

Jessie A. Key, David W. Ball

1ST CANADIAN EDITION

Introductory Chemistry



Introductory Chemistry - 1st Canadian Edition

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David W. Ball and Jessie A. Key

BCCAMPUS
VICTORIA, B.C.



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Chapter 6, Gases, now includes the following sections:

- Kinetic Molecular Theory
- Molecular Effusion and Diffusion
- Real Gases

Chapter 9, Chemical Bonds, now includes the following content:

- Additional periodic trends such as shielding and Z-effective
- Atomic orbital energy diagrams
- Molecular Orbitals

Chapter 10, Solids and Liquids, now includes the following content:

- Phase diagrams and explanations

Chapter 16, Organic Chemistry, now includes the following content;

- Current IUPAC nomenclature

Chapter 17, Kinetics, now includes the following sections:

- Factors that affect the rate of reactions
- Reaction rates
- Rate laws
- Temperature and Rate
- Reaction Mechanisms
- Catalysts

New material: Chapter 18, Chemical Thermodynamics, which includes the following sections:

- Spontaneity
- Entropy and the second law of thermodynamics
- Molecular Entropy and the third law of thermodynamics
- Entropy Changes in Chemical Reactions
- Gibbs Free Energy
- Free Energy and Temperature

- Free Energy and the Equilibrium Constant

Glossary and Appendices added.

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Contents

Acknowledgments Jessie A. Key	xiii
Dedication	xv
About BCcampus Open Education	xvii
 <u>Chapter 1. What is Chemistry</u>	
Some Basic Definitions	3
Chemistry as a Science	15
 <u>Chapter 2. Measurements</u>	
Expressing Numbers	25
Significant Figures	33
Converting Units	45
Other Units: Temperature and Density	59
Expressing Units	69
End-of-Chapter Material	77
 <u>Chapter 3. Atoms, Molecules, and Ions</u>	
Acids	81
Ions and Ionic Compounds	85
Masses of Atoms and Molecules	101
Molecules and Chemical Nomenclature	109
Atomic Theory	117
End-of-Chapter Material	127
 <u>Chapter 4. Chemical Reactions and Equations</u>	
The Chemical Equation	133
Types of Chemical Reactions: Single- and Double-Displacement Reactions	137
Ionic Equations: A Closer Look	147
Composition, Decomposition, and Combustion Reactions	155
Oxidation-Reduction Reactions	161

Neutralization Reactions	169
End-of-Chapter Material	175

Chapter 5. Stoichiometry and the Mole

Stoichiometry	181
The Mole	187
Mole-Mass and Mass-Mass Calculations	195
Limiting Reagents	201
The Mole in Chemical Reactions	207
Yields	211
End-of-Chapter Material	215

Chapter 6. Gases

Pressure	219
Gas Laws	223
Other Gas Laws	233
The Ideal Gas Law and Some Applications	239
Gas Mixtures	249
Kinetic Molecular Theory of Gases	259
Molecular Effusion and Diffusion	267
Real Gases	271
End-of-Chapter Material	277

Chapter 7. Energy and Chemistry

Formation Reactions	281
Energy	295
Stoichiometry Calculations Using Enthalpy	299
Enthalpy and Chemical Reactions	305
Work and Heat	313
Hess's Law	321
End-of-Chapter Material	327

Chapter 8. Electronic Structure

Light	331
Quantum Numbers for Electrons	337

Organization of Electrons in Atoms	347
Electronic Structure and the Periodic Table	359
Periodic Trends	371
End-of-Chapter Material	387

Chapter 9. Chemical Bonds

Lewis Electron Dot Diagrams	391
Electron Transfer: Ionic Bonds	397
Covalent Bonds	403
Other Aspects of Covalent Bonds	413
Violations of the Octet Rule	421
Molecular Shapes and Polarity	425
Valence Bond Theory and Hybrid Orbitals	435
Jessie A. Key	
Molecular Orbitals	445
Jessie A. Key	
End-of-Chapter Material	455

Chapter 10. Solids and Liquids

Properties of Liquids	461
Solids	473
Phase Transitions: Melting, Boiling, and Subliming	481
Intermolecular Forces	489
End-of-Chapter Material	495

Chapter 11. Solutions

Colligative Properties of Solutions	499
Concentrations as Conversion Factors	509
Quantitative Units of Concentration	515
Colligative Properties of Ionic Solutes	521
Some Definitions	525
Dilutions and Concentrations	529
End-of-Chapter Material	533

Chapter 12. Acids and Bases

Acid-Base Titrations	537
Strong and Weak Acids and Bases and Their Salts	541
Brønsted-Lowry Acids and Bases	549
Arrhenius Acids and Bases	555
Autoionization of Water	559
Buffers	563
The pH Scale	569
End-of-Chapter Material	575

Chapter 13. Chemical Equilibrium

Chemical Equilibrium	579
The Equilibrium Constant	583
Shifting Equilibria: Le Chatelier's Principle	591
Calculating Equilibrium Constant Values	599
Some Special Types of Equilibria	605
End-of-Chapter Material	617

Chapter 14. Oxidation and Reduction

Oxidation-Reduction Reactions	621
Balancing Redox Reactions	629
Applications of Redox Reactions: Voltaic Cells	637
Electrolysis	649
End-of-Chapter Material	653

Chapter 15. Nuclear Chemistry

Units of Radioactivity	659
Uses of Radioactive Isotopes	665
Half-Life	671
Radioactivity	677
Nuclear Energy	683
End-of-Chapter Material	689

Chapter 16. Organic Chemistry

Hydrocarbons	693
Branched Hydrocarbons	707
Alkyl Halides and Alcohols	721
Other Oxygen-Containing Functional Groups	729
Other Functional Groups	739
Polymers	749
End-of-Chapter Material	759

Chapter 17. Kinetics

Factors that Affect the Rate of Reactions	767
Jessie A. Key	
Reaction Rates	771
Jessie A. Key	
Rate Laws	775
Jessie A. Key	
Concentration–Time Relationships: Integrated Rate Laws	781
Jessie A. Key	
Activation Energy and the Arrhenius Equation	789
Jessie A. Key	
Reaction Mechanisms	795
Jessie A. Key	
Catalysis	801
Jessie A. Key	
End-of-Chapter Material	805
Jessie A. Key	

Chapter 18. Chemical Thermodynamics

Spontaneous Change	813
Jessie A. Key	
Entropy and the Second Law of Thermodynamics	815
Jessie A. Key	
Measuring Entropy and Entropy Changes	819
Jessie A. Key	
Gibbs Free Energy	825
Jessie A. Key	

Spontaneity: Free Energy and Temperature	829
Jessie A. Key	
Free Energy under Nonstandard Conditions	833
Jessie A. Key	
End-of-Chapter Material	837
Jessie A. Key	
Appendix A: Periodic Table of the Elements	843
David W. Ball	
Appendix B: Selected Acid Dissociation Constants at 25°C	851
Appendix C: Solubility Constants for Compounds at 25°C	855
Appendix D: Standard Thermodynamic Quantities for Chemical Substances at 25°C	861
Appendix E: Standard Reduction Potentials by Value	883
Glossary	895
Jessie A. Key	
About the Authors	919
Versioning History	921

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September 2014

The decision to write a new textbook from scratch is not one to be taken lightly. The author becomes a saint to some and a sinner to others—and the feedback from the “others” is felt more acutely than the feedback from the “some”! Ultimately, the decision to write a new book comes from the deep feeling that an author can make a positive contribution to the field, and that it is ultimately time well invested.

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David W. Ball

February 2011

Dedication

David W. Ball:

For Gail, with thanks for her support of this book and all the other projects in my life.

I scorn to change my state with kings.

William Shakespeare, Sonnet 29

Jessie A. Key:

For all of the amazing chemistry educators that I have had the privilege of learning from.

“If I have seen further, it is from standing on the shoulders of giants.”

Isaac Newton, 1676

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Chapter 1. What is Chemistry

If you are reading these words, you are likely starting a chemistry course. Get ready for a fantastic journey through a world of wonder, delight, and knowledge. One of the themes of this book is “chemistry is everywhere,” and indeed it is; you would not be alive if it weren’t for chemistry because your body is a big chemical machine. If you don’t believe it, don’t worry. Every chapter in this book contains examples that will show you how chemistry is, in fact, everywhere. So enjoy the ride—and enjoy chemistry.



Figure 1.0 “What is Chemistry?”

What is chemistry? Simply put, **chemistry** is the study of the interactions of matter with other matter and with energy. This seems straightforward enough. However, the definition of chemistry includes a wide range of topics that must be understood to gain a mastery of the topic or even take additional courses in chemistry. In this book, we will lay the foundations of chemistry in a topic-by-topic fashion to provide you with the background you need to successfully understand chemistry.

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Some Basic Definitions

Learning Objectives

1. Learn the basic terms used to describe matter.

The definition of chemistry—the study of the interactions of matter with other matter and with energy—uses some terms that should also be defined. We start the study of chemistry by defining some basic terms.

Matter. A book is matter, a computer is matter, food is matter, and dirt in the ground is matter. Sometimes matter may be difficult to identify. For example, air is matter, but because it is so thin compared to other matter (e.g., a book, a computer, food, and dirt), we sometimes forget that air has mass and takes up space. Things that are not matter include thoughts, ideas, emotions, and hopes.

Example 1.1

Problems

Which of the following is matter and not matter?

1. a hot dog
2. love
3. a tree

Solutions

1. A hot dog has mass and takes up space, so it is matter.
2. Love is an emotion, and emotions are not matter.
3. A tree has mass and takes up space, so it is matter.

Test Yourself

Which of the following is matter and not matter?

1. the moon
2. an idea for a new invention

Answers

1. The moon is matter.
2. The invention itself may be matter, but the idea for it is not.



Figure 1.1 “The Phases of Matter.” Chemistry recognizes three fundamental phases of matter: solid (left), liquid (middle), and gas (right).

To understand matter and how it changes, we need to be able to describe matter. There are two basic ways to describe matter: physical properties and chemical properties. **Physical properties** are characteristics that describe matter as it exists. Some of many physical characteristics of matter are shape, color, size, and temperature. An important physical property is the **phase** (or state) of matter. The three fundamental phases of matter are solid, liquid, and gas (see Figure 1.1 “The Phases of Matter”).



Figure 1.2 “Chemical Properties.” The fact that this match burns is a chemical property of the match.

Chemical properties are characteristics of matter that describe how matter changes form in the presence of other matter. Does a sample of matter burn? Burning is a chemical property. Does it behave violently when put in water? This reaction is a chemical property as well (Figure 1.2 “Chemical Properties”). In the following chapters, we will see how descriptions of physical and chemical properties are important aspects of chemistry.



Figure 1.3 “Physical Changes.” The solid ice melts into liquid water—a physical change.

If matter always stayed the same, chemistry would be rather boring. Fortunately, a major part of chemistry involves change. A **physical change** occurs when a sample of matter changes one or more of its physical properties. For example, a solid may melt (Figure 1.3 “Physical Changes”), or alcohol in a thermometer may change volume as the temperature changes. A physical change does not affect the chemical composition of matter. A **chemical change** is the process of demonstrating a chemical property, such as the burning match in Figure 1.2 “Chemical Properties”. As the matter in the match burns, its chemical composition changes, and new forms of matter with new physical properties are created. Note that chemical changes are frequently accompanied by physical changes, as the new matter will likely have different physical properties from the original matter.

Example 1.2

Problems

Describe each process as a physical change or a chemical change.

1. Water in the air turns into snow.
2. A person’s hair is cut.
3. Bread dough becomes fresh bread in an oven.

Solutions

1. Because the water is going from a gas phase to a solid phase, this is a physical change.
2. Your long hair is being shortened. This is a physical change.
3. Because of the oven’s temperature, chemical changes are occurring in the bread dough to make fresh bread. These are chemical changes. (In fact, a lot of cooking involves chemical changes.)

Test Yourself

Identify each process as a physical change or a chemical change.

1. A fire is raging in a fireplace.

2. Water is warmed to make a cup of coffee.

Answers

1. chemical change
2. physical change

A sample of matter that has the same physical and chemical properties throughout is called a **substance**. Sometimes the phrase *pure substance* is used, but the word *pure* isn't needed. The definition of the term substance is an example of how chemistry has a specific definition for a word that is used in everyday language with a different, vaguer definition. Here, we will use the term substance with its strict chemical definition.

Chemistry recognizes two different types of substances: elements and compounds. An **element** is the simplest type of chemical substance; it cannot be broken down into simpler chemical substances by ordinary chemical means. There are about 115 elements known to science, of which 80 are stable. (The other elements are radioactive, a condition we will consider in [Chapter 15 “Nuclear Chemistry”](#).) Each element has its own unique set of physical and chemical properties. Examples of elements include iron, carbon, and gold.

A **compound** is a combination of more than one element. The physical and chemical properties of a compound are different from the physical and chemical properties of its constituent elements; that is, it behaves as a completely different substance. There are over 50 million compounds known, and more are being discovered daily. Examples of compounds include water, penicillin, and sodium chloride (the chemical name for common table salt).

Elements and compounds are not the only ways in which matter can be present. We frequently encounter objects that are physical combinations of more than one element or compound. Physical combinations of more than one substance are called **mixtures**. There are two types of mixtures. A **heterogeneous mixture** is a mixture composed of two or more substances. It is easy to tell, sometimes by the naked eye, that more than one substance is present. A **homogeneous mixture** is a combination of two or more substances that is so intimately mixed that the mixture behaves as a single substance. Another word for a homogeneous mixture is **solution**. Thus, a combination of salt and steel wool is a heterogeneous mixture because it is easy to see which particles of the matter are salt crystals and which are steel wool. On the other hand, if you take salt crystals and dissolve them in water, it is very difficult to tell that you have more than one substance present just by looking—even if you use a powerful microscope. The salt dissolved in water is a homogeneous mixture, or a solution (Figure 1.4 “Types of Mixtures”).



Figure 1.4 “Types of Mixtures.” On the left, the combination of two substances is a heterogeneous mixture because the particles of the two components look different. On the right, the salt crystals have dissolved in the water so finely that you cannot tell that salt is present. The homogeneous mixture appears like a single substance.

Example 1.3

Problems

Identify the following combinations as heterogeneous mixtures or homogenous mixtures.

1. soda water (Carbon dioxide is dissolved in water.)
2. a mixture of iron metal filings and sulfur powder (Both iron and sulfur are elements.)

Solutions

1. Because carbon dioxide is dissolved in water, we can infer from the behaviour of salt crystals dissolved in water that carbon dioxide dissolved in water is (also) a homogeneous mixture.
2. Assuming that the iron and sulfur are simply mixed together, it should be easy to see what is iron and what is sulfur, so this is a heterogeneous mixture.

Test Yourself

Are the following combinations homogeneous mixtures or heterogeneous mixtures?

1. the human body
2. an amalgam, a combination of some other metals dissolved in a small amount of mercury

Answers

1. heterogeneous mixture
2. homogeneous mixture

There are other descriptors that we can use to describe matter, especially elements. We can usually divide elements into metals and nonmetals, and each set shares certain (but not always all) properties. A **metal** is an element that is solid at room temperature (although mercury is a well-known exception), is shiny and silvery, conducts electricity and heat well, can be pounded into thin sheets (a property

called *malleability*), and can be drawn into thin wires (a property called *ductility*). A **nonmetal** is an element that is brittle when solid, does not conduct electricity or heat very well, and cannot be made into thin sheets or wires (Figure 1.5 “Semimetals”). Nonmetals also exist in a variety of phases and colors at room temperature. Some elements have properties of both metals and nonmetals and are called **semimetals** (or metalloids). We will see later how these descriptions can be assigned rather easily to various elements.



Figure 1.5 “Semimetals.” On the left is some elemental mercury, the only metal that exists as a liquid at room temperature. It has all the other expected properties of a metal. On the right, elemental sulfur is a yellow nonmetal that usually is found as a powder.

Figure 1.6 “Describing Matter” is a flowchart of the relationships among the different ways of describing matter.

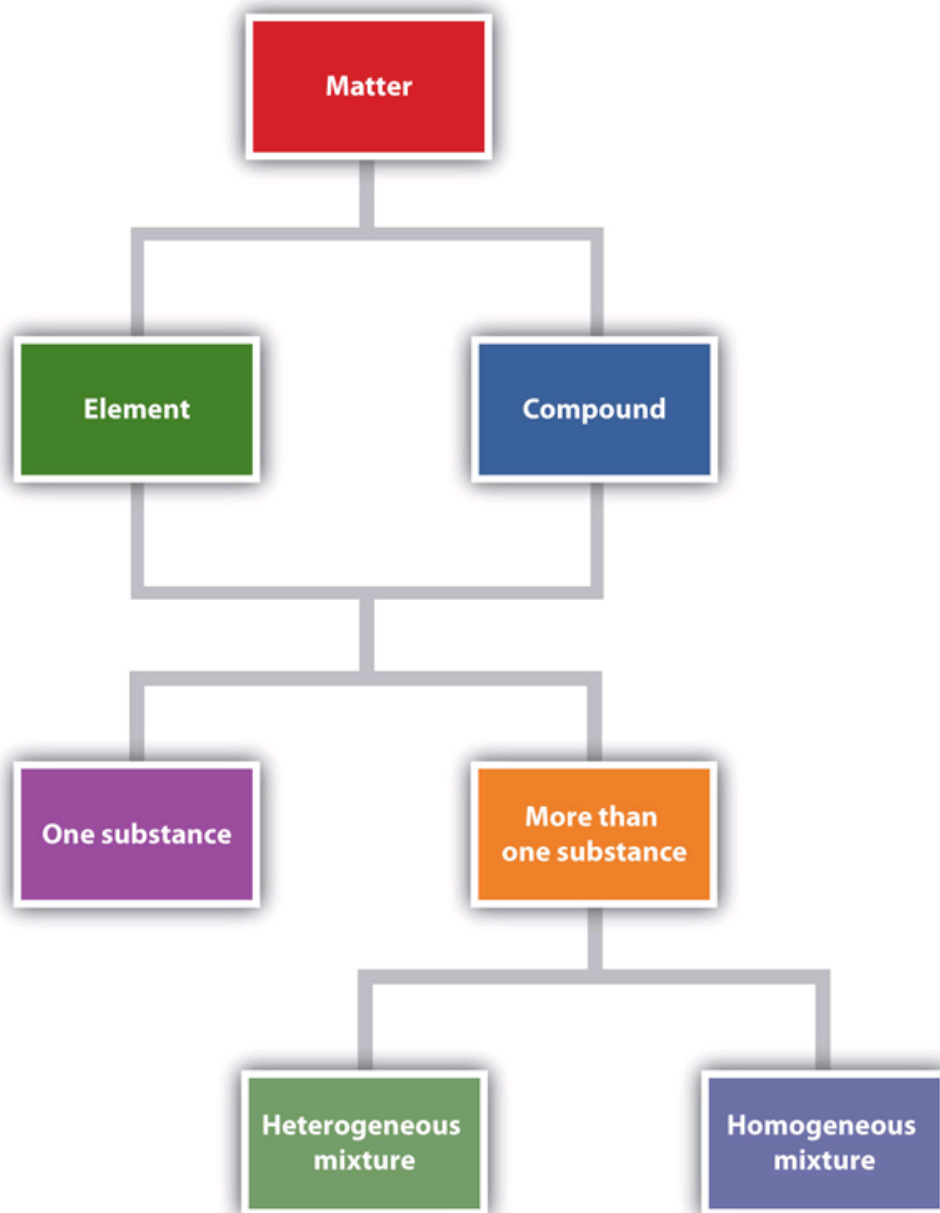


Figure 1.6 “Describing Matter.” This flowchart shows how matter can be described.

Chemistry Is Everywhere: In the Morning

Most people have a morning ritual, a process that they go through every morning to get ready for the day. Chemistry appears in many of these activities.

- If you take a shower or bath in the morning, you probably use soap, shampoo, or both. These items contain chemicals that interact with the oil and dirt on your body and hair to remove them and wash them away. Many of these products also contain chemicals that make you smell good; they are called *fragrances*.
- When you brush your teeth in the morning, you usually use toothpaste, a form of soap, to

clean your teeth. Toothpastes typically contain tiny, hard particles called *abrasives* that physically scrub your teeth. Many toothpastes also contain fluoride, a substance that chemically interacts with the surface of the teeth to help prevent cavities.

- Perhaps you take vitamins, supplements, or medicines every morning. Vitamins and other supplements contain chemicals your body needs in small amounts to function properly. Medicines are chemicals that help combat diseases and promote health.
- Perhaps you make some fried eggs for breakfast. Frying eggs involves heating them enough so that a chemical reaction occurs to cook the eggs.
- After you eat, the food in your stomach is chemically reacted so that the body (mostly the intestines) can absorb food, water, and other nutrients.
- If you drive or take the bus to school or work, you are using a vehicle that probably burns gasoline, a material that burns fairly easily and provides energy to power the vehicle. Recall that burning is a chemical change.

These are just a few examples of how chemistry impacts your everyday life. And we haven't even made it to lunch yet!



Figure 1.7 “Chemistry in Real Life.” Examples of chemistry can be found everywhere—such as in personal hygiene products, food, and motor vehicles.

View the [video “The Chemical World”](#) by Dr. Jessie A. Key for an introduction to the science of chemistry and how it fits into our everyday lives.

Key Takeaways

- Chemistry is the study of matter and its interactions with other matter and energy.
- Matter is anything that has mass and takes up space.
- Matter can be described in terms of physical properties and chemical properties.
- Physical properties and chemical properties of matter can change.
- Matter is composed of elements and compounds.
- Combinations of different substances are called mixtures.
- Elements can be described as metals, nonmetals, and semimetals.

Exercises

Questions

- Identify each as either matter or not matter.
 - a book
 - hate
 - light
 - a car
 - a fried egg
- Give an example of matter in each phase: solid, liquid, or gas.
- Does each statement represent a physical property or a chemical property?
 - Sulfur is yellow.
 - Steel wool burns when ignited by a flame.
 - A gallon of milk weighs over eight pounds.
- Does each statement represent a physical property or a chemical property?
 - A pile of leaves slowly rots in the backyard.
 - In the presence of oxygen, hydrogen can interact to make water.
 - Gold can be stretched into very thin wires.
- Does each statement represent a physical change or a chemical change?
 - Water boils and becomes steam.
 - Food is converted into usable form by the digestive system.
 - The alcohol in many thermometers freezes at about -40 degrees Fahrenheit.
- Does each statement represent a physical change or a chemical change?
 - Graphite, a form of elemental carbon, can be turned into diamond, another form of carbon, at very high temperatures and pressures.
 - The house across the street has been painted a new colour.
 - The elements sodium and chlorine come together to make a new substance called sodium chloride.
- Distinguish between an element and a compound. About how many of each are known?
- What is the difference between a homogeneous mixture and a heterogeneous mixture?
- Identify each as a heterogeneous mixture or a homogeneous mixture.
 - Salt is mixed with pepper.
 - Sugar is dissolved in water.
 - Pasta is cooked in boiling water.
- Identify each as a heterogeneous mixture or a homogeneous mixture.
 - air
 - dirt
 - a television set

11. In Exercise 9, which choices are also solutions?
12. In Exercise 10, which choices are also solutions?
13. Why is iron considered a metal?
14. Why is oxygen considered a nonmetal?
15. Distinguish between a metal and a nonmetal.
16. What properties do semimetals have?
17. Elemental carbon is a black, dull-looking solid that conducts heat and electricity well. It is very brittle and cannot be made into thin sheets or long wires. Of these properties, how does carbon behave as a metal? How does carbon behave as a nonmetal?
18. Pure silicon is shiny and silvery but does not conduct electricity or heat well. Of these properties, how does silicon behave as a metal? How does silicon behave as a nonmetal?

Answers

1.
 - a. matter
 - b. not matter
 - c. not matter
 - d. matter
3.
 - a. physical property
 - b. chemical property
 - c. physical property
5.
 - a. physical change
 - b. chemical change
 - c. physical change
7. An element is a fundamental chemical part of a substance; there are about 115 known elements. A compound is a combination of elements that acts as a different substance; there are over 50 million known substances.
9.
 - a. heterogeneous
 - b. homogeneous
 - c. heterogeneous
11. Choice b is a solution.
13. Iron is a metal because it is solid, is shiny, and conducts electricity and heat well.
15. Metals are typically shiny, conduct electricity and heat well, and are malleable and ductile; nonmetals are a variety of colors and phases, are brittle in the solid phase, and do not conduct heat or electricity well.
17. Carbon behaves as a metal because it conducts heat and electricity well. It is a nonmetal because it is black and brittle and cannot be made into sheets or wires.

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Figure 1.2

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Figure 1.3

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Figure 1.4

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Figure 1.6

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Figure 1.7

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Chemistry as a Science

Learning Objectives

1. Learn what science is and how it works.

Chemistry is a branch of science. Although science itself is difficult to define exactly, the following definition can serve as starting point. **Science** is the process of knowing about the natural universe through observation and experiment. Science is not the only process of knowing (e.g., the ancient Greeks simply sat and *thought*), but it has evolved over more than 350 years into the best process that humanity has devised to date to learn about the universe around us.

The process of science is usually stated as the *scientific method*, which is rather naïvely described as follows: (1) state a hypothesis, (2) test the hypothesis, and (3) refine the hypothesis. Actually, however, the process is not that simple. (For example, I don't go into my lab every day and exclaim, "I am going to state a hypothesis today and spend the day testing it!") The process is not that simple because science and scientists have a body of knowledge that has already been identified as coming from the highest level of understanding, and most scientists build from that body of knowledge.

An educated guess about how the natural universe works is called a **hypothesis**. A scientist who is familiar with how part of the natural universe works—say, a chemist—is interested in furthering that knowledge. That person makes a reasonable guess—a hypothesis—that is designed to see if the universe works in a new way as well. Here's an example of a hypothesis: "if I mix one part of hydrogen with one part of oxygen, I can make a substance that contains both elements."

Most good hypotheses are grounded in previously understood knowledge and represent a testable extension of that knowledge. The scientist then devises ways to test if that guess is or is not correct. That is, the scientist plans experiments. **Experiments** are tests of the natural universe to see if a guess (hypothesis) is correct. An experiment to test our previous hypothesis would be to actually mix hydrogen and oxygen and see what happens. Most experiments include observations of small, well-defined parts of the natural universe designed to see results of the experiments.

Why do we have to do experiments? Why do we have to test? Because the natural universe is not always so obvious, experiments are necessary. For example, it is fairly obvious that if you drop an object from a height, it will fall. Several hundred years ago (coincidentally, near the inception of modern science), the concept of gravity explained that test. However, is it obvious that the entire natural universe is composed of only about 115 fundamental chemical building blocks called elements? This wouldn't seem true if you looked at the world around you and saw all the different forms matter can take. In fact, the concept of *the element* is only about 200 years old, and the last naturally occurring element was identified about 80 years ago. It took decades of tests and millions of experiments to

establish what the elements actually are. These are just two examples; a myriad of such examples exists in chemistry and science in general.

When enough evidence has been collected to establish a general principle of how the natural universe works, the evidence is summarized in a theory. A **theory** is a general statement that explains a large number of observations. “All matter is composed of atoms” is a general statement, a theory, that explains many observations in chemistry. A theory is a very powerful statement in science. There are many statements referred to as “the theory of _____” or the “_____ theory” in science (where the blanks represent a word or concept). When written in this way, theories indicate that science has an overwhelming amount of evidence of its correctness. We will see several theories in the course of this text.

A specific statement that is thought to be never violated by the entire natural universe is called a *law*. A scientific law is the highest understanding of the natural universe that science has and is thought to be inviolate. For example, the fact that all matter attracts all other matter—the law of gravitation—is one such law. Note that the terms theory and law used in science have slightly different meanings from those in common usage; theory is often used to mean hypothesis (“I have a theory...”), whereas a law is an arbitrary limitation that can be broken but with potential consequences (such as speed limits). Here again, science uses these terms differently, and it is important to apply their proper definitions when you use these words in science.

There is an additional phrase in our definition of science: “the natural universe.” Science is concerned only with the natural universe. What is the natural universe? It’s anything that occurs around us, well, naturally. Stars; planets; the appearance of life on earth; and how animals, plants, and other matter function are all part of the natural universe. Science is concerned with that—and only that.

Of course, there are other things that concern us. For example, is the English language part of science? Most of us can easily answer no; English is not science. English is certainly worth knowing (at least for people in predominantly English-speaking countries), but why isn’t it science? English, or any human language, isn’t science because ultimately it is contrived; it is made up. Think of it: the word spelled b-l-u-e represents a certain colour, and we all agree what color that is. But what if we used the word h-a-r-d-n-r-f to describe that colour? (See Figure 1.8 “English Is Not Science”.) That would be fine—as long as everyone agreed. Anyone who has learned a second language must initially wonder why a certain word is used to describe a certain concept; ultimately, the speakers of that language agreed that a particular word would represent a particular concept. It was contrived.

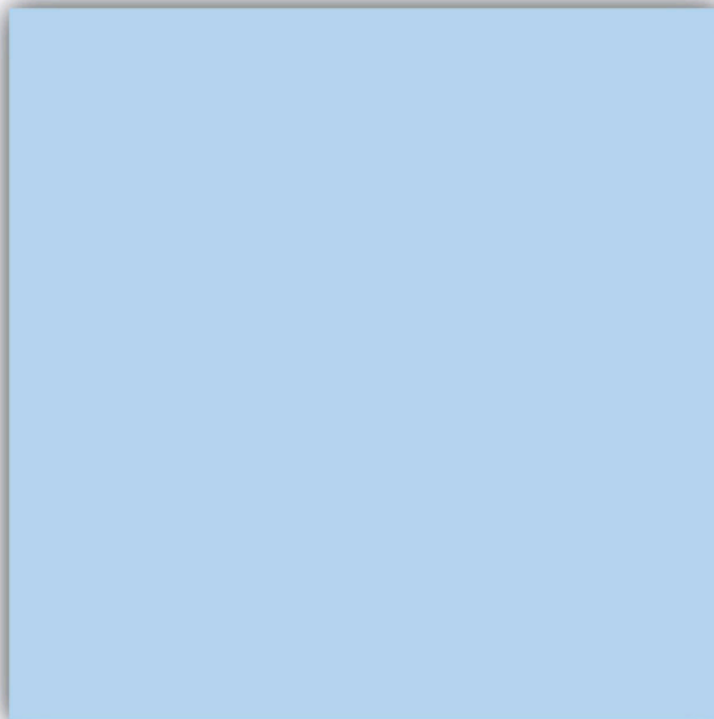


Figure 1.8 “English Is Not Science.”

How would you describe this colour? Blue or hard to describe? Either way, you’re not doing science. That doesn’t mean language isn’t worth knowing. It is very important in society. But it’s not *science*. Science deals only with what occurs naturally.

Example 1.4

Problems

Which of the following fields would be considered science?

1. geology, the study of the earth
2. ethics, the study of morality
3. political science, the study of governance
4. biology, the study of living organisms

Solutions

1. Because the earth is a natural object, the study of it is indeed considered part of science.
2. Ethics is a branch of philosophy that deals with right and wrong. Although these are useful concepts, they are not science.
3. There are many forms of government, but all are created by humans. Despite the fact that the word science appears in its name, political science is not true science.

4. Living organisms are part of the natural universe, so the study of them is part of science.

Test Yourself

Which is part of science, and which is not?

1. dynamics, the study of systems that change over time
2. aesthetics, the concept of beauty

Answers

1. science
2. not science

The field of science has gotten so big that it is common to separate it into more specific fields. First, there is mathematics, the language of science. All scientific fields use mathematics to express themselves—some more than others. Physics and astronomy are scientific fields concerned with the fundamental interactions between matter and energy. Chemistry, as defined previously, is the study of the interactions of matter with other matter and with energy. Biology is the study of living organisms, while geology is the study of the earth. Other sciences can be named as well. Understand that these fields are not always completely separate; the boundaries between scientific fields are not always readily apparent. Therefore, a scientist may be labelled a biochemist if he or she studies the chemistry of biological organisms.

Finally, understand that science can be either qualitative or quantitative. **Qualitative** implies a description of the quality of an object. For example, physical properties are generally qualitative descriptions: sulfur is yellow, your math book is heavy, or that statue is pretty. A **quantitative** description represents the specific amount of something; it means knowing how much of something is present, usually by counting or measuring it. As such, some quantitative descriptions would include 25 students in a class, 650 pages in a book, or a velocity of 66 miles per hour. Quantitative expressions are very important in science; they are also very important in chemistry.

Example 1.5

Problems

Identify each statement as either a qualitative description or a quantitative description.

1. Gold metal is yellow.
2. A ream of paper has 500 sheets in it.
3. The weather outside is snowy.
4. The temperature outside is 24 degrees Fahrenheit.

Solutions

1. Because we are describing a physical property of gold, this statement is qualitative.
2. This statement mentions a specific amount, so it is quantitative.
3. The word snowy is a description of how the day is; therefore, it is a qualitative statement.
4. In this case, the weather is described with a specific quantity—the temperature. Therefore, it is quantitative.

Test Yourself

Are these qualitative or quantitative statements?

1. Roses are red, and violets are blue.
2. Four score and seven years ago....

Answers

1. qualitative
2. quantitative

Food and Drink App: Carbonated Beverages

Some of the simple chemical principles discussed in this chapter can be illustrated with carbonated beverages: sodas, beer, and sparkling wines. Each product is produced in a different way, but they all have one thing in common. They are solutions of carbon dioxide dissolved in water.

Carbon dioxide is a compound composed of carbon and oxygen. Under normal conditions, it is a gas. If you cool it down enough, it becomes a solid known as dry ice. Carbon dioxide is an important compound in the cycle of life on earth.

Even though it is a gas, carbon dioxide can dissolve in water, just like sugar or salt can dissolve in water. When that occurs, we have a homogeneous mixture, or a solution, of carbon dioxide in water. However, very little carbon dioxide can dissolve in water. If the atmosphere were pure carbon dioxide, the solution would be only about 0.07% carbon dioxide. In reality, the air is only about 0.03% carbon dioxide, so the amount of carbon dioxide in water is reduced proportionally.

However, when soda and beer are made, manufacturers do two important things: they use pure carbon dioxide gas, and they use it at very high pressures. With higher pressures, more carbon dioxide can dissolve in the water. When the soda or beer container is sealed, the high pressure of carbon dioxide gas remains inside the package. (Of course, there are more ingredients in soda and beer besides carbon dioxide and water.)

When you open a container of soda or beer, you hear a distinctive *hiss* as the excess carbon dioxide gas escapes. But something else happens as well. The carbon dioxide in the solution comes out of solution as a bunch of tiny bubbles. These bubbles impart a pleasing sensation in the mouth, so much so that the soda industry sold over *225 billion* servings of soda in the United States alone in 2009.

Some sparkling wines are made in the same way—by forcing carbon dioxide into regular wine. Some sparkling wines (including champagne) are made by sealing a bottle of wine with some yeast in it. The yeast *ferments*, a process by which the yeast converts sugars into energy and excess carbon dioxide. The carbon dioxide produced by the yeast dissolves in the wine. Then, when the champagne bottle is opened, the increased pressure of carbon dioxide is released, and the drink bubbles just like an expensive glass of soda.



Figure 1.9 “Carbonation.” Soda, beer, and sparkling wine take advantage of the properties of a solution of carbon dioxide in water.

Key Takeaways

- Science is a process of knowing about the natural universe through observation and experiment.
- Scientists go through a rigorous process to determine new knowledge about the universe; this process is generally referred to as the scientific method.
- Science is broken down into various fields, of which chemistry is one.
- Science, including chemistry, is both qualitative and quantitative.

Exercises

Questions

1. Describe the scientific method.
2. What is the scientific definition of a hypothesis? Why is the phrase a hypothesis is just a guess an inadequate definition?
3. Why do scientists need to perform experiments?
4. What is the scientific definition of a theory? How is this word misused in general conversation?
5. What is the scientific definition of a law? How does it differ from the everyday definition of a law?
6. Name an example of a field that is not considered a science.
7. Which of the following fields are studies of the natural universe?
 - a. biophysics (a mix of biology and physics)
 - b. art
 - c. business
8. Which of the following fields are studies of the natural universe?
 - a. accounting
 - b. geochemistry (a mix of geology and chemistry)
 - c. astronomy (the study of stars and planets [but not the earth])
9. Which of these statements are qualitative descriptions?
 - a. The Titanic was the largest passenger ship build at that time.
 - b. The population of the United States is about 306,000,000 people.
 - c. The peak of Mount Everest is 29,035 feet above sea level.
10. Which of these statements are qualitative descriptions?
 - a. A regular movie ticket in Cleveland costs \$6.00.
 - b. The weather in the Democratic Republic of the Congo is the wettest in all of Africa.
 - c. The deepest part of the Pacific Ocean is the Mariana Trench.
11. Of the statements in Exercise 9, which are quantitative?
12. Of the statements in Exercise 10, which are quantitative?

Answers

1. Simply stated, the scientific method includes three steps: (1) stating a hypothesis, (2) testing the hypothesis, and (3) refining the hypothesis.
3. Scientists perform experiments to test their hypotheses because sometimes the nature of natural universe is not obvious.
5. A scientific law is a specific statement that is thought to be never violated by the entire natural universe. Everyday laws are arbitrary limits that society puts on its members.
7.
 - a. yes
 - b. no
 - c. no

9.
 - a. qualitative
 - b. not qualitative
 - c. not qualitative

11. Statements b and c are quantitative.

Media Attributions

Figure 1.9

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Chapter 2. Measurements

Data suggest that a male child will weigh 50% of his adult weight at about 11 years of age. However, he will reach 50% of his adult height at only 2 years of age. It is obvious, then, that people eventually stop growing up but continue to grow out. Data also suggest that the average human height has been increasing over time. In industrialized countries, the average height of people increased 5.5 inches from 1810 to 1984. Most scientists attribute this simple, basic measurement of the human body to better health and nutrition.

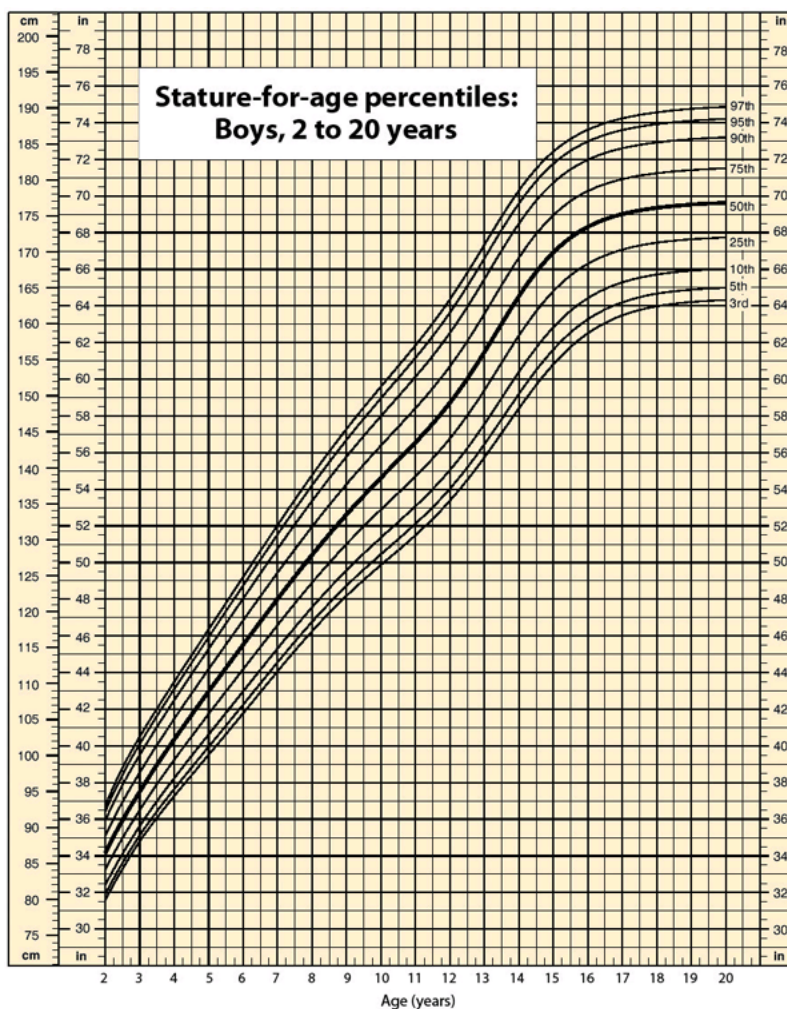


Figure 2.0 Stature-for-age growth chart for boys ages 2 to 20.

In 1983, an Air Canada airplane had to make an emergency landing because it unexpectedly ran out of fuel; ground personnel had filled the fuel tanks with a certain number of pounds of fuel, not kilograms of fuel. In 1999, the Mars Climate Orbiter spacecraft was lost attempting to orbit Mars because the thrusters were programmed in terms of English units, even though the engineers built the spacecraft using metric units. In 1993, a nurse mistakenly administered 23 units of morphine to a patient rather

than the “2–3” units prescribed. (The patient ultimately survived.) These incidents occurred because people weren’t paying attention to quantities.

Chemistry, like all sciences, is quantitative. It deals with *quantities*, things that have amounts and units. Dealing with quantities is very important in chemistry, as is relating quantities to each other. In this chapter, we will discuss how we deal with numbers and units, including how they are combined and manipulated.

Media Attributions

Figure 2.0

- [“Individual growth chart, 2–20 years: Boys stature-for-age”](#) by CDC ([original report, p. 29](#))
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Expressing Numbers

Learning Objectives

1. Learn to express numbers properly.

Quantities have two parts: the number and the unit. The number tells “how many.” It is important to be able to express numbers properly so that the quantities can be communicated properly.

Standard notation is the straightforward expression of a number. Numbers such as 17, 101.5, and 0.00446 are expressed in standard notation. For relatively small numbers, standard notation is fine. However, for very large numbers, such as 306,000,000, or for very small numbers, such as 0.000000419, standard notation can be cumbersome because of the number of zeros needed to place nonzero numbers in the proper position.

Scientific notation is an expression of a number using powers of 10. Powers of 10 are used to express numbers that have many zeros:

Table 2.1 Powers of 10

10^0	= 1
10^1	= 10
10^2	= 100 = 10×10
10^3	= 1,000 = $10 \times 10 \times 10$
10^4	= 10,000 = $10 \times 10 \times 10 \times 10$

and so forth. The raised number to the right of the 10 indicating the number of factors of 10 in the original number is the **exponent**. (Scientific notation is sometimes called *exponential notation*.) The exponent’s value is equal to the number of zeros in the number expressed in standard notation.

Small numbers can also be expressed in scientific notation but with negative exponents:

Table 2.2 Powers of Negative 10

10^{-1}	$= 0.1 = \frac{1}{10}$
10^{-2}	$= 0.01 = \frac{1}{100}$
10^{-3}	$= 0.001 = \frac{1}{1,000}$
10^{-4}	$= 0.0001 = \frac{1}{10,000}$

and so forth. Again, the value of the exponent is equal to the number of zeros in the denominator of the associated fraction. A negative exponent implies a decimal number less than one.

A number is expressed in scientific notation by writing the first nonzero digit, then a decimal point, and then the rest of the digits. The part of a number in scientific notation that is multiplied by a power of 10 is called the **coefficient**. Then determine the power of 10 needed to make that number into the original number and multiply the written number by the proper power of 10. For example, to write 79,345 in scientific notation,

$$79,345 = 7.9345 \times 10,000 = 7.9345 \times 10^4$$

Thus, the number in scientific notation is 7.9345×10^4 . For small numbers, the same process is used, but the exponent for the power of 10 is negative:

$$0.000411 = 4.11 \times \frac{1}{10,000} = 4.11 \times 10^{-4}$$

Typically, the extra zero digits at the end or the beginning of a number are not included.

Example 2.1

Problems

Express these numbers in scientific notation.

1. 306,000
2. 0.00884
3. 2,760,000
4. 0.000000559

Solutions

1. The number 306,000 is 3.06 times 100,000, or 3.06 times 10^5 . In scientific notation, the number is 3.06×10^5 .
2. The number 0.00884 is 8.84 times $\frac{1}{1,000}$, which is 8.84 times 10^{-3} . In scientific notation, the number is 8.84×10^{-3} .
3. The number 2,760,000 is 2.76 times 1,000,000, which is the same as 2.76 times 10^6 . In scientific notation, the number is written as 2.76×10^6 . Note that we omit the zeros at the end of the original number.
4. The number 0.000000559 is 5.59 times $\frac{1}{10,000,000}$, which is 5.59 times 10^{-7} . In scientific notation, the number is written as 5.59×10^{-7} .

Test Yourself

Express these numbers in scientific notation.

1. 23,070
2. 0.0009706

Answers

1. 2.307×10^4
2. 9.706×10^{-4}

Another way to determine the power of 10 in scientific notation is to count the number of places you need to move the decimal point to get a numerical value between 1 and 10. The number of places equals the power of 10. This number is positive if you move the decimal point to the right and negative if you move the decimal point to the left.

Many quantities in chemistry are expressed in scientific notation. When performing calculations, you may have to enter a number in scientific notation into a calculator. Be sure you know how to correctly enter a number in scientific notation into your calculator. Different models of calculators require different actions for properly entering scientific notation. If in doubt, consult your instructor immediately.

Key Takeaways

- Standard notation expresses a number normally.
- Scientific notation expresses a number as a coefficient times a power of 10.
- The power of 10 is positive for numbers greater than 1 and negative for numbers between 0 and 1.



This calculator shows only the coefficient and the power of 10 to represent the number in scientific notation. Thus, the number being displayed is 3.84951×10^{18} , or 3,849,510,000,000,000,000.

Exercises

Questions

- Express these numbers in scientific notation.
 - 56.9
 - 563,100
 - 0.0804
 - 0.00000667
- Express these numbers in scientific notation.
 - 890,000
 - 602,000,000,000
 - 0.0000004099
 - 0.0000000000000011
- Express these numbers in scientific notation.
 - 0.00656
 - 65,600
 - 4,567,000
 - 0.000005507

4. Express these numbers in scientific notation.
- 65
 - 321.09
 - 0.000077099
 - 0.00000000218
5. Express these numbers in standard notation.
- 1.381×10^5
 - 5.22×10^{-7}
 - 9.998×10^4
6. Express these numbers in standard notation.
- 7.11×10^{-2}
 - 9.18×10^2
 - 3.09×10^{-10}
7. Express these numbers in standard notation.
- 8.09×10^0
 - 3.088×10^{-5}
 - -4.239×10^2
8. Express these numbers in standard notation.
- 2.87×10^{-8}
 - 1.78×10^{11}
 - 1.381×10^{-23}
9. These numbers are not written in proper scientific notation. Rewrite them so that they are in proper scientific notation.
- 72.44×10^3
 - $9,943 \times 10^{-5}$
 - $588,399 \times 10^2$
10. These numbers are not written in proper scientific notation. Rewrite them so that they are in proper scientific notation.
- 0.000077×10^{-7}
 - 0.000111×10^8
 - $602,000 \times 10^{18}$
11. These numbers are not written in proper scientific notation. Rewrite them so that they are in proper scientific notation.
- 345.1×10^2
 - 0.234×10^{-3}
 - $1,800 \times 10^{-2}$
12. These numbers are not written in proper scientific notation. Rewrite them so that they are in proper scientific notation.
- $8,099 \times 10^{-8}$
 - 34.5×10^0
 - 0.000332×10^4
13. Write these numbers in scientific notation by counting the number of places the decimal point is moved.

- a. 123,456.78
b. 98,490
c. 0.000000445
14. Write these numbers in scientific notation by counting the number of places the decimal point is moved.
a. 0.000552
b. 1,987
c. 0.0000000887
15. Use your calculator to evaluate these expressions. Express the final answer in proper scientific notation.
a. $456 \times (7.4 \times 10^8) = ?$
b. $(3.02 \times 10^5) \div (9.04 \times 10^{15}) = ?$
c. $0.0044 \times 0.000833 = ?$
16. Use your calculator to evaluate these expressions. Express the final answer in proper scientific notation.
a. $98,000 \times 23,000 = ?$
b. $98,000 \div 23,000 = ?$
c. $(4.6 \times 10^{-5}) \times (2.09 \times 10^3) = ?$
17. Use your calculator to evaluate these expressions. Express the final answer in proper scientific notation.
a. $45 \times 132 \div 882 = ?$
b. $[(6.37 \times 10^4) \times (8.44 \times 10^{-4})] \div (3.2209 \times 10^{15}) = ?$
18. Use your calculator to evaluate these expressions. Express the final answer in proper scientific notation.
a. $(9.09 \times 10^8) \div [(6.33 \times 10^9) \times (4.066 \times 10^{-7})] = ?$
b. $9,345 \times 34.866 \div 0.00665 = ?$

Answers

1. a. 5.69×10^1
b. 5.631×10^5
c. 8.04×10^{-2}
d. 6.67×10^{-6}
3. a. 6.56×10^{-3}
b. 6.56×10^4
c. 4.567×10^6
d. 5.507×10^{-6}
5. a. 138,100
b. 0.000000522
c. 99,980
7. a. 8.09
b. 0.00003088
c. -423.9
9. a. 7.244×10^4

11. b. 9.943×10^{-2}
 c. 5.88399×10^7
- a. 3.451×10^4
 b. 2.34×10^{-4}
 c. 1.8×10^1
13. a. 1.2345678×10^5
 b. 9.849×10^4
 c. 4.45×10^{-7}
15. a. 3.3744×10^{11}
 b. 3.3407×10^{-11}
 c. 3.665×10^{-6}
17. a. 6.7346×10^0
 b. 1.6691×10^{-14}

Media Attributions

Key Takeaways

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Significant Figures

Learning Objectives

1. Apply the concept of significant figures to limit a measurement to the proper number of digits.
2. Recognize the number of significant figures in a given quantity.
3. Limit mathematical results to the proper number of significant figures.

If you use a calculator to evaluate the expression $\frac{337}{217}$, you will get the following:

$$\frac{337}{217} = 1.55299539171 \dots$$

and so on for many more digits. Although this answer is correct, it is somewhat presumptuous. You start with two values that each have three digits, and the answer has *twelve* digits? That does not make much sense from a strict numerical point of view.

Consider using a ruler to measure the width of an object, as shown in Figure 2.1 “Expressing Width”. The object is definitely more than 1 cm long, so we know that the first digit in our measurement is 1. We see by counting the tick marks on the ruler that the object is at least three ticks after the 1. If each tick represents 0.1 cm, then we know the object is at least 1.3 cm wide. But our ruler does not have any more ticks between the 0.3 and the 0.4 marks, so we can’t know exactly how much the next decimal place is. But with a practiced eye we can estimate it. Let us estimate it as about six-tenths of the way between the third and fourth tick marks, which estimates our hundredths place as 6, so we identify a measurement of 1.36 cm for the width of the object.

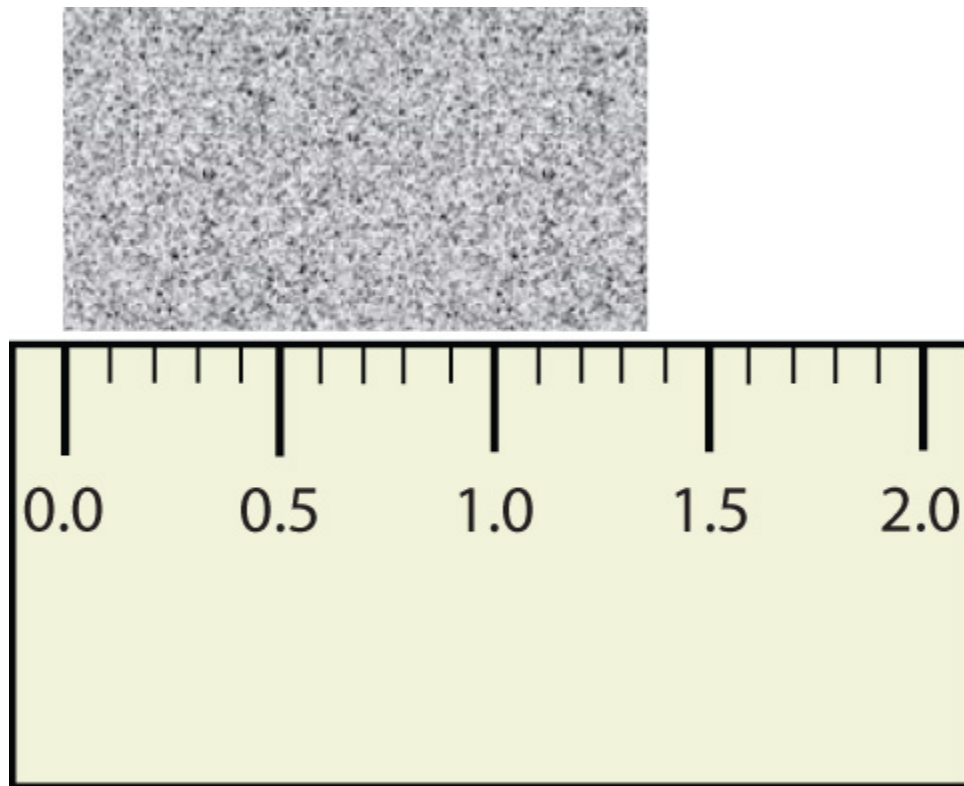


Figure 2.1 “Expressing Width.”

What is the proper way to express the width of this object?

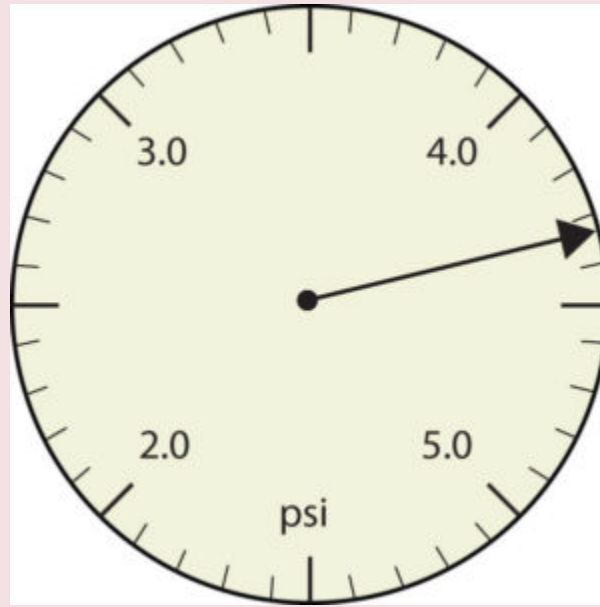
Does it make any sense to try to report a thousandths place for the measurement? No, it doesn't; we are not exactly sure of the hundredths place (after all, it was an estimate only), so it would be fruitless to estimate a thousandths place. Our best measurement, then, stops at the hundredths place, and we report 1.36 cm as proper measurement.

This concept of reporting the proper number of digits in a measurement or a calculation is called **significant figures**. Significant figures (sometimes called significant digits) represent the limits of what values of a measurement or a calculation we are sure of. The convention for a measurement is that the quantity reported should be all known values and the first estimated value. The conventions for calculations are discussed as follows.

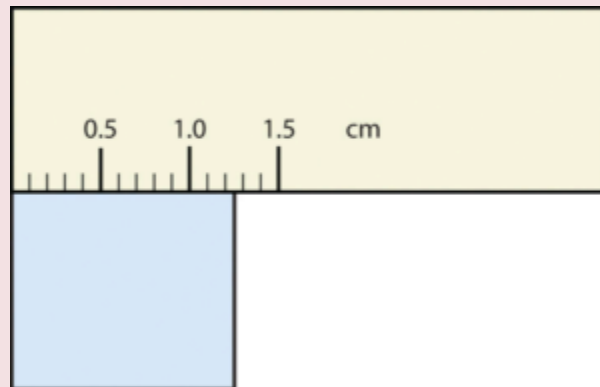
Example 2.2

Problems

Use each diagram to report a measurement to the proper number of significant figures.



1.



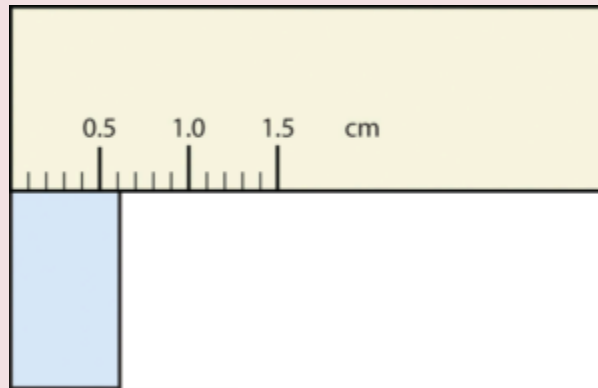
2.

Solutions

1. The arrow is between 4.0 and 5.0, so the measurement is at least 4.0. The arrow is between the third and fourth small tick marks, so it's at least 0.3. We will have to estimate the last place. It looks like about one-third of the way across the space, so let us estimate the hundredths place as 3. Combining the digits, we have a measurement of 4.33 psi (psi stands for "pounds per square inch" and is a unit of pressure, like air in a tire). We say that the measurement is reported to three significant figures.
2. The rectangle is at least 1.0 cm wide but certainly not 2.0 cm wide, so the first significant digit is 1. The rectangle's width is past the second tick mark but not the third; if each tick mark represents 0.1, then the rectangle is at least 0.2 in the next significant digit. We have to estimate the next place because there are no markings to guide us. It appears to be about halfway between 0.2 and 0.3, so we will estimate the next place to be a 5. Thus, the measured width of the rectangle is 1.25 cm. Again, the measurement is reported to three significant figures.

