OT - PDX - METRO: GENERAL CHEMISTRY I



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OT - PDX - Metro: General Chemistry I

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TABLE OF CONTENTS

Licensing

1: Matter and Measurement

- 1.1: Atoms and Molecules
- 1.2: The Scientific Approach to Knowledge
- 1.3: The Classification of Matter
- 1.4: Measurements
  - 1.4.1: Practice Problems on Measurements
- 1.5: The Mole is a Measure of Amount
- 1.6: Accuracy and Precision
- 1.7: Dimensional Analysis
  - 1.7.1: Practice Problems on Dimensional Analysis
- 1.8: Significant Digits

2: Atoms and Elements

- 2.1: A History of Atomic Theory
  - 2.1.1: Practice Problems- A History of Atomic Theory (Optional)
- 2.2: The Structure of the Atom and How We Represent It
  - 2.2.1: Practice Problems- The Structure of the Atom and How We Represent It (Optional)
- 2.3: Calculating Atomic Masses
  - 2.3.1: Practice Problems- Calculating Atomic Masses (Optional)
- 2.4: Using the Mole in Calculations
- 2.5: Highlighting Instrumentation- Mass Spectrometry

3: Nuclei, Ions and the Periodic Table

- 3.1: Nuclear Chemistry and Radioactive Decay
  - 3.1.1: Practice Problems- Nuclear Chemistry and Radioactive Decay (Optional)
- 3.2: Organization of the Periodic Table
  - 3.2.1: Practice Problems - Organization of the Periodic Table (Optional)
- 3.3: Predicting Charges of Ions

4: Molecules, Compounds and Quantifying Chemicals

- 4.1: Chemical Bonding
- 4.2: Representing ionic and molecular compounds and molecules
  - 4.2.1: Practice Problems- Writing Chemical Formulas
- 4.3: Chemical Nomenclature
  - 4.3.1: Practice Problems- Molecular and Ionic Compounds
- 4.4: Formula Mass, Percent Composition, and the Mole
  - 4.4.1: Practice Problems- Formula Mass, Percent Composition, and the Mole
- 4.5: Empirical and Molecular Formulas

- 4.5.1: Practice Problems- Empirical and Molecular Formulas

5: Transformations of Matter

- 5.1: Writing and Balancing Chemical Equations
  - 5.1.1: Practice Problems- Writing and Balancing Chemical Equations
- 5.2: Reaction Stoichiometry
  - 5.2.1: Practice Problems- Reaction Stoichiometry
- 5.3: Calculating Reaction Yields
  - 5.3.1: Practice Problems Calculating Reaction Yields

6: Aqueous Reactions

- 6.1: Solutions and Solution Concentration
  - 6.1.1: Practice Problems- Solution Concentration
- 6.2: Solutions Chemistry
  - 6.2.1: Practice Problems- Solutions Chemistry
- 6.3: Precipitation Reactions
  - 6.3.1: Practice Problems- Precipitation
- 6.4: Classifying Chemical Reactions (Acids and Bases)
  - 6.4.1: Classifying Chemical Reactions (Acids and Bases) (Problems)
- 6.5: Classifying Chemical Reactions (Redox)
  - 6.5.1: Classifying Chemical Reactions (Redox) (Problems)

7: Ideal Gas Behavior

- 7.1: Temperature and Pressure
  - 7.1.1: Practice Problems- Temperature and Pressure
- 7.2: The Gas Laws
  - 7.2.1: Practice Problems- The Gas Laws
- 7.3: Applications of the Ideal Gas Law and Partial Pressures
  - 7.3.1: Practice Problems- Applications of the Ideal Gas Law
- 7.4: Kinetic Molecular Theory
  - 7.4.1: Practice Problems- Kinetic Molecular Theory

8: Thermochemistry

- 8.1: The Basics of Energy
  - 8.1.1: Practice Problems- The Basics of Energy
- 8.2: Measuring Energy and Heat Capacity
- 8.3: Enthalpy
- 8.4: Standard Enthalpy and Hess' Law
  - 8.4.1: Practice Problems- Enthalpy and Hess' Law
- 8.5: Calorimetry
  - 8.5.1: Practice Problems- Calorimetry

