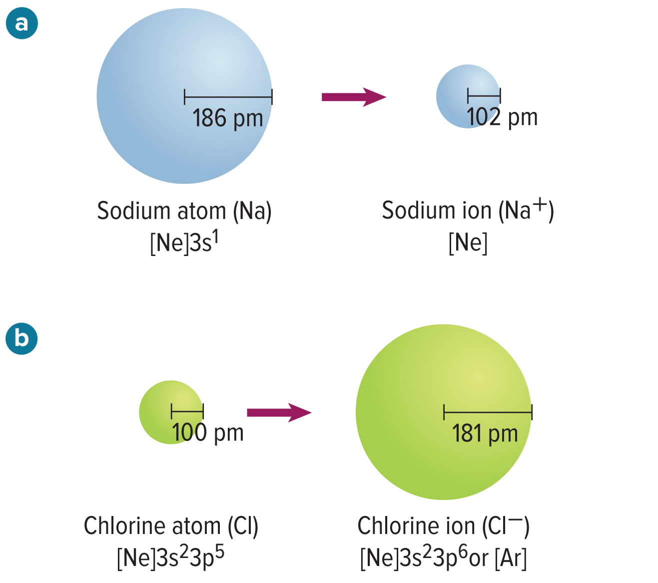


Figure 14 The size of atoms varies greatly when they form ions.

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.

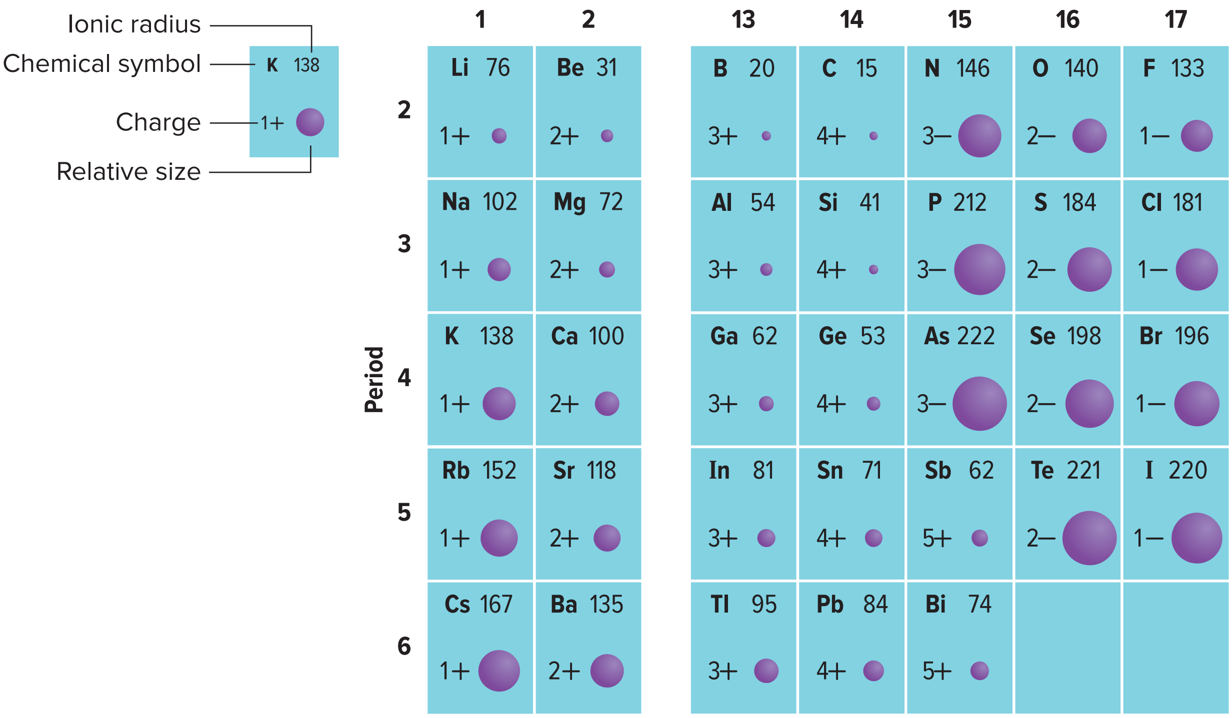


Figure 15 The ionic radii of most of the representative elements are shown in picometers (10^{-12} m).
Explain *why the ionic radii increase for both positive and negative ions as you move down a group.*

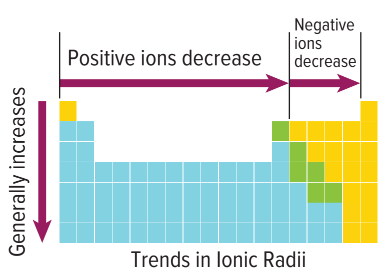
Trends within periods

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.

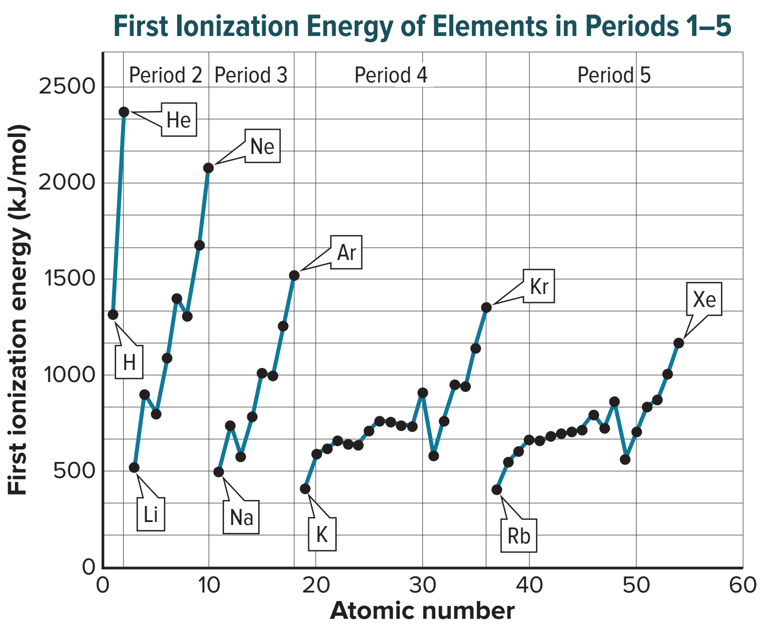


Figure 17 The first ionization energies for elements in periods 1 through 5 are shown as a function of the atomic number.