Test Yourself

What would be the reported width of this rectangle?

**Answer**

0.63 cm

In many cases, you will be given a measurement. How can you tell by looking what digits are significant? For example, the reported population of the United States is 306,000,000. Does that mean that it is *exactly* three hundred six million or is some estimation occurring?

The following conventions dictate which numbers in a reported measurement are significant and which are not significant:

1. Any nonzero digit is significant.
2. Any zeros between nonzero digits (i.e., embedded zeros) are significant.
3. Zeros at the end of a number without a decimal point (i.e., trailing zeros) are not significant; they serve only to put the significant digits in the correct positions. However, zeros at the end of any number with a decimal point are significant.
4. Zeros at the beginning of a decimal number (i.e., leading zeros) are not significant; again, they serve only to put the significant digits in the correct positions.

So, by these rules, the population figure of the United States has only three significant figures: the 3, the 6, and the zero between them. The remaining six zeros simply put the 306 in the millions position.

Example 2.3

Problems

Give the number of significant figures in each measurement.

1. 36.7 m
2. 0.006606 s
3. 2,002 kg

4. 306,490,000 people

Solutions

1. By rule 1, all nonzero digits are significant, so this measurement has three significant figures.
2. By rule 4, the first three zeros are not significant, but by rule 2 the zero between the sixes is; therefore, this number has four significant figures.
3. By rule 2, the two zeros between the twos are significant, so this measurement has four significant figures.
4. The four trailing zeros in the number are not significant, but the other five numbers are, so this number has five significant figures.

Test Yourself

Give the number of significant figures in each measurement.

1. 0.000601 m
2. 65.080 kg

Answers

1. three significant figures
2. five significant figures

How are significant figures handled in calculations? It depends on what type of calculation is being performed. If the calculation is an addition or a subtraction, the rule is as follows: limit the reported answer to the rightmost column that all numbers have significant figures in common. For example, if you were to add 1.2 and 4.41, we note that the first number stops its significant figures in the tenths column, while the second number stops its significant figures in the hundredths column. We therefore limit our answer to the tenths column.

$$\begin{array}{r} 1.2 \\ + \underline{4.41} \\ \hline 5.61 \end{array}$$

limit final answer to the tenths column: 5.6

We drop the last digit—the 1—because it is not significant to the final answer.

The dropping of positions in sums and differences brings up the topic of rounding. Although there are several conventions, in this text we will adopt the following rule: the final answer should be rounded up if the first dropped digit is 5 or greater and rounded down if the first dropped digit is less than 5.

$$\begin{array}{r} 77.2 \\ + \underline{10.46} \\ \hline 87.66 \end{array}$$

limit final answer to the tenths column and round up: 87.7

Example 2.4

Problems

Express the final answer to the proper number of significant figures.

1. $101.2 + 18.702 = ?$
2. $202.88 - 1.013 = ?$

Solutions

1. If we use a calculator to add these two numbers, we would get 119.902. However, most calculators do not understand significant figures, and we need to limit the final answer to the tenths place. Thus, we drop the 02 and report a final answer of 119.9 (rounding down).
2. A calculator would answer 201.867. However, we have to limit our final answer to the hundredths place. Because the first number being dropped is 7, which is greater than 6, we round up and report a final answer of 201.87.

Test Yourself

Express the answer for $3.445 + 90.83 - 72.4$ to the proper number of significant figures.

Answer

21.9

If the operations being performed are multiplication or division, the rule is as follows: limit the answer to the number of significant figures that the data value with the *least* number of significant figures has. So if we are dividing 23 by 448, which have two and three significant figures each, we should limit the final reported answer to two significant figures (the lesser of two and three significant figures):

$$23/448 = 0.051339286\dots = 0.051$$

The same rounding rules apply in multiplication and division as they do in addition and subtraction.

Example 2.5

Problems

Express the final answer to the proper number of significant figures.

1. $76.4 \times 180.4 = ?$
2. $934.9 \div 0.00455 = ?$

Solutions

1. The first number has three significant figures, while the second number has four significant figures. Therefore, we limit our final answer to three significant figures: $76.4 \times 180.4 = 13,782.56 = 13,800$.
2. The first number has four significant figures, while the second number has three significant figures. Therefore we limit our final answer to three significant figures: $934.9 \div 0.00455 = 205,472.5275\dots = 205,000$.

Test Yourself

Express the final answer to the proper number of significant figures.

1. $22.4 \times 8.314 = ?$
2. $1.381 \div 6.02 = ?$

Answers

1. 186
2. 0.229

As you have probably realized by now, the biggest issue in determining the number of significant figures in a value is the zero. Is the zero significant or not? One way to unambiguously determine whether a zero is significant or not is to write a number in scientific notation. Scientific notation will include zeros in the coefficient of the number *only if they are significant*. Thus, the number 8.666×10^6 has four significant figures. However, the number 8.6660×10^6 has five significant figures. That last zero is significant; if it were not, it would not be written in the coefficient. So when in doubt about expressing the number of significant figures in a quantity, use scientific notation and include the number of zeros that are truly significant.

View this [video on significant figures](#) by Dr. Jessie A. Key for more information.

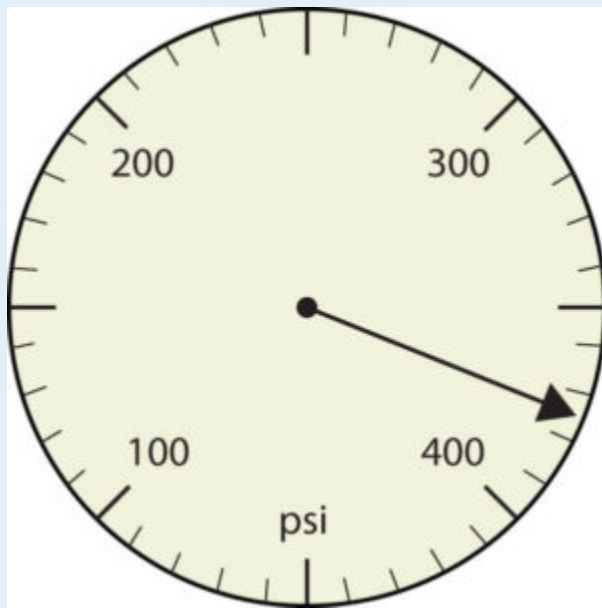
Key Takeaways

- Significant figures in a quantity indicate the number of known values plus one place that is estimated.
- There are rules for which numbers in a quantity are significant and which are not significant.
- In calculations involving addition and subtraction, limit significant figures based on the rightmost place that all values have in common.
- In calculations involving multiplication and division, limit significant figures to the least number of significant figures in all the data values.

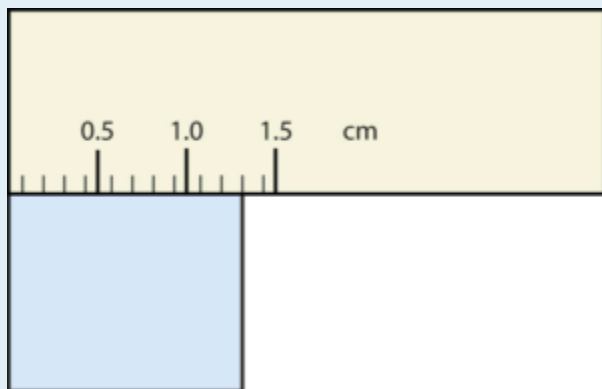
Exercises

Questions

1. Express each measurement to the correct number of significant figures.

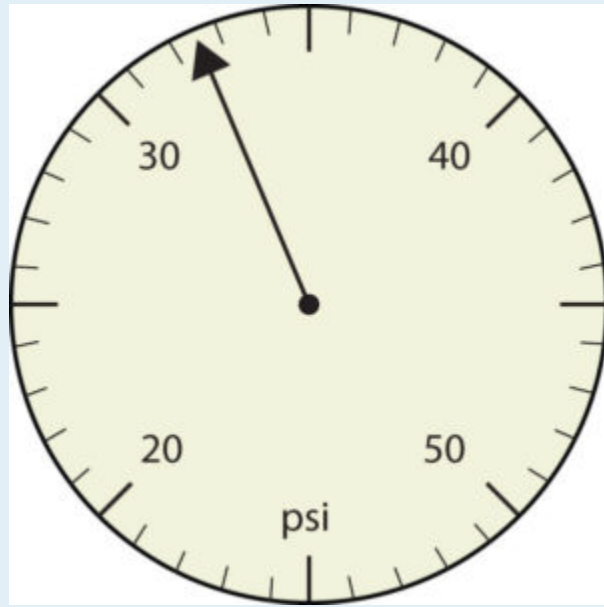


a.

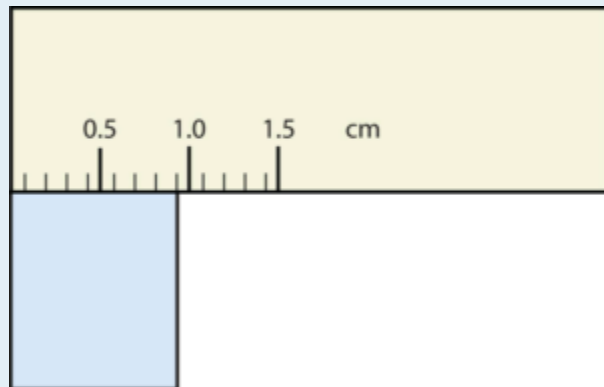


b.

2. Express each measurement to the correct number of significant figures.



a.



b.

3. How many significant figures do these numbers have?
- 23
 - 23.0
 - 0.00023
 - 0.0002302
4. How many significant figures do these numbers have?
- 5.44×10^8
 - 1.008×10^{-5}
 - 43.09
 - 0.0000001381
5. How many significant figures do these numbers have?
- 765,890
 - 765,890.0
 - 1.2000×10^5
 - 0.0005060
6. How many significant figures do these numbers have?

- a. 0.009
 - b. 0.0000009
 - c. 65,444
 - d. 65,040
7. Compute and express each answer with the proper number of significant figures, rounding as necessary.
- a. $56.0 + 3.44 = ?$
 - b. $0.00665 + 1.004 = ?$
 - c. $45.99 - 32.8 = ?$
 - d. $45.99 - 32.8 + 75.02 = ?$
8. Compute and express each answer with the proper number of significant figures, rounding as necessary.
- a. $1.005 + 17.88 = ?$
 - b. $56,700 - 324 = ?$
 - c. $405,007 - 123.3 = ?$
 - d. $55.5 + 66.66 - 77.777 = ?$
9. Compute and express each answer with the proper number of significant figures, rounding as necessary.
- a. $56.7 \times 66.99 = ?$
 - b. $1.000 \div 77 = ?$
 - c. $1.000 \div 77.0 = ?$
 - d. $6.022 \times 1.89 = ?$
10. Compute and express each answer with the proper number of significant figures, rounding as necessary.
- a. $0.000440 \times 17.22 = ?$
 - b. $203,000 \div 0.044 = ?$
 - c. $67 \times 85.0 \times 0.0028 = ?$
 - d. $999,999 \div 3,310 = ?$
11. Write the number 87,449 in scientific notation with four significant figures.
12. Write the number 0.000066600 in scientific notation with five significant figures.
13. Write the number 306,000,000 in scientific notation to the proper number of significant figures.
14. Write the number 0.0000558 in scientific notation with two significant figures.
15. Perform each calculation and limit each answer to three significant figures.
- a. $67,883 \times 0.004321 = ?$
 - b. $(9.67 \times 10^3) \times 0.0055087 = ?$
16. Perform each calculation and limit each answer to four significant figures.
- a. $18,900 \times 76.33 \div 0.00336 = ?$
 - b. $0.77604 \div 76,003 \times 8.888 = ?$

Answers

1.
 - a. 375 psi
 - b. 1.30 cm

3. a. two
 b. three
 c. two
 d. four
5. a. five
 b. seven
 c. five
 d. four
7. a. 59.4
 b. 1.011
 c. 13.2
 d. 88.2
9. a. 3.80×10^3
 b. 0.013
 c. 0.0130
 d. 11.4
11. a. 8.745×10^4
 b. 6.6600×10^{-5}
13. a. 293
 b. 53.3

Converting Units

Learning Objectives

1. Convert from one unit to another unit of the same type.

In the section [Expressing Units](#), we showed some examples of how to replace initial units with other units of the same type to get a numerical value that is easier to comprehend. In this section, we will formalize the process.

Consider a simple example: how many feet are there in 4 yards? Most people will almost automatically answer that there are 12 feet in 4 yards. How did you make this determination? Well, if there are 3 feet in 1 yard and there are 4 yards, then there are $4 \times 3 = 12$ feet in 4 yards.

This is correct, of course, but it is informal. Let us formalize it in a way that can be applied more generally. We know that 1 yard (yd) equals 3 feet (ft):

$$1 \text{ yd} = 3 \text{ ft}$$

In math, this expression is called an *equality*. The rules of algebra say that you can change (i.e., multiply or divide or add or subtract) the equality (as long as you don't divide by zero) and the new expression will still be an equality. For example, if we divide both sides by 2, we get:

$$\frac{1}{2} \text{ yd} = \frac{3}{2} \text{ ft}$$

We see that one-half of a yard equals $\frac{3}{2}$, or one and a half, feet—something we also know to be true, so the above equation is still an equality. Going back to the original equality, suppose we divide both sides of the equation by 1 yard (number *and* unit):

$$\frac{1 \text{ yd}}{1 \text{ yd}} = \frac{3 \text{ ft}}{1 \text{ yd}}$$

The expression is still an equality, by the rules of algebra. The left fraction equals 1. It has the same quantity in the numerator and the denominator, so it must equal 1. The quantities in the numerator and denominator cancel, both the number *and* the unit:

$$\frac{\cancel{1 \text{ yd}}}{\cancel{1 \text{ yd}}} = \frac{3 \text{ ft}}{1 \text{ yd}}$$

When everything cancels in a fraction, the fraction reduces to 1:

$$1 = \frac{3 \text{ ft}}{1 \text{ yd}}$$

We have an expression, $\frac{3 \text{ ft}}{1 \text{ yd}}$, that equals 1. This is a strange way to write 1, but it makes sense: 3 ft equals 1 yd, so the quantities in the numerator and denominator are the same quantity, just expressed with different units. The expression $\frac{3 \text{ ft}}{1 \text{ yd}}$ is called a **conversion factor**, and it is used to formally change the unit of a quantity into another unit. (The process of converting units in such a formal fashion is sometimes called *dimensional analysis* or the *factor label method*.)

To see how this happens, let us start with the original quantity:

$$4 \text{ yd}$$

Now let us multiply this quantity by 1. When you multiply anything by 1, you don't change the value of the quantity. Rather than multiplying by just 1, let us write 1 as $\frac{3 \text{ ft}}{1 \text{ yd}}$:

$$4 \text{ yd} \times \frac{3 \text{ ft}}{1 \text{ yd}}$$

The 4 yd term can be thought of as $\frac{4 \text{ yd}}{1}$; that is, it can be thought of as a fraction with 1 in the denominator. We are essentially multiplying fractions. If the same thing appears in the numerator and denominator of a fraction, they cancel. In this case, what cancels is the unit *yard*:

$$\cancel{4 \text{ yd}} \times \frac{3 \text{ ft}}{\cancel{1 \text{ yd}}}$$

That is all that we can cancel. Now, multiply and divide all the numbers to get the final answer:

$$\frac{4 \times 3 \text{ ft}}{1} = \frac{12 \text{ ft}}{1} = 12 \text{ ft}$$

Again, we get an answer of 12 ft, just as we did originally. But in this case, we used a more formal procedure that is applicable to a variety of problems.

How many millimetres are in 14.66 m? To answer this, we need to construct a conversion factor between millimetres and metres and apply it correctly to the original quantity. We start with the definition of a millimetre, which is:

$$1 \text{ mm} = \frac{1}{1,000} \text{ m}$$

The $\frac{1}{1,000}$ is what the prefix *milli-* means. Most people are more comfortable working without fractions, so we will rewrite this equation by bringing the 1,000 into the numerator of the other side of the equation:

$$1,000 \text{ mm} = 1 \text{ m}$$

Now we construct a conversion factor by dividing one quantity into both sides. But now a question arises: which quantity do we divide by? It turns out that we have two choices, and the two choices will give us different conversion factors, both of which equal 1:

$$\frac{1000 \text{ mm}}{1000 \text{ mm}} = \frac{1 \text{ m}}{1000 \text{ mm}} \quad \text{or} \quad \frac{1000 \text{ mm}}{1 \text{ m}} = \frac{1 \text{ m}}{1 \text{ m}}$$

$$1 = \frac{1 \text{ m}}{1000 \text{ mm}} \quad \text{or} \quad \frac{1000 \text{ mm}}{1 \text{ m}} = 1$$

Which conversion factor do we use? The answer is based on *what unit you want to get rid of in your initial quantity*. The original unit of our quantity is metres, which we want to convert to millimetres. Because the original unit is assumed to be in the numerator, to get rid of it, we want the metre unit in the *denominator*; then they will cancel. Therefore, we will use the second conversion factor. Cancelling units and performing the mathematics, we get:

$$14.66 \cancel{\text{ m}} \times \frac{1000 \text{ mm}}{1 \cancel{\text{ m}}} = 14,660 \text{ mm}$$

Note how m cancels, leaving mm, which is the unit of interest.

The ability to construct and apply proper conversion factors is a very powerful mathematical technique in chemistry. You need to master this technique if you are going to be successful in this and future courses.

Example 2.6

Problems

1. Convert 35.9 kL to litres.
2. Convert 555 nm to metres.

Solutions

1. We will use the fact that 1 kL = 1,000 L. Of the two conversion factors that can be defined, the one that will work is $\frac{1,000 \text{ L}}{1 \text{ kL}}$. Applying this conversion factor, we get:

$$35.9 \cancel{\text{ kL}} \times \frac{1000 \text{ L}}{1 \cancel{\text{ kL}}} = 35,900 \text{ L}$$

2. We will use the fact that $1 \text{ nm} = \frac{1}{1,000,000,000} \text{ m}$, which we will rewrite as $1,000,000,000 \text{ nm} = 1 \text{ m}$, or $10^9 \text{ nm} = 1 \text{ m}$. Of the two possible conversion factors, the appropriate one has the nm unit in the denominator: $\frac{1 \text{ m}}{10^9 \text{ nm}}$. Applying this conversion factor, we get:

$$555 \cancel{\text{ nm}} \times \frac{1 \text{ m}}{10^9 \cancel{\text{ nm}}} = 0.000000555 \text{ m} = 5.55 \times 10^{-7} \text{ m}$$

In the final step, we expressed the answer in scientific notation.

Test Yourself

1. Convert 67.08 μL to litres.
2. Convert 56.8 m to kilometres.

Answers

1. 6.708×10^{-5} L
2. 5.68×10^{-2} km

What if we have a derived unit that is the product of more than one unit, such as m^2 ? Suppose we want to convert square metres to square centimetres? The key is to remember that m^2 means $\text{m} \times \text{m}$, which means we have *two* metre units in our derived unit. That means we have to include *two* conversion factors, one for each unit. For example, to convert 17.6 m^2 to square centimetres, we perform the conversion as follows:

$$\begin{aligned} 17.6 \text{ m}^2 &= 17.6(\cancel{\text{m}} \times \cancel{\text{m}}) \times \frac{100 \text{ cm}}{1 \cancel{\text{m}}} \times \frac{100 \text{ cm}}{1 \cancel{\text{m}}} \\ &= 176,000 \text{ cm} \times \text{cm} \\ &= 1.76 \times 10^5 \text{ cm}^2 \end{aligned}$$

Example 2.7

Problem

How many cubic centimetres are in 0.883 m^3 ?

Solution

With an exponent of 3, we have three length units, so by extension we need to use three conversion factors between metres and centimetres. Thus, we have:

$$0.833 \cancel{\text{m}}^3 \times \frac{100 \text{ cm}}{1 \cancel{\text{m}}} \times \frac{100 \text{ cm}}{1 \cancel{\text{m}}} \times \frac{100 \text{ cm}}{1 \cancel{\text{m}}} = 883,000 \text{ cm}^3 = 8.83 \times 10^5 \text{ cm}^3$$

You should demonstrate to yourself that the three metre units do indeed cancel.

Test Yourself

How many cubic millimetres are present in 0.0923 m^3 ?

Answer

$$9.23 \times 10^7 \text{ mm}^3$$

Suppose the unit you want to convert is in the denominator of a derived unit; what then? Then, in the conversion factor, the unit you want to remove must be in the *numerator*. This will cancel with the original unit in the denominator and introduce a new unit in the denominator. The following example illustrates this situation.

Example 2.8

Problem

Convert 88.4 m/min to metres/second.

Solution

We want to change the unit in the denominator from minutes to seconds. Because there are 60 seconds in 1 minute (60 s = 1 min), we construct a conversion factor so that the unit we want to remove, minutes, is in the numerator: $\frac{1 \text{ min}}{60 \text{ s}}$. Apply and perform the math:

$$\frac{88.4 \text{ m}}{\cancel{\text{min}}} \times \frac{1 \cancel{\text{min}}}{60 \text{ s}} = 1.47 \text{ m/s}$$

Notice how the 88.4 automatically goes in the numerator. That's because any number can be thought of as being in the numerator of a fraction divided by 1.

Test Yourself

Convert 0.203 m/min to metres/second.

Answer

$$0.00338 \text{ m/s or } 3.38 \times 10^{-3} \text{ m/s}$$



A common garden snail moves at a rate of about 0.2 m/min, which is about 0.003 m/s, which is 3 mm/s!

Sometimes there will be a need to convert from one unit with one numerical prefix to another unit with a different numerical prefix. How do we handle those conversions? Well, you could memorize the conversion factors that interrelate all numerical prefixes. Or you can go the easier route: first convert the quantity to the base unit, the unit with no numerical prefix, using the definition of the original prefix. Then convert the quantity in the base unit to the desired unit using the definition of the second prefix. You can do the conversion in two separate steps or as one long algebraic step. For example, to convert 2.77 kg to milligrams:

$$\text{convert to the base unit of grams} \quad 2.77 \cancel{\text{kg}} \times \frac{1000 \text{ g}}{1 \cancel{\text{kg}}} = 2770 \text{ g}$$

$$\text{convert to the desired unit} \quad 2770 \text{ g} \times \frac{1000 \text{ mg}}{1 \text{ g}} = 2,770,000 \text{ mg} = 2.77 \times 10^6 \text{ mg}$$

Alternatively, it can be done in a single multistep process:

$$2.77 \cancel{\text{kg}} \times \frac{1000 \text{ g}}{1 \cancel{\text{kg}}} \times \frac{1000 \text{ mg}}{1 \text{ g}} = 2,770,000 \text{ mg} = 2.77 \times 10^6 \text{ mg}$$

You get the same answer either way.

Example 2.9

Problem

How many nanoseconds are in 368.09 μs ?

Solution

You can either do this as a one-step conversion from microseconds to nanoseconds or convert to the base unit first and then to the

final desired unit. We will use the second method here, showing the two steps in a single line. Using the definitions of the prefixes *micro-* and *nano-*, we get:

$$368.09 \mu\text{s} \times \frac{1 \cancel{\mu}}{10^6 \cancel{\mu}\text{s}} \times \frac{10^9 \text{ ns}}{1 \cancel{\text{s}}} = 368,090 \text{ ns} = 3.6809 \times 10^5 \text{ ns}$$

Test Yourself

How many millilitres are in 607.8 kL?

Answer

6.078×10^8 mL

When considering the significant figures of a final numerical answer in a conversion, there is one important case where a number does not impact the number of significant figures in a final answer—the so-called **exact number**. An exact number is a number from a defined relationship, not a measured one. For example, the prefix *kilo-* means 1,000—*exactly* 1,000, no more or no less. Thus, in constructing the conversion factor

$$\frac{1000 \text{ g}}{1 \text{ kg}}$$

neither the 1,000 nor the 1 enter into our consideration of significant figures. The numbers in the numerator and denominator are defined exactly by what the prefix *kilo-* means. Another way of thinking about it is that these numbers can be thought of as having an infinite number of significant figures, such as:

$$\frac{1000.0000000000 \dots \text{ g}}{1.0000000000 \dots \text{ kg}}$$

The other numbers in the calculation will determine the number of significant figures in the final answer.

Example 2.10

Problem

A rectangular plot in a garden has the dimensions 36.7 cm by 128.8 cm. What is the area of the garden plot in square metres? Express your answer in the proper number of significant figures.

Solution

Area is defined as the product of the two dimensions, which we then have to convert to square metres and express our final answer to the correct number of significant figures, which in this case will be three.

$$36.7 \cancel{\text{cm}} \times 128.8 \cancel{\text{cm}} \times \frac{1 \text{ m}}{100 \cancel{\text{cm}}} \times \frac{1 \text{ m}}{100 \cancel{\text{cm}}} = 0.472696 \text{ m}^2 = 0.473 \text{ m}^2$$

The 1 and 100 in the conversion factors do not affect the determination of significant figures because they are exact numbers, defined by the *centi*- prefix.

Test Yourself

What is the volume of a block in cubic metres whose dimensions are 2.1 cm × 34.0 cm × 118 cm?

Answer

0.0084 m³

Chemistry Is Everywhere: The Gimli Glider

On July 23, 1983, an Air Canada Boeing 767 jet had to glide to an emergency landing at Gimli Industrial Park Airport in Gimli, Manitoba, because it unexpectedly ran out of fuel during flight. There was no loss of life in the course of the emergency landing, only some minor injuries associated in part with the evacuation of the craft after landing. For the remainder of its operational life (the plane was retired in 2008), the aircraft was nicknamed “the Gimli Glider.”



Figure 2.2 “The Gimli Glider.” The Gimli Glider is the Boeing 767 that ran out of fuel and glided to safety at Gimli Airport. The aircraft ran out of fuel because of confusion over the units used to express the amount of fuel.

The 767 took off from Montreal on its way to Ottawa, ultimately heading for Edmonton, Canada. About halfway through the flight, all the engines on the plane began to shut down because of a lack of fuel. When the final engine cut off, all electricity (which was generated by the engines) was lost; the plane became, essentially, a powerless glider. Captain Robert Pearson was an experienced glider pilot, although he had never flown a glider the size of a 767. First Officer Maurice Quintal quickly determined that the aircraft would not be able to make it to Winnipeg, the next large airport. He suggested his old Royal Air Force base at Gimli Station, one of whose runways was still being used as a community airport. Between the efforts of the pilots and the flight crew, they managed to get the airplane safely on the ground (although with buckled landing gear) and all passengers off safely.

What happened? At the time, Canada was transitioning from the older English system to the metric system. The Boeing 767s were the first aircraft whose gauges were calibrated in the metric system of units (litres and kilograms) rather than the English system of units (gallons and pounds). Thus, when the fuel gauge read 22,300, the gauge meant kilograms, but the ground crew mistakenly fuelled the plane with 22,300 *pounds* of fuel. This ended up being just less than half of the fuel needed to make the trip, causing the engines to quit about halfway to Ottawa. Quick thinking and extraordinary skill saved the lives of 61 passengers and 8 crew members—an incident that would not have occurred if people were watching their units.

View this [video on unit conversion](#) by Dr. Jessie A. Key for a lecture on unit conversion.

Key Takeaways

- Units can be converted to other units using the proper conversion factors.
- Conversion factors are constructed from equalities that relate two different units.

- Conversions can be a single step or multistep.
- Unit conversion is a powerful mathematical technique in chemistry that must be mastered.
- Exact numbers do not affect the determination of significant figures.

Exercises

Questions

1. Write the two conversion factors that exist between the two given units.
 - a. millilitres and litres
 - b. microseconds and seconds
 - c. kilometres and metres
2. Write the two conversion factors that exist between the two given units.
 - a. kilograms and grams
 - b. milliseconds and seconds
 - c. centimetres and metres
3. Perform the following conversions.
 - a. 5.4 km to metres
 - b. 0.665 m to millimetres
 - c. 0.665 m to kilometres
4. Perform the following conversions.
 - a. 90.6 mL to litres
 - b. 0.00066 ML to litres
 - c. 750 L to kilolitres
5. Perform the following conversions.
 - a. 17.8 μg to grams
 - b. 7.22×10^2 kg to grams
 - c. 0.00118 g to nanograms
6. Perform the following conversions.
 - a. 833 ns to seconds
 - b. 5.809 s to milliseconds
 - c. 2.77×10^6 s to megaseconds
7. Perform the following conversions.
 - a. 9.44 m^2 to square centimetres
 - b. $3.44 \times 10^8 \text{ mm}^3$ to cubic metres
8. Perform the following conversions.

- a. 0.00444 cm^3 to cubic metres
 - b. $8.11 \times 10^2 \text{ m}^2$ to square nanometres
9. Why would it be inappropriate to convert square centimetres to cubic metres?
10. Why would it be inappropriate to convert from cubic metres to cubic seconds?
11. Perform the following conversions.
- a. 45.0 m/min to metres/second
 - b. 0.000444 m/s to micrometres/second
 - c. 60.0 km/h to kilometres/second
12. Perform the following conversions.
- a. $3.4 \times 10^2 \text{ cm/s}$ to centimetres/minute
 - b. 26.6 mm/s to millimetres/hour
 - c. 13.7 kg/L to kilograms/millilitres
13. Perform the following conversions.
- a. 0.674 kL to millilitres
 - b. $2.81 \times 10^{12} \text{ mm}$ to kilometres
 - c. 94.5 kg to milligrams
14. Perform the following conversions.
- a. $6.79 \times 10^{-6} \text{ kg}$ to micrograms
 - b. 1.22 mL to kilolitres
 - c. $9.508 \times 10^{-9} \text{ ks}$ to milliseconds
15. Perform the following conversions.
- a. $6.77 \times 10^{14} \text{ ms}$ to kiloseconds
 - b. 34,550,000 cm to kilometres
16. Perform the following conversions.
- a. $4.701 \times 10^{15} \text{ mL}$ to kilolitres
 - b. $8.022 \times 10^{-11} \text{ ks}$ to microseconds
17. Perform the following conversions. Note that you will have to convert units in both the numerator and the denominator.
- a. 88 ft/s to miles/hour (Hint: use 5,280 ft = 1 mi.)
 - b. 0.00667 km/h to metres/second
18. Perform the following conversions. Note that you will have to convert units in both the numerator and the denominator.
- a. $3.88 \times 10^2 \text{ mm/s}$ to kilometres/hour
 - b. 1.004 kg/L to grams/millilitre
19. What is the area in square millimetres of a rectangle whose sides are $2.44 \text{ cm} \times 6.077 \text{ cm}$? Express the answer to the proper number of significant figures.
20. What is the volume in cubic centimetres of a cube with sides of 0.774 m? Express the answer to the proper number of significant figures.
21. The formula for the area of a triangle is $\frac{1}{2} \times \text{base} \times \text{height}$. What is the area of a triangle in square centimetres if its base is 1.007 m and its height is 0.665 m? Express the answer to the proper number of significant figures.
22. The formula for the area of a triangle is $\frac{1}{2} \times \text{base} \times \text{height}$. What is the area of a triangle in square metres if its base is

166 mm and its height is 930.0 mm? Express the answer to the proper number of significant figures.

Answers

1.
 - a. $\frac{1,000 \text{ mL}}{1 \text{ L}}$ and $\frac{1 \text{ L}}{1,000 \text{ mL}}$
 - b. $\frac{1,000,000 \mu\text{s}}{1 \text{ s}}$ and $\frac{1 \text{ s}}{1,000,000 \mu\text{s}}$
 - c. $\frac{1,000 \text{ m}}{1 \text{ km}}$ and $\frac{1 \text{ km}}{1,000 \text{ m}}$

3.
 - a. 5,400 m
 - b. 665 mm
 - c. $6.65 \times 10^{-4} \text{ km}$

5.
 - a. $1.78 \times 10^{-5} \text{ g}$
 - b. $7.22 \times 10^5 \text{ g}$
 - c. $1.18 \times 10^6 \text{ ng}$

7.
 - a. $94,400 \text{ cm}^2$
 - b. 0.344 m^3

9. One is a unit of area, and the other is a unit of volume.

11.
 - a. 0.75 m/s
 - b. $444 \mu\text{m/s}$
 - c. $1.666 \times 10^{-2} \text{ km/s}$

13.
 - a. 674,000 mL
 - b. $2.81 \times 10^6 \text{ km}$
 - c. $9.45 \times 10^7 \text{ mg}$

15.
 - a. $6.77 \times 10^8 \text{ ks}$
 - b. 345.5 km

17.
 - a. $6.0 \times 10^1 \text{ mi/h}$
 - b. 0.00185 m/s

19. $1.48 \times 10^3 \text{ mm}^2$
21. $3.35 \times 10^3 \text{ cm}^2$

Media Attributions

Example 2.8

- [“Grapevine snail”](#) by [Jürgen Schoner](#) © [CC BY-SA \(Attribution ShareAlike\)](#)

Figure 2.2

- [“The Gimli Glider”](#) by [Adrian Pingstone](#) © [Public Domain](#)

Other Units: Temperature and Density

Learning Objectives

1. Learn about the various temperature scales that are commonly used in chemistry.
2. Define density and use it as a conversion factor.

There are other units in chemistry that are important, and we will cover others in the course of the entire book. One of the fundamental quantities in science is temperature. **Temperature** is a measure of the average amount of energy of motion, or *kinetic energy*, a system contains. Temperatures are expressed using scales that use units called **degrees**, and there are several temperature scales in use. In the United States, the commonly used temperature scale is the *Fahrenheit scale* (symbolized by °F and spoken as “degrees Fahrenheit”). On this scale, the freezing point of liquid water (the temperature at which liquid water turns to solid ice) is 32°F, and the boiling point of water (the temperature at which liquid water turns to steam) is 212°F.

Science also uses other scales to express temperature. The Celsius scale (symbolized by °C and spoken as “degrees Celsius”) is a temperature scale where 0°C is the freezing point of water and 100°C is the boiling point of water; the scale is divided into 100 divisions between these two landmarks and extended higher and lower. By comparing the Fahrenheit and Celsius scales, a conversion between the two scales can be determined:

$$^{\circ}\text{C} = (^{\circ}\text{F} - 32) \times \frac{5}{9}$$

$$^{\circ}\text{F} = \left(^{\circ}\text{C} \times \frac{9}{5}\right) + 32$$

Using these formulas, we can convert from one temperature scale to another. The number 32 in the formulas is exact and does not count in significant figure determination.

Example 2.11

Problems

1. What is 98.6°F in degrees Celsius?
2. What is 25.0°C in degrees Fahrenheit?

Solutions

1. Using the first formula from above, we have: $^{\circ}\text{C} = (98.6 - 32) \times \frac{5}{9} = 66.6 \times \frac{5}{9} = 37.0^{\circ}\text{C}$

2. Using the second formula from above, we have: $^{\circ}\text{F} = \left(25.0 \times \frac{9}{5}\right) + 32 = 45.0 + 32 = 77.0^{\circ}\text{F}$

Test Yourself

1. Convert 0°F to degrees Celsius.
2. Convert 212°C to degrees Fahrenheit.

Answers

1. -17.8°C
2. 414°F

The fundamental unit of temperature (another fundamental unit of science, bringing us to four) in SI is the **kelvin** (K). The Kelvin temperature scale (note that the name of the scale capitalizes the word *Kelvin*, but the unit itself is lowercase) uses degrees that are the same size as the Celsius degree, but the numerical scale is shifted up by 273.15 units. That is, the conversion between the Kelvin and Celsius scales is as follows:

$$\begin{aligned} \text{K} &= ^{\circ}\text{C} + 273.15 \\ ^{\circ}\text{C} &= \text{K} - 273.15 \end{aligned}$$

For most purposes, it is acceptable to use 273 instead of 273.15. Note that the Kelvin scale does not use the word *degrees*; a temperature of 295 K is spoken of as “two hundred ninety-five kelvins” and not “two hundred ninety-five degrees Kelvin.”

The reason that the Kelvin scale is defined this way is because there exists a minimum possible temperature called **absolute zero**. The Kelvin temperature scale is set so that 0 K is absolute zero, and temperature is counted upward from there. Normal room temperature is about 295 K, as seen in the following example.

Example 2.12

Problem

If normal room temperature is 72.0°F , what is room temperature in degrees Celsius and kelvins?

Solution

First, we use the formula to determine the temperature in degrees Celsius:

$$^{\circ}\text{C} = (72.0 - 32) \times \frac{5}{9} = 40.0 \times \frac{5}{9} = 22.2^{\circ}\text{C}$$

Then we use the appropriate formula above to determine the temperature in the Kelvin scale:

$$\text{K} = 22.2^{\circ}\text{C} + 273.15 = 295.4 \text{ K}$$

So, room temperature is about 295 K.

Test Yourself

What is 98.6°F on the Kelvin scale?

Answer

310.2 K

Figure 2.3 “Fahrenheit, Celsius, and Kelvin Temperatures” compares the three temperature scales. Note that science uses the Celsius and Kelvin scales almost exclusively; virtually no practicing chemist expresses laboratory-measured temperatures with the Fahrenheit scale. In fact, the United States is one of the few countries in the world that still uses the Fahrenheit scale on a daily basis. The other two countries are Liberia and Myanmar (formerly Burma).

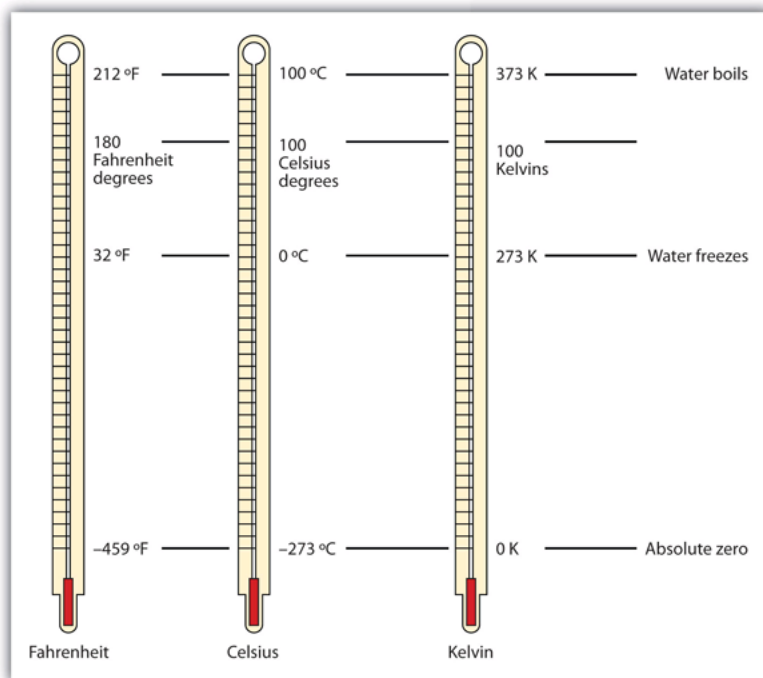


Figure 2.3 “Fahrenheit, Celsius, and Kelvin Temperatures.” A comparison of the three temperature scales. [\[Long Description\]](#)

Density is a physical property that is defined as a substance’s mass divided by its volume:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \text{ or } d = \frac{m}{V}$$

Density is usually a measured property of a substance, so its numerical value affects the significant figures in a calculation. Notice that density is defined in terms of two dissimilar units, mass and volume. That means that density overall has derived units, just like velocity. Common units for density include g/mL, g/cm³, g/L, kg/L, and even kg/m³. Densities for some common substances are listed in Table 2.3 “Densities of Some Common Substances”.

Table 2.3 Densities of Some Common Substances

Substance	Density (g/mL or g/cm ³)
water	1.0
gold	19.3
mercury	13.6
air	0.0012
cork	0.22–0.26
aluminum	2.7
iron	7.87

Because of how it is defined, density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm³. How can you determine what mass of aluminum you have without measuring it? You can use the volume to calculate it. If you multiply the given volume by the known density (from Table 2.3 “Densities of Some Common Substances”), the volume units will cancel and leave you with mass units, telling you the mass of the sample is

$$7.88 \text{ cm}^3 \times 2.7 \text{ g/cm}^3 = 21 \text{ g of aluminum}$$

where we have limited our answer to two significant figures.

Example 2.13

Problem

What is the mass of 44.6 mL of mercury?

Solution

Use the density from Table 2.3 “Densities of Some Common Substances” as a conversion factor to go from volume to mass:

$$44.6 \text{ mL} \times 13.6 \text{ g/mL} = 607 \text{ g}$$

The mass of the mercury is 607 g.

Test Yourself

What is the mass of 25.0 cm³ of iron?

Answer

197 g

Density can also be used as a conversion factor to convert mass to volume—but care must be taken. We have already demonstrated that the number that goes with density normally goes in the numerator when density is written as a fraction. Take the density of gold, for example:

$$d = 19.3 \frac{\text{g}}{\text{mL}} = \frac{19.3 \text{ g}}{1 \text{ mL}}$$

Although this was not previously pointed out, it can be assumed that there is a 1 in the denominator.

That is, the density value tells us that we have 19.3 grams for every 1 millilitre of volume, and the 1 is an exact number. When we want to use density to convert from mass to volume, the numerator and denominator of density need to be switched—that is, we must take the *reciprocal* of the density. In so doing, we move not only the units but also the numbers:

$$\frac{1}{d} = \frac{1 \text{ mL}}{19.3 \text{ g}}$$

This reciprocal density is still a useful conversion factor, but now the mass unit will cancel and the volume unit will be introduced. Thus, if we want to know the volume of 45.9 g of gold, we would set up the conversion as follows:

$$45.9 \text{ g} \times \frac{1 \text{ mL}}{19.3 \text{ g}} = 2.38 \text{ mL}$$

Note how the mass units cancel, leaving the volume unit, which is what we're looking for.

Example 2.14

Problem

A cork stopper from a bottle of wine has a mass of 3.78 g. If the density of cork is 0.22 g/cm³, what is the volume of the cork?

Solution

To use density as a conversion factor, we need to take the reciprocal so that the mass unit of density is in the denominator. Taking the reciprocal, we find:

$$\frac{1}{d} = \frac{1 \text{ cm}^3}{0.22 \text{ g}}$$

We can use this expression as the conversion factor. So:

$$3.78 \text{ g} \times \frac{1 \text{ cm}^3}{0.22 \text{ g}} = 17.2 \text{ cm}^3$$

Test Yourself

What is the volume of 3.78 g of gold?

Answer

0.196 cm³

Care must be used with density as a conversion factor. Make sure the mass units are the same, or the volume units are the same, before using density to convert to a different unit. Often, the unit of the given quantity must be first converted to the appropriate unit before applying density as a conversion factor.

Food and Drink App: Cooking Temperatures

Because degrees Fahrenheit is the common temperature scale in the United States, kitchen appliances, such as ovens, are calibrated in that scale. A cool oven may be only 150°F, while a cake may be baked at 350°F and a chicken roasted at 400°F. The broil setting on many ovens is 500°F, which is typically the highest temperature setting on a household oven.

People who live at high altitudes, typically 2,000 ft above sea level or higher, are sometimes urged to use slightly different cooking instructions on some products, such as cakes and bread, because water boils at a lower temperature the higher in altitude you go, meaning that foods cook slower. For example, in Cleveland water typically boils at 212°F (100°C), but in Denver, the Mile-High City, water boils at about 200°F (93.3°C), which can significantly lengthen cooking times. Good cooks need to be aware of this.

At the other end is pressure cooking. A pressure cooker is a closed vessel that allows steam to build up additional pressure, which increases the temperature at which water boils. A good pressure cooker can get to temperatures as high as 252°F (122°C); at these temperatures, food cooks much faster than it normally would. Great care must be used with pressure cookers because of the high pressure and high temperature. (When a pressure cooker is used to sterilize medical instruments, it is called an *autoclave*.)

Other countries use the Celsius scale for everyday purposes. Therefore, oven dials in their kitchens are marked in degrees Celsius. It can be confusing for US cooks to use ovens abroad—a 425°F oven in the United States is equivalent to a 220°C oven in other countries. These days, many oven thermometers are marked with both temperature scales.

Key Takeaways

- Chemistry uses the Celsius and Kelvin scales to express temperatures.
- A temperature on the Kelvin scale is the Celsius temperature plus 273.15.
- The minimum possible temperature is absolute zero and is assigned 0 K on the Kelvin scale.
- Density relates a substance's mass and volume.
- Density can be used to calculate volume from a given mass or mass from a given volume.

Exercises

Questions

1. Perform the following conversions.
 - a. 255°F to degrees Celsius
 - b. -255°F to degrees Celsius
 - c. 50.0°C to degrees Fahrenheit
 - d. -50.0°C to degrees Fahrenheit
2. Perform the following conversions.

- a. $1,065^{\circ}\text{C}$ to degrees Fahrenheit
 - b. -222°C to degrees Fahrenheit
 - c. 400.0°F to degrees Celsius
 - d. 200.0°F to degrees Celsius
3. Perform the following conversions.
- a. 100.0°C to kelvins
 - b. -100.0°C to kelvins
 - c. 100 K to degrees Celsius
 - d. 300 K to degrees Celsius
4. Perform the following conversions.
- a. $1,000.0\text{ K}$ to degrees Celsius
 - b. 50.0 K to degrees Celsius
 - c. 37.0°C to kelvins
 - d. -37.0°C to kelvins
5. Convert 0 K to degrees Celsius. What is the significance of the temperature in degrees Celsius?
6. Convert 0 K to degrees Fahrenheit. What is the significance of the temperature in degrees Fahrenheit?
7. The hottest temperature ever recorded on the surface of the earth was 136°F in Libya in 1922. What is the temperature in degrees Celsius and in kelvins?
8. The coldest temperature ever recorded on the surface of the earth was -128.6°F in Vostok, Antarctica, in 1983. What is the temperature in degrees Celsius and in kelvins?
9. Give at least three possible units for density.
10. What are the units when density is inverted? Give three examples.
11. A sample of iron has a volume of 48.2 cm^3 . What is its mass?
12. A sample of air has a volume of 1,015 mL. What is its mass?
13. The volume of hydrogen used by the Hindenburg, the German airship that exploded in New Jersey in 1937, was $2.000 \times 10^9\text{ L}$. If hydrogen gas has a density of 0.0899 g/L , what mass of hydrogen was used by the airship?
14. The volume of an Olympic-sized swimming pool is $2.50 \times 10^9\text{ cm}^3$. If the pool is filled with alcohol ($d = 0.789\text{ g/cm}^3$), what mass of alcohol is in the pool?
15. A typical engagement ring has 0.77 cm^3 of gold. What mass of gold is present?
16. A typical mercury thermometer has 0.039 mL of mercury in it. What mass of mercury is in the thermometer?
17. What is the volume of 100.0 g of lead if lead has a density of 11.34 g/cm^3 ?
18. What is the volume of 255.0 g of uranium if uranium has a density of 19.05 g/cm^3 ?
19. What is the volume in litres of 222 g of neon if neon has a density of 0.900 g/L ?
20. What is the volume in litres of 20.5 g of sulfur hexafluoride if sulfur hexafluoride has a density of 6.164 g/L ?
21. Which has the greater volume, 100.0 g of iron ($d = 7.87\text{ g/cm}^3$) or 75.0 g of gold ($d = 19.3\text{ g/cm}^3$)?
22. Which has the greater volume, 100.0 g of hydrogen gas ($d = 0.0000899\text{ g/cm}^3$) or 25.0 g of argon gas ($d = 0.00178\text{ g/cm}^3$)?

Answers

1.
 - a. 124°C
 - b. -159°C
 - c. 122°F

- d. -58°F
3. a. 373 K
b. 173 K
c. -173°C
d. 27°C
5. -273°C . This is the lowest possible temperature in degrees Celsius.
7. 57.8°C ; 331 K
9. g/mL, g/L, and kg/L (answers will vary)
11. 379 g
13. 1.80×10^7 g
15. 15 g
17. 8.818 cm^3
19. 247 L
21. The 100.0 g of iron has the greater volume.

Long Descriptions

Figure 2.3 long description: Fahrenheit, Celsius, and Kelvin temperature scales. Absolute zero in each scale is -459°F , -273°C , and 0 K. Water freezes at 32°F , 0°C , and 273 K. 180°F equals 100°C , which equals 100 kelvins. Water boils at 212°F , 100°C , and 373 K. [\[Return to Figure 2.3\]](#)

Media Attributions

Figure 2.3

- “Fahrenheit, Celsius, and Kelvin Temperatures” by David W. Ball © [CC BY-NC-SA \(Attribution NonCommercial ShareAlike\)](#)

Expressing Units

Learning Objectives

1. Learn the units that go with various quantities.
2. Express units using their abbreviations.
3. Make new units by combining numerical prefixes with units.

A number indicates “how much,” but the unit indicates “of what.” The “of what” is important when communicating a quantity. For example, if you were to ask a friend how close you are to Lake Erie and your friend says “six,” then your friend isn’t giving you complete information. Six *what*? Six miles? Six inches? Six city blocks? The actual distance to the lake depends on what units you use.

Chemistry, like most sciences, uses the International System of Units, or SI for short. (The letters *SI* stand for the French “le Système International d’unités.”) SI specifies certain units for various types of quantities, based on seven **fundamental units** for various quantities. We will use most of the fundamental units in chemistry. Initially, we will deal with three fundamental units. The metre (m) is the SI unit of length. It is a little longer than a yard (see Figure 2.4 “The Metre”). The SI unit of mass is the kilogram (kg), which is about 2.2 pounds (lb). The SI unit of time is the second (s).



Figure 2.4 “The Metre.” The SI standard unit of length, the metre, is a little longer than a yard.

To express a quantity, you need to combine a number with a unit. If you have a length that is 2.4 m,

then you express that length as simply 2.4 m. A time of 15,000 s can be expressed as 1.5×10^4 s in scientific notation.

Sometimes, a given unit is not an appropriate size to easily express a quantity. For example, the width of a human hair is very small, and it doesn't make much sense to express it in metres. SI also defines a series of **numerical prefixes** that refer to multiples or fractions of a fundamental unit to make a unit more conveniently sized for a specific quantity. Table 2.4 "Multiplicative Prefixes for SI Units" lists the prefixes, their abbreviations, and their multiplicative factors. Some of the prefixes, such as kilo-, mega-, and giga-, represent more than one of the fundamental unit, while other prefixes, such as centi-, milli-, and micro-, represent fractions of the original unit. Note, too, that once again we are using powers of 10. Each prefix is a multiple of or fraction of a power of 10.

Table 2.4 Multiplicative Prefixes for SI Units

Prefix	Abbreviation	Multiplicative Amount
giga-	G	$1,000,000,000 \times$
mega-	M	$1,000,000 \times$
kilo-	k	$1,000 \times$
deci-	d	$\frac{1}{10} \times$
centi-	c	$\frac{1}{100} \times$
milli-	m	$\frac{1}{1000} \times$
micro-	μ^1	$\frac{1}{1,000,000} \times$
nano-	n	$\frac{1}{1,000,000,000} \times$
pico-	p	$\frac{1}{1,000,000,000,000} \times$

To use the fractions to generate new units, simply combine the prefix with the unit itself; the abbreviation for the new unit is the combination of the abbreviation for the prefix and the abbreviation of the unit. For example, the kilometre (km) is $1,000 \times$ metre, or 1,000 m. Thus, 5 kilometres (5 km) is equal to 5,000 m. Similarly, a millisecond (ms) is $\frac{1}{1,000} \times$ second, or one one-thousandth of a second.

Thus, 25 ms is 25 thousandths of a second. You will need to become proficient in combining prefixes and units. (You may recognize that one of our fundamental units, the kilogram, automatically has a prefix-unit combination, the kilogram. The word *kilogram* means 1,000 g.)

1. The letter μ is the Greek letter lowercase equivalent to an m and is called "mu" (pronounced "myoo").

In addition to the fundamental units, SI also allows for **derived units** based on a fundamental unit or units. There are many derived units used in science. For example, the derived unit for area comes from the idea that area is defined as width times height. Because both width and height are lengths, they both have the fundamental unit of metre, so the unit of area is metre \times metre, or metre² (m²). This is sometimes spoken as “square metres.” A unit with a prefix can also be used to derive a unit for area, so we can also have cm², mm², or km² as acceptable units for area.

Volume is defined as length times width times height, so it has units of metre \times metre \times metre or metre³ (m³), sometimes spoken as “cubic metres.” The cubic metre is a rather large unit, however, so another unit is defined that is somewhat more manageable: the litre (L). A litre is one one-thousandth of a cubic metre and is a little more than 1 quart in volume (see Figure 2.5 “The Litre”). Prefixes can also be used with the litre unit, so we can speak of millilitres (one one-thousandth of a litre; mL) and kilolitres (1,000 L; kL).

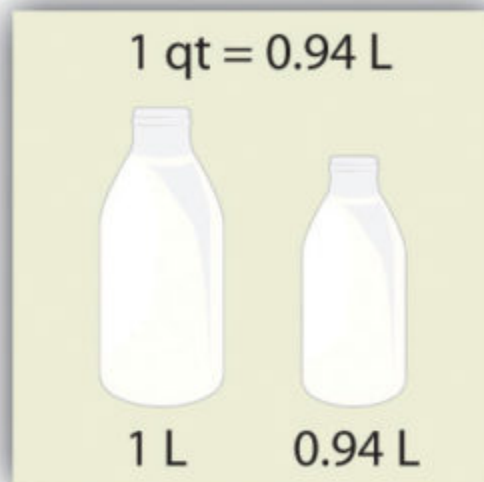


Figure 2.5 “The Litre.” The SI unit of volume, the litre, is slightly larger than 1 quart.

Another definition of a litre is one-tenth of a metre cubed. Because one-tenth of a metre is 10 cm, then a litre is equal to 1,000 cm³ (Figure 2.6 “The Size of 1 Litre”). Because 1 L equals 1,000 mL, we conclude that 1 mL equals 1 cm³; thus, these units are interchangeable.

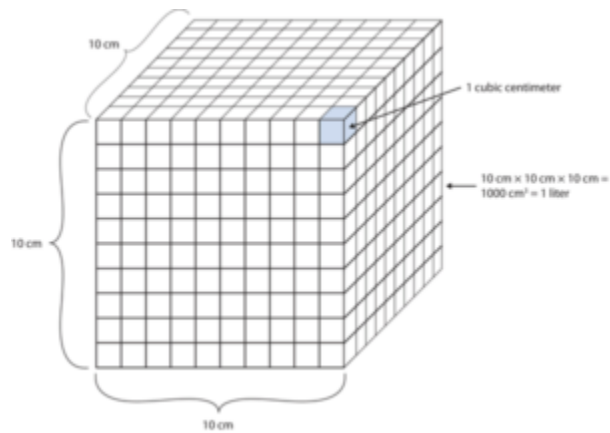


Figure 2.6 “The Size of 1 Litre.” One litre equals $1,000 \text{ cm}^3$, so 1 cm^3 is the same as 1 mL.

Units not only are multiplied together but also can be divided. For example, if you are travelling at one metre for every second of time elapsed, your velocity is 1 metre per second, or 1 m/s. The word *per* implies division, so velocity is determined by dividing a distance quantity by a time quantity. Other units for velocity include kilometres per hour (km/h) or even micrometres per nanosecond ($\mu\text{m}/\text{ns}$). Later, we will see other derived units that can be expressed as fractions.

Example 2.15

Problems

1. A human hair has a diameter of about 6.0×10^{-5} m. Suggest an appropriate unit for this measurement and write the diameter of a human hair in terms of that unit.
2. What is the velocity of a car if it goes 25 m in 5.0 s?

Solutions

1. The scientific notation 10^{-5} is close to 10^{-6} , which defines the micro- prefix. Let us use micrometres as the unit for hair diameter. The number 6.0×10^{-5} can be written as 60×10^{-6} , and a micrometre is 10^{-6} m, so the diameter of a human hair is about 60 μm .
2. If velocity is defined as a distance quantity divided by a time quantity, then velocity is 25 metres/5.0 seconds. Dividing the numbers gives us $25 \div 5.0 = 5.0$, and dividing the units gives us metres/second, or m/s. The velocity is 5.0 m/s.

Test Yourself

1. Express the volume of an Olympic-sized swimming pool, 2,500,000 L, in more appropriate units.
2. A common garden snail moves about 6.1 m in 30 min. What is its velocity in metres per minute (m/min)?

Answers

1. 2.5 ML
2. 0.203 m/min

Key Takeaways

- Numbers tell “how much,” and units tell “of what.”
- Chemistry uses a set of fundamental units and derived units from SI units.
- Chemistry uses a set of prefixes that represent multiples or fractions of units.
- Units can be multiplied and divided to generate new units for quantities.