Index

Glossary

Detailed Licensing

Licensing

*A detailed breakdown of this resource's licensing can be found in **Back Matter/Detailed Licensing**.*

CHAPTER OVERVIEW

1: Matter and Measurement

1.1: Atoms and Molecules
1.2: The Scientific Approach to Knowledge
1.3: The Classification of Matter
1.4: Measurements
1.4.1: Practice Problems on Measurements
1.5: The Mole is a Measure of Amount
1.6: Accuracy and Precision
1.7: Dimensional Analysis
1.7.1: Practice Problems on Dimensional Analysis
1.8: Significant Digits

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1.1: Atoms and Molecules

Skills to Develop

- Outline the historical development of chemistry
- Provide examples of the importance of chemistry in everyday life
- Provide examples illustrating macroscopic, microscopic, and symbolic domains

Throughout human history, people have tried to convert matter into more useful forms. Our Stone Age ancestors chipped pieces of flint into useful tools and carved wood into statues and toys. These endeavors involved changing the shape of a substance without changing the substance itself. But as our knowledge increased, humans began to change the composition of the substances as well—clay was converted into pottery, hides were cured to make garments, copper ores were transformed into copper tools and weapons, and grain was made into bread.

Humans began to practice chemistry when they learned to control fire and use it to cook, make pottery, and smelt metals. Subsequently, they began to separate and use specific components of matter. A variety of drugs such as aloe, myrrh, and opium were isolated from plants. Dyes, such as indigo and Tyrian purple, were extracted from plant and animal matter. Metals were combined to form alloys—for example, copper and tin were mixed together to make bronze—and more elaborate smelting techniques produced iron. Alkalis were extracted from ashes, and soaps were prepared by combining these alkalis with fats. Alcohol was produced by fermentation and purified by distillation.

Attempts to understand the behavior of matter extend back for more than 2500 years. As early as the sixth century BC, Greek philosophers discussed a system in which water was the basis of all things. You may have heard of the Greek postulate that matter consists of four elements: earth, air, fire, and water. Subsequently, an amalgamation of chemical technologies and philosophical speculations were spread from Egypt, China, and the eastern Mediterranean by alchemists, who endeavored to transform “base metals” such as lead into “noble metals” like gold, and to create elixirs to cure disease and extend life (Figure **1.1.1**).



Figure **1.1.1**: This portrayal shows an alchemist’s workshop circa 1580. Although alchemy made some useful contributions to how to manipulate matter, it was not scientific by modern standards. (credit: Chemical Heritage Foundation).

From alchemy came the historical progressions that led to modern chemistry: the isolation of drugs from natural sources, metallurgy, and the dye industry. Today, chemistry continues to deepen our understanding and improve our ability to harness and control the behavior of matter. This effort has been so successful that many people do not realize either the central position of chemistry among the sciences or the importance and universality of chemistry in daily life.

Chemistry: The Central Science

Chemistry is sometimes referred to as “the central science” due to its interconnectedness with a vast array of other STEM disciplines (STEM stands for areas of study in the science, technology, engineering, and math fields). Chemistry and the language of chemists play vital roles in biology, medicine, materials science, forensics, environmental science, and many other fields (Figure **1.1.2**). The basic principles of physics are essential for understanding many aspects of chemistry, and there is extensive overlap between many subdisciplines within the two fields, such as chemical physics and nuclear chemistry. Mathematics, computer science, and information theory provide important tools that help us calculate, interpret, describe, and generally make sense of the chemical world. Biology and chemistry converge in biochemistry, which is crucial to understanding the many complex factors and processes that keep living organisms (such as us) alive. Chemical engineering, materials science, and nanotechnology combine

chemical principles and empirical findings to produce useful substances, ranging from gasoline to fabrics to electronics. Agriculture, food science, veterinary science, and brewing and wine making help provide sustenance in the form of food and drink to the world's population. Medicine, pharmacology, biotechnology, and botany identify and produce substances that help keep us healthy. Environmental science, geology, oceanography, and atmospheric science incorporate many chemical ideas to help us better understand and protect our physical world. Chemical ideas are used to help understand the universe in astronomy and cosmology.