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.

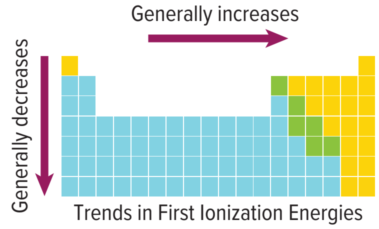


Figure 18 Ionization energies generally increase from left to right in a period and generally decrease as you move down a group.

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.

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.

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.

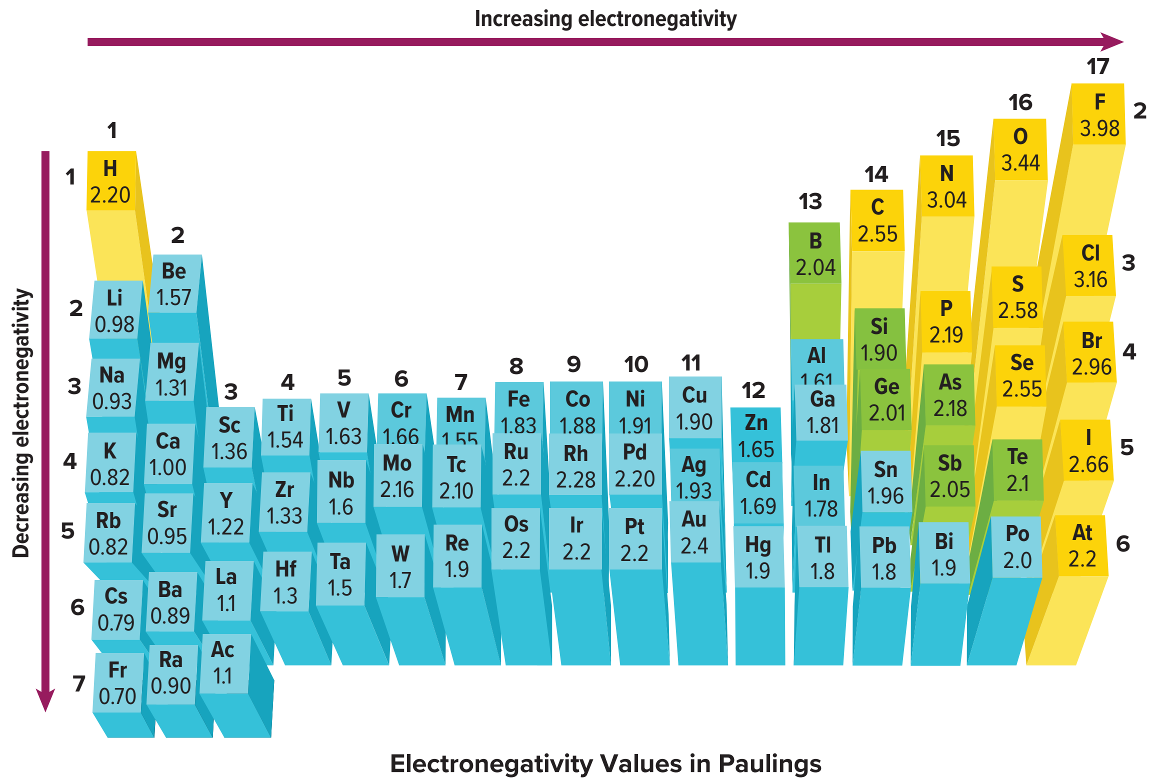


Figure 19 The electronegativity values for most of the elements are shown. The values are given in Paulings, a unit named after American scientist Linus Pauling (1901–1994).

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.

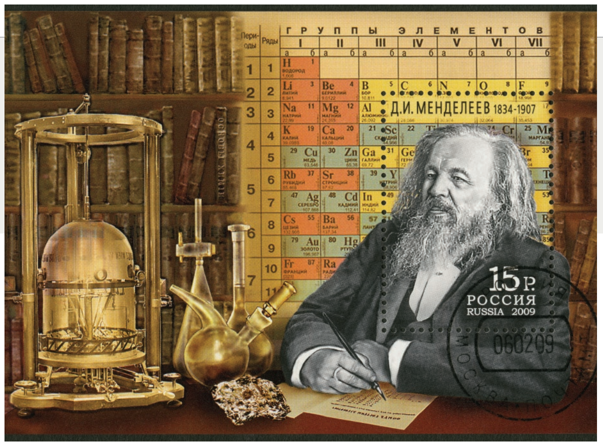
Demonstrate Understanding

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.

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

NATURE OF SCIENCE

The Evolving Periodic Table



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

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.

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.

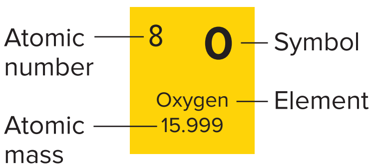
MODULE 5

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.

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.

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	?	?	?