Exercises

Questions

1. Identify the unit in each quantity.
 - a. 2 boxes of crayons
 - b. 3.5 grams of gold
2. Identify the unit in each quantity.
 - a. 32 oz of cheddar cheese
 - b. 0.045 cm³ of water
3. Identify the unit in each quantity.
 - a. 9.58 s (the current world record in the 100 m dash)
 - b. 6.14 m (the current world record in the pole vault)
4. Identify the unit in each quantity.
 - a. 2 dozen eggs
 - b. 2.4 km/s (the escape velocity of the moon, which is the velocity you need at the surface to escape the moon's gravity)
5. Indicate what multiplier each prefix represents.
 - a. k
 - b. m
 - c. M
6. Indicate what multiplier each prefix represents.
 - a. c
 - b. G
 - c. μ
7. Give the prefix that represents each multiplier.
 - a. $\frac{1}{1,000} \times$
 - b. $1,000 \times$

c. $1,000,000,000 \times$

8. Give the prefix that represents each multiplier.

a. $\frac{1}{1,000,000,000} \times$

b. $\frac{1}{100} \times$

c. $1,000,000 \times$

9. Complete the following table with the missing information.

Unit	Abbreviation
kilosecond	
	mL
	Mg
centimetre	

10. Complete the following table with the missing information.

Unit	Abbreviation
kilometre per second	
second	
	cm^3
	μL
nanosecond	

11. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. $3.44 \times 10^{-6} \text{ s}$

b. 3,500 L

c. 0.045 m

12. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. 0.000066 m/s (Hint: you need consider only the unit in the numerator.)

b. $4.66 \times 10^6 \text{ s}$

c. 7,654 L

13. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. 43,600 mL

b. 0.0000044 m

c. 1,438 ms

14. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. 0.000000345 m^3

b. 47,000,000 mm^3

c. 0.00665 L

15. Multiplicative prefixes are used for other units as well, such as computer memory. The basic unit of computer memory is the byte (b). What is the unit for one million bytes?
16. You may have heard the terms microscale or nanoscale to represent the sizes of small objects. What units of length do you think are useful at these scales? What fractions of the fundamental unit of length are these units?
17. Acceleration is defined as a change in velocity per time. Propose a unit for acceleration in terms of the fundamental SI units.
18. Density is defined as the mass of an object divided by its volume. Propose a unit of density in terms of the fundamental SI units.

Answers

1. a. boxes of crayons
b. grams of gold
3. a. seconds
b. metres
5. a. $1,000 \times$
b. $\frac{1}{1,000} \times$
c. $1,000,000 \times$
7. a. milli-
b. kilo-
c. giga-

9.

Unit	Abbreviation
kilosecond	ks
millilitre	mL
megagram	Mg
centimetre	cm

11. a. $3.44 \mu\text{s}$
b. 3.5 kL
c. 4.5 cm
13. a. 43.6 L
b. $4.4 \mu\text{m}$
c. 1.438 s
15. megabytes (Mb)
17. metres/second²

End-of-Chapter Material

Additional Exercises

1. Evaluate $0.00000000552 \times 0.000000006188$ and express the answer in scientific notation. You may have to rewrite the original numbers in scientific notation first.
2. Evaluate $333,999,500,000 \div 0.0000000003396$ and express the answer in scientific notation. You may need to rewrite the original numbers in scientific notation first.
3. Express the number 6.022×10^{23} in standard notation.
4. Express the number 6.626×10^{-34} in standard notation.
5. When powers of 10 are multiplied together, the powers are added together. For example, $10^2 \times 10^3 = 10^{2+3} = 10^5$. With this in mind, can you evaluate $(4.506 \times 10^4) \times (1.003 \times 10^2)$ without entering scientific notation into your calculator?
6. When powers of 10 are divided into each other, the bottom exponent is subtracted from the top exponent. For example, $10^5 \div 10^3 = 10^{5-3} = 10^2$. With this in mind, can you evaluate $(8.552 \times 10^6) \div (3.129 \times 10^3)$ without entering scientific notation into your calculator?
7. Consider the quantity two dozen eggs. Is the number in this quantity “two” or “two dozen”? Justify your choice.
8. Consider the quantity two dozen eggs. Is the unit in this quantity “eggs” or “dozen eggs”? Justify your choice.
9. Fill in the blank: $1 \text{ km} = \underline{\hspace{2cm}} \mu\text{m}$.
10. Fill in the blank: $1 \text{ Ms} = \underline{\hspace{2cm}} \text{ ns}$.
11. Fill in the blank: $1 \text{ cL} = \underline{\hspace{2cm}} \text{ ML}$.
12. Fill in the blank: $1 \text{ mg} = \underline{\hspace{2cm}} \text{ kg}$.
13. Express 67.3 km/h in metres/second.
14. Express 0.00444 m/s in kilometres/hour.
15. Using the idea that $1.602 \text{ km} = 1.000 \text{ mi}$, convert a speed of 60.0 mi/h into kilometres/hour.
16. Using the idea that $1.602 \text{ km} = 1.000 \text{ mi}$, convert a speed of 60.0 km/h into miles/hour.
17. Convert 52.09 km/h into metres/second.
18. Convert 2.155 m/s into kilometres/hour.
19. Use the formulas for converting degrees Fahrenheit into degrees Celsius to determine the relative size of the Fahrenheit degree over the Celsius degree.
20. Use the formulas for converting degrees Celsius into kelvins to determine the relative size of the Celsius degree over kelvins.
21. What is the mass of 12.67 L of mercury?
22. What is the mass of 0.663 m^3 of air?
23. What is the volume of 2.884 kg of gold?
24. What is the volume of 40.99 kg of cork? Assume a density of 0.22 g/cm^3 .

Answers

1. 3.42×10^{-18}
3. 602,200,000,000,000,000,000,000
5. 4.520×10^6
7. The quantity is two; dozen is the unit.
9. 1,000,000,000
11. $\frac{1}{100,000,000}$
13. 18.7 m/s
15. 96.1 km/h
17. 14.47 m/s
19. One Fahrenheit degree is nine-fifths the size of a Celsius degree.
21. 1.72×10^5 g
23. 149 mL

Chapter 3. Atoms, Molecules, and Ions

Although not an SI unit, the angstrom (\AA) is a useful unit of length. It is one ten-billionth of a metre, or 10^{-10} m. Why is it a useful unit? The ultimate particles that compose all matter are about 10^{-10} m in size, or about 1 \AA . This makes the angstrom a natural—though not approved—unit for describing these particles.

The angstrom unit is named after Anders Jonas Ångström, a nineteenth-century Swedish physicist. Ångström's research dealt with light being emitted by glowing objects, including the sun. Ångström studied the brightness of the different colors of light that the sun emitted and was able to deduce that the sun is composed of the same kinds of matter that are present on the earth. By extension, we now know that all matter throughout the universe is similar to the matter that exists on our own planet.

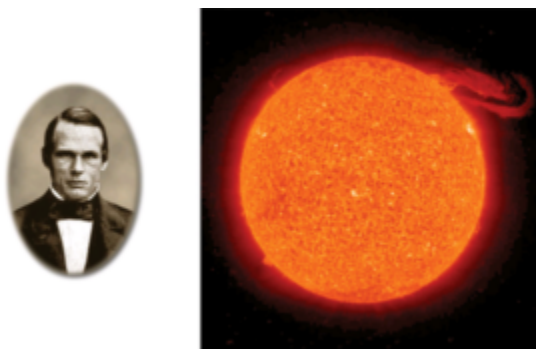


Figure 3.0 “The Sun.” Anders Jonas Ångström, a Swedish physicist, studied the light coming from the sun. His contributions to science were sufficient to have a tiny unit of length named after him, the angstrom, which is one ten-billionth of a metre.

The basic building block of all matter is the atom. Curiously, the idea of atoms was first proposed in the fifth century BCE, when the Greek philosophers Leucippus and Democritus proposed their existence in a surprisingly modern fashion. However, their ideas never took hold among their contemporaries, and it wasn't until the early 1800s that evidence amassed to make scientists reconsider the idea. Today, the concept of the atom is central to the study of matter.

Media Attributions

Figure 3.0

- [“Anders Jonas Ångström”](#) © [Public Domain](#)
- [“Solar prominence from STEREO spacecraft September 29, 2008”](#) by NASA © [Public Domain](#)

Acids

Learning Objectives

1. Define *acid*.
2. Name a simple acid.

There is one other group of compounds that is important to us—acids—and these compounds have interesting chemical properties. Initially, we will define an **acid** as an ionic compound of the H^+ cation dissolved in water. (We will expand on this definition in [Chapter 12 “Acids and Bases”](#).) To indicate that something is dissolved in water, we will use the phase label (aq) next to a chemical formula (where aq stands for “aqueous,” a word that describes something dissolved in water). If the formula does not have this label, then the compound is treated as a molecular compound rather than an acid.

Acids have their own nomenclature system. If an acid is composed of only hydrogen and one other element, the name is *hydro-* + the stem of the other element + *-ic acid*. For example, the compound HCl(aq) is hydrochloric acid, while $\text{H}_2\text{S(aq)}$ is hydrosulfuric acid. (If these acids were not dissolved in water, the compounds would be called hydrogen chloride and hydrogen sulfide, respectively. Both of these substances are well known as molecular compounds; when dissolved in water, however, they are treated as acids.)

If a compound is composed of hydrogen ions and a polyatomic anion, then the name of the acid is derived from the stem of the polyatomic ion’s name. Typically, if the anion name ends in *-ate*, the name of the acid is the stem of the anion name plus *-ic acid*; if the related anion’s name ends in *-ite*, the name of the corresponding acid is the stem of the anion name plus *-ous acid*. Table 3.1 “Names and Formulas of Acids” lists the formulas and names of a variety of acids that you should be familiar with. You should recognize most of the anions in the formulas of the acids.

Table 3.1 Names and Formulas of Acids¹

Formula	Name
HC ₂ H ₃ O ₂	acetic acid
HClO ₃	chloric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydriodic acid
HF	hydrofluoric acid
HNO ₃	nitric acid
H ₂ C ₂ O ₄	oxalic acid
HClO ₄	perchloric acid
H ₃ PO ₄	phosphoric acid
H ₂ SO ₄	sulfuric acid
H ₂ SO ₃	sulfurous acid

Example 3.1

Problems

Name each acid without consulting Table 3.1 “Names and Formulas of Acids”.

1. HBr
2. H₂SO₄

Solutions

1. As a binary acid, the acid’s name is *hydro-* + stem name + *-ic* acid. Because this acid contains a bromine atom, the name is hydrobromic acid.
2. Because this acid is derived from the sulfate ion, the name of the acid is the stem of the anion name + *-ic* acid. The name of this acid is sulfuric acid.

Test Yourself

Name each acid.

1. HF
2. HNO₂

1. Note: The “aq” label is omitted throughout for clarity.

Answers

1. hydrofluoric acid
2. nitrous acid

All acids have some similar properties. For example, acids have a sour taste; in fact, the sour taste of some of our foods, such as citrus fruits and vinegar, is caused by the presence of acids in food. Many acids react with some metallic elements to form metal ions and elemental hydrogen. Acids make certain plant pigments change colours; indeed, the ripening of some fruits and vegetables is caused by the formation or destruction of excess acid in the plant. In [Chapter 12 “Acids and Bases”](#), we will explore the chemical behaviour of acids.

Acids are very prevalent in the world around us. We have already mentioned that citrus fruits contain acid; among other compounds, they contain citric acid, $\text{H}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq})$. Oxalic acid, $\text{H}_2\text{C}_2\text{O}_4(\text{aq})$, is found in spinach and other green leafy vegetables. Hydrochloric acid not only is found in the stomach (stomach acid) but also can be bought in hardware stores as a cleaner for concrete and masonry. Phosphoric acid is an ingredient in some soft drinks.

Key Takeaways

- An acid is a compound of the H^+ ion dissolved in water.
- Acids have their own naming system.
- Acids have certain chemical properties that distinguish them from other compounds.

Exercises**Questions**

1. Give the formula for each acid.
 - a. perchloric acid
 - b. hydriodic acid
2. Give the formula for each acid.
 - a. hydrosulfuric acid
 - b. phosphorous acid
3. Name each acid.
 - a. $\text{HF}(\text{aq})$
 - b. $\text{HNO}_3(\text{aq})$

c. $\text{H}_2\text{C}_2\text{O}_4(\text{aq})$

4. Name each acid.

a. $\text{H}_2\text{SO}_4(\text{aq})$

b. $\text{H}_3\text{PO}_4(\text{aq})$

c. $\text{HCl}(\text{aq})$

5. Name an acid found in food.

6. Name some properties that acids have in common.

Answers

1. a. $\text{HClO}_4(\text{aq})$

b. $\text{HI}(\text{aq})$

3. a. hydrofluoric acid

b. nitric acid

c. oxalic acid

5. oxalic acid (answers will vary)

Ions and Ionic Compounds

Learning Objectives

1. Know how ions form.
2. Learn the characteristic charges that ions have.
3. Construct a proper formula for an ionic compound.
4. Generate a proper name for an ionic compound.

So far, we have discussed elements and compounds that are electrically neutral. They have the same number of electrons as protons, so the negative charges of the electrons is balanced by the positive charges of the protons. However, this is not always the case. Electrons can move from one atom to another; when they do, species with overall electric charges are formed. Such species are called **ions**. Species with overall positive charges are termed **cations**, while species with overall negative charges are called **anions**. Remember that ions are formed only when *electrons* move from one atom to another; a proton never moves from one atom to another. Compounds formed from positive and negative ions are called **ionic compounds**.

Individual atoms can gain or lose electrons. When they do, they become *monatomic* ions. When atoms gain or lose electrons, they usually gain or lose a characteristic number of electrons and so take on a characteristic overall charge. Table 3.2 “Monatomic Ions of Various Charges” lists some common ions in terms of how many electrons they lose (making cations) or gain (making anions). There are several things to notice about the ions in Table 3.2. First, each element that forms cations is a metal, except for one (hydrogen), while each element that forms anions is a nonmetal. This is actually one of the chemical properties of metals and nonmetals: metals tend to form cations, while nonmetals tend to form anions. Second, most atoms form ions of a single characteristic charge. When sodium atoms form ions, they always form a 1+ charge, never a 2+ or 3+ or even 1- charge. Thus, if you commit the information in Table 3.2 to memory, you will always know what charges most atoms form. (In [Chapter 9 “Chemical Bonds”](#), we will discuss *why* atoms form the charges they do.)

Table 3.2 Monatomic Ions of Various Charges

Ions formed by losing a single electron	H ⁺ Na ⁺ K ⁺ Rb ⁺ Ag ⁺ Au ⁺
Ions formed by losing two electrons	Mg ²⁺ Ca ²⁺ Sr ²⁺ Fe ²⁺ Co ²⁺ Ni ²⁺ Cu ²⁺ Zn ²⁺ Sn ²⁺ Hg ²⁺ Pb ²⁺
Ions formed by losing three electrons	Sc ³⁺ Fe ³⁺ Co ³⁺ Ni ³⁺ Au ³⁺ Al ³⁺ Cr ³⁺
Ions formed by losing four electrons	Ti ⁴⁺ Sn ⁴⁺ Pb ⁴⁺

Ions formed by gaining a single electron	F^{-} Cl^{-} Br^{-} I^{-}
Ions formed by gaining two electrons	O^{2-} S^{2-} Se^{2-}
Ions formed by gaining three electrons	N^{3-} P^{3-}

Third, there are some exceptions to the previous point. A few elements, all metals, can form more than one possible charge. For example, iron atoms can form 2+ cations or 3+ cations. Cobalt is another element that can form more than one possible charged ion (2+ and 3+), while lead can form 2+ or 4+ cations. Unfortunately, there is little understanding which two charges a metal atom may take, so it is best to just memorize the possible charges a particular element can have.

Note the convention for indicating an ion. The magnitude of the charge is listed as a right superscript next to the symbol of the element. If the charge is a single positive or negative one, the number 1 is not written; if the magnitude of the charge is greater than 1, then the number is written *before* the + or - sign. An element symbol without a charge written next to it is assumed to be the uncharged atom.

Naming an ion is straightforward. For a cation, simply use the name of the element and add the word *ion* (or if you want to be more specific, add *cation*) after the element's name. So Na^{+} is the sodium ion; Ca^{2+} is the calcium ion. If the element has more than one possible charge, the value of the charge comes after the element name and before the word *ion*. Thus, Fe^{2+} is the iron two ion, while Fe^{3+} is the iron three ion. In print, we use roman numerals in parentheses to represent the charge on the ion, so these two iron ions would be represented as the iron(II) cation and the iron(III) cation, respectively.

For a monatomic anion, use the stem of the element name and append the suffix *-ide* to it, and then add *ion*. This is similar to how we named molecular compounds. Thus, Cl^{-} is the chloride ion, and N^{3-} is the nitride ion.

Example 3.2

Problems

Name each species.

1. O^{2-}
2. Co
3. Co^{2+}

Solutions

1. This species has a 2⁻ charge on it, so it is an anion. Anions are named using the stem of the element name with the suffix *-ide* added. This is the oxide anion.
2. Because this species has no charge, it is an atom in its elemental form. This is cobalt.
3. In this case, there is a 2⁺ charge on the atom, so it is a cation. We note from [Table 3.2 “Monatomic Ions of Various Charges”](#) that cobalt cations can have two possible charges, so the name of the ion must specify which charge the ion has. This is the cobalt(II) cation.

Test Yourself

Name each species.

1. P^{3-}
2. Sr^{2+}

Answers

1. the phosphide anion
2. the strontium cation

Chemical formulas for ionic compounds are called ionic formulas. A proper ionic formula has a cation and an anion in it; an ionic compound is never formed between two cations only or two anions only. The key to writing proper ionic formulas is simple: the total positive charge must balance the total negative charge. Because the charges on the ions are characteristic, sometimes we have to have more than one of a cation or an anion to balance the overall positive and negative charges. It is conventional to use the lowest ratio of ions that are needed to balance the charges.

For example, consider the ionic compound between Na^+ and Cl^- . Each ion has a single charge, one positive and one negative, so we need only one ion of each to balance the overall charge. When writing the ionic formula, we follow two additional conventions: (1) write the formula for the cation first and the formula for the anion next, but (2) do not write the charges on the ions. Thus, for the compound between Na^+ and Cl^- , we have the ionic formula NaCl (Figure 3.1 “NaCl = Table Salt”). The formula Na_2Cl_2 also has balanced charges, but the convention is to use the lowest ratio of ions, which would be one of each. (Remember from our conventions for writing formulas that we don’t write a 1 subscript if there is only one atom of a particular element present.) For the ionic compound between magnesium cations (Mg^{2+}) and oxide anions (O^{2-}), again we need only one of each ion to balance the charges. By convention, the formula is MgO.



Figure 3.1 “NaCl = Table Salt.” The ionic compound NaCl is very common.

For the ionic compound between Mg^{2+} ions and Cl^- ions, we now consider the fact that the charges have different magnitudes, 2+ on the magnesium ion and 1- on the chloride ion. To balance the charges with the lowest number of ions possible, we need to have two chloride ions to balance the charge on the one magnesium ion. Rather than write the formula MgClCl , we combine the two chloride ions and write it with a 2 subscript: MgCl_2 .

What is the formula MgCl_2 telling us? There are two chloride ions in the formula. Although chlorine as an element is a diatomic molecule, Cl_2 , elemental chlorine is not part of this ionic compound. The chlorine is in the form of a negatively charged *ion*, not the neutral *element*. The 2 subscript is in the ionic formula because we need two Cl^- ions to balance the charge on one Mg^{2+} ion.

Example 3.3

Problems

Write the proper ionic formula for each of the two given ions.

1. Ca^{2+} and Cl^-
2. Al^{3+} and F^-
3. Al^{3+} and O^{2-}

Solutions

1. We need two Cl^- ions to balance the charge on one Ca^{2+} ion, so the proper ionic formula is CaCl_2 .
2. We need three F^- ions to balance the charge on the Al^{3+} ion, so the proper ionic formula is AlF_3 .
3. With Al^{3+} and O^{2-} , note that neither charge is a perfect multiple of the other. This means we have to go to a least common multiple, which in this case will be six. To get a total of 6+, we need two Al^{3+} ions; to get 6-, we need three O^{2-} ions. Hence the proper ionic formula is Al_2O_3 .

Test Yourself

Write the proper ionic formulas for each of the two given ions.

1. Fe^{2+} and S^{2-}
2. Fe^{3+} and S^{2-}

Answers

1. FeS
2. Fe_2S_3

Naming ionic compounds is simple: combine the name of the cation and the name of the anion, in both cases omitting the word *ion*. *Do not use numerical prefixes if there is more than one ion necessary to balance the charges.* NaCl is sodium chloride, a combination of the name of the cation (sodium) and the anion (chloride). MgO is magnesium oxide. MgCl_2 is magnesium chloride—*not* magnesium dichloride.

In naming ionic compounds whose cations can have more than one possible charge, we must also include the charge, in parentheses and in roman numerals, as part of the name. Hence FeS is iron(II) sulfide, while Fe_2S_3 is iron(III) sulfide. Again, no numerical prefixes appear in the name. The number of ions in the formula is dictated by the need to balance the positive and negative charges.

Example 3.4

Problems

Name each ionic compound.

1. CaCl_2
2. AlF_3
3. Co_2O_3

Solutions

1. Using the names of the ions, this ionic compound is named calcium chloride. *It is not calcium(II) chloride* because calcium forms only one cation when it forms an ion, and it has a characteristic charge of 2+.
2. The name of this ionic compound is aluminum fluoride.
3. We know that cobalt can have more than one possible charge; we just need to determine what it is. Oxide always has a 2- charge, so with three oxide ions, we have a total negative charge of 6-. This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. Therefore, the proper name for this ionic compound is cobalt(III) oxide.

Test Yourself

Name each ionic compound.

1. Sc_2O_3
2. AgCl

Answers

1. scandium oxide
2. silver chloride

How do you know whether a formula—and by extension, a name—is for a molecular compound or for an ionic compound? Molecular compounds form between nonmetals and nonmetals, while ionic compounds form between metals and nonmetals. The periodic table ([Figure 3.5 “A Simple Periodic Table”](#)) can be used to determine which elements are metals and nonmetals.

There also exists a group of ions that contain more than one atom. These are called polyatomic ions. Table 3.3 “Common Polyatomic Ions” lists the formulas, charges, and names of some common polyatomic ions. Only one of them, the ammonium ion, is a cation; the rest are anions. Most of them also contain oxygen atoms, so sometimes they are referred to as *oxyanions*. Some of them, such as nitrate and nitrite, and sulfate and sulfite, have very similar formulas and names, so care must be taken to get the formulas and names correct. Note that the -ite polyatomic ion has one less oxygen atom in its formula than the -ate ion but with the same ionic charge.

Table 3.3 Common Polyatomic Ions

Name	Formula and Charge
ammonium	NH_4^+
acetate	$\text{C}_2\text{H}_3\text{O}_2^-$, or CH_3COO^-
bicarbonate (hydrogen carbonate)	HCO_3^-
bisulfate (hydrogen sulfate)	HSO_4^-
carbonate	CO_3^{2-}
chlorate	ClO_3^-
chromate	CrO_4^{2-}
cyanide	CN^-
dichromate	$\text{Cr}_2\text{O}_7^{2-}$
hydroxide	OH^-
nitrate	NO_3^-
nitrite	NO_2^-
peroxide	O_2^{2-}
perchlorate	ClO_4^-
phosphate	PO_4^{3-}
sulfate	SO_4^{2-}
sulfite	SO_3^{2-}
triiodide	I_3^-

The naming of ionic compounds that contain polyatomic ions follows the same rules as the naming for other ionic compounds: simply combine the name of the cation and the name of the anion. Do not use numerical prefixes in the name if there is more than one polyatomic ion; the only exception to this is if the name of the ion itself contains a numerical prefix, such as dichromate or triiodide.

Writing the formulas of ionic compounds has one important difference. If more than one polyatomic ion is needed to balance the overall charge in the formula, enclose the formula of the polyatomic ion in parentheses and write the proper numerical subscript to the right and *outside* the parentheses. Thus, the formula between calcium ions, Ca^{2+} , and nitrate ions, NO_3^- , is properly written $\text{Ca}(\text{NO}_3)_2$, not CaNO_3_2 or CaN_2O_6 . Use parentheses where required. The name of this ionic compound is simply calcium nitrate.

Example 3.5

Problems

Write the proper formula and give the proper name for each ionic compound formed between the two listed ions.

1. NH_4^{4+} and S^{2-}
2. Al^{3+} and PO_4^{3-}
3. Fe^{2+} and PO_4^{3-}

Solutions

1. Because the ammonium ion has a 1+ charge and the sulfide ion has a 2- charge, we need two ammonium ions to balance the charge on a single sulfide ion. Enclosing the formula for the ammonium ion in parentheses, we have $(\text{NH}_4)_2\text{S}$. The compound's name is ammonium sulfide.
2. Because the ions have the same magnitude of charge, we need only one of each to balance the charges. The formula is AlPO_4 , and the name of the compound is aluminum phosphate.
3. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, we need three iron(II) ions, and to get 6-, we need two phosphate ions. The proper formula is $\text{Fe}_3(\text{PO}_4)_2$, and the compound's name is iron(II) phosphate.

Test Yourself

Write the proper formula and give the proper name for each ionic compound formed between the two listed ions.

1. NH_4^+ and PO_4^{3-}
2. Co^{3+} and NO_2^-

Answers

1. $(\text{NH}_4)_3\text{PO}_4$, ammonium phosphate
2. $\text{Co}(\text{NO}_2)_3$, cobalt(III) nitrite

Food and Drink App: Sodium in Your Food

The element sodium, at least in its ionic form as Na^+ , is a necessary nutrient for humans to live. In fact, the human body is approximately 0.15% sodium, with the average person having one-twentieth to one-tenth of a kilogram in their body at any given time, mostly in fluids outside cells and in other bodily fluids.

Sodium is also present in our diet. The common table salt we use on our foods is an ionic sodium compound. Many processed foods also contain significant amounts of sodium added to them as a variety of ionic compounds. Why are sodium compounds used so much? Usually sodium compounds are inexpensive, but, more importantly, most ionic sodium compounds dissolve easily. This allows processed food manufacturers to add sodium-containing substances to food mixtures and know that the compound will dissolve and distribute evenly throughout the food. Simple ionic compounds such as sodium nitrite (NaNO_2) are added to cured meats, such as bacon and deli-style meats, while a compound called sodium benzoate is added to many packaged foods as a preservative. Table 3.4 "Some Sodium Compounds Added to Food" is a

partial list of some sodium additives used in food. Some of them you may recognize after reading this chapter. Others you may not recognize, but they are all ionic sodium compounds with some negatively charged ion also present.

Table 3.4 Some Sodium Compounds Added to Food

Sodium Compound	Use in Food
Sodium acetate	preservative, acidity regulator
Sodium adipate	food acid
Sodium alginate	thickener, vegetable gum, stabilizer, gelling agent, emulsifier
Sodium aluminum phosphate	acidity regulator, emulsifier
Sodium aluminosilicate	anticaking agent
Sodium ascorbate	antioxidant
Sodium benzoate	preservative
Sodium bicarbonate	mineral salt
Sodium bisulfite	preservative, antioxidant
Sodium carbonate	mineral salt
Sodium carboxymethylcellulose	emulsifier
Sodium citrates	food acid
Sodium dehydroacetate	preservative
Sodium erythorbate	antioxidant
Sodium erythorbin	antioxidant
Sodium ethyl para-hydroxybenzoate	preservative
Sodium ferrocyanide	anticaking agent
Sodium formate	preservative
Sodium fumarate	food acid
Sodium gluconate	stabilizer
Sodium hydrogen acetate	preservative, acidity regulator
Sodium hydroxide	mineral salt
Sodium lactate	food acid
Sodium malate	food acid
Sodium metabisulfite	preservative, antioxidant, bleaching agent
Sodium methyl para-hydroxybenzoate	preservative
Sodium nitrate	preservative, color fixative
Sodium nitrite	preservative, color fixative
Sodium orthophenyl phenol	preservative
Sodium propionate	preservative
Sodium propyl para-hydroxybenzoate	preservative
Sodium sorbate	preservative
Sodium stearyl lactylate	emulsifier
Sodium succinates	acidity regulator, flavour enhancer

Sodium Compound	Use in Food
Sodium salts of fatty acids	emulsifier, stabilizer, anticaking agent
Sodium sulfite	mineral salt, preservative, antioxidant
Sodium sulfite	preservative, antioxidant
Sodium tartrate	food acid
Sodium tetraborate	preservative

The use of so many sodium compounds in prepared and processed foods has alarmed some physicians and nutritionists. They argue that the average person consumes too much sodium from his or her diet. The average person needs only about 500 mg of sodium every day; most people consume more than this—up to 10 times as much. Some studies have implicated increased sodium intake with high blood pressure; newer studies suggest that the link is questionable. However, there has been a push to reduce the amount of sodium most people ingest every day: avoid processed and manufactured foods, read labels on packaged foods (which include an indication of the sodium content), don't oversalt foods, and use other herbs and spices besides salt in cooking.

Nutrition Facts	
Serving Size 8 oz (227 g/8 oz)	
Servings Per Container About 3	
Amount Per Serving	
Calories 180	Calories from Fat 60
% Daily Value*	
Total Fat 6g	10%
Saturated Fat 1g	5%
<i>Trans</i> Fat 0g	
Cholesterol 5mg	2%
Sodium 75mg	3%
Total Carbohydrate 26g	9%
Dietary Fiber 5g	19%
Sugars 11g	
Protein 8g	
Vitamin A 60%	• Vitamin C 70%
Calcium 8%	• Iron 10%
* Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs.	
	Calories 2,000 2,500
Total Fat	Less than 65g 80g
Sat Fat	Less than 20g 25g
Cholesterol	Less than 300mg 300mg
Sodium	Less than 2,400mg 2,400mg
Total Carbohydrate	300g 375g
Dietary Fiber	25g 30g
Calories per gram:	
Fat 9	• Carbohydrate 4 • Protein 4

Figure 3.2 “Nutrition Facts.” Food labels include the amount of sodium per serving. This particular label shows that there are 75 mg of sodium in one serving of this particular food item.

Key Takeaways

- Ions form when atoms lose or gain electrons.
- Ionic compounds have positive ions and negative ions.
- Ionic formulas balance the total positive and negative charges.

- Ionic compounds have a simple system of naming.
- Groups of atoms can have an overall charge and make ionic compounds.

Exercises

Questions

1. Explain how cations form.
2. Explain how anions form.
3. Give the charge each atom takes when it forms an ion. If more than one charge is possible, list both.
 - a. K
 - b. O
 - c. Co
4. Give the charge each atom takes when it forms an ion. If more than one charge is possible, list both.
 - a. Ca
 - b. I
 - c. Fe
5. Give the charge each atom takes when it forms an ion. If more than one charge is possible, list both.
 - a. Ag
 - b. Au
 - c. Br
6. Give the charge each atom takes when it forms an ion. If more than one charge is possible, list both.
 - a. S
 - b. Na
 - c. H
7. Name the ions from Exercise 3.
8. Name the ions from Exercise 4.
9. Name the ions from Exercise 5.
10. Name the ions from Exercise 6.
11. Give the formula and name for each ionic compound formed between the two listed ions.
 - a. Mg^{2+} and Cl^-
 - b. Fe^{2+} and O^{2-}
 - c. Fe^{3+} and O^{2-}
12. Give the formula and name for each ionic compound formed between the two listed ions.
 - a. K^+ and S^{2-}
 - b. Ag^+ and Br^-
 - c. Sr^{2+} and N^{3-}

13. Give the formula and name for each ionic compound formed between the two listed ions.
- Cu^{2+} and F^-
 - Ca^{2+} and O^{2-}
 - K^+ and P^{3-}
14. Give the formula and name for each ionic compound formed between the two listed ions.
- Na^+ and N^{3-}
 - Co^{2+} and I^-
 - Au^{3+} and S^{2-}
15. Give the formula and name for each ionic compound formed between the two listed ions.
- K^+ and SO_4^{2-}
 - NH_4^+ and S^{2-}
 - NH_4^+ and PO_4^{3-}
16. Give the formula and name for each ionic compound formed between the two listed ions.
- Ca^{2+} and NO_3^-
 - Ca^{2+} and NO_2^-
 - Sc^{3+} and $\text{C}_2\text{H}_3\text{O}_2^-$
17. Give the formula and name for each ionic compound formed between the two listed ions.
- Pb^{4+} and SO_4^{2-}
 - Na^+ and I_3^-
 - Li^+ and $\text{Cr}_2\text{O}_7^{2-}$
18. Give the formula and name for each ionic compound formed between the two listed ions.
- NH_4^+ and N^{3-}
 - Mg^{2+} and CO_3^{2-}
 - Al^{3+} and OH^-
19. Give the formula and name for each ionic compound formed between the two listed ions.
- Ag^+ and SO_3^{2-}
 - Na^+ and HCO_3^-
 - Fe^{3+} and ClO_3^-
20. Give the formula and name for each ionic compound formed between the two listed ions.
- Rb^+ and O_2^{2-}
 - Au^{3+} and HSO_4^-
 - Sr^{2+} and NO_2^-
21. What is the difference between SO_3 and SO_3^{2-} ?
22. What is the difference between NO_2 and NO_2^- ?

Answers

1. Cations form by losing electrons.

3.
 - a. 1+
 - b. 2-
 - c. 2+, 3+
5.
 - a. 1+
 - b. 1+, 3+
 - c. 1-
7.
 - a. the potassium ion
 - b. the oxide ion
 - c. the cobalt(II) and cobalt(III) ions, respectively
9.
 - a. the silver ion
 - b. the gold(I) and gold(III) ions, respectively
 - c. the bromide ion
11.
 - a. magnesium chloride, MgCl_2
 - b. iron(II) oxide, FeO
 - c. iron(III) oxide, Fe_2O_3
13.
 - a. copper(II) fluoride, CuF_2
 - b. calcium oxide, CaO
 - c. potassium phosphide, K_3P
15.
 - a. potassium sulfate, K_2SO_4
 - b. ammonium sulfide, $(\text{NH}_4)_2\text{S}$
 - c. ammonium phosphate, $(\text{NH}_4)_3\text{PO}_4$
17.
 - a. lead(IV) sulfate, $\text{Pb}(\text{SO}_4)_2$
 - b. sodium triiodide, NaI_3
 - c. lithium dichromate, $\text{Li}_2\text{Cr}_2\text{O}_7$
19.
 - a. silver sulfite, Ag_2SO_3
 - b. sodium hydrogen carbonate, NaHCO_3
 - c. iron(III) chlorate, $\text{Fe}(\text{ClO}_3)_3$
21. SO_3 is sulfur trioxide, while SO_3^{2-} is the sulfite ion.

Media Attributions

Figure 3.1

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Masses of Atoms and Molecules

Learning Objectives

1. Express the masses of atoms and molecules.

Because matter is defined as anything that has mass and takes up space, it should not be surprising to learn that atoms and molecules have mass.

Individual atoms and molecules, however, are very small, and the masses of individual atoms and molecules are also very small. For macroscopic objects, we use units such as grams and kilograms to state their masses, but these units are much too big to comfortably describe the masses of individual atoms and molecules. Another scale is needed.

The **atomic mass unit** (u; some texts use amu, but this older style is no longer accepted) is defined as one-twelfth of the mass of a carbon-12 atom, an isotope of carbon that has six protons and six neutrons in its nucleus. By this scale, the mass of a proton is 1.00728 u, the mass of a neutron is 1.00866 u, and the mass of an electron is 0.000549 u. There will not be much error if you estimate the mass of an atom by simply counting the total number of protons and neutrons in the nucleus (i.e., identify its mass number) and ignore the electrons. Thus, the mass of carbon-12 is about 12 u, the mass of oxygen-16 is about 16 u, and the mass of uranium-238 is about 238 u. More exact masses are found in scientific references—for example, the exact mass of uranium-238 is 238.050788 u, so you can see that we are not far off by using the whole-number value as the mass of the atom.

What is the mass of an element? This is somewhat more complicated because most elements exist as a mixture of isotopes, each of which has its own mass. Thus, although it is easy to speak of the mass of an atom, when talking about the mass of an element, we must take the isotopic mixture into account.

The **atomic mass** of an element is a weighted average of the masses of the isotopes that compose an element. What do we mean by a weighted average? Well, consider an element that consists of two isotopes, 50% with mass 10 u and 50% with mass 11 u. A weighted average is found by multiplying each mass by its fractional occurrence (in decimal form) and then adding all the products. The sum is the weighted average and serves as the formal atomic mass of the element. In this example, we have the following:

$$\begin{array}{r}
 0.50 \times 10 \text{ u} = 5.0 \text{ u} \\
 + 0.50 \times 11 \text{ u} = 5.5 \text{ u} \\
 \hline
 \text{sum} = 10.5 \text{ u} = \text{the atomic mass of our element}
 \end{array}$$

Note that no atom in our hypothetical element has a mass of 10.5 u; rather, that is the average mass of the atoms, weighted by their percent occurrence.

This example is similar to a real element. Boron exists as about 20% boron-10 (five protons and five neutrons in the nuclei) and about 80% boron-11 (five protons and six neutrons in the nuclei). The atomic mass of boron is calculated similarly to what we did for our hypothetical example, but the percentages are different:

$$\begin{array}{r} 0.20 \times 10 \text{ u} = 2.0 \text{ u} \\ + 0.80 \times 11 \text{ u} = 8.8 \text{ u} \\ \hline \text{sum} = 10.8 \text{ u} = \text{the atomic mass of boron} \end{array}$$

Thus, we use 10.8 u for the atomic mass of boron.

Virtually all elements exist as mixtures of isotopes, so atomic masses may vary significantly from whole numbers. Table 3.5 “Selected Atomic Masses of Some Elements” lists the atomic masses of some elements; a more expansive table is in [“Appendix A: Periodic Table of the Elements”](#). The atomic masses in Table 3.5 “Selected Atomic Masses of Some Elements” are listed to three decimal places where possible, but in most cases, only one or two decimal places are needed. Note that many of the atomic masses, especially the larger ones, are not very close to whole numbers. This is, in part, the effect of an increasing number of isotopes as the atoms increase in size. (The record number is 10 isotopes for tin.)

Table 3.5 Selected Atomic Masses of Some Elements¹

Element Name	Atomic Mass (u)
Aluminum	26.981
Argon	39.948
Arsenic	74.922
Barium	137.327
Beryllium	9.012
Bismuth	208.980
Boron	10.811
Bromine	79.904
Calcium	40.078
Carbon	12.011
Chlorine	35.453
Cobalt	58.933
Copper	63.546
Fluorine	18.998
Gallium	69.723
Germanium	72.64
Gold	196.967
Helium	4.003
Hydrogen	1.008
Iodine	126.904
Iridium	192.217
Iron	55.845
Krypton	83.798
Lead	207.2
Lithium	6.941
Magnesium	24.305
Manganese	54.938
Mercury	200.59

Element Name	Atomic Mass (u)
Molybdenum	95.94
Neon	20.180
Nickel	58.693
Nitrogen	14.007
Oxygen	15.999
Palladium	106.42
Phosphorus	30.974
Platinum	195.084
Potassium	39.098
Radium	n/a
Radon	n/a
Rubidium	85.468
Scandium	44.956
Selenium	78.96
Silicon	28.086
Silver	107.868
Sodium	22.990
Strontium	87.62
Sulfur	32.065
Tantalum	180.948
Tin	118.710
Titanium	47.867
Tungsten	183.84
Uranium	238.029
Xenon	131.293
Zinc	65.409
Zirconium	91.224

1. Note: Atomic mass is given to three decimal places, if known.

Now that we understand that atoms have mass, it is easy to extend the concept to the mass of molecules. The **molecular mass** is the sum of the masses of the atoms in a molecule. This may seem like a trivial extension of the concept, but it is important to count the number of each type of atom in the molecular formula. Also, although each atom in a molecule is a particular isotope, we use the weighted average, or atomic mass, for each atom in the molecule.

For example, if we were to determine the molecular mass of dinitrogen trioxide, N_2O_3 , we would need to add the atomic mass of nitrogen two times with the atomic mass of oxygen three times:

$$\begin{array}{r}
 2 \text{ N masses} = 2 \times 14.007 \text{ u} = 28.014 \text{ u} \\
 + 3 \text{ O masses} = 3 \times 15.999 \text{ u} = 47.997 \text{ u} \\
 \hline
 \text{total} = 76.011 \text{ u} = \text{the molecular mass of } \text{N}_2\text{O}_3
 \end{array}$$

We would not be far off if we limited our numbers to one or even two decimal places.

Example 3.6

Problems

What is the molecular mass of each substance?

1. NBr_3
2. C_2H_6

Solutions

1. Add one atomic mass of nitrogen and three atomic masses of bromine:

$$\begin{array}{r}
 1\text{N mass} = 1 \times 14.007 \text{ u} = 14.007 \text{ u} \\
 + 3\text{Br masses} = 3 \times 79.904 \text{ u} = 239.712 \text{ u} \\
 \hline
 \text{total} = 253.719 \text{ u} = \text{the molecular mass of } \text{NBr}_3
 \end{array}$$

2. Add two atomic masses of carbon and six atomic masses of hydrogen:

$$\begin{array}{r}
 2\text{C masses} = 2 \times 12.011 \text{ u} = 24.022 \text{ u} \\
 + 6\text{H masses} = 6 \times 1.008 \text{ u} = 6.048 \text{ u} \\
 \hline
 \text{total} = 30.070 \text{ u} = \text{the molecular mass of } \text{C}_2\text{H}_6
 \end{array}$$

The compound C_2H_6 also has a common name—ethane.

Test Yourself

What is the molecular mass of each substance?

1. SO_2
2. PF_3

Answers

1. 64.063 u
2. 87.968 u

Chemistry Is Everywhere: Sulfur Hexafluoride

On March 20, 1995, the Japanese terrorist group Aum Shinrikyo (Sanskrit for “Supreme Truth”) released some sarin gas in the Tokyo subway system; twelve people were killed, and thousands were injured (part (a) in the accompanying figure). Sarin (molecular formula $C_4H_{10}FPO_2$) is a nerve toxin that was first synthesized in 1938. It is regarded as one of the most deadly toxins known, estimated to be about 500 times more potent than cyanide. Scientists and engineers who study the spread of chemical weapons such as sarin (yes, there are such scientists) would like to have a less dangerous chemical, indeed one that is nontoxic, so they are not at risk themselves.

Sulfur hexafluoride is used as a model compound for sarin. SF_6 (a molecular model of which is shown in part (b) in the accompanying figure) has a similar molecular mass (about 146 u) as sarin (about 140 u), so it has similar physical properties in the vapour phase. Sulfur hexafluoride is also very easy to accurately detect, even at low levels, and it is not a normal part of the atmosphere, so there is little potential for contamination from natural sources. Consequently, SF_6 is also used as an aerial tracer for ventilation systems in buildings. It is nontoxic and very chemically inert, so workers do not have to take special precautions other than watching for asphyxiation.

Sulfur hexafluoride also has another interesting use: a spark suppressant in high-voltage electrical equipment. High-pressure SF_6 gas is used in place of older oils that may have contaminants that are environmentally unfriendly (part (c) in the accompanying figure).

Key Takeaways

- The atomic mass unit (u) is a unit that describes the masses of individual atoms and molecules.
- The atomic mass is the weighted average of the masses of all isotopes of an element.
- The molecular mass is the sum of the masses of the atoms in a molecule.

Exercises**Questions**

1. Define *atomic mass unit*. What is its abbreviation?
2. Define *atomic mass*. What is its unit?
3. Estimate the mass, in whole numbers, of each isotope.
 - a. hydrogen-1

- b. hydrogen-3
 - c. iron-56
4. Estimate the mass, in whole numbers, of each isotope.
 - a. phosphorus-31
 - b. carbon-14
 - c. americium-241
5. Determine the atomic mass of each element, given the isotopic composition.
 - a. lithium, which is 92.4% lithium-7 (mass 7.016 u) and 7.60% lithium-6 (mass 6.015 u)
 - b. oxygen, which is 99.76% oxygen-16 (mass 15.995 u), 0.038% oxygen-17 (mass 16.999 u), and 0.205% oxygen-18 (mass 17.999 u)
6. Determine the atomic mass of each element, given the isotopic composition.
 - a. neon, which is 90.48% neon-20 (mass 19.992 u), 0.27% neon-21 (mass 20.994 u), and 9.25% neon-22 (mass 21.991 u)
 - b. uranium, which is 99.27% uranium-238 (mass 238.051 u) and 0.720% uranium-235 (mass 235.044 u)
7. How far off would your answer be from Exercise 5a if you used whole-number masses for individual isotopes of lithium?
8. How far off would your answer be from Exercise 6b if you used whole-number masses for individual isotopes of uranium?
9. Find the atomic or molecular mass of oxygen in each of these forms:
 - a. An oxygen atom
 - b. Oxygen in its elemental form
10. Find the atomic or molecular mass of bromine in each of these forms:
 - a. A bromine atom
 - b. Bromine in its elemental form
11. Determine the mass of each substance.
 - a. F_2
 - b. CO
 - c. CO_2
12. Determine the mass of each substance.
 - a. Kr
 - b. KrF_4
 - c. PF_5
13. Determine the mass of each substance.
 - a. Na
 - b. B_2O_3
 - c. S_2Cl_2
14. Determine the mass of each substance.
 - a. IBr_3
 - b. N_2O_5

c. CCl_4

15. Determine the mass of each substance.

a. GeO_2

b. IF_3

c. XeF_6

16. Determine the mass of each substance.

a. NO

b. N_2O_4

c. Ca

Answers

1. The atomic mass unit is defined as one-twelfth of the mass of a carbon-12 atom. Its abbreviation is u.

3. a. 1

b. 3

c. 56

5. a. 6.940 u

b. 16.000 u

7. We would get 6.924 u.

9. a. 15.999 u

b. 31.998 u

11. a. 37.996 u

b. 28.010 u

c. 44.009 u

13. a. 22.990 u

b. 69.619 u

c. 135.036 u

15. a. 104.64 u

b. 183.898 u

c. 245.281 u

Molecules and Chemical Nomenclature

Learning Objectives

1. Define *molecule*.
2. Name simple molecules based on their formulas.
3. Determine a formula of a molecule based on its name.

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multiatom combinations are called molecules. A molecule is the smallest part of a substance that has the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules. Other elements also exist naturally as diatomic molecules (see the list [“Elements That Exist as Diatomic Molecules”](#)). As with any molecule, these elements are labelled with a molecular formula, a formal listing of what and how many atoms are in a molecule. (Sometimes only the word *formula* is used, and its meaning is inferred from the context.) For example, the molecular formula for elemental hydrogen is H_2 , with H being the symbol for hydrogen and the subscript 2 implying that there are two atoms of this element in the molecule. Other diatomic elements have similar formulas: O_2 , N_2 , and so forth. Other elements exist as molecules—for example, sulfur normally exists as an eight-atom molecule, S_8 , while phosphorus exists as a four-atom molecule, P_4 (see Figure 3.3 “Molecular Art of S”). Otherwise, we will assume that elements exist as individual atoms, rather than molecules. It is assumed that there is only one atom in a formula if there is no numerical subscript on the right side of an element’s symbol.

Elements that Exist as Diatomic Molecules

The following is a list of elements that exist as diatomic molecules:

1. hydrogen
2. oxygen
3. nitrogen
4. fluorine
5. chlorine
6. bromine

7. iodine



Figure 3.3 “Molecular Art of S_8 and P_4 Molecules.”
 If each green ball represents a sulfur atom, then the diagram on the left represents an S_8 molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure 3.3 “Molecular Art of S” shows two examples of how we will be representing molecules in this text. An atom is represented by a small ball or sphere, which generally indicates where the nucleus is in the molecule. A cylindrical line connecting the balls represents the connection between the atoms that make this collection of atoms a molecule. This connection is called a chemical bond. In [Chapter 9 “Chemical Bonds”](#), we will explore the origin of chemical bonds. You will see other examples of this “ball and cylinder” representation of molecules throughout this book.

Many compounds exist as molecules. In particular, when nonmetals connect with other nonmetals, the compound typically exists as molecules. (Compounds between a metal and a nonmetal are different and are considered in [“Ions and Ionic Compounds”](#).) Furthermore, in some cases there are many different kinds of molecules that can be formed between any given elements, with all the different molecules having different chemical and physical properties. How do we tell them apart?

The answer is a very specific system of naming compounds, called chemical nomenclature. By following the rules of nomenclature, each and every compound has its own unique name, and each name refers to one and only one compound. Here, we will start with relatively simple molecules that have only two elements in them, the so-called *binary compounds*:

1. Identify the elements in the molecule from its formula. This is why you need to know the names and symbols of the elements in [Table 3.8 “Names and Symbols of Common Elements”](#).
2. Begin the name with the element name of the first element. If there is more than one atom of this element in the molecular formula, use a numerical prefix to indicate the number of atoms, as listed in Table 3.6 “Numerical Prefixes Used in Naming Molecular Compounds”.
Do not use the prefix mono- if there is only one atom of the first element.

Table 3.6 Numerical Prefixes Used in Naming Molecular Compounds

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

3. Name the second element by using three pieces:
 1. a numerical prefix indicating the number of atoms of the second element, plus
 2. the stem of the element name (e.g., *ox* for oxygen, *chlor* for chlorine, etc.), plus
 3. the suffix *-ide*.
4. Combine the two words, leaving a space between them.

Let us see how these steps work for a molecule whose molecular formula is SO_2 , which has one sulfur atom and two oxygen atoms—this completes step 1. According to step 2, we start with the name of the first element—sulfur. Remember, we don't use the *mono-* prefix for the first element. Now for step 3, we combine the numerical prefix *di-* (see Table 3.6 “Numerical Prefixes Used in Naming Molecular Compounds”) with the stem *ox-* and the suffix *-ide*, to make *dioxide*. Bringing these two words together, we have the unique name for this compound—sulfur dioxide.

Why all this trouble? There is another common compound consisting of sulfur and oxygen whose molecular formula is SO_3 , so the compounds need to be distinguished. SO_3 has three oxygen atoms in it, so it is a different compound with different chemical and physical properties. The system of chemical nomenclature is designed to *give this compound its own unique name*. Its name, if you go through all the steps, is sulfur trioxide. Different compounds have different names.

In some cases, when a prefix ends in *a* or *o* and the element name begins with *o* we drop the *a* or *o* on the prefix. So we see *monoxide* or *pentoxide* rather than *monooxide* or *pentaoxide* in molecule names.

One great thing about this system is that it works both ways. From the name of a compound, you should be able to determine its molecular formula. Simply list the element symbols, with a numerical subscript if there is more than one atom of that element, in the order of the name (we do not use a subscript 1 if there is only one atom of the element present; 1 is implied). From the name *nitrogen*

trichloride, you should be able to get NCl_3 as the formula for this molecule. From the name *diphosphorus pentoxide*, you should be able to get the formula P_2O_5 (note the numerical prefix on the first element, indicating there is more than one atom of phosphorus in the formula).

Example 3.7

Problems

Name each molecule.

1. PF_3
2. CO
3. Se_2Br_2

Solutions

1. A molecule with a single phosphorus atom and three fluorine atoms is called phosphorus trifluoride.
2. A compound with one carbon atom and one oxygen atom is properly called carbon monoxide, not carbon monooxide.
3. There are two atoms of each element, selenium and bromine. According to the rules, the proper name here is *diselenium dibromide*.

Test Yourself

Name each molecule.

1. SF_4
2. P_2S_5

Answers

1. sulfur tetrafluoride
2. diphosphorus pentasulfide

Example 3.8

Problems

Give the formula for each molecule.

1. carbon tetrachloride
2. silicon dioxide
3. trisilicon tetranitride

Solutions

1. The name carbon tetrachloride implies one carbon atom and four chlorine atoms, so the formula is CCl_4 .
2. The name silicon dioxide implies one silicon atom and two oxygen atoms, so the formula is SiO_2 .
3. We have a name that has numerical prefixes on both elements. Tri- means three, and tetra- means four, so the formula of this compound is Si_3N_4 .

Test Yourself

Give the formula for each molecule.

1. disulfur difluoride
2. iodine pentabromide

Answers

1. S_2F_2
2. IBr_5

Some simple molecules have common names that we use as part of the formal system of chemical nomenclature. For example, H_2O is given the name *water*, not *dihydrogen monoxide*. NH_3 is called *ammonia*, while CH_4 is called *methane*. We will occasionally see other molecules that have common names; we will point them out as they occur.

Key Takeaways

- Molecules are groups of atoms that behave as a single unit.
- Some elements exist as molecules: hydrogen, oxygen, sulfur, and so forth.
- There are rules that can express a unique name for any given molecule, and a unique formula for any given name.

Exercises

Questions

1. Which of these formulas represent molecules? State how many atoms are in each molecule.
 - a. Fe
 - b. PCl_3
 - c. P_4
 - d. Ar

2. Which of these formulas represent molecules? State how many atoms are in each molecule.
 - a. I_2
 - b. He
 - c. H_2O
 - d. Al
3. What is the difference between CO and Co?
4. What is the difference between H_2O and H_2O_2 (hydrogen peroxide)?
5. Give the proper formula for each diatomic element.
6. In 1986, when Halley's comet last passed the earth, astronomers detected the presence of S_2 in their telescopes. Why is sulfur not considered a diatomic element?
7. What is the stem of fluorine used in molecule names? CF_4 is one example.
8. What is the stem of selenium used in molecule names? $SiSe_2$ is an example.
9. Give the proper name for each molecule.
 - a. PF_3
 - b. $TeCl_2$
 - c. N_2O_3
10. Give the proper name for each molecule.
 - a. NO
 - b. CS_2
 - c. As_2O_3
11. Give the proper name for each molecule.
 - a. XeF_2
 - b. O_2F_2
 - c. SF_6
12. Give the proper name for each molecule.
 - a. P_4O_{10}
 - b. B_2O_3
 - c. P_2S_3
13. Give the proper name for each molecule.
 - a. N_2O
 - b. N_2O_4
 - c. N_2O_5
14. Give the proper name for each molecule.
 - a. SeO_2
 - b. Cl_2O
 - c. XeF_6
15. Give the proper formula for each name.
 - a. dinitrogen pentoxide
 - b. tetraboron tricarbide

- c. phosphorus pentachloride
16. Give the proper formula for each name.
- nitrogen triiodide
 - diarsenic trisulfide
 - iodine trichloride
17. Give the proper formula for each name.
- dioxygen dichloride
 - dinitrogen trisulfide
 - xenon tetrafluoride
18. Give the proper formula for each name.
- chlorine dioxide
 - selenium dibromide
 - dinitrogen trioxide
19. Give the proper formula for each name.
- iodine trifluoride
 - xenon trioxide
 - disulfur decafluoride
20. Give the proper formula for each name.
- germanium dioxide
 - carbon disulfide
 - diselenium dibromide

Answers

1. a. not a molecule
 b. a molecule; four atoms total
 c. a molecule; four atoms total
 d. not a molecule
3. CO is a compound of carbon and oxygen; Co is the element cobalt.
5. H₂, O₂, N₂, F₂, Cl₂, Br₂, I₂
7. *fluor-*
9. a. phosphorus trifluoride
 b. tellurium dichloride
 c. dinitrogen trioxide
11. a. xenon difluoride
 b. dioxygen difluoride
 c. sulfur hexafluoride
13. a. dinitrogen monoxide

- b. dinitrogen tetroxide
 - c. dinitrogen pentoxide
- 15.
- a. N_2O_5
 - b. B_4C_3
 - c. PCl_5
- 17.
- a. O_2Cl_2
 - b. N_2S_3
 - c. XeF_4
- 19.
- a. IF_3
 - b. XeO_3
 - c. S_2F_{10}

Media Attributions

Figure 3.3

- “Molecular Art of S_8 and P_4 Molecules” by David W. Ball © [CC BY-NC-SA \(Attribution NonCommercial ShareAlike\)](#)

Atomic Theory

Learning Objectives

1. State the modern atomic theory.
2. Learn how atoms are constructed.

The smallest piece of an element that maintains the identity of that element is called an **atom**. Individual atoms are extremely small. It would take about fifty million atoms in a row to make a line that is 1 cm long. The period at the end of a printed sentence has several million atoms in it. Atoms are so small that it is difficult to believe that all matter is made from atoms—but it is.

The concept that atoms play a fundamental role in chemistry is formalized by the modern **atomic theory**, first stated by John Dalton, an English scientist, in 1808. It consists of three parts:

1. All matter is composed of atoms.
2. Atoms of the same element are the same; atoms of different elements are different.
3. Atoms combine in whole-number ratios to form compounds.

These concepts form the basis of chemistry.

Although the word *atom* comes from a Greek word that means “indivisible,” we understand now that atoms themselves are composed of smaller parts called *subatomic particles*. The first part to be discovered was the **electron**, a tiny subatomic particle with a negative charge. It is often represented as e^- , with the right superscript showing the negative charge. Later, two larger particles were discovered. The **proton** is a more massive (but still tiny) subatomic particle with a positive charge, represented as p^+ . The **neutron** is a subatomic particle with about the same mass as a proton but no charge. It is represented as either n or n^0 . We now know that all atoms of all elements are composed of electrons, protons, and (with one exception) neutrons. Table 3.7 “Properties of the Three Subatomic Particles” summarizes the properties of these three subatomic particles.

Table 3.7 Properties of the Three Subatomic Particles

Name	Symbol	Mass (approx. in kg)	Charge
Proton	p^+	1.6×10^{-27}	1+
Neutron	n, n^0	1.6×10^{-27}	none
Electron	e^-	9.1×10^{-31}	1-

How are these particles arranged in atoms? They are not arranged at random. Experiments by Ernest Rutherford in England in the 1910s pointed to a **nuclear model** of the atom. The relatively massive protons and neutrons are collected in the centre of an atom, in a region called the **nucleus** of the atom (plural *nuclei*). The electrons are outside the nucleus and spend their time orbiting in space about the nucleus. (See Figure 3.4 “The Structure of the Atom”.)