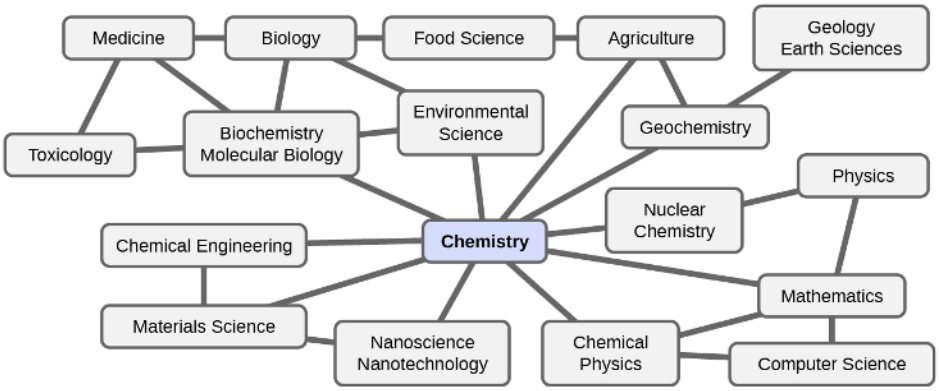


Figure **1.1.2**: Knowledge of chemistry is central to understanding a wide range of scientific disciplines. This diagram shows just some of the interrelationships between chemistry and other fields.

What are some changes in matter that are essential to daily life? Digesting and assimilating food, synthesizing polymers that are used to make clothing, containers, cookware, and credit cards, and refining crude oil into gasoline and other products are just a few examples. As you proceed through this course, you will discover many different examples of changes in the composition and structure of matter, how to classify these changes and how they occurred, their causes, the changes in energy that accompany them, and the principles and laws involved. As you learn about these things, you will be learning chemistry, the study of the composition, properties, and interactions of matter. The practice of chemistry is not limited to chemistry books or laboratories: It happens whenever someone is involved in changes in matter or in conditions that may lead to such changes.

The Domains of Chemistry

Chemists study and describe the behavior of matter and energy in three different domains: macroscopic, microscopic, and symbolic. These domains provide different ways of considering and describing chemical behavior.

Macro is a Greek word that means "large." The macroscopic domain is familiar to us: It is the realm of everyday things that are large enough to be sensed directly by human sight or touch. In daily life, this includes the food you eat and the breeze you feel on your face. The macroscopic domain includes everyday and laboratory chemistry, where we observe and measure physical and chemical properties, or changes such as density, solubility, and flammability.

The microscopic domain of chemistry is almost always visited in the imagination. *Micro* also comes from Greek and means "small." Some aspects of the microscopic domains are visible through a microscope, such as a magnified image of graphite or bacteria. Viruses, for instance, are too small to be seen with the naked eye, but when we're suffering from a cold, we're reminded of how real they are.

However, most of the subjects in the microscopic domain of chemistry—such as atoms and molecules—are too small to be seen even with standard microscopes and often must be pictured in the mind. Other components of the microscopic domain include ions and electrons, protons and neutrons, and chemical bonds, each of which is far too small to see. This domain includes the individual metal atoms in a wire, the ions that compose a salt crystal, the changes in individual molecules that result in a color change, the conversion of nutrient molecules into tissue and energy, and the evolution of heat as bonds that hold atoms together are created.

The symbolic domain contains the specialized language used to represent components of the macroscopic and microscopic domains. Chemical symbols (such as those used in the periodic table), chemical formulas, and chemical equations are part of the symbolic domain, as are graphs and drawings. We can also consider calculations as part of the symbolic domain. These symbols play an important role in chemistry because they help interpret the behavior of the macroscopic domain in terms of the components

of the microscopic domain. One of the challenges for students learning chemistry is recognizing that the same symbols can represent different things in the macroscopic and microscopic domains, and one of the features that makes chemistry fascinating is the use of a domain that must be imagined to explain behavior in a domain that can be observed.