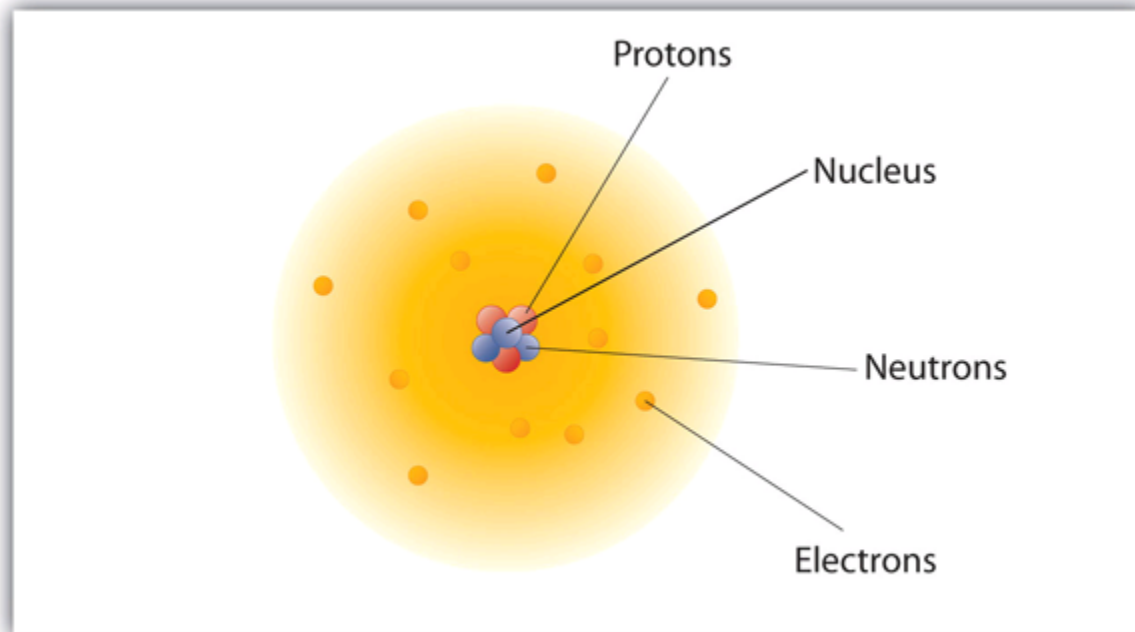


Figure 3.4 “The Structure of the Atom.” Atoms have protons and neutrons in the centre, making the nucleus, while the electrons orbit the nucleus.

The modern atomic theory states that atoms of one element are the same, while atoms of different elements are different. What makes atoms of different elements different? The fundamental characteristic that all atoms of the same element share is the *number of protons*. All atoms of hydrogen have one and only one proton in the nucleus; all atoms of iron have 26 protons in the nucleus. This number of protons is so important to the identity of an atom that it is called the **atomic number** of the element. Thus, hydrogen has an atomic number of 1, while iron has an atomic number of 26. Each element has its own characteristic atomic number.

Atoms of the same element can have different numbers of neutrons, however. Atoms of the same element (i.e., atoms with the same number of protons) with different numbers of neutrons are called **isotopes**. Most naturally occurring elements exist as isotopes. For example, most hydrogen atoms have a single proton in their nucleus. However, a small number (about one in a million) of hydrogen atoms have a proton and a neutron in their nuclei. This particular isotope of hydrogen is called deuterium. A very rare form of hydrogen has one proton and two neutrons in the nucleus; this isotope of hydrogen is called tritium. The sum of the number of protons and neutrons in the nucleus is called the **mass number** of the isotope.

Neutral atoms have the same number of electrons as they have protons, so their overall charge is zero. However, as we shall see later, this will not always be the case.

Example 3.9

Problems

1. The most common carbon atoms have six protons and six neutrons in their nuclei. What are the atomic number and the mass number of these carbon atoms?
2. An isotope of uranium has an atomic number of 92 and a mass number of 235. What are the number of protons and neutrons in the nucleus of this atom?

Solutions

1. If a carbon atom has six protons in its nucleus, its atomic number is 6. If it also has six neutrons in the nucleus, then the mass number is $6 + 6$, or 12.
2. If the atomic number of uranium is 92, then that is the number of protons in the nucleus. Because the mass number is 235, then the number of neutrons in the nucleus is $235 - 92$, or 143.

Test Yourself

The number of protons in the nucleus of a tin atom is 50, while the number of neutrons in the nucleus is 68. What are the atomic number and the mass number of this isotope?

Answer

Atomic number = 50, mass number = 118

When referring to an atom, we simply use the element's name: the term *sodium* refers to the element as well as an atom of sodium. But it can be unwieldy to use the name of elements all the time. Instead, chemistry defines a symbol for each element. The **atomic symbol** is a one- or two-letter abbreviation of the name of the element. By convention, the first letter of an element's symbol is always capitalized, while the second letter (if present) is lowercase. Thus, the symbol for hydrogen is H, the symbol for sodium is Na, and the symbol for nickel is Ni. Most symbols come from the English name of the element, although some symbols come from an element's Latin name. (The symbol for sodium, Na, comes from its Latin name, *natrium*.) Table 3.8 "Names and Symbols of Common Elements" lists some common elements and their symbols. You should memorize the symbols in Table 3.8 "Names and Symbols of Common Elements", as this is how we will be representing elements throughout chemistry.

Table 3.8 Names and Symbols of Common Elements

Element Name	Symbol
Aluminum	Al
Argon	Ar
Arsenic	As
Barium	Ba
Beryllium	Be
Bismuth	Bi
Boron	B
Bromine	Br
Calcium	Ca
Carbon	C
Chlorine	Cl
Chromium	Cr
Cobalt	Co
Copper	Cu
Fluorine	F
Gallium	Ga
Germanium	Ge
Gold	Au
Helium	He
Hydrogen	H
Iodine	I
Iridium	Ir
Iron	Fe
Krypton	Kr
Lead	Pb
Lithium	Li
Magnesium	Mg
Manganese	Mn

Element Name	Symbol
Mercury	Hg
Molybdenum	Mo
Neon	Ne
Nickel	Ni
Nitrogen	N
Oxygen	O
Palladium	Pd
Phosphorus	P
Platinum	Pt
Potassium	K
Radium	Ra
Radon	Rn
Rubidium	Rb
Scandium	Sc
Selenium	Se
Silicon	Si
Silver	Ag
Sodium	Na
Strontium	Sr
Sulfur	S
Tantalum	Ta
Tin	Sn
Titanium	Ti
Tungsten	W
Uranium	U
Xenon	Xe
Zinc	Zn
Zirconium	Zr

The elements are grouped together in a special chart called the **periodic table**. A simple periodic table is shown in Figure 3.5 “A Simple Periodic Table”, while a more extensive one is presented in “[Appendix A: Periodic Table of the Elements](#)”. The elements on the periodic table are listed in order of ascending atomic number. The periodic table has a special shape that will become important to us when we consider the organization of electrons in atoms (see [Chapter 8 “Electronic Structure”](#)). One immediate use of the periodic table helps us identify metals and nonmetals. Nonmetals are in the upper right corner of the periodic table, on one side of the heavy line splitting the right-hand part of the chart. All other elements are metals.

1 H 1.00794																	1 H 1.00794	2 He 4.002602
3 Li 6.941	4 Be 9.012182											5 B 10.811	6 C 12.0107	7 N 14.00674	8 O 15.9994	9 F 18.9984032	10 Ne 20.1797	
11 Na 22.989770	12 Mg 24.3050											13 Al 26.981538	14 Si 28.0855	15 P 30.973761	16 S 32.066	17 Cl 35.4527	18 Ar 39.948	
19 K 39.0983	20 Ca 40.078	21 Sc 44.955910	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.938049	26 Fe 55.845	27 Co 58.933200	28 Ni 58.6534	29 Cu 63.545	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.504	36 Kr 83.80	
37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.90550	46 Pd 106.42	47 Ag 196.56655	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.90447	54 Xe 131.29	
55 Cs 132.90545	56 Ba 137.327	57 La 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.078	79 Au 196.56655	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.58038	84 Po (209)	85 At (210)	86 Rn (222)	
87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (277)	114 (289) (287)		116 (289)		118 (293)		

58 Ce 140.116	59 Pr 140.50765	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.964	64 Gd 157.25	65 Tb 158.92534	66 Dy 162.50	67 Ho 164.93032	68 Er 167.26	69 Tm 168.93421	70 Yb 173.04	71 Lu 174.967
90 Th 232.0381	91 Pa 231.035888	92 U 238.0289	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Figure 3.5 “A Simple Periodic Table.” View an [accessible periodic table online](#).

There is an easy way to represent isotopes using the atomic symbols. We use the construction:



where X is the symbol of the element, A is the mass number, and Z is the atomic number. Thus, for the isotope of carbon that has 6 protons and 6 neutrons, the symbol is:



where C is the symbol for the element, 6 represents the atomic number, and 12 represents the mass number.

Example 3.10

Problems

1. What is the symbol for an isotope of uranium that has an atomic number of 92 and a mass number of 235?
2. How many protons and neutrons are in ${}_{26}^{56}\text{Fe}$?

Solutions

1. The symbol for this isotope is ${}_{92}^{235}\text{U}$.
2. This iron atom has 26 protons and $56 - 26 = 30$ neutrons.

Test Yourself

How many protons are in ${}_{11}^{23}\text{Na}$?

Answer

11 protons

It is also common to state the mass number after the name of an element to indicate a particular isotope. *Carbon-12* represents an isotope of carbon with 6 protons and 6 neutrons, while *uranium-238* is an isotope of uranium that has 146 neutrons.

Key Takeaways

- Chemistry is based on the modern atomic theory, which states that all matter is composed of atoms.
- Atoms themselves are composed of protons, neutrons, and electrons.
- Each element has its own atomic number, which is equal to the number of protons in its nucleus.
- Isotopes of an element contain different numbers of neutrons.
- Elements are represented by an atomic symbol.
- The periodic table is a chart that organizes all the elements.

Exercises

Questions

1. List the three statements that make up the modern atomic theory.

2. Explain how atoms are composed.
3. Which is larger, a proton or an electron?
4. Which is larger, a neutron or an electron?
5. What are the charges for each of the three subatomic particles?
6. Where is most of the mass of an atom located?
7. Sketch a diagram of a boron atom, which has five protons and six neutrons in its nucleus.
8. Sketch a diagram of a helium atom, which has two protons and two neutrons in its nucleus.
9. Define *atomic number*. What is the atomic number for a boron atom?
10. What is the atomic number of helium?
11. Define *isotope* and give an example.
12. What is the difference between deuterium and tritium?
13. Which pair represents isotopes?
 - a. ${}^4_2\text{He}$ and ${}^3_2\text{He}$
 - b. ${}^{56}_{26}\text{Fe}$ and ${}^{56}_{25}\text{Mn}$
 - c. ${}^{28}_{14}\text{Si}$ and ${}^{31}_{15}\text{P}$
14. Which pair represents isotopes?
 - a. ${}^{40}_{20}\text{Ca}$ and ${}^{40}_{19}\text{K}$
 - b. ${}^{56}_{26}\text{Fe}$ and ${}^{58}_{26}\text{Fe}$
 - c. ${}^{238}_{92}\text{U}$ and ${}^{235}_{92}\text{U}$
15. Give complete symbols of each atom, including the atomic number and the mass number.
 - a. an oxygen atom with 8 protons and 8 neutrons
 - b. a potassium atom with 19 protons and 20 neutrons
 - c. a lithium atom with 3 protons and 4 neutrons
16. Give complete symbols of each atom, including the atomic number and the mass number.
 - a. a magnesium atom with 12 protons and 12 neutrons
 - b. a magnesium atom with 12 protons and 13 neutrons
 - c. a xenon atom with 54 protons and 77 neutrons
17. Americium-241 is an isotope used in smoke detectors. What is the complete symbol for this isotope?
18. Carbon-14 is an isotope used to perform radioactive dating tests on previously living material. What is the complete symbol for this isotope?
19. Give atomic symbols for each element.
 - a. sodium
 - b. argon
 - c. nitrogen
 - d. radon
20. Give atomic symbols for each element.
 - a. silver
 - b. gold

- c. mercury
- d. iodine

21. Give the name of the element.

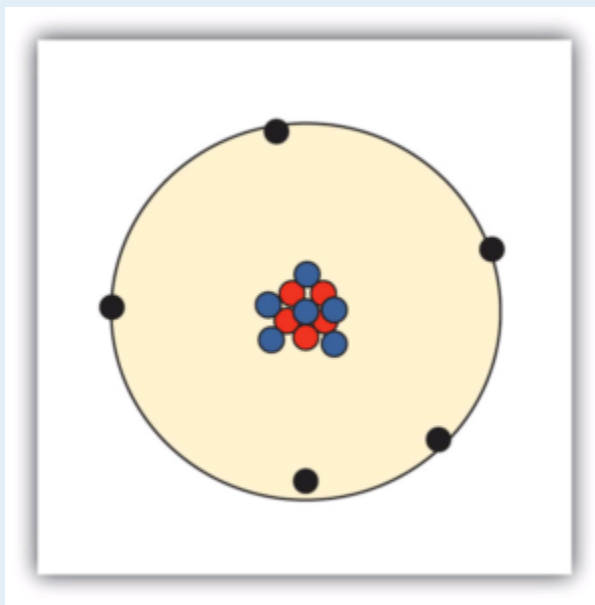
- a. Si
- b. Mn
- c. Fe
- d. Cr

22. Give the name of the element.

- a. F
- b. Cl
- c. Br
- d. I

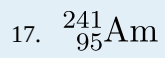
Answers

1. All matter is composed of atoms; atoms of the same element are the same, and atoms of different elements are different; atoms combine in whole-number ratios to form compounds.
3. A proton is larger than an electron.
5. proton: 1+; electron: 1-; neutron: 0



- 7.
9. The atomic number is the number of protons in a nucleus. Boron has an atomic number of five.
11. Isotopes are atoms of the same element but with different numbers of neutrons. ${}^1_1\text{H}$ and ${}^2_1\text{H}$ are examples.
13.
 - a. isotopes
 - b. not isotopes
 - c. not isotopes

15. a. $^{16}_8\text{O}$
 b. $^{39}_{19}\text{K}$
 c. ^7_3Li



19. a. Na
 b. Ar
 c. N
 d. Rn

21. a. silicon
 b. manganese
 c. iron
 d. chromium

Media Attributions

Exercises

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End-of-Chapter Material

Additional Exercises

- How many electrons does it take to make the mass of one proton?
- How many protons does it take to make the mass of a neutron?
- Dalton's initial version of the modern atomic theory says that all atoms of the same element are the same. Is this actually correct? Why or why not?
- How are atoms of the same element the same? How are atoms of the same element different?
- Give complete atomic symbols for the three known isotopes of hydrogen.
- A rare isotope of helium has a single neutron in its nucleus. Write the complete atomic symbol of this isotope.
- Use its place on the periodic table to determine if indium, In, atomic number 49, is a metal or a nonmetal.
- Only a few atoms of astatine, At, atomic number 85, have been detected. On the basis of its position on the periodic table, would you expect it to be a metal or a nonmetal?
- Americium-241 is a crucial part of many smoke detectors. How many neutrons are present in its nucleus?
- Potassium-40 is a radioactive isotope of potassium that is present in the human body. How many neutrons are present in its nucleus?
- Determine the atomic mass of ruthenium from the given abundance and mass data.

Isotope	Abundance	Mass
Ruthenium-96	5.54%	95.907 u
Ruthenium-98	1.87%	97.905 u
Ruthenium-99	12.76%	98.906 u
Ruthenium-100	12.60%	99.904 u
Ruthenium-101	17.06%	100.906 u
Ruthenium-102	31.55%	101.904 u
Ruthenium-104	18.62%	103.905 u

- Determine the atomic mass of tellurium from the given abundance and mass data.

Isotope	Abundance	Mass
Tellurium-120	0.09%	119.904 u
Tellurium-122	2.55%	121.903 u
Tellurium-123	0.89%	122.904 u
Tellurium-124	4.74%	123.903 u
Tellurium-125	7.07%	124.904 u
Tellurium-126	18.84%	125.903 u
Tellurium-128	31.74%	127.904 u
Tellurium-130	34.08%	129.906 u

13. One atomic mass unit has a mass of 1.6605×10^{-24} g. What is the mass of one atom of sodium?
14. One atomic mass unit has a mass of 1.6605×10^{-24} g. What is the mass of one atom of uranium?
15. One atomic mass unit has a mass of 1.6605×10^{-24} g. What is the mass of one molecule of H_2O ?
16. One atomic mass unit has a mass of 1.6605×10^{-24} g. What is the mass of one molecule of PF_5 ?
17. From their positions on the periodic table, will Cu and I form a molecular compound or an ionic compound?
18. From their positions on the periodic table, will N and S form a molecular compound or an ionic compound?
19. Mercury is an unusual element in that when it takes a 1+ charge as a cation, it always exists as the diatomic ion.
 - a. Propose a formula for the mercury(I) ion.
 - b. What is the formula of mercury(I) chloride?
20. Propose a formula for hydrogen peroxide, a substance used as a bleaching agent. (Curiously, this compound does not behave as an acid, despite its formula. It behaves more like a classic nonmetal-nonmetal, molecular compound.)
21. The uranyl cation has the formula UO_2^{2+} . Propose formulas and names for the ionic compounds between the uranyl cation and F^- , SO_4^{2-} , and PO_4^{3-} .
22. The permanganate anion has the formula MnO_4^- . Propose formulas and names for the ionic compounds between the permanganate ion and K^+ , Ca^{2+} , and Fe^{3+} .

Answers

1. About 1,800 electrons
3. It is not strictly correct because of the existence of isotopes.
5. H_{11} , H_{12} , and H_{13}
7. It is a metal.
9. 146 neutrons
11. 101.065 u
13. 3.817×10^{-23} g
15. 2.991×10^{-23} g
17. ionic
19.
 - a. Hg_2^{2+}
 - b. Hg_2Cl_2
21. uranyl fluoride, UO_2F_2 ; uranyl sulfate, UO_2SO_4 ; uranyl phosphate, $(UO_2)_3(PO_4)_2$

Chapter 4. Chemical Reactions and Equations

The space shuttle — and any other rocket-based system — uses chemical reactions to propel itself into space and manoeuvre itself when it gets into orbit. The rockets that lift the orbiter are of two different types. The three main engines are powered by reacting liquid hydrogen with liquid oxygen to generate water. Then there are the two solid rocket boosters, which use a solid fuel mixture that contains mainly ammonium perchlorate and powdered aluminum. The chemical reaction between these substances produces aluminum oxide, water, nitrogen gas, and hydrogen chloride. Although the solid rocket boosters each have a significantly lower mass than the liquid oxygen and liquid hydrogen tanks, they provide over 80% of the lift needed to put the shuttle into orbit — all because of chemical reactions.



Figure 4.0 “Delta IV Medium Rocket DSCS.”

Chemistry is largely about chemical changes. Indeed, if there were no chemical changes, chemistry as such would not exist! Chemical changes are a fundamental part of chemistry. Because chemical changes are so central, it may be no surprise that chemistry has developed some special ways of presenting them.

Media Attributions

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The Chemical Equation

Learning Objectives

1. Define *chemical equation*.
2. Identify the parts of a chemical equation.

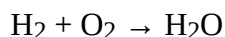
A chemical reaction expresses a chemical change. For example, one chemical property of hydrogen is that it will react with oxygen to make water. We can write that as follows:

hydrogen reacts with oxygen to make water

We can represent this chemical change more succinctly as:

hydrogen + oxygen → water

where the + sign means that the two substances interact chemically with each other and the → symbol implies that a chemical reaction takes place. But substances can also be represented by chemical formulas. Remembering that hydrogen and oxygen both exist as diatomic molecules, we can rewrite our chemical change as:

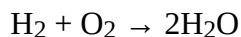


This is an example of a **chemical equation**, which is a concise way of representing a chemical reaction. The initial substances are called **reactants**, and the final substances are called **products**.

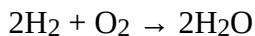
Unfortunately, it is also an *incomplete* chemical equation. The law of conservation of matter says that matter cannot be created or destroyed. In chemical equations, the number of atoms of each element in the reactants must be the same as the number of atoms of each element in the products. If we count the number of hydrogen atoms in the reactants and products, we find two hydrogen atoms. But if we count the number of oxygen atoms in the reactants and products, we find that there are two oxygen atoms in the reactants but only one oxygen atom in the products.

What can we do? Can we change the subscripts in the formula for water so that it has two oxygen atoms in it? No; you *cannot* change the formulas of individual substances because the chemical formula for a given substance is characteristic of that substance. What you *can* do, however, is to change the number of molecules that react or are produced. We do this one element at a time, going from one side of the reaction to the other, changing the number of molecules of a substance until all elements have the same number of atoms on each side.

To accommodate the two oxygen atoms as reactants, let us assume that we have two water molecules as products:



The 2 in front of the formula for water is called a **coefficient**. Now there is the same number of oxygen atoms in the reactants as there are in the product. But in satisfying the need for the same number of oxygen atoms on both sides of the reaction, we have also changed the number of hydrogen atoms on the product side, so the number of hydrogen atoms is no longer equal. No problem—simply go back to the reactant side of the equation and add a coefficient in front of the H_2 . The coefficient that works is 2:



There are now four hydrogen atoms in the reactants and also four atoms of hydrogen in the product. There are two oxygen atoms in the reactants and two atoms of oxygen in the product. The law of conservation of matter has been satisfied. When the reactants and products of a chemical equation have the same number of atoms of all elements present, we say that an equation is **balanced**. All proper chemical equations are balanced. If a substance does not have a coefficient written in front of it, it is assumed to be 1. Also, the convention is to use all whole numbers when balancing chemical equations. This sometimes makes us do a bit more “back and forth” work when balancing a chemical equation.

Example 4.1

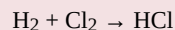
Problems

Write and balance the chemical equation for each given chemical reaction.

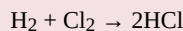
1. Hydrogen and chlorine react to make HCl.
2. Ethane, C_2H_6 , reacts with oxygen to make carbon dioxide and water.

Solutions

1. Let us start by simply writing a chemical equation in terms of the formulas of the substances, remembering that both elemental hydrogen and chlorine are diatomic:



There are two hydrogen atoms and two chlorine atoms in the reactants and one of each atom in the product. We can fix this by including the coefficient 2 on the product side:

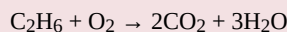


Now there are two hydrogen atoms and two chlorine atoms on both sides of the chemical equation, so it is balanced.

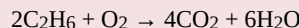
2. Start by writing the chemical equation in terms of the substances involved:



We have two carbon atoms on the left, so we need two carbon dioxide molecules on the product side, so that each side has two carbon atoms; that element is balanced. We have six hydrogen atoms in the reactants, so we need six hydrogen atoms in the products. We can get this by having three water molecules:



Now we have seven oxygen atoms in the products (four from the CO_2 and three from the H_2O). That means we need seven oxygen atoms in the reactants. However, because oxygen is a diatomic molecule, we can only get an even number of oxygen atoms at a time. We can achieve this by multiplying the other coefficients by 2:



By multiplying everything else by 2, we don't unbalance the other elements, and we now get an even number of oxygen atoms in the product—14. We can get 14 oxygen atoms on the reactant side by having 7 oxygen molecules:

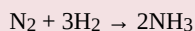


As a check, recount everything to determine that each side has the same number of atoms of each element. This chemical equation is now balanced.

Test Yourself

Write and balance the chemical equation that represents nitrogen and hydrogen reacting to produce ammonia, NH_3 .

Answer



Many chemical equations also include phase labels for the substances: (s) for solid, (ℓ) for liquid, (g) for gas, and (aq) for aqueous (i.e., dissolved in water). Special conditions, such as temperature, may also be listed above the arrow. For example:



Key Takeaways

- A chemical equation is a concise description of a chemical reaction.
- Proper chemical equations are balanced.

Exercises

Questions

1. From the statement “nitrogen and hydrogen react to produce ammonia,” identify the reactants and the products.
2. From the statement “sodium metal reacts with water to produce sodium hydroxide and hydrogen,” identify the reactants and the products.
3. From the statement “magnesium hydroxide reacts with nitric acid to produce magnesium nitrate and water,” identify the reactants and the products.
4. From the statement “propane reacts with oxygen to produce carbon dioxide and water,” identify the reactants and the products.
5. Write and balance the chemical equation described by Exercise 1.
6. Write and balance the chemical equation described by Exercise 2.
7. Write and balance the chemical equation described by Exercise 3.

8. Write and balance the chemical equation described by Exercise 4. The formula for propane is C_3H_8 .
9. Balance: $__NaClO_3 \rightarrow __NaCl + __O_2$
10. Balance: $__N_2 + __H_2 \rightarrow __N_2H_4$
11. Balance: $__Al + __O_2 \rightarrow __Al_2O_3$
12. Balance: $__C_2H_4 + __O_2 \rightarrow __CO_2 + __H_2O$
13. How would you write the balanced chemical equation in Exercise 10 if all substances were gases?
14. How would you write the balanced chemical equation in Exercise 12 if all the substances except water were gases and water itself were a liquid?

Answers

1. reactants: nitrogen and hydrogen; product: ammonia
3. reactants: magnesium hydroxide and nitric acid; products: magnesium nitrate and water
5. $N_2 + 3H_2 \rightarrow 2NH_3$
7. $Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$
9. $2NaClO_3 \rightarrow 2NaCl + 3O_2$
11. $4Al + 3O_2 \rightarrow 2Al_2O_3$
13. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

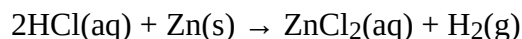
Types of Chemical Reactions: Single- and Double-Displacement Reactions

Learning Objectives

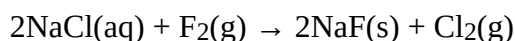
1. Recognize chemical reactions as single-replacement reactions and double-replacement reactions.
2. Use the periodic table, an activity series, or solubility rules to predict whether single-replacement reactions or double-replacement reactions will occur.

Up to now, we have presented chemical reactions as a topic, but we have not discussed how the products of a chemical reaction can be predicted. Here we will begin our study of certain types of chemical reactions that allow us to predict what the products of the reaction will be.

A **single-replacement reaction** is a chemical reaction in which one element is substituted for another element in a compound, generating a new element and a new compound as products. For example:

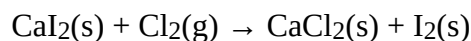


is an example of a single-replacement reaction. The hydrogen atoms in HCl are replaced by Zn atoms, and in the process a new element—hydrogen—is formed. Another example of a single-replacement reaction is:

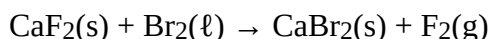


Here the negatively charged ion changes from chloride to fluoride. A typical characteristic of a single-replacement reaction is that there is one element as a reactant and another element as a product.

Not all proposed single-replacement reactions will occur between two given reactants. This is most easily demonstrated with fluorine, chlorine, bromine, and iodine. Collectively, these elements are called the *halogens* and are in the next-to-last column on the periodic table (see Figure 4.1 “Halogens on the Periodic Table”). The elements on top of the column will replace the elements below them on the periodic table but not the other way around. Thus, the reaction represented by:



This reaction will occur, but the reaction



will not because bromine is below fluorine on the periodic table. This is just one of many ways the periodic table helps us understand chemistry.

Chemical reactivity trends are easy to predict when replacing anions in simple ionic compounds—simply use their relative positions on the periodic table. However, when replacing the cations, the trends are not as straightforward. This is partly because there are so many elements that can form cations; an element in one column on the periodic table may replace another element nearby, or it may not. A list called the **activity series** does the same thing the periodic table does for halogens: it lists the elements that will replace elements below them in single-replacement reactions. A simple activity series is shown below.

Activity Series for Cation Replacement in Single-Replacement Reactions

- Li
- K
- Ba
- Sr
- Ca
- Na
- Mg
- Al
- Mn
- Zn
- Cr
- Fe
- Ni
- Sn
- Pb
- H₂
- Cu
- Hg
- Ag
- Pd
- Pt
- Au

Using the activity series is similar to using the positions of the halogens on the periodic table. An element on top will replace an element below it in compounds undergoing a single-replacement reaction. Elements will not replace elements above them in compounds.

Example 4.3

Problems

Use the activity series to predict the products, if any, of each equation.

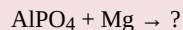
1. $\text{FeCl}_2 + \text{Zn} \rightarrow ?$
2. $\text{HNO}_3 + \text{Au} \rightarrow ?$

Solutions

1. Because zinc is above iron in the activity series, it will replace iron in the compound. The products of this single-replacement reaction are ZnCl_2 and Fe.
2. Gold is below hydrogen in the activity series. As such, it will not replace hydrogen in a compound with the nitrate ion. No reaction is predicted.

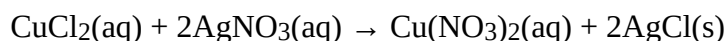
Test Yourself

Use the activity series to predict the products, if any, of this equation.

**Answer**

$\text{Mg}_3(\text{PO}_4)_2$ and Al

A **double-replacement reaction** occurs when parts of two ionic compounds are exchanged, making two new compounds. A characteristic of a double-replacement equation is that there are two compounds as reactants and two different compounds as products. An example is:



There are two equivalent ways of considering a double-replacement equation: either the cations are swapped, or the anions are swapped. (You cannot swap both; you would end up with the same substances you started with.) Either perspective should allow you to predict the proper products, as long as you pair a cation with an anion and not a cation with a cation or an anion with an anion.

Example 4.4

Problem

Predict the products of this double-replacement equation: $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow ?$

Solution

Thinking about the reaction as either switching the cations or switching the anions, we would expect the products to be BaSO_4 and NaCl .

Test Yourself

Predict the products of this double-replacement equation: $\text{KBr} + \text{AgNO}_3 \rightarrow ?$

Answer

KNO_3 and AgBr

Predicting whether a double-replacement reaction occurs is somewhat more difficult than predicting a single-replacement reaction. However, there is one type of double-replacement reaction that we can predict: the precipitation reaction. A **precipitation reaction** occurs when two ionic compounds are dissolved in water and form a new ionic compound that does not dissolve; this new compound falls out of solution as a solid **precipitate**. The formation of a solid precipitate is the driving force that makes the reaction proceed.

To judge whether double-replacement reactions will occur, we need to know what kinds of ionic compounds form precipitates. For this, we use **solubility rules**, which are general statements that predict which ionic compounds dissolve (are soluble) and which do not (are not soluble or insoluble). Tables 4.1a and b “Some Useful Solubility Rules” list some general solubility rules. We need to consider each ionic compound (both the reactants and the possible products) in light of the solubility rules in Table 4.1a and b. If a compound is soluble, we use the (aq) label with it, indicating it dissolves. If a compound is not soluble, we use the (s) label with it and assume that it will precipitate out of solution. If everything is soluble, then no reaction will be expected.

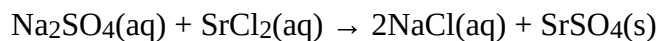
Table 4.1a Some Useful Solubility Rules

These compounds generally dissolve in water (are soluble):	Exceptions:
All compounds of Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , and NH_4^+	None
All compounds of NO_3^- and $\text{C}_2\text{H}_3\text{O}_2^-$	None
Compounds of Cl^- , Br^- , I^-	Ag^+ , Hg_2^{2+} , Pb^{2+}
Compounds of SO_4^{2-}	Hg_2^{2+} , Pb^{2+} , Sr^{2+} , Ba^{2+}

Table 4.1b Some More Useful Solubility Rules

These compounds generally do not dissolve in water (are insoluble):	Exceptions:
Compounds of CO_3^{2-} and PO_4^{3-}	Compounds of Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , and NH_4^+
Compounds of OH^-	Compounds of Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , NH_4^+ , Sr^{2+} , and Ba^{2+}

For example, consider the possible double-replacement reaction between Na_2SO_4 and SrCl_2 . The solubility rules say that all ionic sodium compounds are soluble and all ionic chloride compounds are soluble except for Ag^+ , Hg_2^{2+} , and Pb^{2+} , which are not being considered here. Therefore, Na_2SO_4 and SrCl_2 are both soluble. The possible double-replacement reaction products are NaCl and SrSO_4 . Are these soluble? NaCl is (by the same rule we just quoted), but what about SrSO_4 ? Compounds of the sulfate ion are generally soluble, but Sr^{2+} is an exception: we expect it to be insoluble—a precipitate. Therefore, we expect a reaction to occur, and the balanced chemical equation would be:



You would expect to see a visual change corresponding to SrSO_4 precipitating out of solution (Figure 4.2 “Double-Replacement Reactions”).



Figure 4.2 “Double-Replacement Reactions.” Some double-replacement reactions are obvious because you can see a solid precipitate coming out of the solution.

Example 4.5

Problems

Will a double-replacement reaction occur? If so, identify the products.

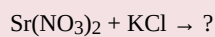
1. $\text{Ca}(\text{NO}_3)_2 + \text{KBr} \rightarrow ?$
2. $\text{NaOH} + \text{FeCl}_2 \rightarrow ?$

Solutions

1. According to the solubility rules, both $\text{Ca}(\text{NO}_3)_2$ and KBr are soluble. Now we consider what the double-replacement products would be by switching the cations (or the anions)—namely, CaBr_2 and KNO_3 . However, the solubility rules predict that these two substances would also be soluble, so no precipitate would form. Thus, we predict no reaction in this case.
2. According to the solubility rules, both NaOH and FeCl_2 are expected to be soluble. If we assume that a double-replacement reaction may occur, we need to consider the possible products, which would be NaCl and $\text{Fe}(\text{OH})_2$. NaCl is soluble, but, according to the solubility rules, $\text{Fe}(\text{OH})_2$ is not. Therefore, a reaction would occur, and $\text{Fe}(\text{OH})_2(\text{s})$ would precipitate out of solution. The balanced chemical equation is: $2\text{NaOH}(\text{aq}) + \text{FeCl}_2(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{Fe}(\text{OH})_2(\text{s})$

Test Yourself

Will a double-replacement equation occur? If so, identify the products.

**Answer**

No reaction; all possible products are soluble.

Key Takeaways

- A single-replacement reaction replaces one element for another in a compound.
- The periodic table or an activity series can help predict whether single-replacement reactions occur.
- A double-replacement reaction exchanges the cations (or the anions) of two ionic compounds.
- A precipitation reaction is a double-replacement reaction in which one product is a solid precipitate.
- Solubility rules are used to predict whether some double-replacement reactions will occur.

Exercises

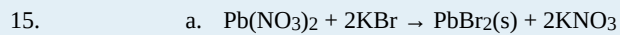
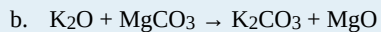
Questions

1. What are the general characteristics that help you recognize single-replacement reactions?
2. What are the general characteristics that help you recognize double-replacement reactions?
3. Assuming that each single-replacement reaction occurs, predict the products and write each balanced chemical equation.
 - a. $\text{Zn} + \text{Fe}(\text{NO}_3)_2 \rightarrow ?$
 - b. $\text{F}_2 + \text{FeI}_3 \rightarrow ?$
4. Assuming that each single-replacement reaction occurs, predict the products and write each balanced chemical equation.
 - a. $\text{Li} + \text{MgSO}_4 \rightarrow ?$
 - b. $\text{NaBr} + \text{Cl}_2 \rightarrow ?$
5. Assuming that each single-replacement reaction occurs, predict the products and write each balanced chemical equation.
 - a. $\text{Sn} + \text{H}_2\text{SO}_4 \rightarrow ?$
 - b. $\text{Al} + \text{NiBr}_2 \rightarrow ?$
6. Assuming that each single-replacement reaction occurs, predict the products and write each balanced chemical equation.
 - a. $\text{Mg} + \text{HCl} \rightarrow ?$
 - b. $\text{HI} + \text{Br}_2 \rightarrow ?$
7. Use the periodic table or the activity series to predict if each single-replacement reaction will occur and, if so, write a balanced chemical equation.
 - a. $\text{FeCl}_2 + \text{Br}_2 \rightarrow ?$
 - b. $\text{Fe}(\text{NO}_3)_3 + \text{Al} \rightarrow ?$
8. Use the periodic table or the activity series to predict if each single-replacement reaction will occur and, if so, write a balanced chemical equation.
 - a. $\text{Zn} + \text{Fe}_3(\text{PO}_4)_2 \rightarrow ?$
 - b. $\text{Ag} + \text{HNO}_3 \rightarrow ?$
9. Use the periodic table or the activity series to predict if each single-replacement reaction will occur and, if so, write a balanced chemical equation.
 - a. $\text{NaI} + \text{Cl}_2 \rightarrow ?$
 - b. $\text{AgCl} + \text{Au} \rightarrow ?$
10. Use the periodic table or the activity series to predict if each single-replacement reaction will occur and, if so, write a balanced chemical equation.
 - a. $\text{Pt} + \text{H}_3\text{PO}_4 \rightarrow ?$
 - b. $\text{Li} + \text{H}_2\text{O} \rightarrow ?$ (Hint: treat H_2O as if it were composed of H^+ and OH^- ions.)
11. Assuming that each double-replacement reaction occurs, predict the products and write each balanced chemical equation.
 - a. $\text{Zn}(\text{NO}_3)_2 + \text{NaOH} \rightarrow ?$
 - b. $\text{HCl} + \text{Na}_2\text{S} \rightarrow ?$
12. Assuming that each double-replacement reaction occurs, predict the products and write each balanced chemical equation.

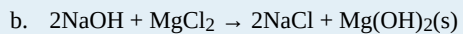
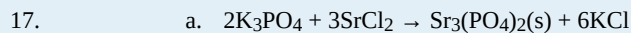
- a. $\text{Ca}(\text{C}_2\text{H}_3\text{O}_2)_2 + \text{HNO}_3 \rightarrow ?$
 b. $\text{Na}_2\text{CO}_3 + \text{Sr}(\text{NO}_2)_2 \rightarrow ?$
13. Assuming that each double-replacement reaction occurs, predict the products and write each balanced chemical equation.
- a. $\text{Pb}(\text{NO}_3)_2 + \text{KBr} \rightarrow ?$
 b. $\text{K}_2\text{O} + \text{MgCO}_3 \rightarrow ?$
14. Assuming that each double-replacement reaction occurs, predict the products and write each balanced chemical equation.
- a. $\text{Sn}(\text{OH})_2 + \text{FeBr}_3 \rightarrow ?$
 b. $\text{CsNO}_3 + \text{KCl} \rightarrow ?$
15. Use the solubility rules to predict if each double-replacement reaction will occur and, if so, write a balanced chemical equation.
- a. $\text{Pb}(\text{NO}_3)_2 + \text{KBr} \rightarrow ?$
 b. $\text{K}_2\text{O} + \text{Na}_2\text{CO}_3 \rightarrow ?$
16. Use the solubility rules to predict if each double-replacement reaction will occur and, if so, write a balanced chemical equation.
- a. $\text{Na}_2\text{CO}_3 + \text{Sr}(\text{NO}_2)_2 \rightarrow ?$
 b. $(\text{NH}_4)_2\text{SO}_4 + \text{Ba}(\text{NO}_3)_2 \rightarrow ?$
17. Use the solubility rules to predict if each double-replacement reaction will occur and, if so, write a balanced chemical equation.
- a. $\text{K}_3\text{PO}_4 + \text{SrCl}_2 \rightarrow ?$
 b. $\text{NaOH} + \text{MgCl}_2 \rightarrow ?$
18. Use the solubility rules to predict if each double-replacement reaction will occur and, if so, write a balanced chemical equation.
- a. $\text{KC}_2\text{H}_3\text{O}_2 + \text{Li}_2\text{CO}_3 \rightarrow ?$
 b. $\text{KOH} + \text{AgNO}_3 \rightarrow ?$

Answers

1. One element replaces another element in a compound.
3. a. $\text{Zn} + \text{Fe}(\text{NO}_3)_2 \rightarrow \text{Zn}(\text{NO}_3)_2 + \text{Fe}$
 b. $3\text{F}_2 + 2\text{FeI}_3 \rightarrow 3\text{I}_2 + 2\text{FeF}_3$
5. a. $\text{Sn} + \text{H}_2\text{SO}_4 \rightarrow \text{SnSO}_4 + \text{H}_2$
 b. $2\text{Al} + 3\text{NiBr}_2 \rightarrow 2\text{AlBr}_3 + 3\text{Ni}$
7. a. No reaction occurs.
 b. $\text{Fe}(\text{NO}_3)_3 + \text{Al} \rightarrow \text{Al}(\text{NO}_3)_3 + \text{Fe}$
9. a. $2\text{NaI} + \text{Cl}_2 \rightarrow 2\text{NaCl} + \text{I}_2$
 b. No reaction occurs.
11. a. $\text{Zn}(\text{NO}_3)_2 + 2\text{NaOH} \rightarrow \text{Zn}(\text{OH})_2 + 2\text{NaNO}_3$
 b. $2\text{HCl} + \text{Na}_2\text{S} \rightarrow 2\text{NaCl} + \text{H}_2\text{S}$
13. a. $\text{Pb}(\text{NO}_3)_2 + 2\text{KBr} \rightarrow \text{PbBr}_2 + 2\text{KNO}_3$



b. No reaction occurs.



Media Attributions

Figure 4.1

- “Halogens on the Periodic Table” by David W. Ball © [CC BY-NC-SA \(Attribution NonCommercial ShareAlike\)](#)

Figure 4.2

- [“Copper solution”](#) by Choij © [Public Domain](#)

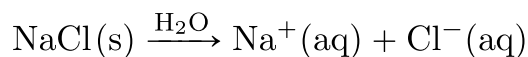
Ionic Equations: A Closer Look

Learning Objectives

1. Write ionic equations for chemical reactions between ionic compounds.
2. Write net ionic equations for chemical reactions between ionic compounds.

For single-replacement and double-replacement reactions, many of the reactions included ionic compounds: compounds between metals and nonmetals or compounds that contained recognizable polyatomic ions. Now we take a closer look at reactions that include ionic compounds.

One important aspect about ionic compounds that differs from molecular compounds has to do with dissolving in a liquid, such as water. When molecular compounds, such as sugar, dissolve in water, the individual molecules drift apart from each other. When ionic compounds dissolve, *the ions physically separate from each other*. We can use a chemical equation to represent this process—for example, with NaCl:



When NaCl dissolves in water, the ions separate and go their own way in solution; the ions are now written with their respective charges, and the (aq) phase label emphasizes that they are dissolved (Figure 4.3 “Ionic Solutions”). This process is called dissociation; we say that the ions *dissociate*.

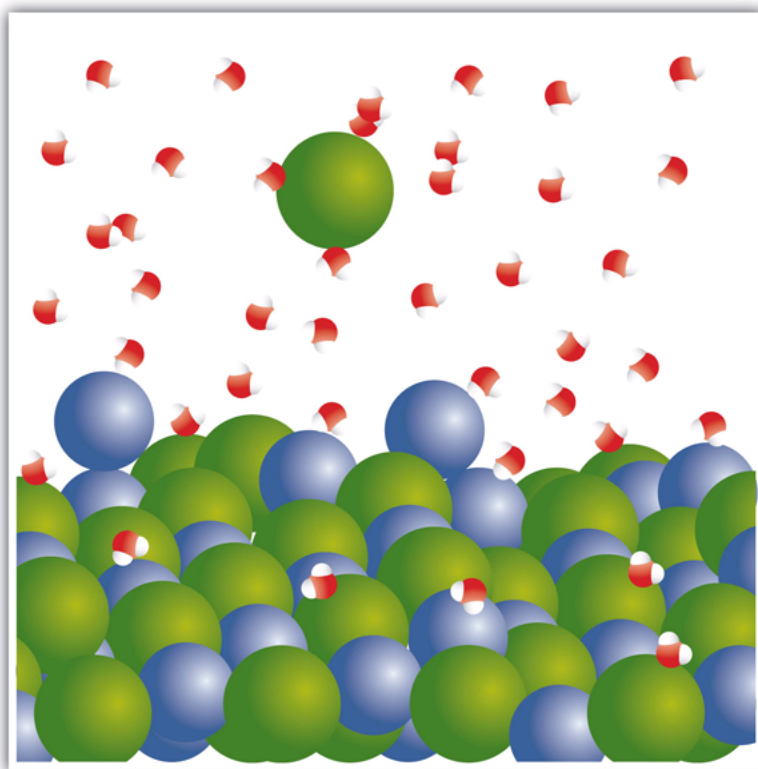
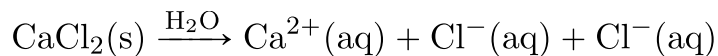
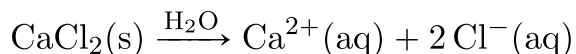


Figure 4.3 “Ionic Solutions.” When an ionic compound dissociates in water, water molecules surround each ion and separate it from the rest of the solid. Each ion goes its own way in solution.

All ionic compounds that dissolve behave this way. (This behaviour was first suggested by the Swedish chemist Svante August Arrhenius (1859–1927) as part of his PhD dissertation in 1884. Interestingly, his PhD examination team had a hard time believing that ionic compounds would behave like this, so they gave Arrhenius a barely passing grade. Later, this work was cited when Arrhenius was awarded the Nobel Prize in Chemistry.) Keep in mind that when the ions separate, *all* the ions separate. Thus, when CaCl_2 dissolves, the one Ca^{2+} ion and the two Cl^- ions separate from each other:



or



That is, the two chloride ions go off on their own. They do not remain as Cl_2 (that would be elemental chlorine; these are chloride ions); they do not stick together to make Cl_2^{-} or Cl_2^{2-} . They become dissociated ions in their own right. Polyatomic ions also retain their overall identity when they are dissolved.

Example 4.6

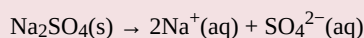
Problems

Write the chemical equation that represents the dissociation of each ionic compound.

1. KBr
2. Na_2SO_4

Solutions

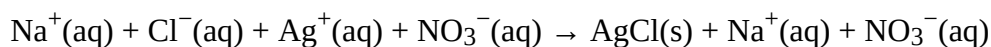
1. $\text{KBr}(\text{s}) \rightarrow \text{K}^{+}(\text{aq}) + \text{Br}^{-}(\text{aq})$
2. Not only do the two sodium ions go their own way, but the sulfate ion stays together as the sulfate ion. The dissolving equation is:

**Test Yourself**

Write the chemical equation that represents the dissociation of $(\text{NH}_4)_2\text{S}$.

Answer

When chemicals in solution react, the proper way of writing the chemical formulas of the dissolved ionic compounds is in terms of the dissociated ions, not the complete ionic formula. A **complete ionic equation** is a chemical equation in which the dissolved ionic compounds are written as separated ions. Solubility rules are very useful in determining which ionic compounds are dissolved and which are not. For example, when $\text{NaCl}(\text{aq})$ reacts with $\text{AgNO}_3(\text{aq})$ in a double-replacement reaction to precipitate $\text{AgCl}(\text{s})$ and form $\text{NaNO}_3(\text{aq})$, the complete ionic equation includes NaCl , AgNO_3 , and NaNO_3 written as separated ions:



This is more representative of what is occurring in the solution.

Example 4.7**Problems**

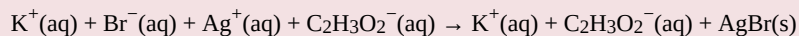
Write the complete ionic equation for each chemical reaction.

1. $\text{KBr}(\text{aq}) + \text{AgC}_2\text{H}_3\text{O}_2(\text{aq}) \rightarrow \text{KC}_2\text{H}_3\text{O}_2(\text{aq}) + \text{AgBr}(\text{s})$
2. $\text{MgSO}_4(\text{aq}) + \text{Ba}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{Mg}(\text{NO}_3)_2(\text{aq}) + \text{BaSO}_4(\text{s})$

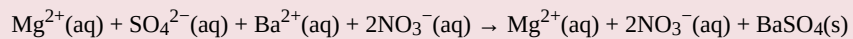
Solutions

For any ionic compound that is aqueous, we will write the compound as separated ions.

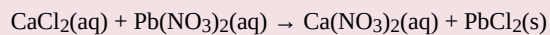
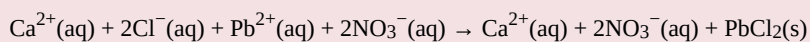
1. The complete ionic equation is:



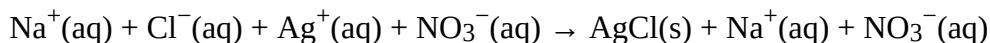
2. The complete ionic equation is:

**Test Yourself**

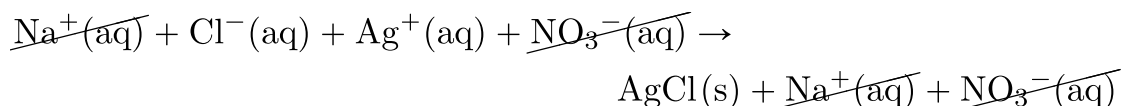
Write the complete ionic equation for:

**Answer**

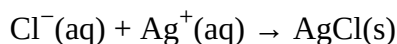
You may notice that in a complete ionic equation, some ions do not change their chemical form; they stay exactly the same on the reactant and product sides of the equation. For example, in



the $\text{Ag}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ ions become $\text{AgCl}(\text{s})$, but the $\text{Na}^+(\text{aq})$ ions and the $\text{NO}_3^-(\text{aq})$ ions stay as $\text{Na}^+(\text{aq})$ ions and $\text{NO}_3^-(\text{aq})$ ions. These two ions are examples of **spectator ions**, ions that do nothing in the overall course of a chemical reaction. They are present, but they do not participate in the overall chemistry. It is common to cancel spectator ions (something also done with algebraic quantities) on the opposite sides of a chemical equation:



What remains when the spectator ions are removed is called the **net ionic equation**, which represents the actual chemical change occurring between the ionic compounds:



It is important to reiterate that the spectator ions are still present in solution, but they don't experience any net chemical change, so they are not written in a net ionic equation.

Example 4.8

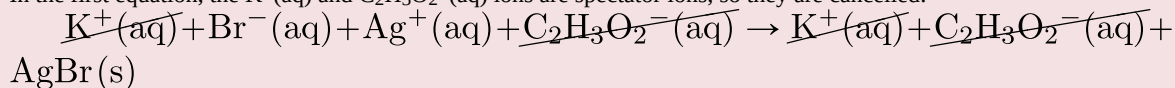
Problems

Write the net ionic equation for each chemical reaction.

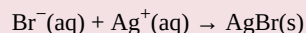
- $\text{K}^+(\text{aq}) + \text{Br}^-(\text{aq}) + \text{Ag}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{K}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) + \text{AgBr}(\text{s})$
- $\text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{BaSO}_4(\text{s})$

Solutions

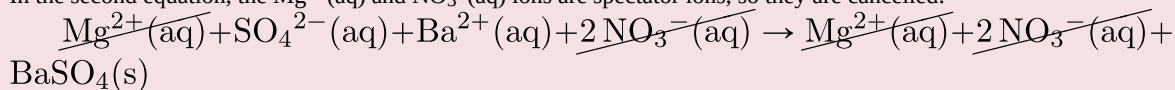
- In the first equation, the $\text{K}^+(\text{aq})$ and $\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$ ions are spectator ions, so they are cancelled:



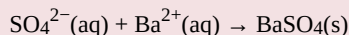
The net ionic equation is:



- In the second equation, the $\text{Mg}^{2+}(\text{aq})$ and $\text{NO}_3^-(\text{aq})$ ions are spectator ions, so they are cancelled:

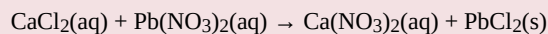


The net ionic equation is:

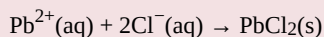


Test Yourself

Write the net ionic equation for:



Answer



Chemistry Is Everywhere: Soluble and Insoluble Ionic Compounds

The concept of solubility versus insolubility in ionic compounds is a matter of degree. Some ionic compounds are very soluble, some are only moderately soluble, and some are soluble so little that they are considered insoluble. For most ionic compounds, there is also a limit to the amount of compound can be dissolved in a sample of water. For example, you can dissolve a maximum of 36.0 g of NaCl in 100 g of water at room temperature, but you can dissolve only 0.00019 g of AgCl in 100 g of water. We consider NaCl soluble but AgCl insoluble.

One place where solubility is important is in the tank-type water heater found in many homes in the United States. Domestic water frequently contains small amounts of dissolved ionic compounds, including calcium carbonate (CaCO_3). However, CaCO_3 has the relatively unusual property of being less soluble in hot water than in cold water. So as the water heater operates by heating water, CaCO_3 can precipitate if there is enough of it in the water. This precipitate, called *limescale*, can also contain magnesium compounds, hydrogen carbonate compounds, and phosphate compounds. The problem is that too much limescale can impede the function of a water heater, requiring more energy to heat water to a specific temperature or even blocking water pipes into or out of the water heater, causing dysfunction.

Another place where solubility versus insolubility is an issue is the Grand Canyon. We usually think of rock as insoluble. But it is actually ever so slightly soluble. This means that over a period of about two billion years, the Colorado River carved rock from the surface by slowly dissolving it, eventually generating a spectacular series of gorges and canyons. And all because of solubility!

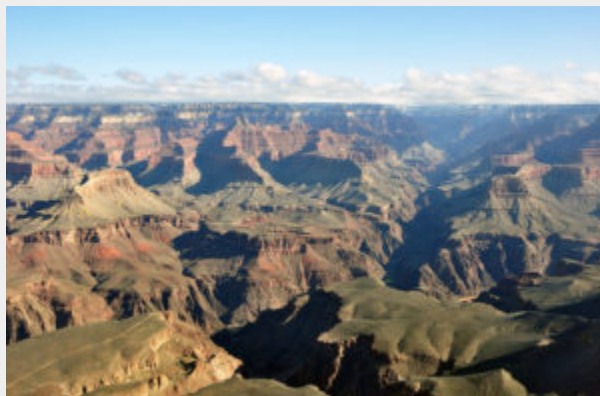


Figure 4.4 “The Grand Canyon.” The Grand Canyon was formed by water running through rock for billions of years, very slowly dissolving it. Note the Colorado River is still present in the lower part of the photo.

Key Takeaways

- Ionic compounds that dissolve separate into individual ions.
- Complete ionic equations show dissolved ionic solids as separated ions.
- Net ionic equations show only the ions and other substances that change in a chemical reaction.

Exercises

Questions

1. Write a chemical equation that represents NaBr(s) dissociating in water.
2. Write a chemical equation that represents SrCl₂(s) dissociating in water.
3. Write a chemical equation that represents (NH₄)₃PO₄(s) dissociating in water.
4. Write a chemical equation that represents Fe(C₂H₃O₂)₃(s) dissociating in water.
5. Write the complete ionic equation for the reaction of FeCl₂(aq) and AgNO₃(aq). You may have to consult the solubility rules.
6. Write the complete ionic equation for the reaction of BaCl₂(aq) and Na₂SO₄(aq). You may have to consult the solubility rules.
7. Write the complete ionic equation for the reaction of KCl(aq) and NaC₂H₃O₂(aq). You may have to consult the solubility rules.
8. Write the complete ionic equation for the reaction of Fe₂(SO₄)₃(aq) and Sr(NO₃)₂(aq). You may have to consult the solubility rules.
9. Write the net ionic equation for the reaction of FeCl₂(aq) and AgNO₃(aq). You may have to consult the solubility rules.
10. Write the net ionic equation for the reaction of BaCl₂(aq) and Na₂SO₄(aq). You may have to consult the solubility rules.
11. Write the net ionic equation for the reaction of KCl(aq) and NaC₂H₃O₂(aq). You may have to consult the solubility rules.
12. Write the net ionic equation for the reaction of Fe₂(SO₄)₃(aq) and Sr(NO₃)₂(aq). You may have to consult the solubility rules.
13. Identify the spectator ions in Exercises 9 and 10.
14. Identify the spectator ions in Exercises 11 and 12.

Answers

1. $\text{NaBr(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{Br}^-(\text{aq})$
3. $(\text{NH}_4)_3\text{PO}_4(\text{s}) \rightarrow 3\text{NH}_4^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$
5. $\text{Fe}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) + 2\text{Ag}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{AgCl(s)}$
7. $\text{K}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{K}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$
9. $2\text{Cl}^-(\text{aq}) + 2\text{Ag}^+(\text{aq}) \rightarrow 2\text{AgCl(s)}$
11. There is no overall reaction.
13. In Exercise 9, Fe²⁺(aq) and NO₃⁻(aq) are spectator ions; in Exercise 10, Na⁺(aq) and Cl⁻(aq) are spectator ions.

Media Attributions

Figure 4.3

- “Ionic Solutions” by David W. Ball © [CC BY-NC-SA \(Attribution NonCommercial ShareAlike\)](#)

Figure 4.4

- [“Grand Canyon, Yavapai Point, 2010”](#) by [Chensiyuan](#) © [CC BY-SA \(Attribution ShareAlike\)](#)

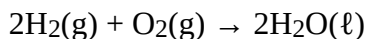
Composition, Decomposition, and Combustion Reactions

Learning Objectives

1. Recognize composition, decomposition, and combustion reactions.
2. Predict the products of a combustion reaction.

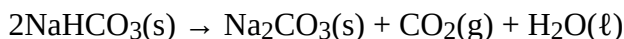
Three classifications of chemical reactions will be reviewed in this section. Predicting the products in some of them may be difficult, but the reactions are still easy to recognize.

A **composition reaction** (sometimes also called a *combination reaction* or a *synthesis reaction*) produces a single substance from multiple reactants. A single substance as a product is the key characteristic of the composition reaction. There may be a coefficient other than one for the substance, but if the reaction has only a single substance as a product, it can be called a composition reaction. In the reaction



water is produced from hydrogen and oxygen. Although there are two molecules of water being produced, there is only one substance—water—as a product. So this is a composition reaction.

A **decomposition reaction** starts from a single substance and produces more than one substance; that is, it decomposes. One substance as a reactant and more than one substance as the products is the key characteristic of a decomposition reaction. For example, if we look at the decomposition of sodium hydrogen carbonate (also known as sodium bicarbonate):



We can see that sodium carbonate, carbon dioxide, and water are produced from the single substance sodium hydrogen carbonate.

Composition and decomposition reactions are difficult to predict; however, they should be easy to recognize.

Example 4.9

Problems

Identify each equation as a composition reaction, a decomposition reaction, or neither.

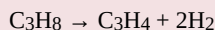
1. $\text{Fe}_2\text{O}_3 + 3\text{SO}_3 \rightarrow \text{Fe}_2(\text{SO}_4)_3$
2. $\text{NaCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{NaNO}_3$
3. $(\text{NH}_4)_2\text{Cr}_2\text{O}_7 \rightarrow \text{Cr}_2\text{O}_3 + 4\text{H}_2\text{O} + \text{N}_2$

Solutions

1. In this equation, two substances combine to make a single substance. This is a composition reaction.
2. Two different substances react to make two new substances. This does not fit the definition of either a composition reaction or a decomposition reaction, so it is neither. In fact, you may recognize this as a double-replacement reaction.
3. A single substance reacts to make multiple substances. This is a decomposition reaction.

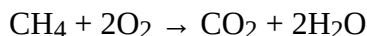
Test Yourself

Identify the equation as a composition reaction, a decomposition reaction, or neither.

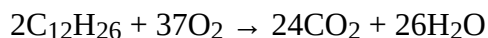
**Answer**

Decomposition.

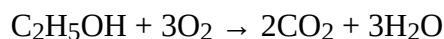
A combustion reaction occurs when a reactant combines with oxygen, many times from the atmosphere, to produce oxides of all other elements as products; any nitrogen in the reactant is converted to elemental nitrogen, N_2 . Many reactants, called *fuels*, contain mostly carbon and hydrogen atoms, reacting with oxygen to produce CO_2 and H_2O . For example, the balanced chemical equation for the combustion of methane, CH_4 , is as follows:



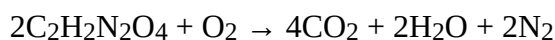
Kerosene can be approximated with the formula $\text{C}_{12}\text{H}_{26}$, and its combustion equation is



Sometimes fuels contain oxygen atoms, which must be counted when balancing the chemical equation. One common fuel is ethanol, $\text{C}_2\text{H}_5\text{OH}$, whose combustion equation is



If nitrogen is present in the original fuel, it is converted to N_2 , not to a nitrogen-oxygen compound. Thus, for the combustion of the fuel dinitroethylene, whose formula is $\text{C}_2\text{H}_2\text{N}_2\text{O}_4$, we have



Example 4.10

Problems

Complete and balance each combustion equation.

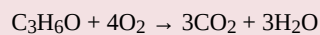
1. the combustion of propane, C_3H_8
2. the combustion of ammonia, NH_3

Solutions

1. The products of the reaction are CO_2 and H_2O , so our unbalanced equation is $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$. Balancing (and you may have to go back and forth a few times to balance this), we get $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$.
2. The nitrogen atoms in ammonia will react to make N_2 , while the hydrogen atoms will react with O_2 to make H_2O , thus $NH_3 + O_2 \rightarrow N_2 + H_2O$. To balance this equation without fractions (which is the convention), we get $4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$.

Test Yourself

Complete and balance the combustion equation for cyclopropanol, C_3H_6O .

Answer

Propane is a fuel used to provide heat for some homes. Propane is stored in large tanks like that shown here.

Key Takeaways

- A composition reaction produces a single substance from multiple reactants.
- A decomposition reaction produces multiple products from a single reactant.
- Combustion reactions are the combination of some compound with oxygen to make oxides of the other elements as products (although nitrogen atoms react to make N₂).

Exercises

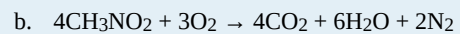
Questions

1. Which is a composition reaction and which is not?
 - a. $\text{NaCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{NaNO}_3$
 - b. $\text{CaO} + \text{CO}_2 \rightarrow \text{CaCO}_3$
2. Which is a composition reaction and which is not?
 - a. $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$
 - b. $2\text{HBr} + \text{Cl}_2 \rightarrow 2\text{HCl} + \text{Br}_2$
3. Which is a composition reaction and which is not?
 - a. $2\text{SO}_2 + \text{O}_2 \rightarrow 2\text{SO}_3$
 - b. $6\text{C} + 3\text{H}_2 \rightarrow \text{C}_6\text{H}_6$
4. Which is a composition reaction and which is not?
 - a. $4\text{Na} + 2\text{C} + 3\text{O}_2 \rightarrow 2\text{Na}_2\text{CO}_3$
 - b. $\text{Na}_2\text{CO}_3 \rightarrow \text{Na}_2\text{O} + \text{CO}_2$
5. Which is a decomposition reaction and which is not?
 - a. $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
 - b. $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
6. Which is a decomposition reaction and which is not?
 - a. $3\text{O}_2 \rightarrow 2\text{O}_3$
 - b. $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$
7. Which is a decomposition reaction and which is not?
 - a. $\text{Na}_2\text{O} + \text{CO}_2 \rightarrow \text{Na}_2\text{CO}_3$
 - b. $\text{H}_2\text{SO}_3 \rightarrow \text{H}_2\text{O} + \text{SO}_2$
8. Which is a decomposition reaction and which is not?
 - a. $2\text{C}_7\text{H}_5\text{N}_3\text{O}_6 \rightarrow 3\text{N}_2 + 5\text{H}_2\text{O} + 7\text{CO} + 7\text{C}$
 - b. $\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}$

9. Which is a combustion reaction and which is not?
- $\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}$
 - $2\text{Fe}_2\text{S}_3 + 9\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3 + 6\text{SO}_2$
10. Which is a combustion reaction and which is not?
- $\text{CH}_4 + 2\text{F}_2 \rightarrow \text{CF}_4 + 2\text{H}_2$
 - $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
11. Which is a combustion reaction and which is not?
- $\text{P}_4 + 5\text{O}_2 \rightarrow 2\text{P}_2\text{O}_5$
 - $2\text{Al}_2\text{S}_3 + 9\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3 + 6\text{SO}_2$
12. Which is a combustion reaction and which is not?
- $\text{C}_2\text{H}_4 + \text{O}_2 \rightarrow \text{C}_2\text{H}_4\text{O}_2$
 - $\text{C}_2\text{H}_4 + \text{Cl}_2 \rightarrow \text{C}_2\text{H}_4\text{Cl}_2$
13. Is it possible for a composition reaction to also be a combustion reaction? Give an example to support your case.
14. Is it possible for a decomposition reaction to also be a combustion reaction? Give an example to support your case.
15. Complete and balance each combustion equation.
- $\text{C}_4\text{H}_9\text{OH} + \text{O}_2 \rightarrow ?$
 - $\text{CH}_3\text{NO}_2 + \text{O}_2 \rightarrow ?$
16. Complete and balance each combustion equation.
- $\text{B}_2\text{H}_6 + \text{O}_2 \rightarrow ?$ (The oxide of boron formed is B_2O_3 .)
 - $\text{Al}_2\text{S}_3 + \text{O}_2 \rightarrow ?$ (The oxide of sulfur formed is SO_2 .)
 - $\text{Al}_2\text{S}_3 + \text{O}_2 \rightarrow ?$ (The oxide of sulfur formed is SO_3 .)

Answers

- not composition
 - composition
- composition
 - composition
- not decomposition
 - decomposition
- not decomposition
 - decomposition
- combustion
 - combustion
- combustion
 - combustion
13. Yes; $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ (answers will vary)
- $\text{C}_4\text{H}_9\text{OH} + 6\text{O}_2 \rightarrow 4\text{CO}_2 + 5\text{H}_2\text{O}$



Media Attributions

Example 4.10

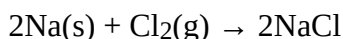
- [“flowers and propane”](#) by [vistavision](#) © [CC BY-NC-ND \(Attribution NonCommercial NoDerivs\)](#)

Oxidation-Reduction Reactions

Learning Objectives

1. Define *oxidation* and *reduction*.
2. Assign oxidation numbers to atoms in simple compounds.
3. Recognize a reaction as an oxidation-reduction reaction.

Consider this chemical reaction:



The reactants are elements, and it is assumed that they are electrically neutral; they have the same number of electrons as protons. The product, however, is ionic; it is composed of Na^+ and Cl^- ions. Somehow, the individual sodium atoms as reactants had to lose an electron to make the Na^+ ion, while the chlorine atoms as reactants had to each gain an electron to become the Cl^- ion. This reaction involves the *transfer of electrons* between atoms.

In reality, electrons are lost by some atoms and gained by other atoms simultaneously. However, mentally we can separate the two processes. **Oxidation** is defined as the loss of one or more electrons by an atom. **Reduction** is defined as the gain of one or more electrons by an atom. In reality, oxidation and reduction always occur together; it is only mentally that we can separate them. Chemical reactions that involve the transfer of electrons are called **oxidation-reduction (or redox) reactions**.

Redox reactions require that we keep track of the electrons assigned to each atom in a chemical reaction. How do we do that? We use an artificial count called the **oxidation number** to keep track of electrons in atoms. Oxidation numbers are assigned to atoms based on a series of rules. Oxidation numbers are not necessarily equal to the charge on the atom; we must keep the concepts of charge and oxidation numbers separate.

The rules for assigning oxidation numbers to atoms are as follows:

1. Atoms in their elemental state are assigned an oxidation number of 0.
2. Atoms in monatomic (i.e., one-atom) ions are assigned an oxidation number equal to their charge. Oxidation numbers are usually written with the sign first, then the magnitude, which differentiates them from charges.
3. In compounds, fluorine is assigned a -1 oxidation number; oxygen is usually assigned a -2 oxidation number (except in peroxide compounds [where it is -1] and in binary compounds with fluorine [where it is positive]); and hydrogen is usually assigned a $+1$ oxidation number (except when it exists as the hydride ion, H^- , in which case rule 2 prevails).

- In compounds, all other atoms are assigned an oxidation number so that the sum of the oxidation numbers on all the atoms in the species equals the charge on the species (which is zero if the species is neutral).

Let us work through a few examples for practice. In H_2 , both hydrogen atoms have an oxidation number of 0, by rule 1. In NaCl , sodium has an oxidation number of +1, while chlorine has an oxidation number of -1, by rule 2. In H_2O , the hydrogen atoms each have an oxidation number of +1, while the oxygen has an oxidation number of -2, even though hydrogen and oxygen do not exist as ions in this compound as per rule 3. By contrast, by rule 3 in hydrogen peroxide (H_2O_2), each hydrogen atom has an oxidation number of +1, while each oxygen atom has an oxidation number of -1. We can use rule 4 to determine oxidation numbers for the atoms in SO_2 . Each oxygen atom has an oxidation number of -2; for the sum of the oxidation numbers to equal the charge on the species (which is zero), the sulfur atom is assigned an oxidation number of +4. Does this mean that the sulfur atom has a 4+ charge on it? No, it only means that the sulfur atom is assigned a +4 oxidation number by our rules of apportioning electrons among the atoms in a compound.

Example 4.11

Problems

Assign oxidation numbers to the atoms in each substance.

- Br_2
- SiO_2
- $\text{Ba}(\text{NO}_3)_2$

Solutions

- Br_2 is the elemental form of bromine. Therefore, by rule 1, each atom has an oxidation number of 0.
- By rule 3, oxygen is normally assigned an oxidation number of -2. For the sum of the oxidation numbers to equal the charge on the species (which is zero), the silicon atom is assigned an oxidation number of +4.
- The compound barium nitrate can be separated into two parts: the Ba^{2+} ion and the nitrate ion. Considering these separately, the Ba^{2+} ion has an oxidation number of +2 by rule 2. Now consider the NO_3^- ion. Oxygen is assigned an oxidation number of -2, and there are three oxygens. According to rule 4, the sum of the oxidation number on all atoms must equal the charge on the species, so we have the simple algebraic equation $x + 3(-2) = -1$, where x is the oxidation number of the nitrogen atom and -1 represents the charge on the species. Evaluating,

$$\begin{aligned} x + (-6) &= -1 \\ x &= 5 \end{aligned}$$

Thus, the oxidation number on the N atom in the nitrate ion is +5.

Test Yourself

Assign oxidation numbers to the atoms in H_3PO_4 .

Answer

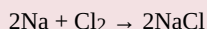
H = +1, O = -2, P = +5

All redox reactions occur with a simultaneous change in the oxidation numbers of some atoms. At least two elements must change their oxidation numbers. When an oxidation number of an atom is increased in the course of a redox reaction, that atom is being *oxidized*. When an oxidation number of an atom is decreased in the course of a redox reaction, that atom is being *reduced*. Oxidation and reduction are thus also defined in terms of increasing or decreasing oxidation numbers, respectively.

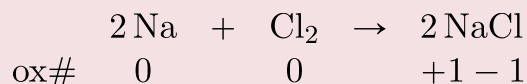
Example 4.12

Problem

Identify what is being oxidized and reduced in this redox equation:

**Solution**

Consider the reactants. Because both reactants are the elemental forms of their atoms, the Na and Cl atoms as reactants have oxidation numbers of 0. In the ionic product, the sodium ions have an oxidation number of +1, while the chloride ions have an oxidation number of -1:

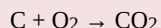


We note that the sodium is increasing its oxidation number from 0 to +1, so it is being oxidized; chlorine is decreasing its oxidation number from 0 to -1, so it is being reduced.

Because oxidation numbers are changing, this is a redox reaction. Note that the total number of electrons lost by the sodium (two, one lost from each atom) is gained by the chlorine atoms (two, one gained for each atom).

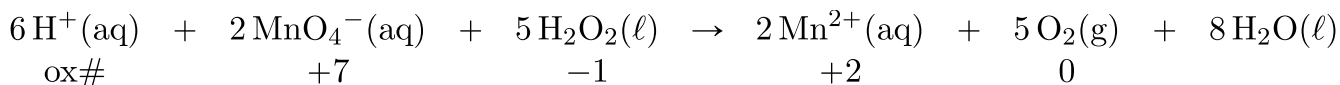
Test Yourself

Identify what is being oxidized and reduced in this redox equation.

**Answer**

C is being oxidized from 0 to +4; O is being reduced from 0 to -2.

In this introduction to oxidation-reduction reactions, we are using straightforward examples. However, oxidation reactions can become quite complex; the following equation represents a redox reaction:



To demonstrate that this is a redox reaction, the oxidation numbers of the species being oxidized and reduced are listed; can you determine what is being oxidized and what is being reduced? This is also an example of a net ionic reaction; spectator ions that do not change oxidation numbers are not displayed in the equation.

Food and Drink App: Acids in Foods

Many foods and beverages contain acids. Acids impart a sour note to the taste of foods, which may add some pleasantness to the food. For example, orange juice contains citric acid, $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$. Note how this formula shows hydrogen atoms in two places; the first hydrogen atoms written are the hydrogen atoms that can form H^+ ions, while the second hydrogen atoms written are part of the citrate ion, $\text{C}_6\text{H}_5\text{O}_7^{3-}$. Lemons and limes contain much more citric acid—about 60 times as much—which accounts for these citrus fruits being more sour than most oranges. Vinegar is essentially a ~5% solution of acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) in water. Apples contain malic acid ($\text{H}_2\text{C}_4\text{H}_4\text{O}_5$; the name *malic acid* comes from the apple's botanical genus name, *malus*), while lactic acid ($\text{HC}_3\text{H}_5\text{O}_3$) is found in wine and sour milk products, such as yogurt and some cottage cheeses.

Table 4.2 “Various Acids Found in Food and Beverages” lists some acids found in foods, either naturally or as an additive. Frequently, the salts of acid anions are used as additives, such as monosodium glutamate (MSG), which is the sodium salt derived from glutamic acid. As you read the list, you should come to the inescapable conclusion that it is impossible to avoid acids in food and beverages.

Table 4.2 Various Acids Found in Food and Beverages

Acid Name	Acid Formula	Use and Appearance
acetic acid	$\text{HC}_2\text{H}_3\text{O}_2$	flavouring; found in vinegar
adipic acid	$\text{H}_2\text{C}_6\text{H}_8\text{O}_4$	flavouring; found in processed foods and some antacids
alginic acid	various	thickener; found in drinks, ice cream, and weight loss products
ascorbic acid	$\text{HC}_6\text{H}_7\text{O}_6$	antioxidant, also known as vitamin C; found in fruits and vegetables
benzoic acid	$\text{HC}_6\text{H}_5\text{CO}_2$	preservative; found in processed foods
citric acid	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7$	flavouring; found in citrus fruits
dehydroacetic acid	$\text{HC}_8\text{H}_7\text{O}_4$	preservative, especially for strawberries and squash
erythroic acid	$\text{HC}_6\text{H}_7\text{O}_6$	antioxidant; found in processed foods
fatty acids	various	thickener and emulsifier; found in processed foods
fumaric acid	$\text{H}_2\text{C}_4\text{H}_2\text{O}_4$	flavouring; acid reactant in some baking powders
glutamic acid	$\text{H}_2\text{C}_5\text{H}_7\text{NO}_4$	flavouring; found in processed foods and in tomatoes, some cheeses, and soy products
lactic acid	$\text{HC}_3\text{H}_5\text{O}_3$	flavouring; found in wine, yogurt, cottage cheese, and other sour milk products
malic acid	$\text{H}_2\text{C}_4\text{H}_4\text{O}_5$	flavouring; found in apples and unripe fruit
phosphoric acid	H_3PO_4	flavouring; found in some colas
propionic acid	$\text{HC}_3\text{H}_5\text{O}_2$	preservative; found in baked goods
sorbic acid	$\text{HC}_6\text{H}_7\text{O}_2$	preservative; found in processed foods
stearic acid	$\text{HC}_{18}\text{H}_{35}\text{O}_2$	anticaking agent; found in hard candies
succinic acid	$\text{H}_2\text{C}_4\text{H}_4\text{O}_4$	flavouring; found in wine and beer
tartaric acid	$\text{H}_2\text{C}_4\text{H}_4\text{O}_6$	flavouring; found in grapes, bananas, and tamarinds

Key Takeaways

- Oxidation-reduction (redox) reactions involve the transfer of electrons from one atom to another.
- Oxidation numbers are used to keep track of electrons in atoms.
- There are rules for assigning oxidation numbers to atoms.
- Oxidation is an increase of oxidation number (a loss of electrons); reduction is a decrease in oxidation number (a gain of electrons).

Exercises

Questions

1. Is the reaction $2\text{K}(s) + \text{Br}_2(\ell) \rightarrow 2\text{KBr}(s)$ an oxidation-reduction reaction? Explain your answer.
2. Is the reaction $\text{NaCl}(aq) + \text{AgNO}_3(aq) \rightarrow \text{NaNO}_3(aq) + \text{AgCl}(s)$ an oxidation-reduction reaction? Explain your answer.
3. In the reaction $2\text{Ca}(s) + \text{O}_2(g) \rightarrow 2\text{CaO}$, indicate what has lost electrons and what has gained electrons.
4. In the reaction $16\text{Fe}(s) + 3\text{S}_8(s) \rightarrow 8\text{Fe}_2\text{S}_3(s)$, indicate what has lost electrons and what has gained electrons.
5. In the reaction $2\text{Li}(s) + \text{O}_2(g) \rightarrow \text{Li}_2\text{O}_2(s)$, indicate what has been oxidized and what has been reduced.
6. In the reaction $2\text{Ni}(s) + 3\text{I}_2(s) \rightarrow 2\text{NiI}_3(s)$, indicate what has been oxidized and what has been reduced.
7. What are two different definitions of oxidation?
8. What are two different definitions of reduction?
9. Assign oxidation numbers to each atom in each substance.
 - a. P_4
 - b. SO_2
 - c. SO_2^{2-}
 - d. $\text{Ca}(\text{NO}_3)_2$
10. Assign oxidation numbers to each atom in each substance.
 - a. PF_5
 - b. $(\text{NH}_4)_2\text{S}$
 - c. Hg
 - d. Li_2O_2 (lithium peroxide)
11. Assign oxidation numbers to each atom in each substance.
 - a. CO
 - b. CO_2
 - c. NiCl_2
 - d. NiCl_3
12. Assign oxidation numbers to each atom in each substance.
 - a. NaH (sodium hydride)

- b. NO_2
 c. NO_2^-
 d. AgNO_3
13. Assign oxidation numbers to each atom in each substance.
- a. CH_2O
 b. NH_3
 c. Rb_2SO_4
 d. $\text{Zn}(\text{C}_2\text{H}_3\text{O}_2)_2$
14. Assign oxidation numbers to each atom in each substance.
- a. C_6H_6
 b. $\text{B}(\text{OH})_3$
 c. Li_2S
 d. Au
15. Identify what is being oxidized and reduced in the redox equation $2\text{NO} + \text{Cl}_2 \rightarrow 2\text{NOCl}$ by assigning oxidation numbers to the atoms.
16. Identify what is being oxidized and reduced in the redox equation $\text{Fe} + \text{SO}_3 \rightarrow \text{FeSO}_3$ by assigning oxidation numbers to the atoms.
17. Identify what is being oxidized and reduced in the redox equation $2\text{KrF}_2 + 2\text{H}_2\text{O} \rightarrow 2\text{Kr} + 4\text{HF} + \text{O}_2$ by assigning oxidation numbers to the atoms.
18. Identify what is being oxidized and reduced in the redox equation $\text{SO}_3 + \text{SCl}_2 \rightarrow \text{SOCl}_2 + \text{SO}_2$ by assigning oxidation numbers to the atoms.
19. Identify what is being oxidized and reduced in the redox equation $2\text{K} + \text{MgCl}_2 \rightarrow 2\text{KCl} + \text{Mg}$ by assigning oxidation numbers to the atoms.
20. Identify what is being oxidized and reduced in the redox equation $\text{C}_7\text{H}_{16} + 11\text{O}_2 \rightarrow 7\text{CO}_2 + 8\text{H}_2\text{O}$ by assigning oxidation numbers to the atoms.

Answers

1. Yes; both K and Br are changing oxidation numbers.
3. Ca has lost electrons, and O has gained electrons.
5. Li has been oxidized, and O has been reduced.
7. Loss of electrons; increase in oxidation number.
9. a. P: 0
 b. S: +4; O: -2
 c. S: +2; O: -2
 d. Ca: 2+; N: +5; O: -2
11. a. C: +2; O: -2
 b. C: +4; O: -2
 c. Ni: +2; Cl: -1
 d. Ni: +3; Cl: -1
13. a. C: 0; H: +1; O: -2

- b. N: -3; H: +1
- c. Rb: +1; S: +6; O: -2
- d. Zn: +2; C: 0; H: +1; O: -2

- 15. N is being oxidized, and Cl is being reduced.
- 17. O is being oxidized, and Kr is being reduced.
- 19. K is being oxidized, and Mg is being reduced.

Neutralization Reactions

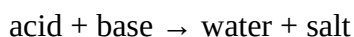
Learning Objectives

1. Identify an acid and a base.
2. Identify a neutralization reaction and predict its products.

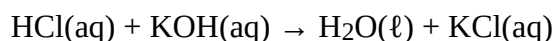
In the section called “[Acids](#)”, we defined an acid as an ionic compound that contains H^+ as the cation. This is slightly incorrect, but until additional concepts were developed, a better definition needed to wait. Now we can redefine an acid: an **acid** is any compound that increases the amount of hydrogen ion (H^+) in an aqueous solution. The chemical opposite of an acid is a base. The equivalent definition of a base is that a **base** is a compound that increases the amount of hydroxide ion (OH^-) in an aqueous solution. These original definitions were proposed by Arrhenius (the same person who proposed ion dissociation) in 1884, so they are referred to as the Arrhenius definitions of an acid and a base, respectively.

You may recognize that, based on the description of a hydrogen atom, an H^+ ion is a hydrogen atom that has lost its lone electron; that is, H^+ is simply a proton. Do we really have bare protons moving about in aqueous solution? No. What is more likely is that the H^+ ion has attached itself to one (or more) water molecule(s). To represent this chemically, we define the **hydronium ion** $\text{H}_3\text{O}^+(\text{aq})$, a water molecule with an extra hydrogen ion attached to it, as H_3O^+ , which represents an additional proton attached to a water molecule. We use the hydronium ion as the more logical way a hydrogen ion appears in an aqueous solution, although in many chemical reactions H^+ and H_3O^+ are treated equivalently.

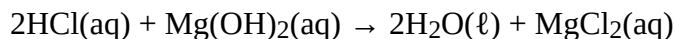
The reaction of an acid and a base is called a **neutralization reaction**. Although acids and bases have their own unique chemistries, the acid and base cancel each other’s chemistry to produce a rather innocuous substance—water. In fact, the general reaction between an acid and a base is



where the term salt is generally used to define any ionic compound (soluble or insoluble) that is formed from a reaction between an acid and a base. (In chemistry, the word *salt* refers to more than just table salt.) For example, the balanced chemical equation for the reaction between $\text{HCl}(\text{aq})$ and $\text{KOH}(\text{aq})$ is



where the salt is KCl . By counting the number of atoms of each element, we find that only one water molecule is formed as a product. However, in the reaction between $\text{HCl}(\text{aq})$ and $\text{Mg}(\text{OH})_2(\text{aq})$, additional molecules of HCl and H_2O are required to balance the chemical equation:



Here, the salt is MgCl_2 . (This is one of several reactions that take place when a type of antacid—a base—is used to treat stomach acid.)

Example 4.13

Problems

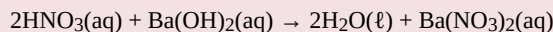
Write the neutralization reactions between each acid and base.

1. $\text{HNO}_3(\text{aq})$ and $\text{Ba}(\text{OH})_2(\text{aq})$
2. $\text{H}_3\text{PO}_4(\text{aq})$ and $\text{Ca}(\text{OH})_2(\text{aq})$

Solutions

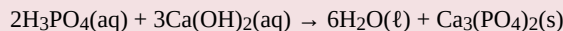
First, we will write the chemical equation with the formulas of the reactants and the expected products; then we will balance the equation.

1. The expected products are water and barium nitrate, so the initial chemical reaction is $\text{HNO}_3(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell) + \text{Ba}(\text{NO}_3)_2(\text{aq})$. To balance the equation, we need to realize that there will be two H_2O molecules, so two HNO_3 molecules are required:



This chemical equation is now balanced.

2. The expected products are water and calcium phosphate, so the initial chemical equation is $\text{H}_3\text{PO}_4(\text{aq}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell) + \text{Ca}_3(\text{PO}_4)_2(\text{s})$. According to the solubility rules, $\text{Ca}_3(\text{PO}_4)_2$ is insoluble, so it has an (s) phase label. To balance this equation, we need two phosphate ions and three calcium ions; we end up with six water molecules to balance the equation:

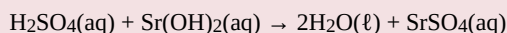


This chemical equation is now balanced.

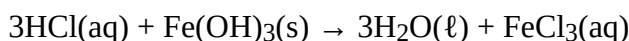
Test Yourself

Write the neutralization reaction between $\text{H}_2\text{SO}_4(\text{aq})$ and $\text{Sr}(\text{OH})_2(\text{aq})$.

Answer

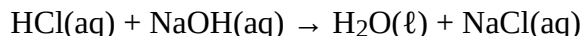


Neutralization reactions are one type of chemical reaction that proceeds even if one reactant is not in the aqueous phase. For example, the chemical reaction between $\text{HCl}(\text{aq})$ and $\text{Fe}(\text{OH})_3(\text{s})$ still proceeds according to the equation:

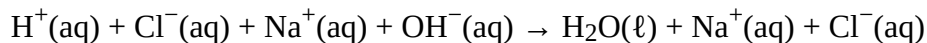


even though $\text{Fe}(\text{OH})_3$ is not soluble. When one realizes that $\text{Fe}(\text{OH})_3(\text{s})$ is a component of rust, this explains why some cleaning solutions for rust stains contain acids—the neutralization reaction produces products that are soluble and wash away. (Washing with acids like HCl is one way to remove rust and rust stains, but HCl must be used with caution!)

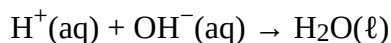
Complete and net ionic reactions for neutralization reactions will depend on whether the reactants and products are soluble, even if the acid and base react. For example, in the reaction of $\text{HCl}(\text{aq})$ and $\text{NaOH}(\text{aq})$:



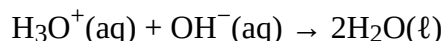
The complete ionic reaction is:



The $\text{Na}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ ions are spectator ions, so we can remove them to have:

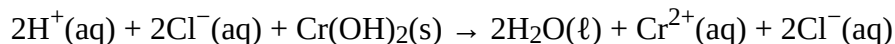


as the net ionic equation. If we wanted to write this in terms of the hydronium ion, $\text{H}_3\text{O}^+(\text{aq})$, we would write it as:

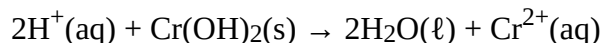


With the exception of the introduction of an extra water molecule, these two net ionic equations are equivalent.

However, for the reaction between $\text{HCl}(\text{aq})$ and $\text{Cr}(\text{OH})_2(\text{s})$, because chromium(II) hydroxide is insoluble, we cannot separate it into ions for the complete ionic equation:



The chloride ions are the only spectator ions here, so the net ionic equation is:



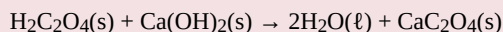
Example 4.14

Problem

Oxalic acid, $\text{H}_2\text{C}_2\text{O}_4(\text{s})$, and $\text{Ca}(\text{OH})_2(\text{s})$ react very slowly. What is the net ionic equation between these two substances if the salt formed is insoluble? (The anion in oxalic acid is the oxalate ion, $\text{C}_2\text{O}_4^{2-}$.)

Solution

The products of the neutralization reaction will be water and calcium oxalate:

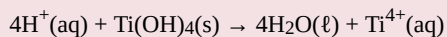


Because nothing is dissolved, there are no substances to separate into ions, so the net ionic equation is the equation of the three solids and one liquid.

Test Yourself

What is the net ionic equation between $\text{HNO}_3(\text{aq})$ and $\text{Ti}(\text{OH})_4(\text{s})$?

Answer



Key Takeaways

- The Arrhenius definition of an acid is a substance that increases the amount of H^+ in an aqueous solution.
- The Arrhenius definition of a base is a substance that increases the amount of OH^- in an aqueous solution.
- Neutralization is the reaction of an acid and a base, which forms water and a salt.
- Net ionic equations for neutralization reactions may include solid acids, solid bases, solid salts, and water.

Exercises

Questions

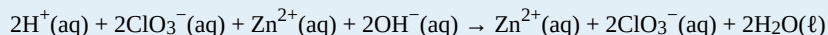
1. What is the Arrhenius definition of an acid?
2. What is the Arrhenius definition of a base?
3. Predict the products of each acid-base combination listed. Assume that a neutralization reaction occurs.
 - a. HCl and KOH
 - b. H_2SO_4 and KOH
 - c. H_3PO_4 and $\text{Ni}(\text{OH})_2$
4. Predict the products of each acid-base combination listed. Assume that a neutralization reaction occurs.
 - a. HBr and $\text{Fe}(\text{OH})_3$
 - b. HNO_2 and $\text{Al}(\text{OH})_3$
 - c. HClO_3 and $\text{Mg}(\text{OH})_2$
5. Write a balanced chemical equation for each neutralization reaction in Exercise 3.
6. Write a balanced chemical equation for each neutralization reaction in Exercise 4.
7. Write a balanced chemical equation for the neutralization reaction between each given acid and base. Include the proper phase labels.
 - a. $\text{HI}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow ?$
 - b. $\text{H}_2\text{SO}_4(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow ?$
8. Write a balanced chemical equation for the neutralization reaction between each given acid and base. Include the proper phase labels.

- a. $\text{HNO}_3(\text{aq}) + \text{Fe}(\text{OH})_3(\text{s}) \rightarrow ?$
 b. $\text{H}_3\text{PO}_4(\text{aq}) + \text{CsOH}(\text{aq}) \rightarrow ?$
- Write the net ionic equation for each neutralization reaction in Exercise 7.
 - Write the net ionic equation for each neutralization reaction in Exercise 8.
 - Write the complete and net ionic equations for the neutralization reaction between $\text{HClO}_3(\text{aq})$ and $\text{Zn}(\text{OH})_2(\text{s})$. Assume the salt is soluble.
 - Write the complete and net ionic equations for the neutralization reaction between $\text{H}_2\text{C}_2\text{O}_4(\text{s})$ and $\text{Sr}(\text{OH})_2(\text{aq})$. Assume the salt is insoluble.
 - Explain why the net ionic equation for the neutralization reaction between $\text{HCl}(\text{aq})$ and $\text{KOH}(\text{aq})$ is the same as the net ionic equation for the neutralization reaction between $\text{HNO}_3(\text{aq})$ and RbOH .
 - Explain why the net ionic equation for the neutralization reaction between $\text{HCl}(\text{aq})$ and $\text{KOH}(\text{aq})$ is different from the net ionic equation for the neutralization reaction between $\text{HCl}(\text{aq})$ and AgOH .
 - Write the complete and net ionic equations for the neutralization reaction between $\text{HCl}(\text{aq})$ and $\text{KOH}(\text{aq})$ using the hydronium ion in place of H^+ . What difference does it make when using the hydronium ion?
 - Write the complete and net ionic equations for the neutralization reaction between $\text{HClO}_3(\text{aq})$ and $\text{Zn}(\text{OH})_2(\text{s})$ using the hydronium ion in place of H^+ . Assume the salt is soluble. What difference does it make when using the hydronium ion?

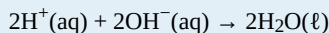
Answers

- An Arrhenius acid increases the amount of H^+ ions in an aqueous solution.
- KCl and H_2O
 - K_2SO_4 and H_2O
 - $\text{Ni}_3(\text{PO}_4)_2$ and H_2O
- $\text{HCl} + \text{KOH} \rightarrow \text{KCl} + \text{H}_2\text{O}$
 - $\text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$
 - $2\text{H}_3\text{PO}_4 + 3\text{Ni}(\text{OH})_2 \rightarrow \text{Ni}_3(\text{PO}_4)_2 + 6\text{H}_2\text{O}$
- $\text{HI}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{KI}(\text{aq}) + \text{H}_2\text{O}(\ell)$
 - $\text{H}_2\text{SO}_4(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{H}_2\text{O}(\ell)$
- $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell)$
 - $2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{H}_2\text{O}(\ell)$

- Complete ionic equation:

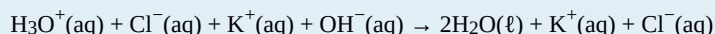


Net ionic equation:

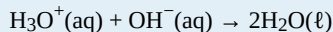


- Because the salts are soluble in both cases, the net ionic reaction is just $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O}(\ell)$.

- Complete ionic equation:



Net ionic equation:



The difference is simply the presence of an extra water molecule as a product.

End-of-Chapter Material

Additional Exercises

1. Chemical equations can also be used to represent physical processes. Write a chemical reaction for the boiling of water, including the proper phase labels.
2. Chemical equations can also be used to represent physical processes. Write a chemical reaction for the freezing of water, including the proper phase labels.
3. Explain why $4\text{Na(s)} + 2\text{Cl}_2\text{(g)} \rightarrow 4\text{NaCl(s)}$ should not be considered a proper chemical equation.
4. Explain why $\text{H}_2\text{(g)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{H}_2\text{O(l)}$ should not be considered a proper chemical equation.
5. Does the chemical reaction represented by $3\text{Zn(s)} + 2\text{Al(NO}_3)_3\text{(aq)} \rightarrow 3\text{Zn(NO}_3)_2\text{(aq)} + 2\text{Al(s)}$ proceed as written? Why or why not?
6. Does the chemical reaction represented by $2\text{Au(s)} + 2\text{HNO}_3\text{(aq)} \rightarrow 2\text{AuNO}_3\text{(aq)} + \text{H}_2\text{(g)}$ proceed as written? Gold is a relatively useful metal for certain applications, such as jewellery and electronics. Does your answer suggest why this is so?
7. Explain what is wrong with the double-replacement reaction $\text{NaCl(aq)} + \text{KBr(aq)} \rightarrow \text{NaK(aq)} + \text{ClBr(aq)}$.
8. Predict the products of and balance the double-replacement reaction $\text{Ag}_2\text{SO}_4\text{(aq)} + \text{SrCl}_2\text{(aq)} \rightarrow ?$
9. Write the complete and net ionic equations for the double-replacement reaction $\text{BaCl}_2\text{(aq)} + \text{Ag}_2\text{SO}_4\text{(aq)} \rightarrow ?$
10. Write the complete and net ionic equations for the double-replacement reaction $\text{Ag}_2\text{SO}_4\text{(aq)} + \text{SrCl}_2\text{(aq)} \rightarrow ?$
11. Identify the spectator ions in the reaction $\text{NaCl(aq)} + \text{KBr(aq)} \rightarrow \text{NaBr(aq)} + \text{KCl(aq)}$. What is the net ionic equation?
12. Complete the reaction $3\text{H}_2\text{SO}_4\text{(aq)} + 2\text{Al(OH)}_3\text{(s)} \rightarrow ?$ and identify the spectator ions. What is the net ionic equation?
13. Can a reaction be a composition reaction and a redox reaction at the same time? Give an example to support your answer.
14. Can a reaction be a combustion reaction and a redox reaction at the same time? Give an example to support your answer.
15. Can a reaction be a decomposition reaction and a redox reaction at the same time? Give an example to support your answer.
16. Can a reaction be a combustion reaction and a double-replacement reaction at the same time? Give an example to support your answer.
17. Why is CH_4 not normally considered an acid?
18. Methyl alcohol has the formula CH_3OH . Why would methyl alcohol not normally be considered a base?
19. What are the oxidation numbers of the nitrogen atoms in these substances?
 - a. N_2
 - b. NH_3
 - c. NO
 - d. N_2O
 - e. NO_2
 - f. N_2O_4
 - g. N_2O_5

- h. NaNO_3
20. What are the oxidation numbers of the sulfur atoms in these substances?
- SF_6
 - Na_2SO_4
 - K_2SO_3
 - SO_3
 - SO_2
 - S_8
 - Na_2S
21. Disproportion is a type of redox reaction in which the same substance is both oxidized and reduced. Identify the element that is disproportionating in $2\text{CuCl}(\text{aq}) \rightarrow \text{CuCl}_2(\text{aq}) + \text{Cu}(\text{s})$ and indicate the initial and final oxidation numbers of that element.
22. Disproportion is a type of redox reaction in which the same substance is both oxidized and reduced. Identify the element that is disproportionating in $3\text{Cl}_2(\text{g}) + 6\text{OH}^-(\text{aq}) \rightarrow 5\text{Cl}^-(\text{aq}) + \text{ClO}_3^-(\text{aq}) + 3\text{H}_2\text{O}(\ell)$ and indicate the initial and final oxidation numbers of that element.

Answers

- $\text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{O}(\text{g})$
 - The coefficients are not in their lowest whole-number ratio.
 - No; zinc is lower in the activity series than aluminum.
 - In the products, the cation is pairing with the cation, and the anion is pairing with the anion.
 - Complete ionic equation: $\text{Ba}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq}) + 2\text{Ag}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{AgCl}(\text{s})$
Net ionic equation: The net ionic equation is the same as the complete ionic equation.
 - Each ion is a spectator ion; there is no overall net ionic equation.
 - Yes; $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$ (answers will vary)
 - Yes; $2\text{HCl} \rightarrow \text{H}_2 + \text{Cl}_2$ (answers will vary)
 - It does not increase the H^+ ion concentration; it is not a compound of H^+ .
- 19.
- 0
 - 3
 - +2
 - +1
 - +4
 - +4
 - +5
 - +5
21. Copper is disproportionating. Initially, its oxidation number is +1; in the products, its oxidation numbers are +2 and 0,

respectively.

Chapter 5. Stoichiometry and the Mole

At Contrived State University in Anytown, Ohio, a new building was dedicated in March 2010 to house the College of Education. The 100,000-square-foot building has enough office space to accommodate 86 full-time faculty members and 167 full-time staff.

In a fit of monetary excess, the university administration offered to buy new furniture (desks and chairs) and computer workstations for all faculty and staff members moving into the new building. However, to save on long-term energy and materials costs, the university offered to buy only 1 laser printer per 10 employees, with the plan to network the printers together.

How many laser printers did the administration have to buy? It is rather simple to show that 26 laser printers are needed for all the employees. However, what if a chemist was calculating quantities for a chemical reaction? Interestingly enough, similar calculations can be performed for chemicals as well as laser printers.

In filling a new office building with furniture and equipment, managers do calculations similar to those performed by scientists doing chemical reactions.

We have already established that quantities are important in science, especially in chemistry. It is important to make accurate measurements of a variety of quantities when performing experiments. However, it is also important to be able to relate one measured quantity to another, unmeasured quantity. In this chapter, we will consider how we manipulate quantities to relate them to each other.

Stoichiometry

Learning Objectives

1. Define *stoichiometry*.
2. Relate quantities in a balanced chemical reaction on a molecular basis.

Consider a classic recipe for pound cake: 1 pound of eggs, 1 pound of butter, 1 pound of flour, and 1 pound of sugar. (That's why it's called "pound cake.") If you have 4 pounds of butter, how many pounds of sugar, flour, and eggs do you need? You would need 4 pounds each of sugar, flour, and eggs.

Now suppose you have 1.00 g H₂. If the chemical reaction follows the balanced chemical equation $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$, then what mass of oxygen do you need to make water?

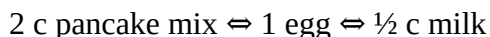
Curiously, this chemical reaction question is very similar to the pound cake question. Both of them involve relating a quantity of one substance to a quantity of another substance or substances. The relating of one chemical substance to another using a balanced chemical reaction is called **stoichiometry**. Using stoichiometry is a fundamental skill in chemistry; it greatly broadens your ability to predict what will occur and, more importantly, how much is produced.

Let us consider a more complicated example. A recipe for pancakes calls for 2 cups (c) of pancake mix, 1 egg, and $\frac{1}{2}$ c of milk. We can write this in the form of a chemical equation:



If you have 9 c of pancake mix, how many eggs and how much milk do you need? It might take a little bit of work, but eventually you will find you need $4\frac{1}{2}$ eggs and $2\frac{1}{4}$ c milk.

How can we formalize this? We can make a conversion factor using our original recipe and use that conversion factor to convert from a quantity of one substance to a quantity of another substance, similar to the way we constructed a conversion factor between feet and yards in [Chapter 2 "Measurements"](#). Because one recipe's worth of pancakes requires 2 c of pancake mix, 1 egg, and $\frac{1}{2}$ c of milk, we actually have the following mathematical relationships that relate these quantities:



where \Leftrightarrow is the mathematical symbol for "is equivalent to." This does not mean that 2 c of pancake mix equals 1 egg. However, *as far as this recipe is concerned*, these are the equivalent quantities needed for a single recipe of pancakes. So, any possible quantities of two or more ingredients must have the same numerical ratio as the ratios in the equivalence.

We can deal with these equivalences in the same way we deal with equalities in unit conversions: we

can make conversion factors that essentially equal 1. For example, to determine how many eggs we need for 9 c of pancake mix, we construct the conversion factor:

$$\frac{1 \text{ egg}}{2 \text{ c pancake mix}}$$

This conversion factor is, in a strange way, equivalent to 1 because the recipe relates the two quantities. Starting with our initial quantity and multiplying by our conversion factor, we get:

$$9 \text{ c } \cancel{\text{pancake mix}} \times \frac{1 \text{ egg}}{2 \text{ c } \cancel{\text{pancake mix}}} = 4.5 \text{ eggs}$$

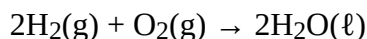
Note how the units *cups pancake mix* cancelled, leaving us with units of *eggs*. This is the formal, mathematical way of getting our amounts to mix with 9 c of pancake mix. We can use a similar conversion factor for the amount of milk:

$$9 \text{ c } \cancel{\text{pancake mix}} \times \frac{\frac{1}{2} \text{ c milk}}{2 \text{ c } \cancel{\text{pancake mix}}} = 2.25 \text{ c milk}$$

Again, units cancel, and new units are introduced.

A balanced chemical equation is nothing more than *a recipe for a chemical reaction*. The difference is that a balanced chemical equation is written in terms of atoms and molecules, not cups, pounds, and eggs.

For example, consider the following chemical equation:



We can interpret this as, literally, “two hydrogen molecules react with one oxygen molecule to make two water molecules.” That interpretation leads us directly to some equivalences, just as our pancake recipe did:



These equivalences allow us to construct conversion factors:

$$\frac{2 \text{ molecules H}_2}{1 \text{ molecule O}_2} \Leftrightarrow \frac{2 \text{ molecules H}_2}{2 \text{ molecules H}_2\text{O}} \Leftrightarrow \frac{1 \text{ molecule O}_2}{2 \text{ molecules H}_2\text{O}}$$

and so forth. These conversions can be used to relate quantities of one substance to quantities of another. For example, suppose we need to know how many molecules of oxygen are needed to react with 16 molecules of H_2 . As we did with converting units, we start with our given quantity and use the appropriate conversion factor:

$$16 \text{ } \cancel{\text{molecules H}_2} \times \frac{1 \text{ molecule O}_2}{2 \text{ } \cancel{\text{molecules H}_2}} = 8 \text{ molecules O}_2$$

Note how the unit *molecules H₂* cancels algebraically, just as any unit does in a conversion like this.

The conversion factor came directly from the coefficients in the balanced chemical equation. This is another reason why a properly balanced chemical equation is important.

Example 5.1

Problem

How many molecules of SO_3 are needed to react with 144 molecules of Fe_2O_3 given the balanced chemical equation $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{SO}_3(\text{g}) \rightarrow \text{Fe}_2(\text{SO}_4)_3$?

Solution

We use the balanced chemical equation to construct a conversion factor between Fe_2O_3 and SO_3 . The number of molecules of Fe_2O_3 goes on the bottom of our conversion factor so it cancels with our given amount, and the molecules of SO_3 go on the top. Thus, the appropriate conversion factor is:

$$\frac{3 \text{ molecules } \text{SO}_3}{1 \text{ molecule } \text{Fe}_2\text{O}_3}$$

Starting with our given amount and applying the conversion factor, the result is:

$$144 \text{ molecules } \cancel{\text{Fe}_2\text{O}_3} \times \frac{3 \text{ molecules } \text{SO}_3}{1 \text{ molecule } \cancel{\text{Fe}_2\text{O}_3}} = 432 \text{ molecules } \text{SO}_3$$

We need 432 molecules of SO_3 to react with 144 molecules of Fe_2O_3 .

Test Yourself

How many molecules of H_2 are needed to react with 29 molecules of N_2 to make ammonia if the balanced chemical equation is $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$?

Answer

87 molecules

Chemical equations also allow us to make conversions regarding the number of atoms in a chemical reaction because a chemical formula lists the number of atoms of each element in a compound. The formula H_2O indicates that there are two hydrogen atoms and one oxygen atom in each molecule, and these relationships can be used to make conversion factors:

$$\frac{2 \text{ atoms H}}{1 \text{ molecule } \text{H}_2\text{O}} \Leftrightarrow \frac{1 \text{ molecule } \text{H}_2\text{O}}{1 \text{ atom O}}$$

Conversion factors like this can also be used in stoichiometry calculations.

Example 5.2

Problem

How many molecules of NH_3 can you make if you have 228 atoms of H_2 ?

Solution

From the formula, we know that one molecule of NH_3 has three H atoms. Use that fact as a conversion factor:

$$228 \text{ atoms H} \times \frac{1 \text{ molecule NH}_3}{3 \text{ atoms H}} = 76 \text{ molecules NH}_3$$

Test Yourself

How many molecules of $\text{Fe}_2(\text{SO}_4)_3$ can you make from 777 atoms of S?

Answer

259 molecules

For a video lecture on stoichiometry, view this [video on stoichiometry](#) by Dr. Jessie A. Key.

Key Takeaways

- Quantities of substances can be related to each other using balanced chemical equations.

Exercises

Questions

1. Think back to the pound cake recipe. What possible conversion factors can you construct relating the components of the recipe?
2. Think back to the pancake recipe. What possible conversion factors can you construct relating the components of the recipe?
3. What are all the conversion factors that can be constructed from the balanced chemical reaction $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$?
4. What are all the conversion factors that can be constructed from the balanced chemical reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$?
5. Given the chemical equation $\text{Na}(\text{s}) + \text{H}_2\text{O}(\ell) \rightarrow \text{NaOH}(\text{aq}) + \text{H}_2(\text{g})$

- a. Balance the equation.
 - b. How many molecules of H_2 are produced when 332 atoms of Na react?
6. Given the chemical equation $\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_3(\text{g})$
- a. Balance the equation.
 - b. How many molecules of O_2 are needed when 38 atoms of S react?
7. For the balanced chemical equation $6\text{H}^+(\text{aq}) + 2\text{MnO}_4^-(\text{aq}) + 5\text{H}_2\text{O}_2(\ell) \rightarrow 2\text{Mn}^{2+}(\text{aq}) + 5\text{O}_2(\text{g}) + 8\text{H}_2\text{O}(\ell)$, how many molecules of H_2O are produced when 75 molecules of H_2O_2 react?
8. For the balanced chemical reaction $2\text{C}_6\text{H}_6(\ell) + 15\text{O}_2(\text{g}) \rightarrow 12\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\ell)$, how many molecules of CO_2 are produced when 56 molecules of C_6H_6 react?
9. Given the balanced chemical equation $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{SO}_3(\text{g}) \rightarrow \text{Fe}_2(\text{SO}_4)_3$, how many molecules of $\text{Fe}_2(\text{SO}_4)_3$ are produced if 321 atoms of S react?
10. For the balanced chemical equation $\text{CuO}(\text{s}) + \text{H}_2\text{S}(\text{g}) \rightarrow \text{CuS} + \text{H}_2\text{O}(\ell)$, how many molecules of CuS are formed if 9,044 atoms of H react?
11. For the balanced chemical equation $\text{Fe}_2\text{O}_3(\text{s}) + 3\text{SO}_3(\text{g}) \rightarrow \text{Fe}_2(\text{SO}_4)_3$, suppose we need to make 145,000 molecules of $\text{Fe}_2(\text{SO}_4)_3$. How many molecules of SO_3 do we need?
12. One way to make sulfur hexafluoride is to react thioformaldehyde, CH_2S , with elemental fluorine, is described by $\text{CH}_2\text{S} + 6\text{F}_2 \rightarrow \text{CF}_4 + 2\text{HF} + \text{SF}_6$. If 45,750 molecules of SF_6 are needed, how many molecules of F_2 are required?
13. Construct the three independent conversion factors possible for these two reactions: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ and $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}_2$. Why are the ratios between H_2 and O_2 different? The conversion factors are different because the stoichiometries of the balanced chemical reactions are different.
14. Construct the three independent conversion factors possible for these two reactions $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$ and $4\text{Na} + 2\text{Cl}_2 \rightarrow 4\text{NaCl}$. What similarities, if any, exist in the conversion factors from these two reactions?

Answers

1. $\frac{1 \text{ pound butter}}{1 \text{ pound flour}}$ or $\frac{1 \text{ pound sugar}}{1 \text{ pound eggs}}$ are two conversion factors that can be constructed from the pound cake recipe. Other conversion factors are also possible.
3. $\frac{2 \text{ molecules H}_2}{1 \text{ molecule O}_2} \Leftrightarrow \frac{1 \text{ molecule O}_2}{2 \text{ molecules H}_2\text{O}} \Leftrightarrow \frac{2 \text{ molecules H}_2}{2 \text{ molecules H}_2\text{O}}$ and their reciprocals are the conversion factors that can be constructed.
5. $2\text{Na}(\text{s}) + 2\text{H}_2\text{O}(\ell) \rightarrow 2\text{NaOH}(\text{aq}) + \text{H}_2(\text{g})$ and 166 molecules
7. 120 molecules
9. 107 molecules
11. 435,000 molecules
13.
 - a. $\frac{2 \text{ molecules H}_2}{1 \text{ molecule O}_2} \Leftrightarrow \frac{1 \text{ molecule O}_2}{2 \text{ molecules H}_2\text{O}} \Leftrightarrow \frac{2 \text{ molecules H}_2}{2 \text{ molecules H}_2\text{O}}$
 - b. $\frac{1 \text{ molecule H}_2}{1 \text{ molecule O}_2} \Leftrightarrow \frac{1 \text{ molecule O}_2}{1 \text{ molecule H}_2\text{O}_2} \Leftrightarrow \frac{1 \text{ molecule H}_2}{1 \text{ molecule H}_2\text{O}_2}$

The Mole

Learning Objectives

1. Describe the unit *mole*.
2. Relate the mole quantity of substance to its mass.

So far, we have been talking about chemical substances in terms of individual atoms and molecules. Yet we don't typically deal with substances an atom or a molecule at a time; we work with millions, billions, and trillions of atoms and molecules at a time. What we need is a way to deal with macroscopic, rather than microscopic, amounts of matter. We need a unit of amount that relates quantities of substances on a scale that we can interact with.

Chemistry uses a unit called mole. A **mole** (mol) is a number of things equal to the number of atoms in exactly 12 g of carbon-12. Experimental measurements have determined that this number is very large:

$$1 \text{ mol} = 6.02214179 \times 10^{23} \text{ things}$$

Understand that a mole means a number of things, just like a dozen means a certain number of things—twelve, in the case of a dozen. But a mole is a much larger number of things. These things can be atoms, or molecules, or eggs; however, in chemistry, we usually use the mole to refer to the amounts of atoms or molecules. Although the number of things in a mole is known to eight decimal places, it is usually fine to use only two or three decimal places in calculations. The numerical value of things in a mole is often called *Avogadro's number* (N_A), which is also known as the *Avogadro constant*, after Amadeo Avogadro, an Italian chemist who first proposed its importance.

Example 5.3

Problem

How many molecules are present in 2.76 mol of H₂O? How many atoms is this?

Solution

The definition of a mole is an equality that can be used to construct a conversion factor. Also, because we know that there are three atoms in each molecule of H₂O, we can also determine the number of atoms in the sample.

$$2.76 \cancel{\text{ mol H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{ molecules H}_2\text{O}}{\cancel{\text{ mol H}_2\text{O}}} = 1.66 \times 10^{24} \text{ molecules H}_2\text{O}$$

To determine the total number of atoms, we have:

$$1.66 \times 10^{24} \text{ molecules } \cancel{\text{H}_2\text{O}} \times \frac{3 \text{ atoms}}{1 \text{ molecule}} = 4.99 \times 10^{24} \text{ atoms}$$

Test Yourself

How many molecules are present in 4.61×10^{-2} mol of O_2 ?

Answer

2.78×10^{22} molecules

How big is a mole? It is very large. Suppose you had a mole of dollar bills that need to be counted. If everyone on earth (about 6 billion people) counted one bill per second, it would take about 3.2 million years to count all the bills. A mole of sand would fill a cube about 32 km on a side. A mole of pennies stacked on top of each other would have about the same diameter as our galaxy, the Milky Way. A mole is a lot of things—but atoms and molecules are very tiny. One mole of carbon atoms would make a cube that is 1.74 cm on a side, small enough to carry in your pocket.

Why is the mole unit so important? It represents the link between the microscopic and the macroscopic, especially in terms of mass. *A mole of a substance has the same mass in grams as one unit (atom or molecules) has in atomic mass units.* The mole unit allows us to express amounts of atoms and molecules in visible amounts that we can understand.

For example, we already know that, by definition, a mole of carbon has a mass of exactly 12 g. This means that exactly 12 g of C has 6.022×10^{23} atoms:

$$12 \text{ g C} = 6.022 \times 10^{23} \text{ atoms C}$$

We can use this equality as a conversion factor between the number of atoms of carbon and the number of grams of carbon. How many grams are there, say, in 1.50×10^{25} atoms of carbon? This is a one-step conversion:

$$1.50 \times 10^{25} \text{ atoms C} \times \frac{12.0000 \text{ g C}}{6.022 \times 10^{23} \text{ atoms C}} = 299 \text{ g C}$$

But it also goes beyond carbon. Previously we defined atomic and molecular masses as the number of atomic mass units per atom or molecule. Now we can do so in terms of grams. The atomic mass of an element is the number of grams in 1 mol of atoms of that element, while the molecular mass of a compound is the number of grams in 1 mol of molecules of that compound. Sometimes these masses are called **molar masses** to emphasize the fact that they are the mass for 1 mol of things. (The term *molar* is the adjective form of mole and has nothing to do with teeth.)

Here are some examples:

- The mass of a hydrogen (H) atom is 1.0079 u; the mass of 1 mol of H atoms is 1.0079 g.
- Elemental hydrogen exists as a diatomic molecule, H_2 . One molecule has a mass of 1.0079

+ 1.0079 = 2.0158 u, while 1 mol H₂ has a mass of 2.0158 g.

- A molecule of H₂O has a mass of about 18.01 u; 1 mol H₂O has a mass of 18.01 g.
- A single unit of NaCl has a mass of 58.45 u; NaCl has a molar mass of 58.45 g.

In each of these moles of substances, there are 6.022×10^{23} units: 6.022×10^{23} atoms of H, 6.022×10^{23} molecules of H₂ and H₂O, 6.022×10^{23} units of NaCl ions. These relationships give us plenty of opportunities to construct conversion factors for simple calculations.

Example 5.4

Problem

What is the molar mass of C₆H₁₂O₆?

Solution

To determine the molar mass, we simply add the atomic masses of the atoms in the molecular formula but express the total in grams per mole, not atomic mass units. The masses of the atoms can be taken from the periodic table or the list of elements in "[Appendix A: Periodic Table of the Elements](#)":

Atom or Molecule	Mass Formula	Mass
6C	6×12.011	72.066
12H	12×1.0079	12.0948
6O	6×15.999	95.994
C ₆ H ₁₂ O ₆	$72.066 + 12.0948 + 95.994$	180.155 g/mol

Per convention, the unit *grams per mole* is written as a fraction.

Test Yourself

What is the molar mass of AgNO₃?

Answer

169.87 g/mol

Knowing the molar mass of a substance, we can calculate the number of moles in a certain mass of a substance and vice versa, as these examples illustrate. The molar mass is used as the conversion factor.

Example 5.5

Problem

What is the mass of 3.56 mol of HgCl_2 ? The molar mass of HgCl_2 is 271.49 g/mol.

Solution

Use the molar mass as a conversion factor between moles and grams. Because we want to cancel the mole unit and introduce the gram unit, we can use the molar mass as given:

$$3.56 \cancel{\text{ mol HgCl}_2} \times \frac{271.49 \text{ g HgCl}_2}{\cancel{\text{ mol HgCl}_2}} = 967 \text{ g HgCl}_2$$

Test Yourself

What is the mass of 33.7 mol of H_2O ?

Answer

607 g

Example 5.6

Problem

How many moles of H_2O are present in 240.0 g of water (about the mass of a cup of water)?

Solution

Use the molar mass of H_2O as a conversion factor from mass to moles. The molar mass of water is $(1.0079 + 1.0079 + 15.999) = 18.015 \text{ g/mol}$. However, because we want to cancel the gram unit and introduce moles, we need to take the reciprocal of this quantity, or $1 \text{ mol}/18.015 \text{ g}$:

$$240.0 \cancel{\text{ g H}_2\text{O}} \times \frac{1 \text{ mol H}_2\text{O}}{18.015 \cancel{\text{ g H}_2\text{O}}} = 13.32 \text{ mol H}_2\text{O}$$

Test Yourself

How many moles are present in 35.6 g of H_2SO_4 (molar mass = 98.08 g/mol)?

Answer

0.363 mol

Other conversion factors can be combined with the definition of mole—density, for example.

Example 5.7

Problem

The density of ethanol is 0.789 g/mL. How many moles are in 100.0 mL of ethanol? The molar mass of ethanol is 46.08 g/mol.

Solution

Here, we use density to convert from volume to mass and then use the molar mass to determine the number of moles.

$$100.0 \text{ mL ethanol} \times \frac{0.789 \text{ g}}{\text{mL}} \times \frac{1 \text{ mol}}{46.08 \text{ g}} = 1.71 \text{ mol ethanol}$$

Test Yourself

If the density of benzene, C₆H₆, is 0.879 g/mL, how many moles are present in 17.9 mL of benzene?

Answer

0.201 mol

Key Takeaways

- The mole is a key unit in chemistry.
- The molar mass of a substance, in grams, is numerically equal to one atom's or molecule's mass in atomic mass units.

Exercises

Questions

1. How many atoms are present in 4.55 mol of Fe?
2. How many atoms are present in 0.0665 mol of K?
3. How many molecules are present in 2.509 mol of H₂S?
4. How many molecules are present in 0.336 mol of acetylene (C₂H₂)?
5. How many moles are present in 3.55×10^{24} Pb atoms?
6. How many moles are present in 2.09×10^{22} Ti atoms?
7. How many moles are present in 1.00×10^{23} PF₃ molecules?

8. How many moles are present in 5.52×10^{25} penicillin molecules?
9. Determine the molar mass of each substance.
 - a. Si
 - b. SiH_4
 - c. K_2O
10. Determine the molar mass of each substance.
 - a. Cl_2
 - b. SeCl_2
 - c. $\text{Ca}(\text{C}_2\text{H}_3\text{O}_2)_2$
11. Determine the molar mass of each substance.
 - a. Al
 - b. Al_2O_3
 - c. CoCl_3
12. Determine the molar mass of each substance.
 - a. O_3
 - b. NaI
 - c. $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
13. What is the mass of 4.44 mol of Rb?
14. What is the mass of 0.311 mol of Xe?
15. What is the mass of 12.34 mol of $\text{Al}_2(\text{SO}_4)_3$?
16. What is the mass of 0.0656 mol of PbCl_2 ?
17. How many moles are present in 45.6 g of CO?
18. How many moles are present in 0.00339 g of LiF?
19. How many moles are present in 1.223 g of SF_6 ?
20. How many moles are present in 48.8 g of BaCO_3 ?
21. How many moles are present in 54.8 mL of mercury if the density of mercury is 13.6 g/mL?
22. How many moles are present in 56.83 mL of O_2 if the density of O_2 is 0.00133 g/mL?

Answers

1. 2.74×10^{24} atoms
3. 1.511×10^{24} molecules
5. 90 mol
7. 0.166 mol
9.
 - a. 28.086 g
 - b. 32.118 g
 - c. 94.195 g
11.
 - a. 26.981 g
 - b. 101.959 g

c. 165.292 g

13. 379 g

15. 4,222 g

17. 1.63 mol

19. 0.008374 mol

21. 3.72 mol

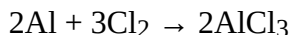
Mole-Mass and Mass-Mass Calculations

Learning Objectives

1. From a given number of moles of a substance, calculate the mass of another substance involved using the balanced chemical equation.
2. From a given mass of a substance, calculate the moles of another substance involved using the balanced chemical equation.
3. From a given mass of a substance, calculate the mass of another substance involved using the balanced chemical equation.

Mole-mole calculations are not the only type of calculations that can be performed using balanced chemical equations. Recall that the molar mass can be determined from a chemical formula and used as a conversion factor. We can add that conversion factor as another step in a calculation to make a **mole-mass calculation**, where we start with a given number of moles of a substance and calculate the mass of another substance involved in the chemical equation, or vice versa.

For example, suppose we have the balanced chemical equation:



Suppose we know we have 123.2 g of Cl_2 . How can we determine how many moles of AlCl_3 we will get when the reaction is complete? First and foremost, *chemical equations are not balanced in terms of grams; they are balanced in terms of moles*. So to use the balanced chemical equation to relate an amount of Cl_2 to an amount of AlCl_3 , we need to convert the given amount of Cl_2 into moles. We know how to do this by simply using the molar mass of Cl_2 as a conversion factor. The molar mass of Cl_2 (which we get from the atomic mass of Cl from the periodic table) is 70.90 g/mol. We must invert this fraction so that the units cancel properly:

$$123.2 \cancel{\text{g Cl}_2} \times \frac{1 \text{ mol Cl}_2}{70.90 \cancel{\text{g Cl}_2}} = 1.738 \text{ mol Cl}_2$$

Now that we have the quantity in moles, we can use the balanced chemical equation to construct a conversion factor that relates the number of moles of Cl_2 to the number of moles of AlCl_3 . The numbers in the conversion factor come from the coefficients in the balanced chemical equation:

$$\frac{2 \text{ mol AlCl}_3}{3 \text{ mol Cl}_2}$$

Using this conversion factor with the molar quantity we calculated above, we get:

$$1.738 \cancel{\text{ mol Cl}_2} \times \frac{2 \text{ mol AlCl}_3}{3 \cancel{\text{ mol Cl}_2}} = 1.159 \text{ mol AlCl}_3$$

So, we will get 1.159 mol of AlCl₃ if we react 123.2 g of Cl₂.

In this last example, we did the calculation in two steps. However, it is mathematically equivalent to perform the two calculations sequentially on one line:

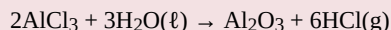
$$123.2 \cancel{\text{ g Cl}_2} \times \frac{1 \cancel{\text{ mol Cl}_2}}{70.90 \cancel{\text{ g Cl}_2}} \times \frac{2 \text{ mol AlCl}_3}{3 \cancel{\text{ mol Cl}_2}} = 1.159 \text{ mol AlCl}_3$$

The units still cancel appropriately, and we get the same numerical answer in the end. Sometimes the answer may be slightly different from doing it one step at a time because of rounding of the intermediate answers, but the final answers should be effectively the same.

Example 5.8

Problem

How many moles of HCl will be produced when 249 g of AlCl₃ are reacted according to this chemical equation?



Solution

We will do this in two steps: convert the mass of AlCl₃ to moles and then use the balanced chemical equation to find the number of moles of HCl formed. The molar mass of AlCl₃ is 133.33 g/mol, which we have to invert to get the appropriate conversion factor:

$$1.87 \cancel{\text{ mol AlCl}_3} \times \frac{6 \text{ mol HCl}}{2 \cancel{\text{ mol AlCl}_3}} = 5.61 \text{ mol HCl}$$

Now we can use this quantity to determine the number of moles of HCl that will form. From the balanced chemical equation, we construct a conversion factor between the number of moles of AlCl₃ and the number of moles of HCl:

$$\frac{6 \text{ mol HCl}}{2 \text{ mol AlCl}_3}$$

Applying this conversion factor to the quantity of AlCl₃, we get:

$$1.87 \cancel{\text{ mol AlCl}_3} \times \frac{6 \text{ mol HCl}}{2 \cancel{\text{ mol AlCl}_3}} = 5.61 \text{ mol HCl}$$

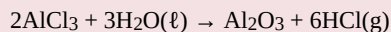
Alternatively, we could have done this in one line:

$$249 \cancel{\text{ g AlCl}_3} \times \frac{1 \cancel{\text{ mol AlCl}_3}}{133.33 \cancel{\text{ g AlCl}_3}} \times \frac{6 \text{ mol HCl}}{2 \cancel{\text{ mol AlCl}_3}} = 5.60 \text{ mol HCl}$$

The last digit in our final answer is slightly different because of rounding differences, but the answer is essentially the same.

Test Yourself

How many moles of Al₂O₃ will be produced when 23.9 g of H₂O are reacted according to this chemical equation?



Answer

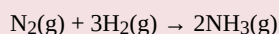
0.442 mol

A variation of the mole-mass calculation is to start with an amount in moles and then determine an amount of another substance in grams. The steps are the same but are performed in reverse order.

Example 5.9

Problem

How many grams of NH_3 will be produced when 33.9 mol of H_2 are reacted according to this chemical equation?

**Solution**

The conversions are the same, but they are applied in a different order. Start by using the balanced chemical equation to convert to moles of another substance and then use its molar mass to determine the mass of the final substance. In two steps, we have:

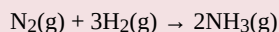
$$33.9 \cancel{\text{ mol H}_2} \times \frac{2 \text{ mol NH}_3}{3 \cancel{\text{ mol H}_2}} = 22.6 \text{ mol NH}_3$$

Now, using the molar mass of NH_3 , which is 17.03 g/mol, we get:

$$22.6 \cancel{\text{ mol NH}_3} \times \frac{17.03 \text{ g NH}_3}{1 \cancel{\text{ mol NH}_3}} = 385 \text{ g NH}_3$$

Test Yourself

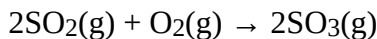
How many grams of N_2 are needed to produce 2.17 mol of NH_3 when reacted according to this chemical equation?

**Answer**

30.4 g (Note: here we go from a product to a reactant, showing that mole-mass problems can begin and end with any substance in the chemical equation.)

It should be a trivial task now to extend the calculations to **mass-mass calculations**, in which we start with a mass of some substance and end with the mass of another substance in the chemical reaction. For this type of calculation, the molar masses of two different substances must be used—be sure to keep track of which is which. Again, however, it is important to emphasize that before the balanced chemical reaction is used, the mass quantity must first be converted to moles. Then the coefficients of the balanced chemical reaction can be used to convert to moles of another substance, which can then be converted to a mass.

For example, let us determine the number of grams of SO₃ that can be produced by the reaction of 45.3 g of SO₂ and O₂:



First, we convert the given amount, 45.3 g of SO₂, to moles of SO₂ using its molar mass (64.06 g/mol):

$$45.3 \text{ g } \cancel{\text{SO}_2} \times \frac{1 \text{ mol } \text{SO}_2}{64.06 \text{ g } \cancel{\text{SO}_2}} = 0.707 \text{ mol } \text{SO}_2$$

Second, we use the balanced chemical reaction to convert from moles of SO₂ to moles of SO₃:

$$0.707 \text{ mol } \cancel{\text{SO}_2} \times \frac{2 \text{ mol } \text{SO}_3}{2 \text{ mol } \cancel{\text{SO}_2}} = 0.707 \text{ mol } \text{SO}_3$$

Finally, we use the molar mass of SO₃ (80.06 g/mol) to convert to the mass of SO₃:

$$0.707 \text{ mol } \cancel{\text{SO}_3} \times \frac{80.06 \text{ g } \text{SO}_3}{1 \text{ mol } \cancel{\text{SO}_3}} = 56.6 \text{ g } \text{SO}_3$$

We can also perform all three steps sequentially, writing them on one line as:

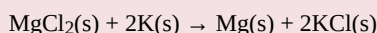
$$45.3 \text{ g } \cancel{\text{SO}_2} \times \frac{1 \text{ mol } \cancel{\text{SO}_2}}{64.06 \text{ g } \cancel{\text{SO}_2}} \times \frac{2 \text{ mol } \cancel{\text{SO}_3}}{2 \text{ mol } \cancel{\text{SO}_2}} \times \frac{80.06 \text{ g } \text{SO}_3}{1 \text{ mol } \cancel{\text{SO}_3}} = 56.6 \text{ g } \text{SO}_3$$

We get the same answer. Note how the initial and all the intermediate units cancel, leaving grams of SO₃, which is what we are looking for, as our final answer.

Example 5.10

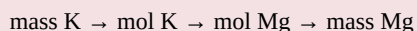
Problem

What mass of Mg will be produced when 86.4 g of K are reacted?



Solution

We will simply follow the steps:

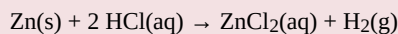


In addition to the balanced chemical equation, we need the molar masses of K (39.09 g/mol) and Mg (24.31 g/mol). In one line,

$$86.4 \text{ g } \cancel{\text{K}} \times \frac{1 \text{ mol } \cancel{\text{K}}}{39.09 \text{ g } \cancel{\text{K}}} \times \frac{1 \text{ mol } \text{Mg}}{2 \text{ mol } \cancel{\text{K}}} \times \frac{24.31 \text{ g } \text{Mg}}{1 \text{ mol } \cancel{\text{Mg}}} = 26.87 \text{ Mg}$$

Test Yourself

What mass of H₂ will be produced when 122 g of Zn are reacted?

**Answer**

3.77 g

Key Takeaways

- Mole quantities of one substance can be related to mass quantities using a balanced chemical equation.
- Mass quantities of one substance can be related to mass quantities using a balanced chemical equation.
- In all cases, quantities of a substance must be converted to moles before the balanced chemical equation can be used to convert to moles of another substance.

Exercises**Questions**

1. What mass of CO₂ is produced by the combustion of 1.00 mol of CH₄?

$$\text{CH}_4\text{(g)} + 2\text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + 2\text{H}_2\text{O(l)}$$
2. What mass of H₂O is produced by the combustion of 1.00 mol of CH₄?

$$\text{CH}_4\text{(g)} + 2\text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + 2\text{H}_2\text{O(l)}$$
3. What mass of HgO is required to produce 0.692 mol of O₂?

$$2\text{HgO(s)} \rightarrow 2\text{Hg(l)} + \text{O}_2\text{(g)}$$
4. What mass of NaHCO₃ is needed to produce 2.659 mol of CO₂?

$$2\text{NaHCO}_3\text{(s)} \rightarrow \text{Na}_2\text{CO}_3\text{(s)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$$
5. How many moles of Al can be produced from 10.87 g of Ag?

$$\text{Al(NO}_3)_3\text{(s)} + 3\text{Ag} \rightarrow \text{Al} + 3\text{AgNO}_3$$
6. How many moles of HCl can be produced from 0.226 g of SOCl₂?

$$\text{SOCl}_2\text{(l)} + \text{H}_2\text{O(l)} \rightarrow \text{SO}_2\text{(g)} + 2\text{HCl(g)}$$
7. How many moles of O₂ are needed to prepare 1.00 g of Ca(NO₃)₂?

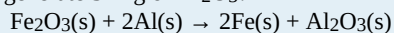
$$\text{Ca(s)} + \text{N}_2\text{(g)} + 3\text{O}_2\text{(g)} \rightarrow \text{Ca(NO}_3)_2\text{(s)}$$
8. How many moles of C₂H₅OH are needed to generate 106.7 g of H₂O?

$$\text{C}_2\text{H}_5\text{OH(l)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{CO}_2\text{(g)} + 3\text{H}_2\text{O(l)}$$
9. What mass of O₂ can be generated by the decomposition of 100.0 g of NaClO₃?

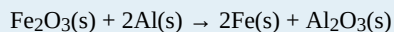
$$2\text{NaClO}_3 \rightarrow 2\text{NaCl(s)} + 3\text{O}_2\text{(g)}$$
10. What mass of Li₂O is needed to react with 1,060 g of CO₂?

$$\text{Li}_2\text{O(aq)} + \text{CO}_2\text{(g)} \rightarrow \text{Li}_2\text{CO}_3\text{(aq)}$$

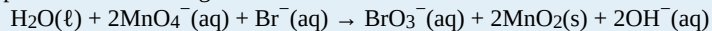
11. What mass of Fe_2O_3 must be reacted to generate 324 g of Al_2O_3 ?



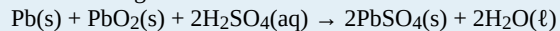
12. What mass of Fe is generated when 100.0 g of Al are reacted?



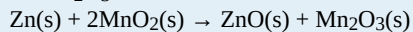
13. What mass of MnO_2 is produced when 445 g of H_2O are reacted?



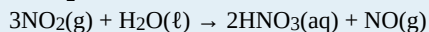
14. What mass of PbSO_4 is produced when 29.6 g of H_2SO_4 are reacted?



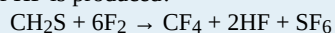
15. If 83.9 g of ZnO are formed, what mass of Mn_2O_3 is formed with it?



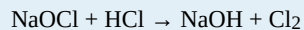
16. If 14.7 g of NO_2 are reacted, what mass of H_2O is reacted with it?



17. If 88.4 g of CH_2S are reacted, what mass of HF is produced?



18. If 100.0 g of Cl_2 are needed, what mass of NaOCl must be reacted?



Answers

1. 44.0 g
3. 3.00×10^2 g
5. 0.0336 mol
7. 0.0183 mol
9. 45.1 g
11. 507 g
13. 4.30×10^3 g
15. 163 g
17. 76.7 g

Limiting Reagents

Learning Objectives

1. Identify a limiting reagent from a set of reactants.
2. Calculate how much product will be produced from the limiting reagent.
3. Calculate how much reactant(s) remains when the reaction is complete.

One additional assumption we have made about chemical reactions—in addition to the assumption that reactions proceed all the way to completion—is that all the reactants are present in the proper quantities to react to products. This is not always the case.

Consider Figure 5.1 “Making Water”. Here we are taking hydrogen atoms and oxygen atoms (left) to make water molecules (right). However, there are not enough oxygen atoms to use up all the hydrogen atoms. We run out of oxygen atoms and cannot make any more water molecules, so the process stops when we run out of oxygen atoms.

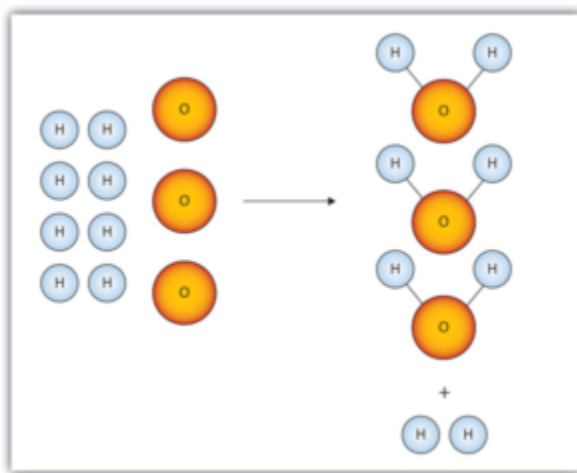
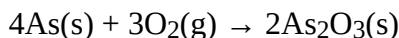


Figure 5.1 “Making Water.” In this scenario for making water molecules, we run out of O atoms before we use up all the H atoms. Similar situations exist for many chemical reactions when one reactant runs out before the other.

A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the **limiting reagent**; the other reactant or reactants are considered to be *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reagent and which is in excess.

The key to recognizing which reactant is the limiting reagent is based on a mole-mass or mass-mass calculation: whichever reactant gives the *lesser* amount of product is the limiting reagent. What we need to do is determine an amount of one product (either moles or mass) assuming all of each reactant reacts. Whichever reactant gives the least amount of that particular product is the limiting reagent. It does not matter which product we use, as long as we use the same one each time. It does not matter whether we determine the number of moles or grams of that product; however, we will see shortly that knowing the final mass of product can be useful.

For example, consider this reaction:



Suppose we start a reaction with 50.0 g of As and 50.0 g of O₂. Which one is the limiting reagent? We need to perform two mole-mass calculations, each assuming that each reactant reacts completely. Then we compare the amount of the product produced by each and determine which is less.

The calculations are as follows:

$$50.0 \text{ g } \cancel{\text{As}} \times \frac{1 \cancel{\text{ mol As}}}{74.92 \text{ g } \cancel{\text{As}}} \times \frac{2 \text{ mol As}_2\text{O}_3}{4 \cancel{\text{ mol As}}} = 0.334 \text{ mol As}_2\text{O}_3$$

$$50.0 \text{ g } \cancel{\text{O}_2} \times \frac{1 \cancel{\text{ mol O}_2}}{32.00 \text{ g } \cancel{\text{O}_2}} \times \frac{2 \text{ mol As}_2\text{O}_3}{3 \cancel{\text{ mol O}_2}} = 1.04 \text{ mol As}_2\text{O}_3$$

Comparing these two answers, it is clear that 0.334 mol of As₂O₃ is less than 1.04 mol of As₂O₃, so arsenic is the limiting reagent. If this reaction is performed under these initial conditions, the arsenic will run out before the oxygen runs out. We say that the oxygen is “in excess.”

Identifying the limiting reagent, then, is straightforward. However, there are usually two associated questions:

1. What mass of product (or products) is then actually formed?
2. What mass of what reactant is left over?

The first question is straightforward to answer: simply perform a conversion from the number of moles of product formed to its mass, using its molar mass. For As₂O₃, the molar mass is 197.84 g/mol; knowing that we will form 0.334 mol of As₂O₃ under the given conditions, we will get:

$$0.334 \cancel{\text{ mol As}_2\text{O}_3} \times \frac{197.84 \text{ g As}_2\text{O}_3}{1 \cancel{\text{ mol As}_2\text{O}_3}} = 66.1 \text{ g As}_2\text{O}_3$$

The second question is somewhat more convoluted to answer. First, we must do a mass-mass calculation relating the limiting reagent (here, As) to the other reagent (O₂). Once we determine the mass of O₂ that reacted, we subtract that from the original amount to determine the amount left over. According to the mass-mass calculation:

$$50.0 \text{ g } \cancel{\text{As}} \times \frac{1 \cancel{\text{ mol As}}}{74.92 \text{ g } \cancel{\text{As}}} \times \frac{3 \cancel{\text{ mol O}_2}}{4 \cancel{\text{ mol As}}} \times \frac{32.00 \text{ g O}_2}{1 \cancel{\text{ mol O}_2}} = 16.0 \text{ g O}_2 \text{ reacted}$$

Because we reacted 16.0 g of our original O₂, we subtract that from the original amount, 50.0 g, to get the mass of O₂ remaining:

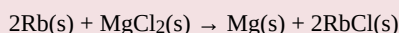
$$50.0 \text{ g O}_2 - 16.0 \text{ g O}_2 \text{ reacted} = 34.0 \text{ g O}_2 \text{ left over}$$

You must remember to perform this final subtraction to determine the amount remaining; a common error is to report the 16.0 g as the amount remaining.

Example 5.11

Problem

A 5.00 g quantity of Rb are combined with 3.44 g of MgCl₂ according to this chemical reaction:



What mass of Mg is formed, and what mass of what reactant is left over?

Solution

Because the question asks what mass of magnesium is formed, we can perform two mass-mass calculations and determine which amount is less.

$$5.00 \text{ g Rb} \times \frac{1 \text{ mol Rb}}{85.47 \text{ g Rb}} \times \frac{1 \text{ mol Mg}}{2 \text{ mol Rb}} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 0.711 \text{ g Mg}$$

$$3.44 \text{ g MgCl}_2 \times \frac{1 \text{ mol MgCl}_2}{95.21 \text{ g MgCl}_2} \times \frac{1 \text{ mol Mg}}{1 \text{ mol MgCl}_2} \times \frac{24.31 \text{ g Mg}}{1 \text{ mol Mg}} = 0.878 \text{ g Mg}$$

The 0.711 g of Mg is the lesser quantity, so the associated reactant — 5.00 g of Rb — is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of MgCl₂ reacted with the 5.00 g of Rb and then subtract the amount reacted from the original amount.

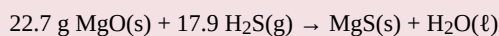
$$5.00 \text{ g Rb} \times \frac{1 \text{ mol Rb}}{85.47 \text{ g Rb}} \times \frac{1 \text{ mol MgCl}_2}{2 \text{ mol Rb}} \times \frac{95.21 \text{ g MgCl}_2}{1 \text{ mol MgCl}_2} = 2.78 \text{ g MgCl}_2 \text{ reacted}$$

Because we started with 3.44 g of MgCl₂, we have:

$$3.44 \text{ g MgCl}_2 - 2.78 \text{ g MgCl}_2 \text{ reacted} = 0.66 \text{ g MgCl}_2 \text{ left}$$

Test Yourself

Given the initial amounts listed, what is the limiting reagent, and what is the mass of the leftover reagent?



Answer

H₂S is the limiting reagent; 1.5 g of MgO are left over.

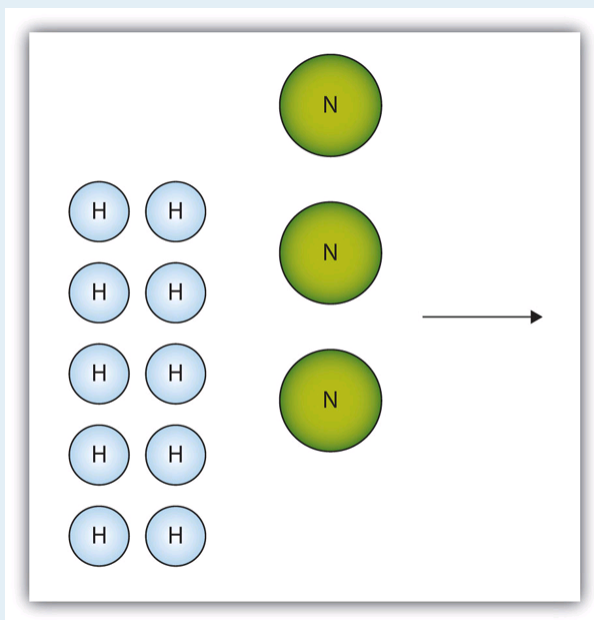
Key Takeaways

- The limiting reagent is that reactant that produces the least amount of product.
- Mass-mass calculations can determine how much product is produced and how much of the other reactants remain.

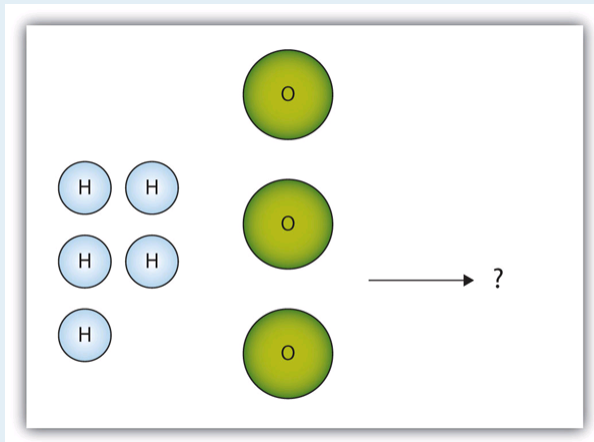
Exercises

Questions

1. The box below shows a group of nitrogen and hydrogen molecules that will react to produce ammonia, NH_3 . What is the limiting reagent?



2. The box below shows a group of hydrogen and oxygen molecules that will react to produce water, H_2O . What is the limiting reagent?



- Given the statement “20.0 g of methane is burned in excess oxygen,” is it obvious which reactant is the limiting reagent?
- Given the statement “the metal is heated in the presence of excess hydrogen,” is it obvious which substance is the limiting reagent despite not specifying any quantity of reactant?
- Acetylene (C_2H_2) is formed by reacting 7.08 g of C and 4.92 g of H_2 .

$$2\text{C}(\text{s}) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_2(\text{g})$$
 What is the limiting reagent? How much of the other reactant is in excess?
- Ethane (C_2H_6) is formed by reacting 7.08 g of C and 4.92 g of H_2 .

$$2\text{C}(\text{s}) + 3\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})$$
 What is the limiting reagent? How much of the other reactant is in excess?
- Given the initial amounts listed, what is the limiting reagent, and how much of the other reactant is in excess?

$$35.6 \text{ g P}_4\text{O}_6(\text{s}) + 4.77 \text{ g 6H}_2\text{O}(\ell) \rightarrow 4\text{H}_3\text{PO}_4$$
- Given the initial amounts listed, what is the limiting reagent, and how much of the other reactant is in excess?

$$377 \text{ g 3NO}_2(\text{g}) + 244 \text{ g H}_2\text{O}(\ell) \rightarrow 2\text{HNO}_3(\text{aq}) + \text{NO}(\text{g})$$
- To form the precipitate PbCl_2 , 2.88 g of NaCl and 7.21 g of $\text{Pb}(\text{NO}_3)_2$ are mixed in solution. How much precipitate is formed? How much of which reactant is in excess?
- In a neutralization reaction, 18.06 g of KOH are reacted with 13.43 g of HNO_3 . What mass of H_2O is produced, and what mass of which reactant is in excess?

Answers

- Nitrogen is the limiting reagent.
- Yes; methane is the limiting reagent.
- C is the limiting reagent; 4.33 g of H_2 are left over.
- H_2O is the limiting reagent; 25.9 g of P_4O_6 are left over.
- 6.06 g of PbCl_2 are formed; 0.33 g of NaCl is left over.

Media Attributions

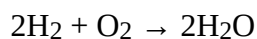
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The Mole in Chemical Reactions

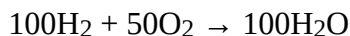
Learning Objectives

1. Balance a chemical equation in terms of moles.
2. Use the balanced equation to construct conversion factors in terms of moles.
3. Calculate moles of one substance from moles of another substance using a balanced chemical equation.

Consider this balanced chemical equation:



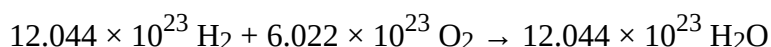
We interpret this as “two molecules of hydrogen react with one molecule of oxygen to make two molecules of water.” The chemical equation is balanced as long as the coefficients are in the ratio 2:1:2. For instance, this chemical equation is also balanced:



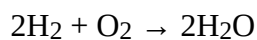
This equation is not conventional—because convention says that we use the lowest ratio of coefficients—but it is balanced. So is this chemical equation:



Again, this is not conventional, but it is still balanced. Suppose we use a much larger number:



These coefficients are also in the ratio of 2:1:2. But these numbers are related to the number of things in a mole: the first and last numbers are two times Avogadro’s number, while the second number is Avogadro’s number. That means that the first and last numbers represent 2 mol, while the middle number is just 1 mol. Well, why not just use the number of moles in balancing the chemical equation?

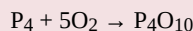


This is the same balanced chemical equation we started with! What this means is that chemical equations are not just balanced in terms of molecules; *they are also balanced in terms of moles*. We can just as easily read this chemical equation as “two moles of hydrogen react with one mole of oxygen to make two moles of water.” All balanced chemical reactions are balanced in terms of moles.

Example 5.12

Problem

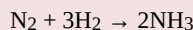
Interpret this balanced chemical equation in terms of moles.

**Solution**

The coefficients represent the number of moles that react, not just molecules. We would speak of this equation as “one mole of molecular phosphorus reacts with five moles of elemental oxygen to make one mole of tetraphosphorus decoxide.”

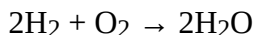
Test Yourself

Interpret this balanced chemical equation in terms of moles.

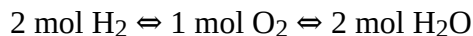
**Answer**

One mole of elemental nitrogen reacts with three moles of elemental hydrogen to produce two moles of ammonia.

In [Chapter 4 “Chemical Reactions and Equations”](#), in the section called [“The Chemical Equation”](#), we stated that a chemical equation is simply a recipe for a chemical reaction. As such, chemical equations also give us equivalences—equivalences between the reactants and the products. However, now we understand that *these equivalences are expressed in terms of moles*. Consider the following chemical equation:



This chemical reaction gives us the following equivalences:



Any two of these quantities can be used to construct a conversion factor that lets us relate the number of moles of one substance to an equivalent number of moles of another substance. If, for example, we want to know how many moles of oxygen will react with 17.6 mol of hydrogen, we construct a conversion factor between 2 mol of H₂ and 1 mol of O₂ and use it to convert from moles of one substance to moles of another:

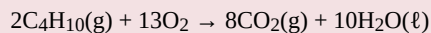
$$17.6 \cancel{\text{ mol H}_2} \times \frac{1 \text{ mol O}_2}{2 \cancel{\text{ mol H}_2}} = 8.80 \text{ mol O}_2$$

Note how the mol H₂ unit cancels, and mol O₂ is the new unit introduced. This is an example of a **mole-mole calculation**, when you start with moles of one substance and convert to moles of another substance by using the balanced chemical equation. The example may seem simple because the numbers are small, but numbers won’t always be so simple!

Example 5.13

Problem

Consider the following balanced chemical equation:



If 154 mol of O_2 are reacted, how many moles of CO_2 are produced?

Solution

We are relating an amount of oxygen to an amount of carbon dioxide, so we need the equivalence between these two substances. According to the balanced chemical equation, the equivalence is:



We can use this equivalence to construct the proper conversion factor. We start with what we are given and apply the conversion factor:

$$154 \cancel{\text{ mol O}_2} \times \frac{8 \text{ mol CO}_2}{13 \cancel{\text{ mol O}_2}} = 94.8 \text{ mol CO}_2$$

The mol O_2 unit is in the denominator of the conversion factor so it cancels. Both the 8 and the 13 are exact numbers, so they don't contribute to the number of significant figures in the final answer.

Test Yourself

Using the above equation, how many moles of H_2O are produced when 154 mol of O_2 react?

Answer

118 mol

It is important to reiterate that balanced chemical equations are balanced in terms of *moles*. Not grams, kilograms, or litres — but moles. Any stoichiometry problem will likely need to work through the mole unit at some point, especially if you are working with a balanced chemical reaction.

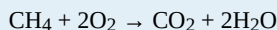
Key Takeaways

- Balanced chemical reactions are balanced in terms of moles.
- A balanced chemical reaction gives equivalences in moles that allow stoichiometry calculations to be performed.

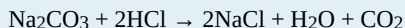
Exercises

Questions

- Express in mole terms what this chemical equation means.



- Express in mole terms what this chemical equation means.



- How many molecules of each substance are involved in the equation in Exercise 1 if it is interpreted in terms of moles?
- How many molecules of each substance are involved in the equation in Exercise 2 if it is interpreted in terms of moles?
- For the chemical equation $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$, what equivalences can you write in terms of moles? Use the \Leftrightarrow sign.
- For the chemical equation $2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3$, what equivalences can you write in terms of moles? Use the \Leftrightarrow sign.
- Write the balanced chemical reaction for the combustion of C_5H_{12} (the products are CO_2 and H_2O) and determine how many moles of H_2O are formed when 5.8 mol of O_2 are reacted.
- Write the balanced chemical reaction for the formation of $\text{Fe}_2(\text{SO}_4)_3$ from Fe_2O_3 and SO_3 and determine how many moles of $\text{Fe}_2(\text{SO}_4)_3$ are formed when 12.7 mol of SO_3 are reacted.
- For the balanced chemical equation $3\text{Cu}(\text{s}) + 2\text{NO}_3^-(\text{aq}) + 8\text{H}^+(\text{aq}) \rightarrow 3\text{Cu}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\ell) + 2\text{NO}(\text{g})$, how many moles of Cu^{2+} are formed when 55.7 mol of H^+ are reacted?
- For the balanced chemical equation $\text{Al}(\text{s}) + 3\text{Ag}^+(\text{aq}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3\text{Ag}(\text{s})$, how many moles of Ag are produced when 0.661 mol of Al are reacted?
- For the balanced chemical reaction $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\ell)$, how many moles of H_2O are produced when 0.669 mol of NH_3 react?
- For the balanced chemical reaction $4\text{NaOH}(\text{aq}) + 2\text{S}(\text{s}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\ell)$, how many moles of Na_2SO_4 are formed when 1.22 mol of O_2 react?
- For the balanced chemical reaction $4\text{KO}_2(\text{s}) + 2\text{CO}_2(\text{g}) \rightarrow 2\text{K}_2\text{CO}_3(\text{s}) + 3\text{O}_2(\text{g})$, determine the number of moles of both products formed when 6.88 mol of KO_2 react.
- For the balanced chemical reaction $2\text{AlCl}_3 + 3\text{H}_2\text{O}(\ell) \rightarrow \text{Al}_2\text{O}_3 + 6\text{HCl}(\text{g})$, determine the number of moles of both products formed when 0.0552 mol of AlCl_3 react.

Answers

- One mole of CH_4 reacts with 2 mol of O_2 to make 1 mol of CO_2 and 2 mol of H_2O .
- 6.022×10^{23} molecules of CH_4 , 1.2044×10^{24} molecules of O_2 , 6.022×10^{23} molecules of CO_2 , and 1.2044×10^{24} molecules of H_2O
- 2 mol of $\text{C}_2\text{H}_6 \Leftrightarrow 7$ mol of $\text{O}_2 \Leftrightarrow 4$ mol of $\text{CO}_2 \Leftrightarrow 6$ mol of H_2O
- $\text{C}_5\text{H}_{12} + 8 \text{O}_2 \rightarrow 5\text{CO}_2 + 6\text{H}_2\text{O}$; 4.4 mol
- 20.9 mol
- 1.00 mol
- 3.44 mol of K_2CO_3 ; 5.16 mol of O_2

Yields

Learning Objectives

1. Define and determine theoretical yields, actual yields, and percent yields.

In all the previous calculations we have performed involving balanced chemical equations, we made two assumptions:

1. The reaction goes exactly as written.
2. The reaction proceeds completely.

In reality, such things as side reactions occur that make some chemical reactions rather messy. For example, in the actual combustion of some carbon-containing compounds, such as methane, some CO is produced as well as CO₂. However, we will continue to ignore side reactions, unless otherwise noted.

The second assumption, that the reaction proceeds completely, is more troublesome. Many chemical reactions do not proceed to completion as written, for a variety of reasons (some of which we will consider in [Chapter 13 “Chemical Equilibrium”](#)). When we calculate an amount of product assuming that all the reactant reacts, we calculate the **theoretical yield**, an amount that is theoretically produced as calculated using the balanced chemical reaction.

In many cases, however, this is not what really happens. In many cases, less—sometimes much less—of a product is made during the course of a chemical reaction. The amount that is actually produced in a reaction is called the **actual yield**. By definition, the actual yield is less than or equal to the theoretical yield. If it is not, then an error has been made.

Both theoretical yields and actual yields are expressed in units of moles or grams. It is also common to see something called a percent yield. The **percent yield** is a comparison between the actual yield and the theoretical yield and is defined as:

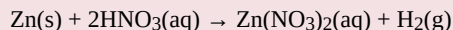
$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

It does not matter whether the actual and theoretical yields are expressed in moles or grams, as long as they are expressed in the same units. However, the percent yield always has units of percent. Proper percent yields are between 0% and 100%—again, if percent yield is greater than 100%, an error has been made.

Example 5.14

Problem

A worker reacts 30.5 g of Zn with nitric acid and evaporates the remaining water to obtain 65.2 g of $\text{Zn}(\text{NO}_3)_2$. What are the theoretical yield, the actual yield, and the percent yield?

**Solution**

A mass-mass calculation can be performed to determine the theoretical yield. We need the molar masses of Zn (65.39 g/mol) and $\text{Zn}(\text{NO}_3)_2$ (189.41 g/mol). In three steps, the mass-mass calculation is:

$$30.5 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \times \frac{1 \text{ mol Zn}(\text{NO}_3)_2}{1 \text{ mol Zn}} \times \frac{189.41 \text{ g Zn}(\text{NO}_3)_2}{1 \text{ mol Zn}(\text{NO}_3)_2} = 88.3 \text{ g Zn}(\text{NO}_3)_2$$

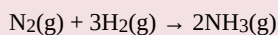
Thus, the theoretical yield is 88.3 g of $\text{Zn}(\text{NO}_3)_2$. The actual yield is the amount that was actually made, which was 65.2 g of $\text{Zn}(\text{NO}_3)_2$. To calculate the percent yield, we take the actual yield and divide it by the theoretical yield and multiply by 100:

$$\frac{65.2 \text{ g Zn}(\text{NO}_3)_2}{88.3 \text{ g Zn}(\text{NO}_3)_2} \times 100\% = 73.8\%$$

The worker achieved almost three-fourths of the possible yield.

Test Yourself

A synthesis produced 2.05 g of NH_3 from 16.5 g of N_2 . What is the theoretical yield and the percent yield?

**Answer**

theoretical yield = 20.1 g; percent yield = 10.2%

Chemistry Is Everywhere: Actual Yields in Drug Synthesis and Purification

Many drugs are the product of several steps of chemical synthesis. Each step typically occurs with less than 100% yield, so the overall percent yield might be very small. The general rule is that the overall percent yield is the product of the percent yields of the individual synthesis steps. For a drug synthesis that has many steps, the overall percent yield can be very tiny, which is one factor in the huge cost of some drugs. For example, if a 10-step synthesis has a percent yield of 90% for each step, the overall yield for the entire synthesis is only 35%. Many scientists work every day trying to improve percent yields of the steps in the synthesis to decrease costs, improve profits, and minimize waste.

Even purifications of complex molecules into drug-quality purity are subject to percent yields. Consider the purification of impure albuterol. Albuterol ($C_{13}H_{21}NO_2$; accompanying figure) is an inhaled drug used to treat asthma, bronchitis, and other obstructive pulmonary diseases. It is synthesized from norepinephrine, a naturally occurring hormone and neurotransmitter. Its initial synthesis makes very impure albuterol that is purified in five chemical steps. The details of the steps do not concern us; only the percent yields do.

Table 5.1 Percent yield of the purification of albuterol

Synthesis	Percent Yield
impure albuterol → intermediate A	70%
intermediate A → intermediate B	100%
intermediate B → intermediate C	40%
intermediate C → intermediate D	72%
intermediate D → purified albuterol	35%
impure albuterol → purified albuterol	$70\% \times 100\% \times 40\% \times 72\% \times 35\% = 7.5\%$

That is, only about *one-fourteenth* of the original material was turned into the purified drug. This gives you one reason why some drugs are so expensive; a lot of material is lost in making a high-purity pharmaceutical.

Key Takeaways

- Theoretical yield is what you calculate the yield will be using the balanced chemical reaction.
- Actual yield is what you actually get in a chemical reaction.
- Percent yield is a comparison of the actual yield with the theoretical yield.

Exercises

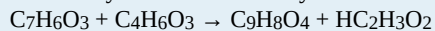
Questions

1. What is the difference between the theoretical yield and the actual yield?
2. What is the difference between the actual yield and the percent yield?

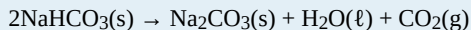
3. A worker isolates 2.675 g of SiF_4 after reacting 2.339 g of SiO_2 with HF. What are the theoretical yield and the actual yield?



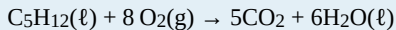
4. A worker synthesizes aspirin, $\text{C}_9\text{H}_8\text{O}_4$, according to this chemical equation. If 12.66 g of $\text{C}_7\text{H}_6\text{O}_3$ are reacted and 12.03 g of aspirin are isolated, what are the theoretical yield and the actual yield?



5. A chemist decomposes 1.006 g of NaHCO_3 and obtains 0.0334 g of Na_2CO_3 . What are the theoretical yield and the actual yield?



6. A chemist combusts a 3.009 g sample of C_5H_{12} and obtains 3.774 g of H_2O . What are the theoretical yield and the actual yield?



7. What is the percent yield in Exercise 3?
 8. What is the percent yield in Exercise 4?
 9. What is the percent yield in Exercise 5?
 10. What is the percent yield in Exercise 6?

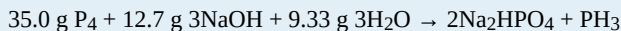
Answers

1. Theoretical yield is what you expect stoichiometrically from a chemical reaction; actual yield is what you actually get from a chemical reaction.
3. theoretical yield = 4.052 g; actual yield = 2.675 g
5. theoretical yield = 0.635 g; actual yield = 0.0334 g
7. 66.02%
9. 5.26%

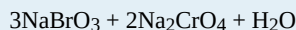
End-of-Chapter Material

Additional Exercises

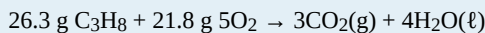
- How many molecules of O_2 will react with 6.022×10^{23} molecules of H_2 to make water? The reaction is $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$.
- How many molecules of H_2 will react with 6.022×10^{23} molecules of N_2 to make ammonia? The reaction is $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$.
- How many moles are present in 6.411 kg of CO_2 ? How many molecules is this?
- How many moles are present in 2.998 mg of SCl_4 ? How many molecules is this?
- What is the mass in milligrams of 7.22×10^{20} molecules of CO_2 ?
- What is the mass in kilograms of 3.408×10^{25} molecules of $SiSi_2$?
- What is the mass in grams of 1 molecule of H_2O ?
- What is the mass in grams of 1 atom of Al?
- What is the volume of 3.44 mol of Ga if the density of Ga is 6.08 g/mL?
- What is the volume of 0.662 mol of He if the density of He is 0.1785 g/L?
- For the chemical reaction $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$, assume that 13.4 g of C_4H_{10} reacts completely to products. The density of CO_2 is 1.96 g/L. What volume in litres of CO_2 is produced?
- For the chemical reaction $2GaCl_3(s) + 3H_2(g) \rightarrow 2Ga(l) + 6HCl(g)$, if 223 g of $GaCl_3$ reacts completely to products and the density of Ga is 6.08 g/mL, what volume in millilitres of Ga is produced?
- Calculate the mass of each product when 100.0 g of $CuCl$ react according to the reaction $2CuCl(aq) \rightarrow CuCl_2(aq) + Cu(s)$. What do you notice about the sum of the masses of the products? What concept is being illustrated here?
- Calculate the mass of each product when 500.0 g of $SnCl_2$ react according to the reaction $2SnCl_2(aq) \rightarrow SnCl_4(aq) + Sn(s)$. What do you notice about the sum of the masses of the products? What concept is being illustrated here?
- What mass of CO_2 is produced from the combustion of 1 gal of gasoline? The chemical formula of gasoline can be approximated as C_8H_{18} . Assume that there are 2,801 g of gasoline per gallon.
- What mass of H_2O is produced from the combustion of 1 gal of gasoline? The chemical formula of gasoline can be approximated as C_8H_{18} . Assume that there are 2,801 g of gasoline per gallon.
- A chemical reaction has a theoretical yield of 19.98 g and a percent yield of 88.40%. What is the actual yield?
- A chemical reaction has an actual yield of 19.98 g and a percent yield of 88.40%. What is the theoretical yield?
- Given the initial amounts listed, what is the limiting reagent, and how much of the other reactants are in excess?



- Given the initial amounts listed, what is the limiting reagent, and how much of the other reactants are in excess?

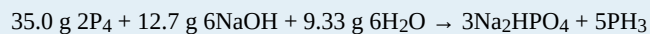


- Verify that it does not matter which product you use to predict the limiting reagent by using both products in this combustion reaction to determine the limiting reagent and the amount of the reactant in excess. Initial amounts of each reactant are given.



- Just in case you suspect Exercise 21 is rigged, do it for another chemical reaction and verify that it does not matter which

product you use to predict the limiting reagent by using both products in this combustion reaction to determine the limiting reagent and the amount of the reactant in excess. Initial amounts of each reactant are given.



Answers

1. 1.2044×10^{24} molecules
3. 145.7 mol; 8.77×10^{25} molecules
5. 52.8 mg
7. 2.99×10^{-23} g
9. 39.4 mL
11. 20.7 L
13. 67.91 g of CuCl_2 ; 32.09 g of Cu. The two masses add to 100.0 g, the initial amount of starting material, demonstrating the law of conservation of matter.
15. 8,632 g
17. 17.66 g
19. The limiting reagent is NaOH; 21.9 g of P_4 and 3.61 g of H_2O are left over.
21. Both products predict that O_2 is the limiting reagent; 20.3 g of C_3H_8 are left over.

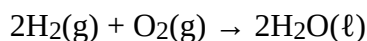
Chapter 6. Gases

Perhaps one of the most spectacular chemical reactions involving a gas occurred on May 6, 1937, when the German airship *Hindenburg* exploded on approach to the Naval Air Station in Lakehurst, New Jersey. The actual cause of the explosion is still unknown, but the entire volume of hydrogen gas used to float the airship, about $200,000 \text{ m}^3$, burned in less than a minute. Thirty-six people, including one on the ground, were killed.



Figure 6.0 “Hindenburg.” The German airship *Hindenburg* (left) was one of the largest airships ever built. However, it was filled with hydrogen gas and exploded in Lakehurst, New Jersey, at the end of a transatlantic voyage in May 1937 (right).

Hydrogen is the lightest known gas. Any balloon filled with hydrogen gas will float in air if its mass is not too great. This makes hydrogen an obvious choice for flying machines based on balloons—airships, dirigibles, and blimps. However, hydrogen also has one obvious drawback: it burns in air according to the well-known chemical equation, which is:



So although hydrogen is an obvious choice, it is also a dangerous choice.

Helium gas is also lighter than air and has 92% of the lifting power of hydrogen. Why, then, was helium not used in the *Hindenburg*? In the 1930s, helium was much more expensive. In addition, the best source of helium at the time was the United States, which banned helium exports to pre-World War II Germany. Today all airships use helium, a legacy of the *Hindenburg* disaster.

Of the three basic phases of matter—solids, liquids, and gases—only one of them has predictable physical properties: gases. In fact, the study of the properties of gases was the beginning of the development of modern chemistry from its alchemical roots. The interesting thing about some of these properties is that they are independent of the identity of the gas. That is, it doesn't matter if the gas is helium gas, oxygen gas, or sulfur vapours; some of their behaviour is predictable and, as we will find, very similar. In this chapter, we will review some of the common behaviours of gases.

Let us start by reviewing some properties of gases. Gases have no definite shape or volume; they tend to fill whatever container they are in. They can compress and expand, sometimes to a great extent. Gases have extremely low densities, one-thousandth or less the density of a liquid or solid. Combinations of gases tend to mix together spontaneously; that is, they form solutions. Air, for

example, is a solution of mostly nitrogen and oxygen. Any understanding of the properties of gases must be able to explain these characteristics.

Media Attributions

Figure 6.0

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- [“Hindenburg burning”](#) by Gus Pasquerella © [Public Domain](#)

Pressure

Learning Objectives

1. Define *pressure*.
2. Learn the units of pressure and how to convert between them

The kinetic theory of gases indicates that gas particles are always in motion and are colliding with other particles and the walls of the container holding them. Although collisions with container walls are elastic (i.e., there is no net energy gain or loss because of the collision), a gas particle does exert a force on the wall during the collision. The accumulation of all these forces distributed over the area of the walls of the container causes something we call pressure. **Pressure** (P) is defined as the force of all the gas particle/wall collisions divided by the area of the wall:

$$\text{pressure} = \frac{\text{force}}{\text{area}}$$

All gases exert pressure; it is one of the fundamental measurable quantities of this phase of matter. Even our atmosphere exerts pressure—in this case, the gas is being “held in” by the earth’s gravity, rather than the gas being in a container. The pressure of the atmosphere is 101,325 Pa.

Pressure has a variety of units. The formal, SI-approved unit of pressure is the *pascal* (Pa), which is defined as 1 N/m^2 (one newton of force over an area of one square metre). However, this is usually too small in magnitude to be useful. A common unit of pressure is the **atmosphere (atm)**, which was originally defined as the average atmospheric pressure at sea level.

However, “average atmospheric pressure at sea level” is difficult to pinpoint because of atmospheric pressure variations. A more reliable and common unit is **millimetre of mercury (mmHg)**, which is the amount of pressure exerted by a column of mercury exactly 1 mm high. An equivalent unit is the **torr**, which equals 1 mmHg. (The torr is named after Evangelista Torricelli, a seventeenth-century Italian scientist who invented the mercury barometer.) With these definitions of pressure, the atmosphere unit is redefined: 1 atm is defined as exactly 760 mmHg, or 760 torr. We thus have the following equivalences:

$$1 \text{ atm} = 760 \text{ mmHg} = 760 \text{ torr}$$

We can use these equivalences as with any equivalences—to perform conversions from one unit to another. Relating these to the formal SI unit of pressure, $1 \text{ atm} = 101,325 \text{ Pa}$.

Example 6.1

Problem

How many atmospheres are there in 595 torr?

Solution

Using the pressure equivalences, we construct a conversion factor between torr and atmospheres:

$$1 \text{ atm} = 760 \text{ torr}$$

Thus:

$$595 \cancel{\text{ torr}} \times \frac{1 \text{ atm}}{760 \cancel{\text{ torr}}} = 0.783 \text{ atm}$$

Because the numbers in the conversion factor are exact, the number of significant figures in the final answer is determined by the initial value of pressure.

Test Yourself

How many atmospheres are there in 1,022 torr?

Answer

1.345 atm

Example 6.2

Problem

The atmosphere on Mars is largely CO₂ at a pressure of 6.01 mmHg. What is this pressure in atmospheres?

Solution

Use the pressure equivalences to construct the proper conversion factor between millimetres of mercury and atmospheres.

$$6.01 \cancel{\text{ mmHg}} \times \frac{1 \text{ atm}}{760 \cancel{\text{ mmHg}}} = 0.00791 \text{ atm} \times 10^{-3} \text{ atm}$$

At the end, we expressed the answer in scientific notation.

Test Yourself

Atmospheric pressure is low in the eye of a hurricane. In a 1979 hurricane in the Pacific Ocean, a pressure of 0.859 atm was reported inside the eye. What is this pressure in torr?

Answer

652 torr

Key Takeaways

- Pressure is a force exerted over an area.
- Pressure has several common units that can be converted

Exercises**Questions**

1. Define *pressure*. What causes it?
2. Define and relate three units of pressure.
3. If a force of 16.7 N is pressed against an area of 2.44 m^2 , what is the pressure in pascals?
4. If a force of 2,546 N is pressed against an area of 0.0332 m^2 , what is the pressure in pascals?
5. Explain why the original definition of atmosphere did not work well.
6. What units of pressure are equal to each other?
7. How many atmospheres are in 889 mmHg?
8. How many atmospheres are in 223 torr?
9. How many torr are in 2.443 atm?
10. How many millimetres of mercury are in 0.334 atm?
11. How many millimetres of mercury are in 334 torr?
12. How many torr are in 0.777 mmHg?
13. How many pascals are in 1 torr?
14. A pressure of 0.887 atm equals how many pascals?

Answers

1. Pressure is force per unit area. It is caused by gas particles hitting the walls of their container.
3. 6.84 Pa
5. Because the atmospheric pressure at sea level is variable, it is not a consistent unit of pressure.
7. 1.17 atm
9. 1,857 torr
11. 334 mmHg

13. 133 Pa

Gas Laws

Learning Objectives

1. Learn what is meant by the term *gas laws*.
2. Learn and apply Boyle's law.
3. Learn and apply Charles's law.

When seventeenth-century scientists began studying the physical properties of gases, they noticed simple relationships between some of the measurable properties of gases. Take pressure (P) and volume (V), for example. Scientists noted that for a given amount of a gas (usually expressed in units of moles [n]), if the temperature (T) of the gas is kept constant, pressure and volume are related: as one increases, the other decreases. As one decreases, the other increases. We say that pressure and volume are *inversely related*.

There is more to it, however: pressure and volume of a given amount of gas at a constant temperature are *numerically* related. If you take the pressure value and multiply it by the volume value, the product is a constant for a given amount of gas at a constant temperature:

$$P \times V = \text{constant at constant } n \text{ and } T$$

If either volume or pressure changes while the amount and temperature stay the same, then the other property must change so that the product of the two properties still equals that same constant. That is, if the original conditions are labelled P_1 and V_1 and the new conditions are labelled P_2 and V_2 , we have:

$$P_1V_1 = \text{constant} = P_2V_2$$

where the properties are assumed to be multiplied together. Leaving out the middle part, we have simply:

$$P_1V_1 = P_2V_2 \text{ at constant } n \text{ and } T$$

This equation is an example of a gas law. A **gas law** is a simple mathematical formula that allows you to model, or predict, the behaviour of a gas. This particular gas law is called **Boyle's law**, after the English scientist Robert Boyle, who first announced it in 1662. Figure 6.1 Boyle's Law shows two representations of what Boyle's law describes.

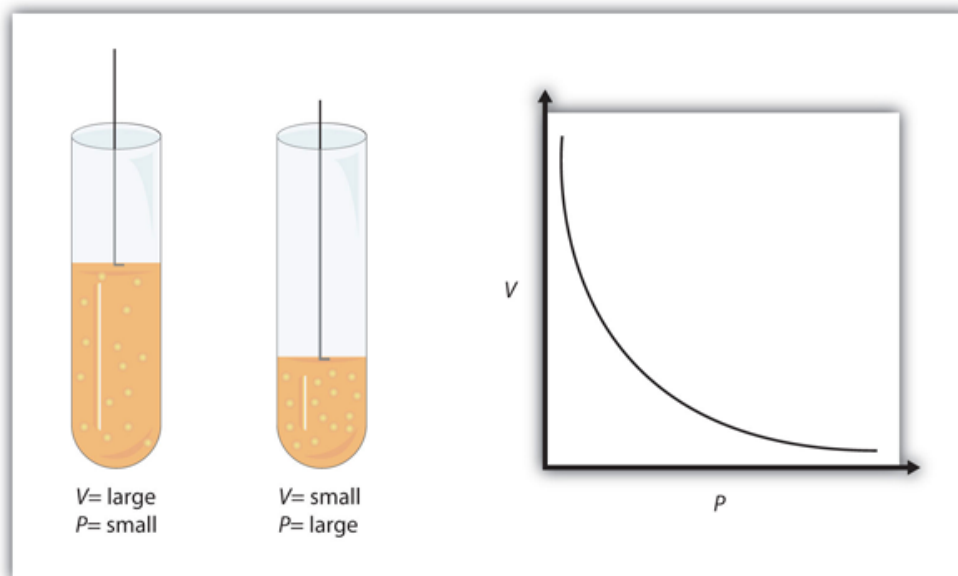


Figure 6.1 “Boyle’s Law.” A piston having a certain pressure and volume (left piston) will have half the volume when its pressure is twice as much (right piston). One can also plot P versus V for a given amount of gas at a certain temperature; such a plot will look like the graph on the right.

A piston having a certain pressure and volume (left piston) will have half the volume when its pressure is twice as much (right piston). One can also plot P versus V for a given amount of gas at a certain temperature; such a plot will look like the graph on the right.

Boyle’s law is an example of a second type of mathematical problem we see in chemistry—one based on a mathematical formula. Tactics for working with mathematical formulas are different from tactics for working with conversion factors. First, most of the questions you will have to answer using formulas are word-type questions, so the first step is to identify what quantities are known and assign them to variables. Second, in most formulas, some mathematical rearrangements (i.e., algebra) must be performed to solve for an unknown variable. The rule is that to find the value of the unknown variable, you must mathematically isolate the unknown variable *by itself and in the numerator* of one side of the equation. Finally, units must be consistent. For example, in Boyle’s law there are two pressure variables, and they must have the same unit. There are also two volume variables; they also must have the same unit. In most cases, it won’t matter *what* the unit is, but the unit must be the *same* on both sides of the equation.

Example 6.3

Problem

A sample of gas has an initial pressure of 2.44 atm and an initial volume of 4.01 L. Its pressure changes to 1.93 atm. What is the new volume if temperature and amount are kept constant?

Solution

First, determine what quantities we are given. We are given an initial pressure and an initial volume, so let these values be P_1 and V_1 :

$$P_1 = 2.44 \text{ atm and } V_1 = 4.01 \text{ L}$$

We are given another quantity, final pressure of 1.93 atm, but not a final volume. This final volume is the variable we will solve for.

$$P_2 = 1.93 \text{ atm and } V_2 = ? \text{ L}$$

Substituting these values into Boyle's law, we get:

$$(2.44 \text{ atm})(4.01 \text{ L}) = (1.93 \text{ atm})V_2$$

To solve for the unknown variable, we isolate it by dividing both sides of the equation by 1.93 atm — both the number *and* the unit:

$$\frac{(2.44 \text{ atm})(4.01 \text{ L})}{1.93 \text{ atm}} = \frac{(1.93 \text{ atm})V_2}{1.93 \text{ atm}}$$

Note that, on the left side of the equation, the unit *atm* is in the numerator and the denominator of the fraction. They cancel algebraically, just as a number would. On the right side, the unit *atm* and the number 1.93 are in the numerator and the denominator, so the entire quantity cancels:

$$\frac{(2.44 \cancel{\text{ atm}})(4.01 \text{ L})}{1.93 \cancel{\text{ atm}}} = \frac{(\cancel{1.93 \text{ atm}})V_2}{\cancel{1.93 \text{ atm}}}$$

What we have left is:

$$\frac{(2.44)(4.01 \text{ L})}{1.93} = V_2$$

Now we simply multiply and divide the numbers together and combine the answer with the L unit, which is a unit of volume. Doing so, we get:

$$V_2 = 5.07 \text{ L}$$

Does this answer make sense? We know that pressure and volume are inversely related; as one decreases, the other increases. Pressure is decreasing (from 2.44 atm to 1.93 atm), so volume should be increasing to compensate, and it is (from 4.01 L to 5.07 L). So the answer makes sense based on Boyle's law.

Test Yourself

If $P_1 = 334 \text{ torr}$, $V_1 = 37.8 \text{ mL}$, and $P_2 = 102 \text{ torr}$, what is V_2 ?

Answer

124 mL

As mentioned, you can use any units for pressure or volume, but both pressures must be expressed in the same units, and both volumes must be expressed in the same units.

Example 6.4

Problem

A sample of gas has an initial pressure of 722 torr and an initial volume of 88.8 mL. Its volume changes to 0.663 L. What is the new pressure?

Solution

We can still use Boyle's law to answer this, but now the two volume quantities have different units. It does not matter which unit we change, as long as we perform the conversion correctly. Let us change the 0.663 L to millilitres:

$$0.663 \cancel{\text{ L}} \times \frac{1000 \text{ mL}}{1 \cancel{\text{ L}}} = 663 \text{ mL}$$

Now that both volume quantities have the same units, we can substitute into Boyle's law:

$$(722 \text{ torr})(88.8 \text{ mL}) = P_2(663 \text{ mL})$$

$$\frac{(722 \text{ torr})(88.8 \text{ mL})}{663 \text{ mL}} = P_2$$

The mL units cancel, and we multiply and divide the numbers to get:

$$P_2 = 96.7 \text{ torr}$$

The volume is increasing, and the pressure is decreasing, which is as expected for Boyle's law.

Test Yourself

If $V_1 = 456 \text{ mL}$, $P_1 = 308 \text{ torr}$, and $P_2 = 1.55 \text{ atm}$, what is V_2 ?

Answer

119 mL

There are other measurable characteristics of a gas. One of them is temperature (T). One could vary the temperature of a gas sample and note what effect it has on the other properties of the gas. Early scientists did just this, discovering that if the amount of a gas and its pressure are kept constant, then changing the temperature changes the volume (V). As temperature increases, volume increases; as temperature decreases, volume decreases. We say that these two characteristics are *directly related*.

A mathematical relationship between V and T should be possible except for one question: What temperature scale should we use? We know from [Chapter 2 "Measurements"](#) that science uses several possible temperature scales. Experiments show that the volume of a gas is related to its *absolute temperature in Kelvin, not its temperature in degrees Celsius*. If the temperature of a gas is expressed in kelvins, then experiments show that the *ratio* of volume to temperature is a constant:

$$\frac{V}{T} = \text{constant}$$

We can modify this equation as we modified Boyle's law: the initial conditions V_1 and T_1 have a certain value, and the value must be the same when the conditions of the gas are changed to some new conditions V_2 and T_2 , as long as pressure and the amount of the gas remain constant. Thus, we have another gas law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ at constant } P \text{ and } n$$

This gas law is commonly referred to as **Charles's law**, after the French scientist Jacques Charles, who performed experiments on gases in the 1780s. The tactics for using this mathematical formula are similar to those for Boyle's law. To determine an unknown quantity, use algebra to isolate the unknown variable by itself and in the numerator; the units of similar variables must be the same. But we add one more tactic: all temperatures must be expressed in the absolute temperature scale (Kelvin). As a reminder, we review the conversion between the absolute temperature scale and the Celsius temperature scale:

$$K = ^\circ C + 273$$

where K represents the temperature in kelvins, and $^\circ C$ represents the temperature in degrees Celsius.

Figure 6.2 "Charles's Law" shows two representations of how Charles's law works.

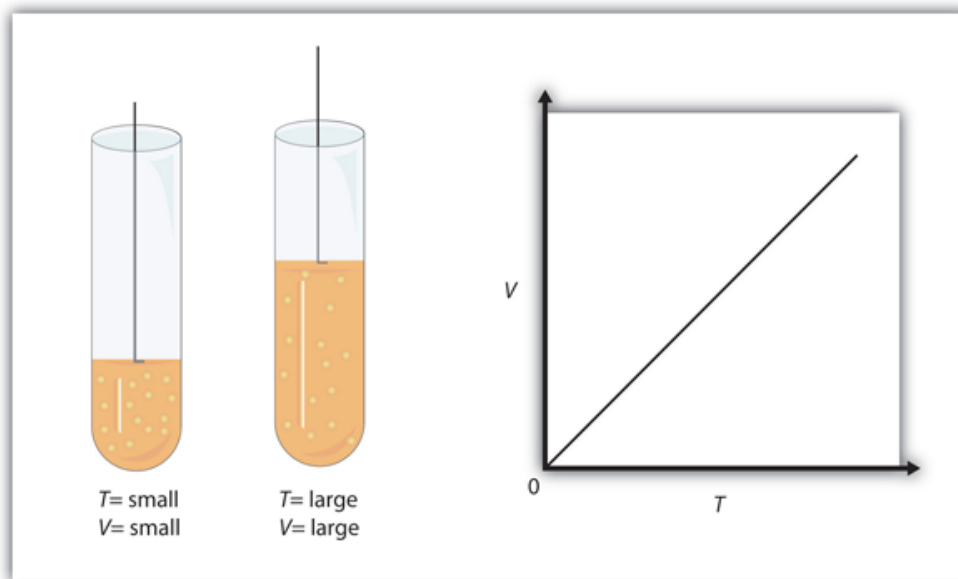


Figure 6.2 "Charles's Law." A piston having a certain volume and temperature (left piston) will have twice the volume when its temperature is twice as much (right piston). One can also plot V versus T for a given amount of gas at a certain pressure; such a plot will look like the graph on the right.

Example 6.5

Problem

A sample of gas has an initial volume of 34.8 mL and an initial temperature of 315 K. What is the new volume if the temperature is increased to 559 K? Assume constant pressure and amount for the gas.

Solution

First, we assign the given values to their variables. The initial volume is V_1 , so $V_1 = 34.8$ mL, and the initial temperature is T_1 , so $T_1 = 315$ K. The temperature is increased to 559 K, so the final temperature $T_2 = 559$ K. We note that the temperatures are already given in kelvins, so we do not need to convert the temperatures. Substituting into the expression for Charles's law yields:

$$\frac{34.8 \text{ mL}}{315 \text{ K}} = \frac{V_2}{559 \text{ K}}$$

We solve for V_2 by algebraically isolating the V_2 variable on one side of the equation. We do this by multiplying both sides of the equation by 559 K (number and unit). When we do this, the temperature unit cancels on the left side, while the entire 559 K cancels on the right side:

$$\frac{(\cancel{559 \text{ K}})(34.8 \text{ mL})}{315 \cancel{\text{ K}}} = \frac{V_2(\cancel{559 \text{ K}})}{\cancel{559 \text{ K}}}$$

The expression simplifies to:

$$\frac{(559)(34.8 \text{ mL})}{315} = V_2$$

By multiplying and dividing the numbers, we see that the only remaining unit is mL, so our final answer is:

$$V_2 = 61.8 \text{ mL}$$

Does this answer make sense? We know that as temperature increases, volume increases. Here, the temperature is increasing from 315 K to 559 K, so the volume should also increase, which it does.

Test Yourself

If $V_1 = 3.77$ L and $T_1 = 255$ K, what is V_2 if $T_2 = 123$ K?

Answer

1.82 L

It is more mathematically complicated if a final temperature must be calculated because the T variable is in the denominator of Charles's law. There are several mathematical ways to work this, but perhaps the simplest way is to take the reciprocal of Charles's law. That is, rather than write it as:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

write the equation as:

$$\frac{T_1}{V_1} = \frac{T_2}{V_2}$$

It is still an equality and a correct form of Charles's law, but now the temperature variable is in the numerator, and the algebra required to predict a final temperature is simpler.

Example 6.6

Problem

A sample of a gas has an initial volume of 34.8 L and an initial temperature of -67°C . What must be the temperature of the gas for its volume to be 25.0 L?

Solution

Here, we are looking for a final temperature, so we will use the reciprocal form of Charles's law. However, the initial temperature is given in degrees Celsius, not kelvins. We must convert the initial temperature to kelvins:

$$-67^\circ\text{C} + 273 = 206 \text{ K}$$

In using the gas law, we must use $T_1 = 206 \text{ K}$ as the temperature. Substituting into the reciprocal form of Charles's law, we get:

$$\frac{206 \text{ K}}{34.8 \text{ L}} = \frac{T_2}{25.0 \text{ L}}$$

Bringing the 25.0 L quantity over to the other side of the equation, we get:

$$\frac{(25.0 \cancel{\text{L}})(206 \text{ K})}{34.8 \cancel{\text{L}}} = T_2$$

The L units cancel, so our final answer is:

$$T_2 = 148 \text{ K}$$

This is also equal to -125°C . As temperature decreases, volume decreases, which it does in this example.

Test Yourself

If $V_1 = 623 \text{ mL}$, $T_1 = 255^\circ\text{C}$, and $V_2 = 277 \text{ mL}$, what is T_2 ?

Answer

235 K, or -38°C

Key Takeaways

- The behaviour of gases can be modelled with gas laws.
- Boyle's law relates a gas's pressure and volume at constant temperature and amount.
- Charles's law relates a gas's volume and temperature at constant pressure and amount.
- In gas laws, temperatures must always be expressed in kelvins.

Exercises

Questions

1. Define *gas law*. What restrictions are there on the units that can be used for the physical properties?
2. What unit of temperature must be used for gas laws?
3. Complete this sentence: Boyle's law relates the [blank] of a gas inversely with the [blank] of that gas.
4. Complete this sentence: Charles's law relates the [blank] of a gas directly with the [blank] of that gas.
5. What properties must be held constant when applying Boyle's law?
6. What properties must be held constant when applying Charles's law?
7. A gas has an initial pressure of 1.445 atm and an initial volume of 1.009 L. What is its new pressure if volume is changed to 0.556 L? Assume temperature and amount are held constant.
8. A gas has an initial pressure of 633 torr and an initial volume of 87.3 mL. What is its new pressure if volume is changed to 45.0 mL? Assume temperature and amount are held constant.
9. A gas has an initial pressure of 4.33 atm and an initial volume of 5.88 L. What is its new volume if pressure is changed to 0.506 atm? Assume temperature and amount are held constant.
10. A gas has an initial pressure of 87.0 torr and an initial volume of 28.5 mL. What is its new volume if pressure is changed to 206 torr? Assume temperature and amount are held constant.
11. A gas has an initial volume of 638 mL and an initial pressure of 779 torr. What is its final volume in litres if its pressure is changed to 0.335 atm? Assume temperature and amount are held constant.
12. A gas has an initial volume of 0.966 L and an initial pressure of 3.07 atm. What is its final pressure in torr if its volume is changed to 3,450 mL? Assume temperature and amount are held constant.
13. A gas has an initial volume of 67.5 mL and an initial temperature of 315 K. What is its new volume if temperature is changed to 244 K? Assume pressure and amount are held constant.
14. A gas has an initial volume of 2.033 L and an initial temperature of 89.3 K. What is its volume if temperature is changed to 184 K? Assume pressure and amount are held constant.
15. A gas has an initial volume of 655 mL and an initial temperature of 295 K. What is its new temperature if volume is changed to 577 mL? Assume pressure and amount are held constant.
16. A gas has an initial volume of 14.98 L and an initial temperature of 238 K. What is its new temperature if volume is changed to 12.33 L? Assume pressure and amount are held constant.
17. A gas has an initial volume of 685 mL and an initial temperature of 29°C. What is its new temperature if volume is changed to 1.006 L? Assume pressure and amount are held constant.
18. A gas has an initial volume of 3.08 L and an initial temperature of -73°C. What is its new volume if its temperature is changed to 104°C? Assume pressure and amount are held constant.

Answers

1. A gas law is a simple mathematical formula that allows one to predict the physical properties of a gas. The units of changing properties (volume, pressure, etc.) must be the same.
3. pressure; volume
5. amount of gas and temperature
7. 2.62 atm
9. 50.3 L
11. 1.95 L

13. 52.3 mL
15. 260 K
17. 444 K, or 171°C

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Other Gas Laws

Learning Objectives

1. Review other simple gas laws.
2. Learn and apply the combined gas law.

You may notice in Boyle's law and Charles's law that we actually refer to four physical properties of a gas: pressure (P), volume (V), temperature (T), and amount (in moles, n). We do this because these are the only four independent physical properties of a gas. There are other physical properties, but they are all related to one (or more) of these four properties.

Boyle's law is written in terms of two of these properties, with the other two being held constant. Charles's law is written in terms of two different properties, with the other two being held constant. It may not be surprising to learn that there are other gas laws that relate other pairs of properties—as long as the other two are held constant. Here we will mention a few.

Gay-Lussac's law relates pressure with absolute temperature. In terms of two sets of data, Gay-Lussac's law is:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ at constant } V \text{ and } n$$

Note that it has a structure very similar to that of Charles's law, only with different variables—pressure instead of volume. **Avogadro's law** introduces the last variable for amount. The original statement of Avogadro's law states that equal volumes of different gases at the same temperature and pressure contain the same number of particles of gas. Because the number of particles is related to the number of moles ($1 \text{ mol} = 6.022 \times 10^{23}$ particles), Avogadro's law essentially states that equal volumes of different gases at the same temperature and pressure contain the same *amount* (moles, particles) of gas. Put mathematically into a gas law, Avogadro's law is:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \text{ at constant } P \text{ and } T$$

(First announced in 1811, it was Avogadro's proposal that volume is related to the number of particles that eventually led to naming the number of things in a mole as Avogadro's number.) Avogadro's law is useful because for the first time we are seeing amount, in terms of the number of moles, as a variable in a gas law.

Example 6.7

Problem

A 2.45 L volume of gas contains 4.5×10^{21} gas particles. How many gas particles are there in 3.87 L if the gas is at constant pressure and temperature?

Solution

We can set up Avogadro's law as follows:

$$\frac{2.45 \text{ L}}{4.5 \times 10^{21} \text{ particles}} = \frac{3.87 \text{ L}}{n_2}$$

We algebraically rearrange to solve for n_2 :

$$n_2 = \frac{(3.87 \cancel{\text{ L}})(4.5 \times 10^{21} \text{ particles})}{2.45 \cancel{\text{ L}}}$$

The L units cancel, so we solve for n_2 :

$$n_2 = 7.1 \times 10^{21} \text{ particles}$$

Test Yourself

A 12.8 L volume of gas contains 3.00×10^{20} gas particles. At constant temperature and pressure, what volume do 8.22×10^{18} gas particles fill?

Answer

0.351 L

The variable n in Avogadro's law can also stand for the number of moles of gas in addition to the number of particles.

One thing we notice about all the gas laws is that, collectively, volume and pressure are always in the numerator, and temperature is always in the denominator. This suggests that we can propose a gas law that combines pressure, volume, and temperature. This gas law is known as the **combined gas law**, and its mathematical form is:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \text{ at constant } n$$

This allows us to follow changes in all three major properties of a gas. Again, the usual warnings apply about how to solve for an unknown algebraically (isolate it on one side of the equation in the numerator), units (they must be the same for the two similar variables of each type), and units of temperature (must be in kelvins).

Example 6.8

Problem

A sample of gas at an initial volume of 8.33 L, an initial pressure of 1.82 atm, and an initial temperature of 286 K simultaneously changes its temperature to 355 K and its volume to 5.72 L. What is the final pressure of the gas?

Solution

We can use the combined gas law directly; all the units are consistent with each other, and the temperatures are given in Kelvin. Substituting,

$$\frac{(1.82 \text{ atm})(8.33 \text{ L})}{286 \text{ K}} = \frac{P_2(5.72 \text{ L})}{355 \text{ K}}$$

We rearrange this to isolate the P_2 variable all by itself. When we do so, certain units cancel:

$$\frac{(1.82 \text{ atm})(8.33 \cancel{\text{ L}})(355 \cancel{\text{ K}})}{(286 \cancel{\text{ K}})(5.72 \cancel{\text{ L}})} = P_2$$

Multiplying and dividing all the numbers, we get:

$$P_2 = 3.29 \text{ atm}$$

Ultimately, the pressure increased, which would have been difficult to predict because two properties of the gas were changing.

Test Yourself

If $P_1 = 662$ torr, $V_1 = 46.7$ mL, $T_1 = 266$ K, $P_2 = 409$ torr, and $T_2 = 371$ K, what is V_2 ?

Answer

105 mL

As with other gas laws, if you need to determine the value of a variable in the denominator of the combined gas law, you can either cross-multiply all the terms or just take the reciprocal of the combined gas law. Remember, the variable you are solving for must be in the numerator and all by itself on one side of the equation.

Key Takeaways

- There are other gas laws that relate any two physical properties of a gas.
- The combined gas law relates pressure, volume, and temperature of a gas.

Exercises

Questions

1. State Gay-Lussac's law.
2. State Avogadro's law.
3. Use Gay-Lussac's law to determine the final pressure of a gas whose initial pressure is 602 torr, initial temperature is 356 K, and final temperature is 277 K. Assume volume and amount are held constant.
4. Use Gay-Lussac's law to determine the final temperature of a gas whose initial pressure is 1.88 atm, initial temperature is 76.3 K, and final pressure is 6.29 atm. Assume volume and amount are held constant.
5. If 3.45×10^{22} atoms of Ar have a volume of 1.55 L at a certain temperature and pressure, what volume do 6.00×10^{23} atoms of Ar have at the same temperature and pressure?
6. If 5.55×10^{22} atoms of He occupy a volume of 2.06 L at 0°C at 1.00 atm pressure, what volume do 2.08×10^{23} atoms of He occupy under the same conditions?
7. Use Avogadro's law to determine the final volume of a gas whose initial volume is 6.72 L, initial amount is 3.88 mol, and final amount is 6.10 mol. Assume pressure and temperature are held constant.
8. Use Avogadro's law to determine the final amount of a gas whose initial volume is 885 mL, initial amount is 0.552 mol, and final volume is 1,477 mL. Assume pressure and temperature are held constant.
9. Use the combined gas law to complete this table. Assume that the amount remains constant in all cases.

V_1	P_1	T_1	V_2	P_2	T_2
56.9 mL	334 torr	266 K		722 torr	334 K
0.976 L	2.33 atm	443 K	1.223 L		355 K
3.66 L	889 torr	23°C	2.19 L	739 torr	

10. Use the combined gas law to complete this table. Assume that the amount remains constant in all cases.

V_1	P_1	T_1	V_2	P_2	T_2
56.7 mL	1.07 atm	-34°C		998 torr	375 K
3.49 L	338 torr	45°C	1,236 mL		392 K
2.09 mL	776 torr	45°C	0.461 mL	0.668 atm	

11. A gas starts at the conditions 78.9 mL, 3.008 atm, and 56°C . Its conditions change to 35.6 mL and 2.55 atm. What is its final temperature?
12. The initial conditions of a sample of gas are 319 K, 3.087 L, and 591 torr. What is its final pressure if volume is changed to 2.222 L and temperature is changed to 299 K?
13. A gas starts with initial pressure of 7.11 atm, initial temperature of 66°C , and initial volume of 90.7 mL. If its conditions change to 33°C and 14.33 atm, what is its final volume?
14. A sample of gas doubles its pressure and doubles its absolute temperature. By what amount does the volume change?

Answers

1. The pressure of a gas is proportional to its absolute temperature.
3. 468 torr

5. 27.0 L

7. 10.6 L

9.

V_1	P_1	T_1	V_2	P_2	T_2
56.9 mL	334 torr	266 K	33.1 mL	722 torr	334 K
0.976 L	2.33 atm	443 K	1.223 L	1.49 atm	355 K
3.66 L	889 torr	23°C	2.19 L	739 torr	147 K, or -126°C

11. 126 K, or -147°C

13. 40.6 mL

The Ideal Gas Law and Some Applications

Learning Objectives

1. Learn the ideal gas law.
2. Apply the ideal gas law to any set of conditions of a gas.
3. Apply the ideal gas law to molar volumes, density, and stoichiometry problems.

So far, the gas laws we have considered have all required that the gas change its conditions; then we predict a resulting change in one of its properties. Are there any gas laws that relate the physical properties of a gas at any given time?

Consider a further extension of the combined gas law to include n . By analogy to Avogadro's law, n is positioned in the denominator of the fraction, opposite the volume. So:

$$\frac{PV}{nT} = \text{constant}$$

Because pressure, volume, temperature, and amount are the only four independent physical properties of a gas, the constant in the above equation is truly a constant; indeed, because we do not need to specify the identity of a gas to apply the gas laws, this constant is the same for all gases. We define this constant with the symbol R , so the previous equation is written as:

$$\frac{PV}{nT} = R$$

which is usually rearranged as:

$$PV = nRT$$

This equation is called the **ideal gas law**. It relates the four independent properties of a gas at any time. The constant R is called the ideal gas law constant. Its value depends on the units used to express pressure and volume. Table 6.1 "Values of the Ideal Gas Law Constant R " lists the numerical values of R .

Table 6.1 Values of the Ideal Gas Law Constant R

Numerical Value	Units
0.08205	$\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$
62.36	$\frac{\text{L} \cdot \text{torr}}{\text{mol} \cdot \text{K}} = \frac{\text{L} \cdot \text{mmHg}}{\text{mol} \cdot \text{K}}$
8.314	$\frac{\text{J}}{\text{mol} \cdot \text{K}}$

The ideal gas law is used like any other gas law, with attention paid to the units and making sure that temperature is expressed in kelvins. However, *the ideal gas law does not require a change in the conditions of a gas sample*. The ideal gas law implies that if you know any three of the physical properties of a gas, you can calculate the fourth property.

Example 6.9

Problem

A 4.22 mol sample of Ar has a pressure of 1.21 atm and a temperature of 34°C. What is its volume?

Solution

The first step is to convert temperature to kelvins:

$$34 + 273 = 307 \text{ K}$$

Now we can substitute the conditions into the ideal gas law:

$$(1.21 \text{ atm})(V) = (4.22 \text{ mol}) \left(0.08205 \cdot \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (307 \text{ K})$$

The atm unit is in the numerator of both sides, so it cancels. On the right side of the equation, the mol and K units appear in the numerator and the denominator, so they cancel as well. The only unit remaining is L, which is the unit of volume we are looking for. We isolate the volume variable by dividing both sides of the equation by 1.21:

$$V = \frac{(4.22)(0.08205)(307)}{1.21} \text{ L}$$

Then solving for volume, we get:

$$V = 87.9 \text{ L}$$

Test Yourself

A 0.0997 mol sample of O₂ has a pressure of 0.692 atm and a temperature of 333 K. What is its volume?

Answer

3.94 L

Example 6.10

Problem

At a given temperature, 0.00332 g of Hg in the gas phase has a pressure of 0.00120 mmHg and a volume of 435 L. What is its temperature?

Solution

We are not given the number of moles of Hg directly, but we are given a mass. We can use the molar mass of Hg to convert to the number of moles.

$$0.00332 \text{ g Hg} \times \frac{1 \text{ mol Hg}}{200.59 \text{ g Hg}} = 0.0000165 \text{ mol} = 1.65 \times 10^{-5} \text{ mol}$$

Pressure is given in units of millimetres of mercury. We can either convert this to atmospheres or use the value of the ideal gas constant that includes the mmHg unit. We will take the second option. Substituting into the ideal gas law:

$$(0.00120 \text{ mmHg})(435 \text{ L}) = (1.65 \times 10^{-5} \text{ mol}) \left(62.36 \cdot \frac{\text{L} \cdot \text{mmHg}}{\text{mol} \cdot \text{K}} \right) T$$

The mmHg, L, and mol units cancel, leaving the K unit, the unit of temperature. Isolating T all by itself on one side, we get:

$$T = \frac{(0.00120)(435)}{(1.65 \times 10^{-5})(62.36)} \text{ K}$$

Then solving for K, we get:

$$T = 507 \text{ K}$$

Test Yourself

For a 0.00554 mol sample of H₂, $P = 23.44$ torr and $T = 557$ K. What is its volume?

Answer

8.21 L

The ideal gas law can also be used in stoichiometry problems.

Example 6.11

Problem

What volume of H₂ is produced at 299 K and 1.07 atm when 55.8 g of Zn metal react with excess HCl?



Solution

Here we have a stoichiometry problem where we need to find the number of moles of H_2 produced. Then we can use the ideal gas law, with the given temperature and pressure, to determine the volume of gas produced. First, the number of moles of H_2 is calculated:

$$55.8 \text{ g } \cancel{\text{Zn}} \times \frac{1 \text{ mol } \cancel{\text{Zn}}}{65.41 \text{ g } \cancel{\text{Zn}}} \times \frac{1 \text{ mol } \text{H}_2}{1 \text{ mol } \cancel{\text{Zn}}} = 0.853 \text{ mol } \text{H}_2$$

Now that we know the number of moles of gas, we can use the ideal gas law to determine the volume, given the other conditions:

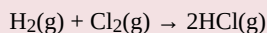
$$(1.07 \text{ atm})V = (0.853 \text{ mol}) \left(0.08205 \cdot \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (299 \text{ K})$$

All the units cancel except for L, for volume, which means:

$$V = 19.6 \text{ L}$$

Test Yourself

What pressure of HCl is generated if 3.44 g of Cl_2 are reacted in 4.55 L at 455 K?

**Answer**

0.796 atm

It should be obvious by now that some physical properties of gases depend strongly on the conditions. What we need is a set of standard conditions so that properties of gases can be properly compared to each other. **Standard temperature and pressure (STP)** is defined as exactly 100 kPa of pressure (0.986 atm) and 273 K (0°C). For simplicity, we will use 1 atm as standard pressure. Defining STP allows us to compare more directly the properties of gases that differ from each other.

One property shared among gases is a molar volume. The **molar volume** is the volume of 1 mol of a gas. At STP, the molar volume of a gas can be easily determined by using the ideal gas law:

$$(1 \text{ atm})V = (1 \text{ mol}) \left(0.08205 \cdot \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (273 \text{ K})$$

All the units cancel except for L, the unit of volume. So:

$$V = 22.4 \text{ L}$$

Note that we have not specified the identity of the gas; we have specified only that the pressure is 1 atm and the temperature is 273 K. This makes for a very useful approximation: *any gas at STP has a volume of 22.4 L per mole of gas*; that is, the molar volume at STP is 22.4 L/mol (Figure 6.3 “Molar Volume”). This molar volume makes a useful conversion factor in stoichiometry problems if the conditions are at STP. If the conditions are not at STP, a molar volume of 22.4 L/mol is not applicable. However, if the conditions are not at STP, the combined gas law can be used to calculate the volume of the gas at STP; then the 22.4 L/mol molar volume can be used.

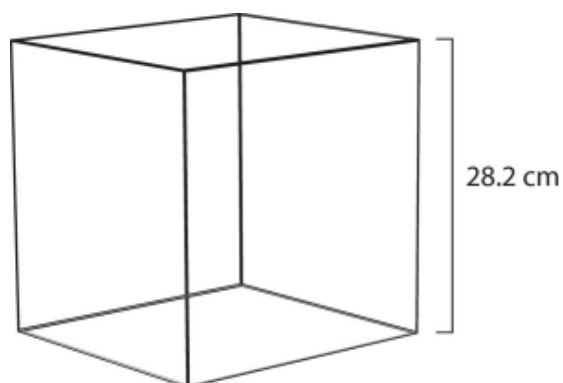


Figure 6.3 “Molar Volume.” A mole of gas at STP occupies 22.4 L, the volume of a cube that is 28.2 cm on a side.

Example 6.12

Problem

How many moles of Ar are present in 38.7 L at STP?

Solution

We can use the molar volume, 22.4 L/mol, as a conversion factor, but we need to reverse the fraction so that the L units cancel and mol units are introduced. It is a one-step conversion:

$$38.7 \cancel{\text{L}} \times \frac{1 \text{ mol}}{22.4 \cancel{\text{L}}} = 1.73 \text{ mol}$$

Test Yourself

What volume does 4.87 mol of Kr have at STP?

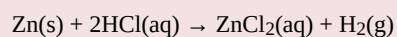
Answer

109 L

Example 6.13

Problem

What volume of H₂ is produced at STP when 55.8 g of Zn metal react with excess HCl?



Solution

This is a stoichiometry problem with a twist: we need to use the molar volume of a gas at STP to determine the final answer. The first part of the calculation is the same as in a previous example:

$$55.8 \text{ g } \cancel{\text{Zn}} \times \frac{1 \cancel{\text{ mol Zn}}}{65.41 \text{ g } \cancel{\text{ Zn}}} \times \frac{1 \text{ mol H}_2}{1 \cancel{\text{ mol Zn}}} = 0.853 \text{ mol H}_2$$

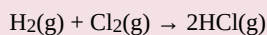
Now we can use the molar volume, 22.4 L/mol, because the gas is at STP:

$$0.853 \cancel{\text{ mol H}_2} \times \frac{22.4 \text{ L}}{1 \cancel{\text{ mol H}_2}} = 19.1 \text{ L H}_2$$

Alternatively, we could have applied the molar volume as a third conversion factor in the original stoichiometry calculation.

Test Yourself

What volume of HCl is generated if 3.44 g of Cl₂ are reacted at STP?

**Answer**

2.17 L

The ideal gas law can also be used to determine the densities of gases. Recall that density is defined as the mass of a substance divided by its volume:

$$d = \frac{m}{V}$$

Assume that you have exactly 1 mol of a gas. If you know the identity of the gas, you can determine the molar mass of the substance. Using the ideal gas law, you can also determine the volume of that mole of gas, using whatever the temperature and pressure conditions are. Then you can calculate the density of the gas by using:

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

Example 6.14

Problem

What is the density of N₂ at 25°C and 0.955 atm?

Solution

First, we must convert the temperature into kelvins:

$$25 + 273 = 298 \text{ K}$$

If we assume exactly 1 mol of N_2 , then we know its mass: 28.0 g. Using the ideal gas law, we can calculate the volume:

$$(0.955 \text{ atm})V = (1 \text{ mol}) \left(0.08205 \cdot \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (298 \text{ K})$$

All the units cancel except for L, the unit of volume. So:

$$V = 25.6 \text{ L}$$

Knowing the molar mass and the molar volume, we can determine the density of N_2 under these conditions:

$$d = \frac{28.0 \text{ g}}{25.6 \text{ L}} = 1.09 \text{ g/L}$$

Test Yourself

What is the density of CO_2 at a pressure of 0.0079 atm and 227 K? (These are the approximate atmospheric conditions on Mars.)

Answer

0.019 g/L

Chemistry Is Everywhere: Breathing

Breathing (more properly called *respiration*) is the process by which we draw air into our lungs so that our bodies can take up oxygen from the air. Let us apply the gas laws to breathing.

Start by considering pressure. We draw air into our lungs because the diaphragm, a muscle underneath the lungs, moves down to reduce pressure in the lungs, causing external air to rush in to fill the lower-pressure volume. We expel air by the diaphragm pushing against the lungs, increasing pressure inside the lungs and forcing the high-pressure air out. What are the pressure changes involved? A quarter of an atmosphere? A tenth of an atmosphere? Actually, under normal conditions, a pressure difference of only 1 or 2 torr makes us breathe in and out.

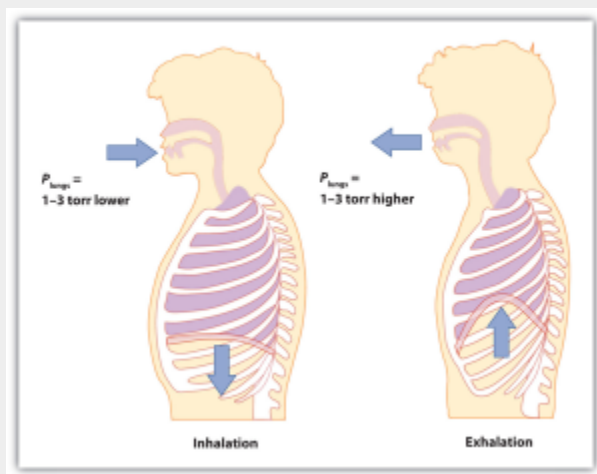


Figure 6.4 “Breathing Mechanics.” Breathing involves pressure differences between the inside of the lungs and the air outside. The pressure differences are only a few torr.

A normal breath is about 0.50 L. If room temperature is about 22°C, then the air has a temperature of about 295 K. With normal pressure being 1.0 atm, how many moles of air do we take in for every breath? The ideal gas law gives us an answer:

$$(1.0 \text{ atm})(0.50 \text{ L}) = n \left(0.08205 \cdot \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (295 \text{ K})$$

Solving for the number of moles, we get:

$$n = 0.021 \text{ mol air}$$

This ends up being about 0.6 g of air per breath—not much, but enough to keep us alive.

Key Takeaways

- The ideal gas law relates the four independent physical properties of a gas at any time.
- The ideal gas law can be used in stoichiometry problems in which chemical reactions involve gases.
- Standard temperature and pressure (STP) are a useful set of benchmark conditions to compare other properties of gases.
- At STP, gases have a volume of 22.4 L per mole.
- The ideal gas law can be used to determine densities of gases.

Exercises

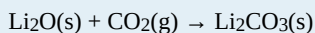
Questions

1. What is the ideal gas law? What is the significance of R ?
2. Why does R have different numerical values (see Table 6.1 “Values of the Ideal Gas Law Constant R ”)?
3. A sample of gas has a volume of 3.91 L, a temperature of 305 K, and a pressure of 2.09 atm. How many moles of gas are present?
4. A 3.88 mol sample of gas has a temperature of 28°C and a pressure of 885 torr. What is its volume?
5. A 0.0555 mol sample of Kr has a temperature of 188°C and a volume of 0.577 L. What pressure does it have?
6. If 1.000 mol of gas has a volume of 5.00 L and a pressure of 5.00 atm, what is its temperature?
7. A sample of 7.55 g of He has a volume of 5,520 mL and a temperature of 123°C. What is its pressure in torr?
8. A sample of 87.4 g of Cl_2 has a temperature of -22°C and a pressure of 993 torr. What is its volume in millilitres?
9. A sample of Ne has a pressure of 0.772 atm and a volume of 18.95 L. If its temperature is 295 K, what mass is present in the sample?
10. A mercury lamp contains 0.0055 g of Hg vapour in a volume of 15.0 mL. If the operating temperature is 2,800 K, what is the pressure of the mercury vapour?
11. Oxygen is a product of the decomposition of mercury(II) oxide:



What volume of O_2 is formed from the decomposition of 3.009 g of HgO if the gas has a pressure of 744 torr and a temperature of 122°C?

12. Lithium oxide is used to absorb carbon dioxide:



What volume of CO_2 can 6.77 g of Li_2O absorb if the CO_2 pressure is 3.5×10^{-4} atm and the temperature is 295 K?

13. What is the volume of 17.88 mol of Ar at STP?
 14. How many moles are present in 334 L of H_2 at STP?
 15. How many litres of CO_2 at STP are produced from 100.0 g of C_8H_{18} , the approximate formula of gasoline?

$$2\text{C}_8\text{H}_{18}(\ell) + 25\text{O}_2(\text{g}) \rightarrow 16\text{CO}_2(\text{g}) + 18\text{H}_2\text{O}(\ell)$$

 16. How many litres of O_2 at STP are required to burn 3.77 g of butane from a disposable lighter?

$$2\text{C}_4\text{H}_{10}(\text{g}) + 13\text{O}_2(\text{g}) \rightarrow 8\text{CO}_2(\text{g}) + 10\text{H}_2\text{O}(\ell)$$

 17. What is the density of each gas at STP?

- He
- Ne
- Ar
- K

18. What is the density of each gas at STP?

- H_2
- N_2
- O_2

19. What is the density of SF_6 at 335 K and 788 torr?
 20. What is the density of He at -200°C and 33.9 torr?

Answers

- The ideal gas law is $PV = nRT$. R is the ideal gas law constant, which relates the other four variables.
- 0.327 mol
- 3.64 atm
- 8,440 torr
- 12.2 g
- 0.230 L
- 401 L
- 157 L
- 0.179 g/L
 - 0.901 g/L
 - 1.78 g/L
 - 3.74 g/L
- 5.51 g/L

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Gas Mixtures

Learning Objectives

1. Learn Dalton's law of partial pressures.

One of the properties of gases is that they mix with each other. When they do so, they become a solution—a homogeneous mixture. Some of the properties of gas mixtures are easy to determine if we know the composition of the gases in the mix.

In gas mixtures, each component in the gas phase can be treated separately. Each component of the mixture shares the same temperature and volume. (Remember that gases expand to fill the volume of their container; gases in a mixture do that as well.) However, each gas has its own pressure. The **partial pressure** of a gas, P_i , is the pressure that an individual gas in a mixture has. Partial pressures are expressed in torr, millimetres of mercury, or atmospheres like any other gas pressure; however, we use the term *pressure* when talking about pure gases and the term *partial pressure* when we are talking about the individual gas components in a mixture.

Dalton's law of partial pressures states that the total pressure of a gas mixture, P_{tot} , is equal to the sum of the partial pressures of the components, P_i :

$$P_{\text{tot}} = P_1 + P_2 + P_3 + \dots = \sum_{\# \text{ of gases}} P_i$$

Although this may seem to be a trivial law, it reinforces the idea that gases behave independently of each other.

Example 6.15

Problem

A mixture of H_2 at 2.33 atm and N_2 at 0.77 atm is in a container. What is the total pressure in the container?

Solution

Dalton's law of partial pressures states that the total pressure is equal to the sum of the partial pressures. We simply add the two pressures together:

$$P_{\text{tot}} = 2.33 \text{ atm} + 0.77 \text{ atm} = 3.10 \text{ atm}$$

Test Yourself

Air can be thought of as a mixture of N_2 and O_2 . In 760 torr of air, the partial pressure of N_2 is 608 torr. What is the partial pressure of O_2 ?

Answer

152 torr

Example 6.16

Problem

A 2.00 L container with 2.50 atm of H_2 is connected to a 5.00 L container with 1.90 atm of O_2 inside. The containers are opened, and the gases mix. What is the final pressure inside the containers?

Solution

Because gases act independently of each other, we can determine the resulting final pressures using Boyle's law and then add the two resulting pressures together to get the final pressure. The total final volume is $2.00\text{ L} + 5.00\text{ L} = 7.00\text{ L}$. First, we use Boyle's law to determine the final pressure of H_2 :

$$(2.50\text{ atm})(2.00\text{ L}) = P_2(7.00\text{ L})$$

Solving for P_2 , we get:

$$P_2 = 0.714\text{ atm} = \text{partial pressure of } H_2$$

Now we do that same thing for the O_2 :

$$(1.90\text{ atm})(5.00\text{ L}) = P_2(7.00\text{ L})$$

$$P_2 = 1.36\text{ atm} = \text{partial pressure of } O_2$$

The total pressure is the sum of the two resulting partial pressures:

$$P_{\text{tot}} = 0.714\text{ atm} + 1.36\text{ atm} = 2.07\text{ atm}$$

Test Yourself

If 0.75 atm of He in a 2.00 L container is connected to a 3.00 L container with 0.35 atm of Ne and the connection between the containers is opened, what is the resulting total pressure?

Answer

0.51 atm

One of the reasons we have to deal with Dalton's law of partial pressures is that gases are frequently collected by bubbling through water. As we will see in [Chapter 10 "Solids and Liquids"](#), liquids are constantly evaporating into a vapour until the vapour achieves a partial pressure characteristic of the substance and the temperature. This partial pressure is called a **vapour pressure**. Table 6.2 "Vapour

Pressure of Water versus Temperature” lists the vapour pressures of H₂O versus temperature. Note that if a substance is normally a gas under a given set of conditions, the term *partial pressure* is used; the term *vapour pressure* is reserved for the partial pressure of a vapour when the liquid is the normal phase under a given set of conditions.

Table 6.2 Vapour Pressure of Water versus Temperature

Temperature (°C)	Vapour Pressure (torr)
5	6.54
10	9.21
15	12.79
20	17.54
21	18.66
22	19.84
23	21.08
24	22.39
25	23.77
30	31.84
35	42.20
40	55.36
50	92.59
60	149.5
70	233.8
80	355.3
90	525.9
100	760.0

Any time a gas is collected over water, the total pressure is equal to the partial pressure of the gas *plus* the vapour pressure of water. This means that the amount of gas collected will be less than the total pressure suggests.

Example 6.17

Problem

Hydrogen gas is generated by the reaction of nitric acid and elemental iron. The gas is collected in an inverted 2.00 L container immersed in a pool of water at 22°C. At the end of the collection, the partial pressure inside the container is 733 torr. How many moles of H₂ gas were generated?

Solution

We need to take into account that the total pressure includes the vapour pressure of water. According to [Table 6.2 “Vapour Pressure of Water versus Temperature,”](#) the vapour pressure of water at 22°C is 19.84 torr. According to Dalton’s law of partial pressures, the total pressure equals the sum of the pressures of the individual gases, so:

$$733 \text{ torr} = P_{\text{H}_2} + P_{\text{H}_2\text{O}} = P_{\text{H}_2} + 19.84 \text{ torr}$$

We solve by subtracting:

$$P_{\text{H}_2} = 713 \text{ torr}$$

Now we can use the ideal gas law to determine the number of moles (remembering to convert the temperature to kelvins, making it 295 K):

$$(713 \text{ torr})(2.00 \text{ L}) = n \left(62.36 \cdot \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (295 \text{ K})$$

All the units cancel except for mol, which is what we are looking for. So:

$$n = 0.0775 \text{ mol H}_2 \text{ collected}$$

Test Yourself

CO₂, generated by the decomposition of CaCO₃, is collected in a 3.50 L container over water. If the temperature is 50°C and the total pressure inside the container is 833 torr, how many moles of CO₂ were generated?

Answer

0.129 mol

Finally, we introduce a new unit that can be useful, especially for gases. The **mole fraction**, χ_i , is the ratio of the number of moles of component i in a mixture divided by the total number of moles in the sample:

$$\chi_i = \frac{\text{moles of component } i}{\text{total number of moles}}$$

(χ is the lowercase Greek letter *chi*.) Note that mole fraction is *not* a percentage; its values range from 0 to 1. For example, consider the combination of 4.00 g of He and 5.0 g of Ne. Converting both to moles, we get:

$$4.00 \text{ g He} \times \frac{1 \text{ mol He}}{4.00 \text{ g He}} = 1.00 \text{ mol He} \text{ and } 5.0 \text{ g Ne} \times \frac{1 \text{ mol Ne}}{20.0 \text{ g Ne}} = 0.25 \text{ mol Ne}$$

The total number of moles is the sum of the two mole amounts:

$$\text{total moles} = 1.00 \text{ mol} + 0.025 \text{ mol} = 1.25 \text{ mol}$$

The mole fractions are simply the ratio of each mole amount to the total number of moles, 1.25 mol:

$$\chi_{\text{He}} = \frac{1.00 \cancel{\text{ mol}}}{1.25 \cancel{\text{ mol}}} = 0.800$$

$$\chi_{\text{Ne}} = \frac{0.25 \cancel{\text{ mol}}}{1.25 \cancel{\text{ mol}}} = 0.200$$

The sum of the mole fractions equals exactly 1.

For gases, there is another way to determine the mole fraction. When gases have the same volume and temperature (as they would in a mixture of gases), the number of moles is proportional to partial pressure, so the mole fractions for a gas mixture can be determined by taking the ratio of partial pressure to total pressure:

$$\chi_i = \frac{P_i}{P_{\text{tot}}}$$

This expression allows us to determine mole fractions without calculating the moles of each component directly.

Example 6.18

Problem

A container has a mixture of He at 0.80 atm and Ne at 0.60 atm. What is the mole fraction of each component?

Solution

According to Dalton's law, the total pressure is the sum of the partial pressures:

$$P_{\text{tot}} = 0.80 \text{ atm} + 0.60 \text{ atm} = 1.40 \text{ atm}$$

The mole fractions are the ratios of the partial pressure of each component to the total pressure:

$$\chi_{\text{He}} = \frac{0.80 \text{ atm}}{1.40 \text{ atm}} = 0.57$$

$$\chi_{\text{Ne}} = \frac{0.60 \text{ atm}}{1.40 \text{ atm}} = 0.43$$

Again, the sum of the mole fractions is exactly 1.

Test Yourself

What are the mole fractions when 0.65 atm of O₂ and 1.30 atm of N₂ are mixed in a container?

Answer

$$\chi_{\text{O}_2} = 0.33; \chi_{\text{N}_2} = 0.67$$

Food and Drink App: Carbonated Beverages

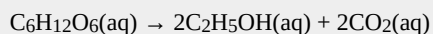
Carbonated beverages—sodas, beer, sparkling wines—have one thing in common: they have CO₂ gas dissolved in them in such sufficient quantities that it affects the drinking experience. Most people find the drinking experience pleasant—indeed, in the United States alone, over 1.5×10^9 gal of soda are consumed each year, which is almost 50 gal per person! This figure does not include other types of carbonated beverages, so the total consumption is probably significantly higher.

All carbonated beverages are made in one of two ways. First, the flat beverage is subjected to a high pressure of CO₂ gas, which forces the gas into solution. The carbonated beverage is then packaged in a tightly sealed package (usually a bottle or a can) and sold. When the container is opened, the CO₂ pressure is released, resulting in the well-known *hiss*, and CO₂ bubbles come out of solution (Figure 6.5 “Opening a Carbonated Beverage”). This must be done with care: if the CO₂ comes out too violently, a mess can occur!



Figure 6.5 “Opening a Carbonated Beverage.” If you are not careful opening a container of a carbonated beverage, you can make a mess as the CO_2 comes out of solution suddenly.

The second way a beverage can become carbonated is by the ingestion of sugar by yeast, which then generates CO_2 as a digestion product. This process is called *fermentation*. The overall reaction is:



When this process occurs in a closed container, the CO_2 produced dissolves in the liquid, only to be released from solution when the container is opened. Most fine sparkling wines and champagnes are turned into carbonated beverages this way. Less-expensive sparkling wines are made like sodas and beer, with exposure to high pressures of CO_2 gas.

Key Takeaways

- The pressure of a gas in a gas mixture is termed the *partial pressure*.
- Dalton’s law of partial pressure states that the total pressure in a gas mixture is the sum of the individual partial pressures.
- Collecting gases over water requires that we take the vapour pressure of water into account.
- Mole fraction is another way to express the amounts of components in a mixture.

Exercises

Questions

1. What is the total pressure of a gas mixture containing these partial pressures: $P_{N_2} = 0.78$ atm, $P_{H_2} = 0.33$ atm, and $P_{O_2} = 1.59$ atm?
2. What is the total pressure of a gas mixture containing these partial pressures: $P_{Ne} = 312$ torr, $P_{He} = 799$ torr, and $P_{Ar} = 831$ torr?
3. In a gas mixture of He and Ne, the total pressure is 335 torr and the partial pressure of He is 0.228 atm. What is the partial pressure of Ne?
4. In a gas mixture of O_2 and N_2 , the total pressure is 2.66 atm and the partial pressure of O_2 is 888 torr. What is the partial pressure of N_2 ?
5. A 3.55 L container has a mixture of 56.7 g of Ar and 33.9 g of He at 33°C. What are the partial pressures of the gases and the total pressure inside the container?
6. A 772 mL container has a mixture of 2.99 g of H_2 and 44.2 g of Xe at 388 K. What are the partial pressures of the gases and the total pressure inside the container?
7. A sample of O_2 is collected over water in a 5.00 L container at 20°C. If the total pressure is 688 torr, how many moles of O_2 are collected?
8. A sample of H_2 is collected over water in a 3.55 L container at 50°C. If the total pressure is 445 torr, how many moles of H_2 are collected?
9. A sample of CO is collected over water in a 25.00 L container at 5°C. If the total pressure is 0.112 atm, how many moles of CO are collected?
10. A sample of NO_2 is collected over water in a 775 mL container at 25°C. If the total pressure is 0.990 atm, how many moles of NO_2 are collected?
11. A sample of NO is collected over water in a 75.0 mL container at 25°C. If the total pressure is 0.495 atm, how many grams of NO are collected?
12. A sample of ClO_2 is collected over water in a 0.800 L container at 15°C. If the total pressure is 1.002 atm, how many grams of ClO_2 are collected?
13. Determine the mole fractions of each component when 44.5 g of He is mixed with 8.83 g of H_2 .
14. Determine the mole fractions of each component when 9.33 g of SO_2 is mixed with 13.29 g of SO_3 .
15. In a container, 4.56 atm of F_2 is combined with 2.66 atm of Cl_2 . What is the mole fraction of each component?
16. In a container, 77.3 atm of SiF_4 are mixed with 33.9 atm of O_2 . What is the mole fraction of each component?

Answers

1. 2.70 atm
3. 162 torr, or 0.213 atm
5. $P_{Ar} = 10.0$ atm; $P_{He} = 59.9$ atm; $P_{tot} = 69.9$ atm
7. 0.183 mol
9. 0.113 mol
11. 0.0440 g
13. $\chi_{He} = 0.718$; $\chi_{H_2} = 0.282$
15. $\chi_{F_2} = 0.632$; $\chi_{Cl_2} = 0.368$

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Kinetic Molecular Theory of Gases

Learning Objectives

1. State the major concepts behind the kinetic molecular theory of gases.
2. Demonstrate the relationship between kinetic energy and molecular speed.
3. Apply the kinetic molecular theory to explain and predict the gas laws.

Gases were among the first substances studied using the modern scientific method, which was developed in the 1600s. It did not take long to recognize that gases all shared certain physical behaviours, suggesting that gases could be described by one all-encompassing theory. The **kinetic molecular theory of gases** is a model that helps us understand the physical properties of gases at the molecular level. It is based on the following concepts:

1. Gases consist of particles (molecules or atoms) that are in constant random motion.
2. Gas particles are constantly colliding with each other and the walls of their container. These collisions are elastic; that is, there is no net loss of energy from the collisions.
3. Gas particles are small and the total volume occupied by gas molecules is negligible relative to the total volume of their container.
4. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas.
5. The average kinetic energy of gas particles is proportional to the absolute temperature of the gas, and all gases at the same temperature have the same average kinetic energy.

Figure 6.6 “The Kinetic Molecular Theory of Gases” shows a representation of how we mentally picture the gas phase.

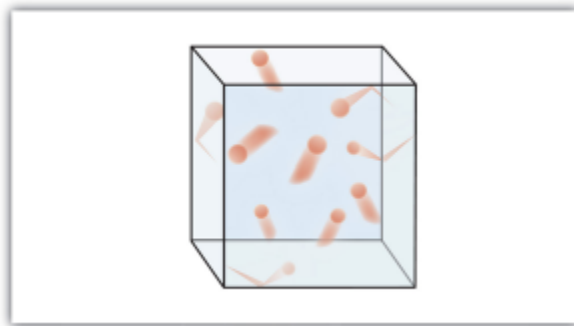


Figure 6.6 “The Kinetic Molecular Theory of Gases.” The kinetic molecular theory of gases describes this state of matter as composed of tiny particles in constant motion with a lot of distance between the particles.

Because most of the volume occupied by a gas is empty space, a gas has a low density and can expand or contract under the appropriate influence. The fact that gas particles are in constant motion means that two or more gases will always mix as the particles from the individual gases move and collide with each other. The number of collisions the gas particles make with the walls of their container and the force with which they collide determine the magnitude of the gas pressure.

Kinetic Energy and Molecular Speed

Gas particles are in constant motion, and any object in motion has **kinetic energy** (E_k). Kinetic energy, for an individual atom, can be calculated by the following equation where m is the mass, and u is the speed.

$$E_k = \frac{1}{2}mu^2$$

Overall, the molecules in a sample of a gas share an average kinetic energy; however, individual molecules exhibit a distribution of kinetic energies because of having a distribution of speeds (Figure 6.7 “Stylized Molecular Speed Distribution”). This distribution of speeds arises from the collisions that occur between molecules in the gas phase. Although these collisions are elastic (there is no net loss of energy), the individual speeds of each molecule involved in the collision may change. For example, in the collision of two molecules, one molecule may be deflected at a slightly higher speed and the other at a slightly lower speed, but the average kinetic energy does not change.

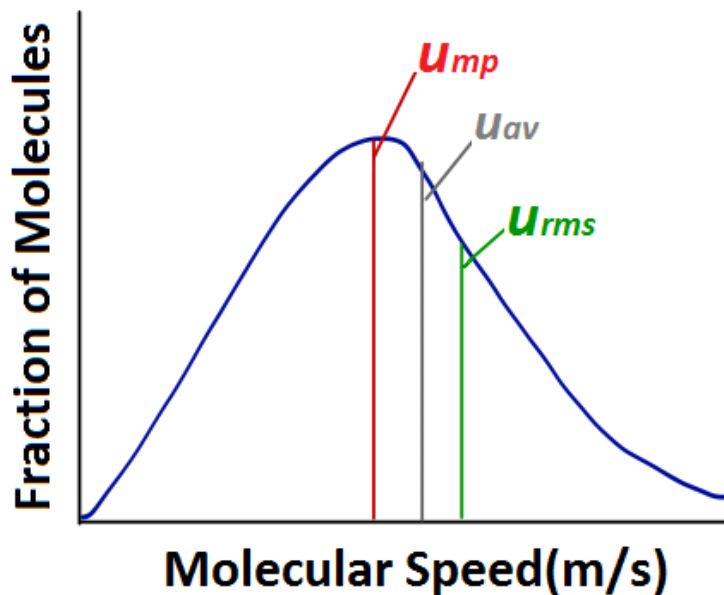


Figure 6.7 “Stylized Molecular Speed Distribution.”

When analyzing a diagram of the distribution of molecular speeds, there are several commonly used terms to be familiar with. The most probable speed (u_{mp}) is the speed of the largest number of molecules, and corresponds to the peak of the distribution. The average speed (u_{av}) is the mean speed of all gas molecules in the sample. The **root-mean-square (rms) speed** (u_{rms}) corresponds to the speed of molecules having exactly the same kinetic energy as the average kinetic energy of the sample.

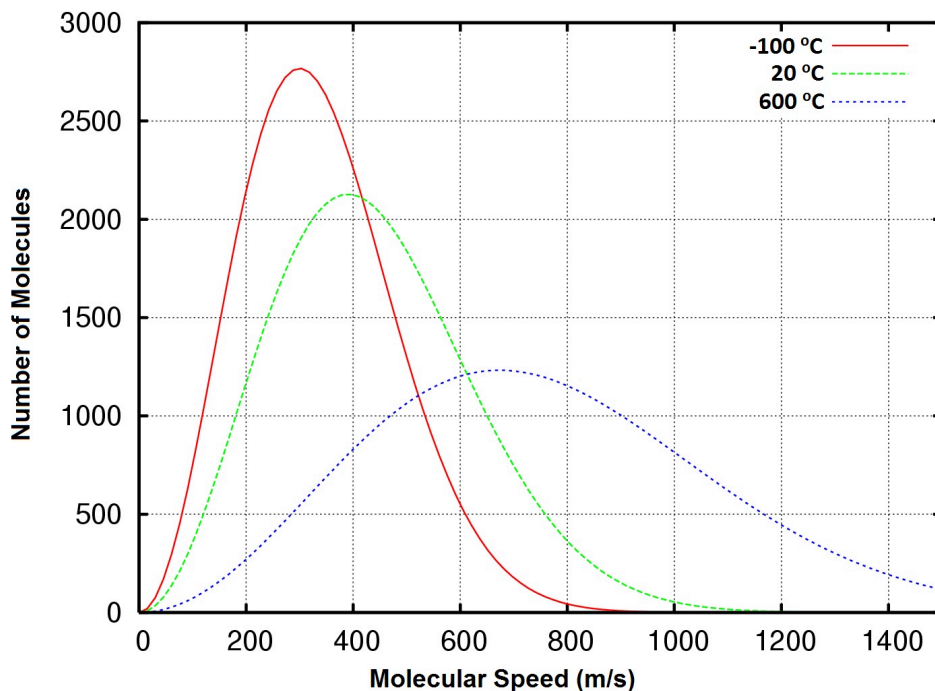


Figure 6.8 “Distribution of the Molecular Speeds of Oxygen Gas at -100 , 20 , and 600°C .”

According to the kinetic molecular theory, the average kinetic energy of gas particles is proportional to

the absolute temperature of the gas. This can be expressed with the following equation where k represents the Boltzmann constant. The Boltzmann constant is simply the gas constant R divided by the Avogadro's constant (N_A). The bar above certain terms indicates they are average values.

$$\overline{E_k} = \frac{3}{2}kT$$

Since average kinetic energy is related both to the absolute temperature and the molecular speed, we can combine the equation above with the previous one to determine the rms speed.

$$\begin{aligned}\overline{E_k} &= \frac{1}{2}m\overline{u^2} = \frac{3}{2}kT \\ &= \sqrt{\overline{u^2}} = \sqrt{\frac{3kT}{m}}\end{aligned}$$

This demonstrates that the rms speed is related to the temperature. We can further manipulate this equation by multiplying the numerator and denominator by Avogadro's constant (N_A) to give us a form using the gas constant (R) and molar mass (M).

$$\sqrt{\overline{u^2}} = \sqrt{\frac{3RT}{M}}$$

This form of the equation demonstrates that the rms speed of gas molecules is also related to the molar mass of the substance. Comparing two gases of different molar mass at the same temperature, we see that despite having the same average kinetic energy, the gas with the smaller molar mass will have a higher rms speed.

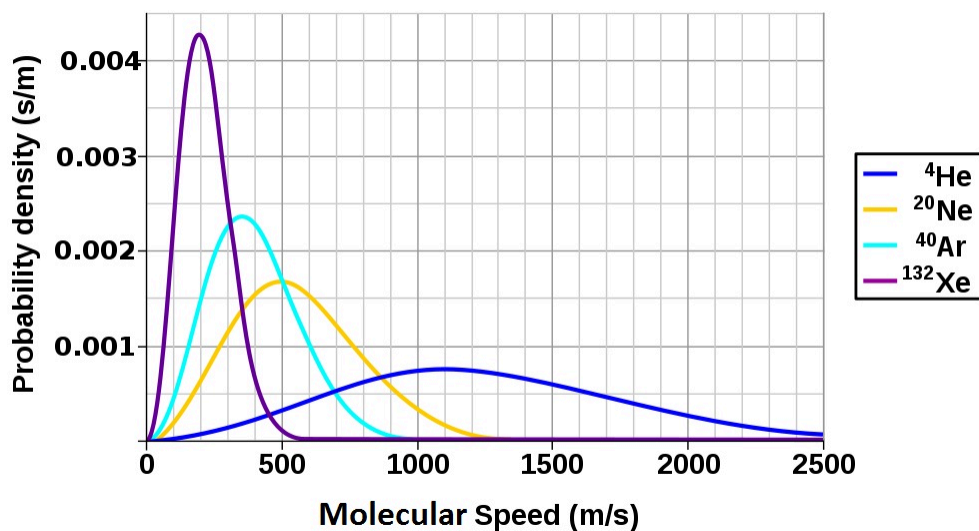


Figure 6.9 “Molecular Speed Distribution of Noble Gases.” (Source: Adapted from MaxwellBoltzmann-en.svg by Pd Bailey/Public Domain.)

Example 6.19

Problem

Calculate the rms speed of nitrogen molecules at 25°C.

Solution

$$\begin{aligned}\sqrt{u^2} &= \sqrt{\frac{3RT}{M}} \\ &= \sqrt{\frac{3(1.38 \times 10^{-23} \text{ J/K})(6.02 \times 10^{23} \text{ mol}^{-1})(298.15 \text{ K})}{28.02 \times 10^{-3} \text{ kg/mol}}} \\ &= \sqrt{2.654 \times 10^5 \text{ J/kg}}\end{aligned}$$

Knowing that $1 \text{ J} = 1 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2}$, we can convert to metres per second:

$$\begin{aligned}\sqrt{u^2} &= \sqrt{2.654 \times 10^5 \text{ J/kg} \times \frac{1 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2}}{1 \text{ J}}} \\ &= \sqrt{2.654 \times 10^5 \text{ m}^2 \cdot \text{s}^{-2}} \\ &= 515.2 \text{ m/s}\end{aligned}$$

Applying the Kinetic Molecular Theory to the Gas Laws

The kinetic molecular theory can be used to explain or predict the experimental trends that were used to generate the gas laws. Let's work through a few scenarios to demonstrate this point.

What will happen to the pressure of a system where the volume is decreased at constant temperature?

This problem can be approached in two ways:

1. The ideal gas law can be rearranged to solve for pressure and estimate the change in pressure:

$$PV = nRT$$

$$P = \frac{nRT}{V}$$

Volume is located in the denominator of the equation, and it is being decreased. This means the rest of the equation is being divided by a smaller number, so that should make the pressure larger.

- The kinetic molecular theory can be used. Since the temperature is remaining constant, the average kinetic energy and the rms speed remain the same as well. The volume of the container has decreased, which means that the gas molecules have to move a shorter distance to have a collision. There will therefore be more collisions per second, causing an increase in pressure.

What will happen to the pressure of a system where the temperature is increased and the volume remains constant?

Again, this type of problem can be approached in two ways:

- The ideal gas law can be rearranged to solve for pressure and estimate the change in pressure.

$$P = \frac{nRT}{V}$$

Temperature is located in the numerator; there is a direct relationship between temperature and pressure. Therefore an increase in temperature should cause an increase in pressure.

- The kinetic molecular theory can be used. Temperature is increased, so the average kinetic energy and the rms speed should also increase. This means that the gas molecules will hit the container walls more frequently and with greater force because they are all moving faster. This should increase the pressure.

Key Takeaways

- The physical behaviour of gases is explained by the kinetic molecular theory of gases.
- The number of collisions that gas particles make with the walls of their container and the force at which they collide determine the magnitude of the gas pressure.
- Temperature is proportional to average kinetic energy.

Exercises

Questions

- State the ideas of the kinetic molecular theory of gases.

2. Calculate the rms speed of CO₂ at 40°C.
3. Using the kinetic molecular theory, explain how an increase in the number of moles of gas at constant volume and temperature affects the pressure.

Answers

1. Gases consist of tiny particles of matter that are in constant motion. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic; that is, there is no net loss of energy from the collisions. Gas particles are separated by large distances. The size of gas particles is tiny compared to the distances that separate them and the volume of the container. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas. The average kinetic energy of gas particles is dependent on the temperature of the gas.
2. 421 m/s
3. Temperature remains the same, so the average kinetic energy and the rms speed should remain the same. Increasing the number of moles of gas means there are more molecules of gas available to collide with the walls of the container at any given time. Therefore pressure should increase.

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Molecular Effusion and Diffusion

Learning Objectives

1. Explore the nature of gas movement: molecular effusion and diffusion.
2. Examine and apply Graham's law of effusion.

Effusion

The movement of gas molecules can be divided into a few different types. **Effusion** is the movement of gas molecules from one container to another via a tiny hole. Typically the container to which the gas is moving is kept under lower pressure.

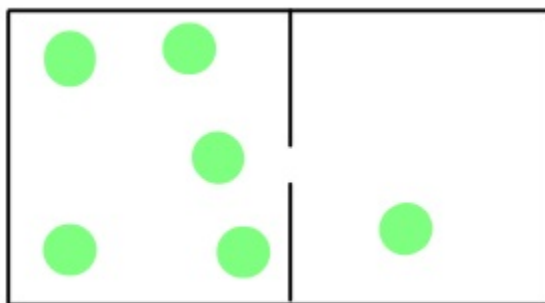


Figure 6.10 "Molecular Effusion."

In 1846, the Scottish chemist Thomas Graham found that the rate of effusion of a gas (the amount of gas transferred between containers in a certain amount of time) is inversely proportional to the square root of its molar mass. This means that gases with a lighter molecular weight have higher effusion rates.

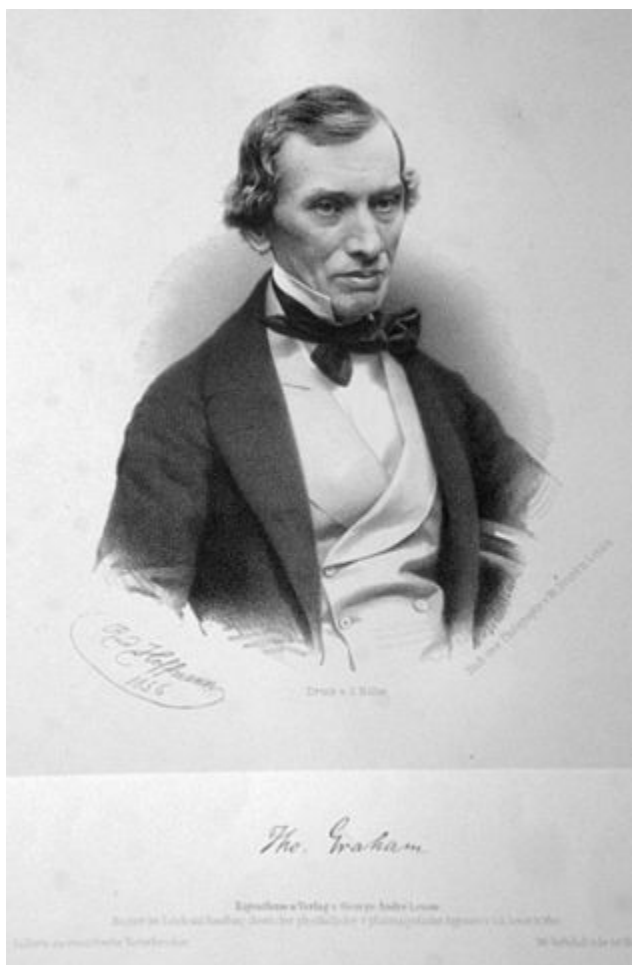


Figure 6.11 “Thomas Graham.” Graham proposed his law of effusion in 1846.

This finding is summarized in **Graham’s law of effusion**:

$$\frac{\text{Rate of Effusion (Gas 1)}}{\text{Rate of Effusion (Gas 2)}} = \sqrt{\frac{M_2}{M_1}}$$

This finding can be rationalized by thinking through the process of effusion on the molecular level. For a gas molecule to successfully move from one container to another, it must hit and pass through the tiny hole present in the container. Gases with higher rms speed are more likely to hit and pass through the hole so effusion is dependent on rms speed:

$$\frac{\text{Rate of Effusion (Gas 1)}}{\text{Rate of Effusion (Gas 2)}} = \frac{\sqrt{\frac{3RT}{M_1}}}{\sqrt{\frac{3RT}{M_2}}} = \sqrt{\frac{M_2}{M_1}}$$

Example 6.20

Problem

An unknown halogen (diatomic) gas effuses at a rate that is approximately 1.89 times the rate of I₂ gas at the same temperature. Determine the molar mass and identity of this unknown gas.

Solution

$$\frac{\text{Rate of Effusion (Gas 1)}}{\text{Rate of Effusion (Gas 2)}} = \sqrt{\frac{M_2}{M_1}}$$

$$\frac{\text{Rate of Effusion (Gas I}_2\text{)}}{\text{Rate of Effusion (Gas 2)}} = \sqrt{\frac{M_2}{MI_2}} = \frac{1}{1.89} = \sqrt{\frac{M_2}{253.80 \text{ g/mol}}}$$

$$0.279 = \frac{M_2}{253.80 \text{ g/mol}}$$

$$M_2 = 71.1 \text{ g/mol}$$

Therefore, the unknown gas is Cl₂.

Diffusion

Another type of gas movement is called **diffusion**; it is the movement of gas molecules through one or more additional types of gas via random molecular motion. Similar to effusion, gases with lower molecular weights (which have a higher rms speed) diffuse faster than gases with higher molecular weights. However, in diffusion, movement is much more complicated as collisions occur between molecules that change the direction and speed of the molecules. As a result of these collisions, the path a molecule travels in diffusion is made up of numerous straight, short segments. The term **mean free path** is used to describe the average distance travelled by a molecule between collisions.

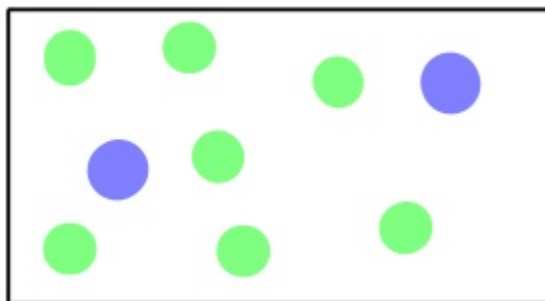


Figure 6.12 “Molecular Diffusion.”



Figure 6.13 “Particle Mean Free Path.” Stylized depiction of the path travelled by a gas particle during diffusion. Other particles have been omitted for clarity.

View this [video on diffusion](#) by Dr. Jessie A. Key for more information.

Key Takeaways

- Effusion is the movement of gas molecules from one container to another through a tiny hole.
- Rates of effusion can be compared at the same temperature using Graham’s law.
- Diffusion is the movement of gas molecules through one or more other types of gas via random molecular motion.
- Both the rates of effusion and diffusion are influenced by the molecular weight of the gas particle

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Real Gases

Learning Objectives

1. To examine the differences between ideal and real gases.
2. To explore the effects of high pressure and low temperature on real gas samples.
3. To apply van der Waals equation to correct the ideal gas law for real gases using the experimentally determined constants a and b .

An **ideal gas** is one that conforms exactly to the tenets of the kinetic molecular theory, where the volume occupied by the gas particles is negligible relative to the total volume of the container, and there are no appreciable intermolecular attractions or repulsions.

Real gases can deviate from ideal behaviour, especially at high pressures and low temperatures. The extent of deviation is measured using the **compressibility factor**. The compressibility factor is obtained by solving for n in the ideal gas law: dividing the product of pressure and volume by the product of the gas constant and temperature $\left(\frac{PV}{RT}\right)$ for one mole of a given substance. Under ideal conditions, this ratio of $\left(\frac{PV}{RT}\right)$ should be exactly equal to 1.

Real Gases at High Pressure

At higher pressures, gas molecules are closer together in a space. As a result of this crowding, gas molecules experience greater attractive intermolecular forces. Intermolecular forces hold molecules together more, lessening the force and frequency of collisions with the container wall and thus lowering the pressure below ideal values. As well, at higher pressure, molecules occupy a larger proportion of the volume of the container. With other gas molecules taking up a larger proportion of the volume of the container, the unoccupied volume of the container available to any one molecule is smaller than in ideal conditions. This decrease in available volume causes an increase in pressure beyond ideal conditions.

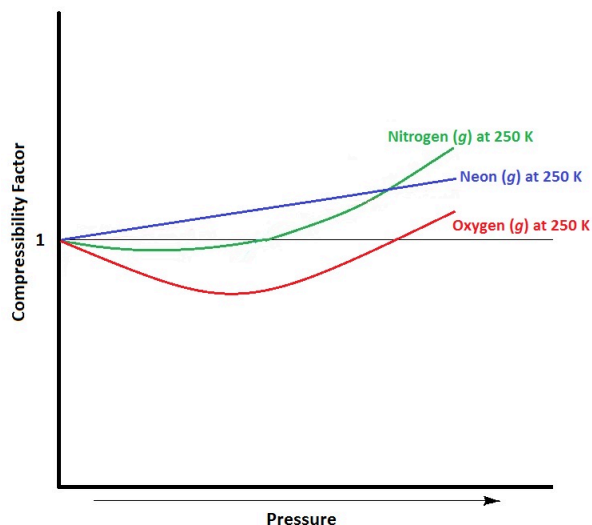


Figure 6.14. “Compressibility Factors.”
Approximate compressibility factors of three gases at 250 K.

Real Gases at Low Temperature

Temperature also influences deviations from ideal gas behaviour (see Figure 6.15). As temperature decreases, the average kinetic energy of the gas particles decreases. A larger proportion of gas molecules therefore have insufficient kinetic energy to overcome attractive intermolecular forces from neighbouring atoms. This means that gas molecules become “stickier” to each other, and collide with the walls of the container with less frequency and force, decreasing pressure below that of ideal values.

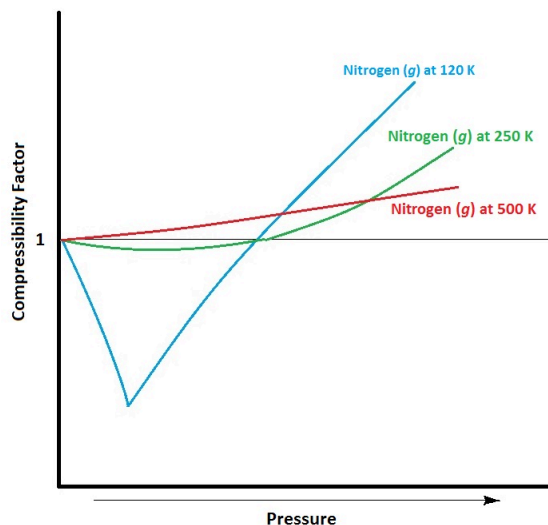


Figure 6.15 “Compressibility Factor of Nitrogen.”
Approximate compressibility factor of nitrogen at different temperatures.

The van der Waals Equation



Figure 6.16 “Johannes Diderik van der Waals.”

In 1873, Dutch scientist Johannes van der Waals developed an equation that compensates for deviations from ideal gas behaviour. The **van der Waals equation** uses two additional experimentally determined constants: a , which is a term to correct for intermolecular forces, and b , which corrects for the volume of the gas molecules (see Table 6.3 “Selected van der Waals Constants for Gas Molecules”). The van der Waals equation appears as:

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

It should be noted that, if the new terms a and b are equal to zero (under ideal conditions), the equation simplifies back to the ideal gas law: $PV = nRT$.

Table 6.3 Selected van der Waals Constants for Gas Molecules¹

Element	$a \left(\frac{L^2 \text{ atm}}{\text{mol}^2} \right)$	$b \left(\frac{L}{\text{mol}} \right)$
Helium	0.03457	0.0237
Neon	0.2135	0.01709
Hydrogen	0.2476	0.02661
Argon	1.355	0.0320
Nitric oxide	1.358	0.02789
Oxygen	1.378	0.03183
Nitrogen	1.408	0.03913
Carbon monoxide	1.505	0.03985
Methane	2.283	0.04278
Krypton	2.349	0.03978
Carbon dioxide	3.640	0.04267
Hydrogen chloride	3.716	0.04081
Nitrous oxide	3.832	0.04415
Ammonia	4.225	0.0371
Xenon	4.250	0.05105

Example 6.21**Problem**

Use the van der Waals equation and Table 6.3 to determine the pressure, in atmospheres, of 2.00 moles of oxygen gas in a 30.00 L flask at 25.0°C.

1. Adapted from [Van der Waals constants \(data page\) on Wikipedia](#) (under a [CC BY-SA 3.0](#) licence).

Solution

$$\left(P + a \left[\frac{n}{V}\right]^2\right) (V - nb) = nRT$$

$$\left(P + 1.378 \frac{\text{L}^2 \text{ atm}}{\text{mol}^2} \left[\frac{2.00 \text{ mol}}{30.00 \text{ L}}\right]\right) \left(30.00 \text{ L} - 2.00 \text{ mol} \left(0.03183 \frac{\text{L}}{\text{mol}}\right)\right) = 2.00 \text{ mol}$$

$$(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1})(298.15 \text{ K}) = P$$

$$P = 1.63 \text{ atm}$$

Key Takeaways

- An ideal gas is one that conforms exactly to the tenets of the kinetic molecular theory, where the volume occupied by the gas particles is negligible relative to the total volume of the container and there are no appreciable intermolecular attractions or repulsions.
- A real gas is one that deviates from ideal behaviour, due to the effects of gas particles occupying a finite volume and the strength of intermolecular forces.
- The van der Waals equation compensates for deviations from ideal gas behaviour.

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End-of-Chapter Material

Additional Exercises

1. What is the pressure in pascals if a force of 4.88 kN is pressed against an area of 235 cm²?
2. What is the pressure in pascals if a force of 3.44×10^4 MN is pressed against an area of 1.09 km²?
3. What is the final temperature of a gas whose initial conditions are 667 mL, 822 torr, and 67°C, and whose final volume and pressure are 1.334 L and 2.98 atm, respectively? Assume the amount remains constant.
4. What is the final pressure of a gas whose initial conditions are 1.407 L, 2.06 atm, and -67°C, and whose final volume and temperature are 608 mL and 449 K, respectively? Assume the amount remains constant.
5. Propose a combined gas law that relates volume, pressure, and amount at constant temperature.
6. Propose a combined gas law that relates amount, pressure, and temperature at constant volume.
7. A sample of 6.022×10^{23} particles of gas has a volume of 22.4 L at 0°C and a pressure of 1.000 atm. Although it may seem silly to contemplate, what volume would one particle of gas occupy?
8. One mole of liquid N₂ has a volume of 34.65 mL at -196°C. At that temperature, 1 mol of N₂ gas has a volume of 6.318 L if the pressure is 1.000 atm. What pressure is needed to compress the N₂ gas to 34.65 mL?
9. Use two values of *R* to determine the ratio between an atmosphere and a torr. Does the number make sense?
10. Use two values of *R* to determine how many joules are in a litre-atmosphere.
11. At an altitude of 40 km above the earth's surface, the atmospheric pressure is 5.00 torr, and the surrounding temperature is -20°C. If a weather balloon is filled with 1.000 mol of He at 760 torr and 22°C, what is its:
 - a. initial volume before ascent?
 - b. final volume when it reaches 40 km in altitude? (Assume the pressure of the gas equals the surrounding pressure.)
12. If a balloon is filled with 1.000 mol of He at 760 torr and 22°C, what is its:
 - a. initial volume before ascent?
 - b. final volume if it descends to the bottom of the Mariana Trench, where the surrounding temperature is 1.4°C and the pressure is 1,060 atm?
13. Air — a mixture of mostly N₂ and O₂ — can be approximated as having a molar mass of 28.8 g/mol. What is the density of air at 1.00 atm and 22°C? (This is approximately sea level.)
14. Air — a mixture of mostly N₂ and O₂ — can be approximated as having a molar mass of 28.8 g/mol. What is the density of air at 0.26 atm and -26°C? (This is approximately the atmospheric condition at the summit of Mount Everest.)
15. On the surface of Venus, the atmospheric pressure is 91.8 atm, and the temperature is 460°C. What is the density of CO₂ under these conditions? (The Venusian atmosphere is composed largely of CO₂.)
16. On the surface of Mars, the atmospheric pressure is 4.50 torr, and the temperature is -87°C. What is the density of CO₂ under these conditions? (The Martian atmosphere, similar to its Venusian counterpart, is composed largely of CO₂.)
17. HNO₃ reacts with iron metal according to $\text{Fe}(s) + 2\text{HNO}_3(aq) \rightarrow \text{Fe}(\text{NO}_3)_2(aq) + \text{H}_2(g)$. In a reaction vessel, 23.8 g of Fe are reacted but only 446 mL of H₂ are collected over water at 25°C and a pressure of 733 torr. What is the percent yield of the reaction?
18. NaHCO₃ is decomposed by heat according to $2\text{NaHCO}_3(s) \rightarrow \text{Na}_2\text{CO}_3(s) + \text{H}_2\text{O}(\ell) + \text{CO}_2(g)$. If you start with 100.0 g of NaHCO₃ and collect 10.06 L of CO₂ over water at 20°C and 0.977 atm, what is the percent yield of the decomposition reaction?
19. Determine if the following actions will cause the pressure of a particular gas sample to increase, decrease, or remain the

same:

- a. decreasing the temperature
 - b. decreasing the molar mass of the gas
 - c. decreasing the volume of the container
20. Under what conditions do gases deviate most from ideal gas behaviour? Explain your answer.
 21. Place the following gases in order from lowest to highest average molecular speed at 25°C: He, Ar, O₂, I₂.
 22. The effusion rate of an unknown noble gas sample is 0.35 times that of neon, at the same temperature. Determine the molecular weight and identity of the unknown noble gas.
 23. Use the van der Waals equation to determine the pressure of 2.00 moles of helium in a 5.00 L balloon at 300.00 K. How does this value compare to what you would obtain with the ideal gas law?

Answers

1. 208,000 Pa
3. 1,874 K
5. $\frac{P_1 V_1}{n_1} = \frac{P_2 V_2}{n_2}$
7. 3.72×10^{-23} L
9. 1 atm = 760 torr
11.
 - a. 24.2 L
 - b. 3155 L
13. 1.19 g/L
15. 67.2 g/L
17. 3.99%
19.
 - a. decreased pressure
 - b. increased pressure
 - c. increased pressure
21. I₂ < Ar < O₂ < He
23. van der Waals: 9.94 atm, ideal: 9.85 atm

Chapter 7. Energy and Chemistry

It takes energy to launch a spaceship into space. If it takes 1 energy unit to warm 0.25 g of water by 1°C, then it takes over 15,100 energy units to put that 0.25 g of water into earth orbit. The most powerful engines designed to lift rockets into space were part of the Saturn V rocket, that was built by the National Aeronautics and Space Administration (NASA). The rocket had three stages, with the first stage having the capability of launching about 3.5 million kg of mass. About 2.3 million kg was the actual fuel for the first stage; rockets in space have the unpleasant task of having to take their own chemicals with them to provide thrust.

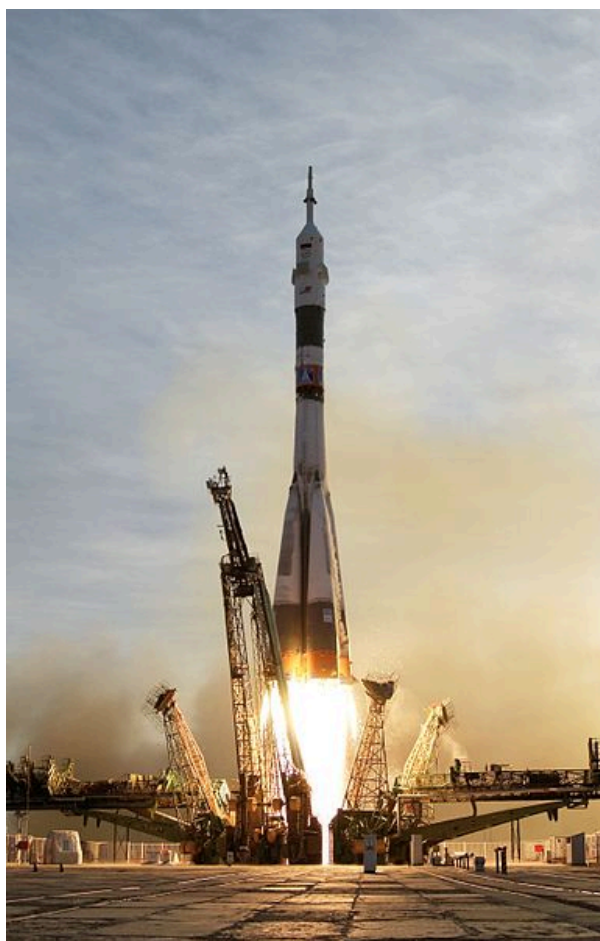


Figure 7.1 “Soyuz TMA-5 Launch.” It takes a lot of energy to launch a rocket into space. The Saturn V rocket used five of the most powerful engines ever built to take its initial step into orbit.

Having to carry its own fuel puts a lot of mass burden on an engine in space. This is why NASA is developing other types of engines to minimize fuel mass. An ion thruster uses xenon atoms that have had at least one electron removed from their atoms. The resulting ions can be accelerated by electric

fields, causing a thrust. Because xenon atoms are very large for atoms, the thrusting efficiency is high even though the actual thrust is low. Because of this, ion engines are useful only in space.

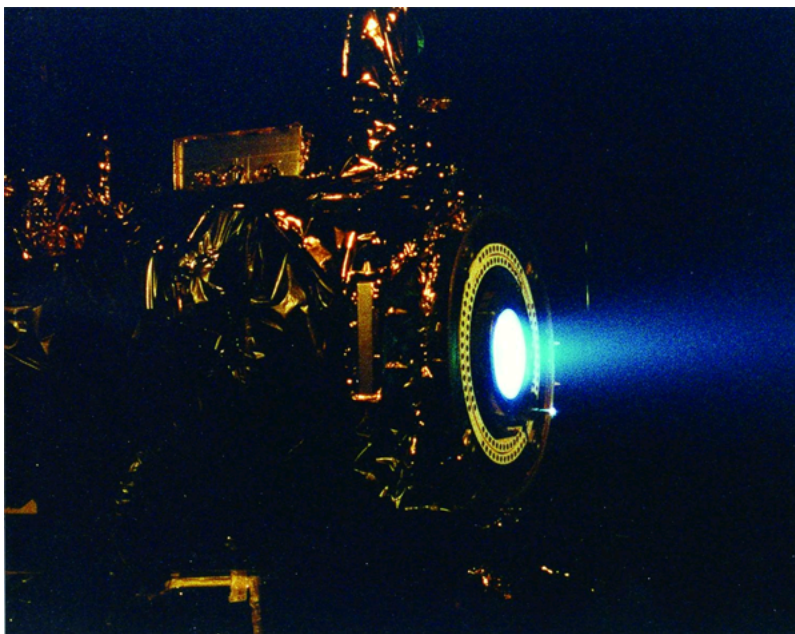


Figure 7.II “Ion Engine Test Firing.”

Ion drives have low thrust but high efficiency. They have already been used on several space missions, including NASA’s *Deep Space 1* spacecraft and Japan’s *Hayabusa* asteroid sampling probe.

Energy is a very important quantity in science and the world around us. Although most of our energy ultimately comes from the sun, much of the energy we use on a daily basis is rooted in chemical reactions. The gasoline in your car, the electricity in your house, the food in your diet—all provide substances for chemical reactions to provide energy (gasoline, food) or are produced from chemical reactions (electricity, about 50% of which is generated by burning coal). As such, it is only natural that the study of chemistry involves energy.

Media Attributions

- [“Soyuz TMA-5 Launch”](#) by Bill Ingalls for NASA © [Public Domain](#)
- [“Ion Engine Test Firing”](#) by NASA © [Public Domain](#)

Formation Reactions

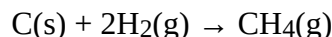
Learning Objectives

1. Define a formation reaction and be able to recognize one.
2. Use enthalpies of formation to determine the enthalpy of reaction.

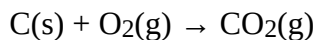
Hess's law allows us to construct new chemical reactions and predict what their enthalpies of reaction will be. This is a very useful tool because now we don't have to measure the enthalpy changes of every possible reaction. We need measure only the enthalpy changes of certain benchmark reactions and then use these reactions to algebraically construct any possible reaction and combine the enthalpies of the benchmark reactions accordingly.

But what are the benchmark reactions? We need to have some agreed-on sets of reactions that provide the central data for any thermochemical equation.

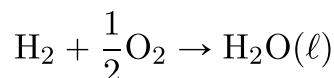
Formation reactions are chemical reactions that form one mole of a substance from its constituent elements in their standard states. By standard states, we mean a diatomic molecule, if that is how the element exists and the proper phase at normal temperatures (typically room temperature). The product is one mole of substance, which may require that coefficients on the reactant side be fractional (a change from our normal insistence that all coefficients be whole numbers). For example, the formation reaction for methane (CH₄) is:



The formation reaction for carbon dioxide (CO₂) is:



In both cases, one of the elements is a diatomic molecule because that is the standard state for that particular element. The formation reaction for H₂O — $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$ — is *not* in a standard state because the coefficient on the product is 2; for a proper formation reaction, only one mole of product is formed. Thus, we have to divide all coefficients by 2:



On a molecular scale, we are using half of an oxygen molecule, which may be problematic to visualize. However, on a molar level, it implies that we are reacting only half of a mole of oxygen molecules, which should be an easy concept for us to understand.

Example 7.1

Problem

Which of the following are proper formation reactions?

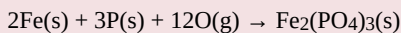
1. $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$
2. $\text{Si}(\text{s}) + 2\text{F}_2(\text{g}) \rightarrow \text{SiF}_4(\text{g})$
3. $\text{CaO}(\text{s}) + \text{CO}_2 \rightarrow \text{CaCO}_3(\text{s})$

Solution

1. In this reaction, two moles of product are produced, so this is not a proper formation reaction.
2. In this reaction, one mole of a substance is produced from its elements in their standard states, so this is a proper formation reaction.
3. One mole of a substance is produced, but it is produced from two other compounds, not its elements. So this is not a proper formation reaction.

Test Yourself

Is this a proper formation reaction? Explain why or why not.

**Answer**

This is not a proper formation reaction because oxygen is not written as a diatomic molecule.

Given the formula of any substance, you should be able to write the proper formation reaction for that substance.

Example 7.2

Problem

Write formation reactions for each of the following.

1. $\text{FeO}(\text{s})$
2. $\text{C}_2\text{H}_6(\text{g})$

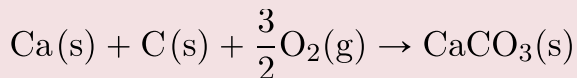
Solution

In both cases, there is one mole of the substance as product, and the coefficients of the reactants may have to be fractional to balance the reaction.

1. $\text{Fe(s)} + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{FeO(s)}$
2. $2\text{C(s)} + 3\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})$

Test Yourself

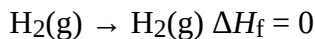
Write the equation for the formation of $\text{CaCO}_3(\text{s})$.

Answer

The enthalpy change for a formation reaction is called the **enthalpy of formation**. The subscript *f* is the clue that the reaction of interest is a formation reaction. Thus, for the formation of $\text{FeO}(\text{s})$:

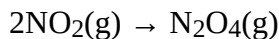


Note that now we are using kJ/mol as the unit because it is understood that the enthalpy change is for one mole of substance. Note, too, by definition, that the enthalpy of formation of an element is exactly zero because making an element from an element is no change. For example:

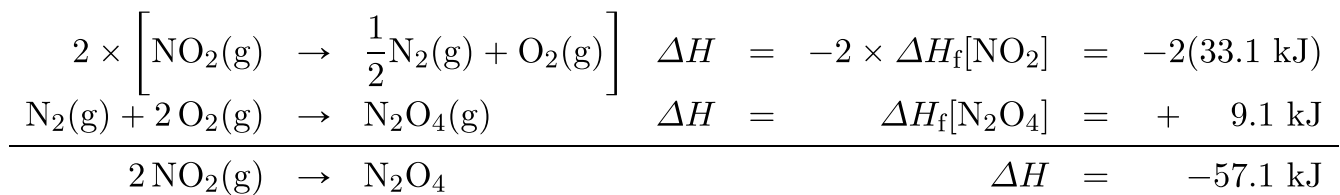


Formation reactions and their enthalpies are important because *these are the thermochemical data that are tabulated* for any chemical reaction. Table 7.1 “Enthalpies of Formation for Various Substances” lists some enthalpies of formation for a variety of substances; in some cases, however, phases can be important (e.g., for H_2O).

It is easy to show that any general chemical equation can be written in terms of the formation reactions of its reactants and products, some of them reversed (which means the sign must change in accordance with Hess’s law). For example, consider:



We can write it in terms of the (reverse) formation reaction of NO_2 and the formation reaction of N_2O_4 :



We must multiply the first reaction by 2 to get the correct overall balanced equation. We are simply using Hess's law in combining the ΔH_f values of the formation reactions.

Table 7.1 Enthalpies of Formation for Various Substances¹

Compound	ΔH_f (kJ/mol)
Ag(s)	0
AgBr(s)	-100.37
AgCl(s)	-127.01
Al(s)	0
Al ₂ O ₃ (s)	-1,675.7
Ar(g)	0
Au(s)	0
BaSO ₄ (s)	-1,473.19
Br ₂ (ℓ)	0
C(s, dia)	1.897
C(s, gra)	0
CCl ₄ (ℓ)	-128.4
CH ₂ O(g)	-115.90
CH ₃ COOH(ℓ)	-483.52
CH ₃ OH(ℓ)	-238.4
CH ₄ (g)	-74.87
CO(g)	-110.5
CO ₂ (g)	-393.51
C ₂ H ₅ OH(ℓ)	-277.0
C ₂ H ₆ (g)	-83.8
C ₆ H ₁₂ (ℓ)	-157.7
C ₆ H ₁₂ O ₆ (s)	-1277
C ₆ H ₁₄ (ℓ)	-198.7
C ₆ H ₅ CH ₃ (ℓ)	12.0
C ₆ H ₆ (ℓ)	48.95
C ₁₀ H ₈ (s)	77.0
C ₁₂ H ₂₂ O ₁₁ (s)	-2,221.2
Ca(s)	0

1. Sources: National Institute of Standards and Technology's Chemistry WebBook, <http://webbook.nist.gov/chemistry>; D. R. Lide, ed., *CRC Handbook of Chemistry and Physics*, 89th ed. (Boca Raton, FL: CRC Press, 2008); J. A. Dean, ed., *Lange's Handbook of Chemistry*, 14th ed. (New York: McGraw-Hill, 1992).

Compound	ΔH_f (kJ/mol)
CaCl ₂ (s)	-795.80
CaCO ₃ (s, arag)	-1,207.1
CaCO ₃ (s, calc)	-1,206.9
Cl ₂ (g)	0
Cr(s)	0
Cr ₂ O ₃ (s)	-1,134.70
Cs(s)	0
Cu(s)	0
F ₂ (g)	0
Fe(s)	0
Fe ₂ (SO ₄) ₃ (s)	-2,583.00
Fe ₂ O ₃ (s)	-825.5
Ga(s)	0
HBr(g)	-36.29
HCl(g)	-92.31
HF(g)	-273.30
HI(g)	26.5
HNO ₂ (g)	-76.73
HNO ₃ (g)	-134.31
H ₂ (g)	0
H ₂ O(g)	-241.8
H ₂ O(l)	-285.83
H ₂ O(s)	-292.72
He(g)	0
Hg(l)	0
Hg ₂ Cl ₂ (s)	-265.37
I ₂ (s)	0
K(s)	0
KBr(s)	-393.8

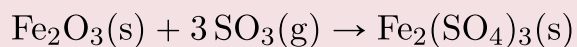
Compound	ΔH_f (kJ/mol)
KCl(s)	-436.5
KF(s)	-567.3
KI(s)	-327.9
Li(s)	0
LiBr(s)	-351.2
LiCl(s)	-408.27
LiF(s)	-616.0
LiI(s)	-270.4
Mg(s)	0
MgO(s)	-601.60
NH ₃ (g)	-45.94
NO(g)	90.29
NO ₂ (g)	33.10
N ₂ (g)	0
N ₂ O(g)	82.05
N ₂ O ₄ (g)	9.08
N ₂ O ₅ (g)	11.30
Na(s)	0
NaBr(s)	-361.1
NaCl(s)	-385.9
NaF(s)	-576.6
NaI(s)	-287.8
NaHCO ₃ (s)	-950.81
NaN ₃ (s)	21.71
Na ₂ CO ₃ (s)	-1,130.77
Na ₂ O(s)	-417.98
Na ₂ SO ₄ (s)	-331.64
Ne(g)	0
Ni(s)	0

Compound	ΔH_f (kJ/mol)
O ₂ (g)	0
O ₃ (g)	142.67
PH ₃ (g)	22.89
Pb(s)	0
PbCl ₂ (s)	-359.41
PbO ₂ (s)	-274.47
PbSO ₄ (s)	-919.97
Pt(s)	0
S(s)	0
SO ₂ (g)	-296.81
SO ₃ (g)	-395.77
SO ₃ (ℓ)	-438
Si(s)	0
U(s)	0
UF ₆ (s)	-2,197.0
UO ₂ (s)	-1,085.0
Xe(g)	0
Zn(s)	0
ZnCl ₂ (s)	-415.05

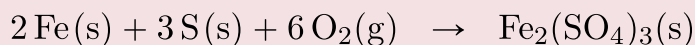
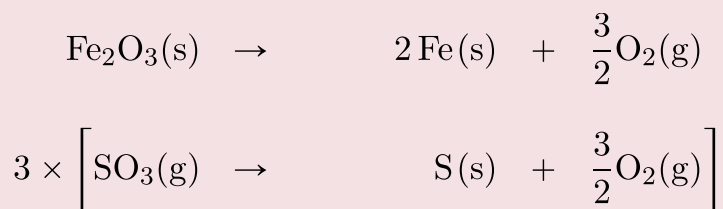
Example 7.3

Problem

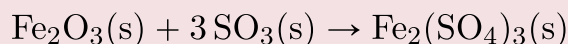
Show that the following reaction can be written as a combination of formation reactions:

**Solution**

There will be three formation reactions. The one for the products will be written as a formation reaction, while the ones for the reactants will be written in reverse. Furthermore, the formation reaction for SO₃ will be multiplied by 3 because there are three moles of SO₃ in the balanced chemical equation. The formation reactions are as follows:

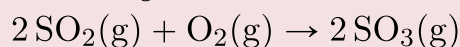


When these three equations are combined and simplified, the overall reaction is:

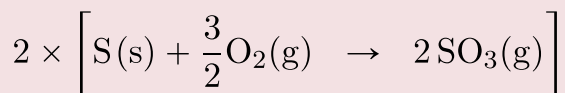
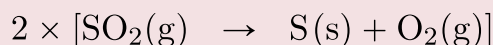


Test Yourself

Write the formation reactions that will yield the following:



Answer



Now that we have established formation reactions as the major type of thermochemical reaction we will be interested in, do we always need to write all the formation reactions when we want to determine the enthalpy change of any random chemical reaction? No. There is an easier way. You may have noticed in all our examples that we change the signs on all the enthalpies of formation of the reactants, and we don't change the signs on the enthalpies of formation of the products. We also multiply the enthalpies of formation of any substance by its coefficient — technically, even when it is just 1. This allows us to make the following statement: *the enthalpy change of any chemical reaction is equal to the sum of the enthalpies of formation of the products minus the sum of the enthalpies of formation of the reactants.* In mathematical terms:

$$\Delta H_{\text{rxn}} = \left[(2 \text{ mol})(-92.3 \text{ kJ/mol}) + (1 \text{ mol})(0 \text{ kJ/mol}) \right]$$

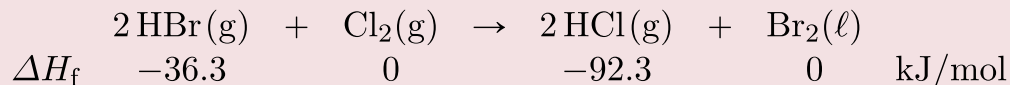
$$- \left[(2 \text{ mol})(-36.3 \text{ kJ/mol}) + (1 \text{ mol})(0 \text{ kJ/mol}) \right]$$

Where n_p and n_r are the number of moles of products and reactants, respectively (even if they are just 1 mol), and $\Delta H_{f,p}$ and $\Delta H_{f,r}$ are the enthalpies of formation of the product and reactant species, respectively. This *products-minus-reactants* scheme is very useful in determining the enthalpy change of any chemical reaction, if the enthalpy of formation data are available. Because the mol units cancel when multiplying the amount by the enthalpy of formation, the enthalpy change of the chemical reaction has units of energy (joules or kilojoules) only.

Example 7.4

Problem

Use the products-minus-reactants approach to determine the enthalpy of reaction for:

**Solution**

The enthalpies of formation are multiplied by the number of moles of each substance in the chemical equation, and the total enthalpy of formation for reactants is subtracted from the total enthalpy of formation of the products:

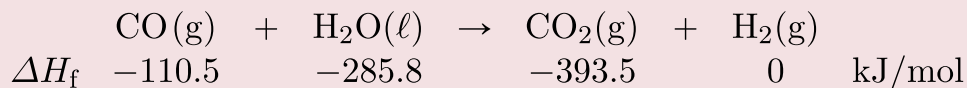
$$\begin{aligned} \Delta H_{\text{rxn}} &= \left[(2 \text{ mol})(-92.3 \text{ kJ/mol}) + (1 \text{ mol})(0 \text{ kJ/mol}) \right] \\ &\quad - \left[(2 \text{ mol})(-36.3 \text{ kJ/mol}) + (1 \text{ mol})(0 \text{ kJ/mol}) \right] \end{aligned}$$

All the mol units cancel. Multiplying and combining all the values, we get:

$$\Delta H_{\text{rxn}} = -112.0 \text{ kJ}$$

Test Yourself

What is the enthalpy of reaction for this chemical equation?

**Answer**

+2.8 kJ

Food and Drink App: Calories and Nutrition

The next section, called “Energy,” mentions the connection between the calorie unit and nutrition: the calorie is the common unit of energy used in nutrition, but we really consider the kilocalorie (spelled Calorie with a capital C). A daily diet of 2,000 Cal is actually 2,000,000 cal, or over 8,000,000 J, of energy.

Nutritionists typically generalize the Calorie content of foods by separating it into the three main food types, or macronutrients: proteins, carbohydrates, and fats. The general rule of thumb is as follows:

Table 7.2 Energy Content of Macronutrients

Macronutrient	Energy Content
protein	4 Cal/g
carbohydrate	4 Cal/g
fat	9 Cal/g

This table is very useful. Assuming a 2,000 Cal daily diet, if our diet consists solely of proteins and carbohydrates, we need only about 500 g of food for sustenance — a little more than a pound. If our diet consists solely of fats, we need only about 220 g of food — less than a half pound. Of course, most of us have a mixture of proteins, carbohydrates, and fats in our diets. Water has no caloric value in the diet, so any water in the diet is calorically useless. (However, it is important for hydration; also, many forms of water in our diet are highly flavoured and sweetened, which bring other nutritional issues to bear.)

When your body works, it uses calories provided by the diet as its energy source. If we eat more calories than our body uses, we gain weight — about 1 lb of weight for every additional 3,500 Cal we ingest. Similarly, if we want to lose weight, we need to expend an extra 3,500 Cal than we ingest to lose 1 lb of weight. No fancy or fad diets are needed; maintaining an ideal body weight is a straightforward matter of thermochemistry — pure and simple.

Key Takeaways

- A formation reaction is the formation of one mole of a substance from its constituent elements.
- Enthalpies of formation are used to determine the enthalpy change of any given reaction.

Exercises

Questions

1. Define *formation reaction* and give an example.
2. Explain the importance of formation reactions in thermochemical equations.
3. Which of the following reactions is a formation reaction? If it is not a formation reaction, explain why.
 - a. $\text{H}_2(\text{g}) + \text{S}(\text{s}) \rightarrow \text{H}_2\text{S}(\text{g})$
 - b. $2\text{HBr}(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g}) + \text{Br}_2(\ell)$
4. Which of the following reactions is a formation reaction? If it is not a formation reaction, explain why.
 - a. $\text{Fe}(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{FeO}(\text{s})$
 - b. $\text{Hg}(\ell) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{HgO}(\text{s})$
5. Which of the following reactions is a formation reaction? If it is not a formation reaction, explain why.

- a. $\text{H}_2(\text{g}) + \text{S}(\text{s}) + 2\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{SO}_4(\ell)$
 b. $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\ell)$
6. Which of the following reactions is a formation reaction? If it is not a formation reaction, explain why.
- a. $\text{Zn}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})$
 b. $2\text{Na}(\text{s}) + \text{C}(\text{s}) + \frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{Na}_2\text{CO}_3(\text{s})$
7. Write a proper formation reaction for each substance.
- a. $\text{H}_3\text{PO}_4(\text{s})$
 b. $\text{Na}_2\text{O}(\text{s})$
 c. $\text{C}_3\text{H}_7\text{OH}(\ell)$
8. Write a proper formation reaction for each substance.
- a. $\text{N}_2\text{O}_5(\text{g})$
 b. $\text{BaSO}_4(\text{s})$
 c. $\text{Fe}(\text{OH})_3(\text{s})$
9. Write a proper formation reaction for each substance.
- a. $\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$
 b. $\text{Zn}(\text{NO}_3)_2(\text{s})$
 c. $\text{Al}(\text{OH})_3(\text{s})$
10. Write a proper formation reaction for each substance.
- a. $\text{O}_3(\text{g})$
 b. $\text{Na}_2\text{O}_2(\text{s})$
 c. $\text{PCl}_5(\text{g})$
11. Write the reaction $\text{MgCO}_3(\text{s}) \rightarrow \text{MgO}(\text{s}) + \text{CO}_2(\text{g})$ in terms of formation reactions.
 12. Write the reaction $2\text{NO} + 4\text{NO}_2 \rightarrow 2\text{N}_2\text{O}_5 + \text{N}_2$ in terms of formation reactions.
 13. Write the reaction $2\text{CuCl}(\text{s}) \rightarrow \text{Cu}(\text{s}) + \text{CuCl}_2(\text{s})$ in terms of formation reactions.
 14. Write the reaction $\text{SiH}_4 + 4\text{F}_2 \rightarrow \text{SiF}_4 + 4\text{HF}$ in terms of formation reactions.
 15. Determine the enthalpy change of the reaction $\text{CH}_2\text{O}(\text{g}) + \text{O}_2 \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$. Data can be found in Table 7.2.
 16. Determine the enthalpy change of the reaction $2\text{AgBr}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{AgCl}(\text{s}) + \text{Br}_2(\ell)$. Data can be found in Table 7.2.
 17. Determine the enthalpy change of the reaction $\text{Mg}(\text{s}) + \text{N}_2\text{O}_5(\text{g}) \rightarrow \text{MgO}(\text{s}) + 2\text{NO}_2(\text{g})$. Data can be found in Table 7.2.
 18. Determine the enthalpy change of the reaction $2\text{C}_6\text{H}_6(\ell) + 15\text{O}_2(\text{g}) \rightarrow 12\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\ell)$. Data can be found in Table 7.2.

Answers

1. A formation reaction is a reaction that produces one mole of a substance from its elements. Example: $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$
3. a. formation reaction
 b. It is not the formation of a single substance, so it is not a formation reaction.
5. a. formation reaction
 b. It is not the formation of a single substance, so it is not a formation reaction.

7. a. $\frac{3}{2}\text{H}_2(\text{g}) + \text{P}(\text{s}) + 2\text{O}_2(\text{g}) \rightarrow \text{H}_3\text{PO}_4(\text{s})$
b. $2\text{Na}(\text{s}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{Na}_2\text{O}(\text{s})$
c. $3\text{C}(\text{s}) + \frac{1}{2}\text{O}_2(\text{g}) + 4\text{H}_2(\text{g}) \rightarrow \text{C}_3\text{H}_7\text{OH}(\ell)$
9. a. $12\text{C}(\text{s}) + 11\text{H}_2(\text{g}) + \frac{11}{2}\text{O}_2(\text{g}) \rightarrow \text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{s})$
b. $\text{Zn}(\text{s}) + \text{N}_2(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow \text{Zn}(\text{NO}_3)_2$
c. $\text{Al}(\text{s}) + \frac{3}{2}\text{O}_2(\text{g}) + \frac{3}{2}\text{H}_2(\text{g}) \rightarrow \text{Al}(\text{OH})_3(\text{s})$
11. a. $\text{MgCO}_3(\text{s}) \rightarrow \text{Mg}(\text{s}) + \text{C}(\text{s}) + \frac{3}{2}\text{O}_2(\text{g})$
b. $\text{Mg}(\text{s}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{MgO}(\text{s})$
c. $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$
13. a. $2 \times \left[\text{CuCl}(\text{s}) \rightarrow \text{Cu}(\text{s}) + \frac{1}{2}\text{Cl}_2(\text{g}) \right]$
b. $\text{Cu}(\text{s}) \rightarrow \text{Cu}(\text{s})$
c. $\text{Cu}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow \text{CuCl}_2(\text{s})$

15. $\Delta H = -563.44 \text{ kJ}$

17. $\Delta H = -546.7 \text{ kJ}$

Energy

Learning Objectives

1. Define *energy*.
2. Know the units of energy.
3. Understand the law of conservation of energy.

Energy is the ability to do work. Think about it: when you have a lot of energy, you can do a lot of work; but if you're low on energy, you don't want to do much work. Work (w) itself is defined as a force (F) operating over a distance (Δx):

$$w = F \times \Delta x$$

In SI, force has units of newtons (N), while distance has units of metres. Therefore, work has units of N·m (Newton-metres). This compound unit is redefined as a **joule** (J):

$$\begin{aligned} 1 \text{ joule} &= 1 \text{ newton} \cdot \text{metre} \\ 1 \text{ J} &= 1 \text{ N} \cdot \text{m} \end{aligned}$$

Because energy is the ability to do work, energy is also measured in joules. This is the primary unit of energy we will use here.

How much is 1 J? It is enough to warm up about one-fourth of a gram of water by 1°C. It takes about 12,000 J to warm a cup of coffee from room temperature to 50°C. So a joule is not a lot of energy. It is not uncommon to measure energies in thousands of joules, so the kilojoule (kJ) is a common unit of energy, with 1 kJ equal to 1,000 J.

An older — but still common — unit of energy is the **calorie**. The calorie (cal) was originally defined in terms of warming up a given quantity of water. The modern definition of calorie equates it to joules:

$$1 \text{ cal} = 4.184 \text{ J}$$

One area where the calorie is used is in nutrition. Energy contents of foods are often expressed in calories. However, the calorie unit used for foods is actually the kilocalorie (kcal). Most foods indicate this by spelling the word with a capital C — Calorie. Figure 7.1 “Calories on Food Labels” shows one example.