A helpful way to understand the three domains is via the essential and ubiquitous substance of water. That water is a liquid at moderate temperatures, will freeze to form a solid at lower temperatures, and boil to form a gas at higher temperatures (Figure **1.1.3**) are macroscopic observations. But some properties of water fall into the microscopic domain—what we cannot observe with the naked eye. The description of water as comprised of two hydrogen atoms and one oxygen atom, and the explanation of freezing and boiling in terms of attractions between these molecules, is within the microscopic arena. The formula H_2O, which can describe water at either the macroscopic or microscopic levels, is an example of the symbolic domain. The abbreviations (*g*) for gas, (*s*) for solid, and (*l*) for liquid are also symbolic.

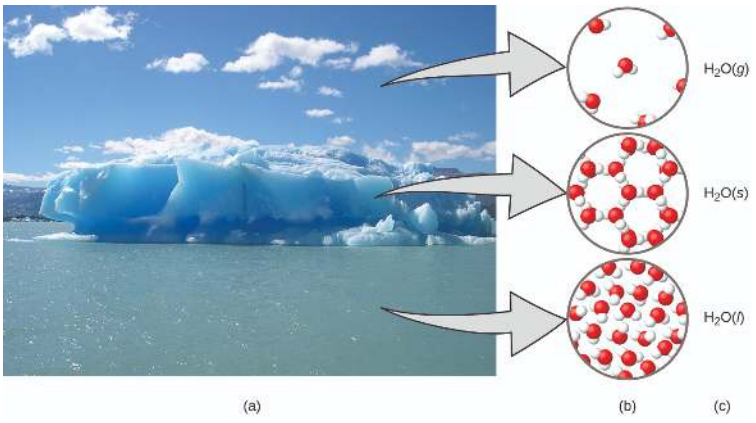


Figure **1.1.3**: (a) Moisture in the air, icebergs, and the ocean represent water in the macroscopic domain. (b) At the molecular level (microscopic domain), gas molecules are far apart and disorganized, solid water molecules are close together and organized, and liquid molecules are close together and disorganized. (c) The formula H_2O symbolizes water, and (g), (s), and (l) symbolize its phases. Note that clouds are actually comprised of either very small liquid water droplets or solid water crystals; gaseous water in our atmosphere is not visible to the naked eye, although it may be sensed as humidity. (credit a: modification of work by "Gorkaazk"/Wikimedia Commons).

Summary

Chemistry deals with the composition, structure, and properties of matter, and the ways by which various forms of matter may be interconverted. Thus, it occupies a central place in the study and practice of science and technology. Chemists use the scientific method to perform experiments, pose hypotheses, and formulate laws and develop theories, so that they can better understand the behavior of the natural world. To do so, they operate in the macroscopic, microscopic, and symbolic domains. Chemists measure, analyze, purify, and synthesize a wide variety of substances that are important to our lives.

Contributors

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1.2: The Scientific Approach to Knowledge

Skills to Develop

- Describe the process of the scientific method

Scientists search for answers to questions and solutions to problems by using a procedure called the **scientific method**. This procedure consists of making observations, formulating hypotheses, and designing experiments, which in turn lead to additional observations, hypotheses, and experiments in repeated cycles (Figure 1.2.1).

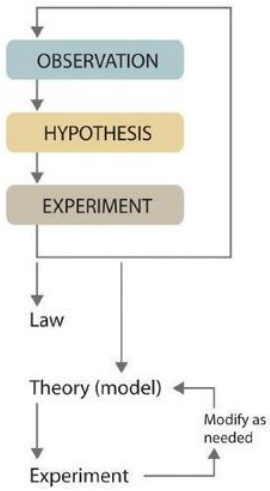


Figure 1.2.1: The Scientific Method. As depicted in this flowchart, the scientific method consists of making observations, formulating hypotheses, and designing experiments. A scientist may enter the cycle at any point.

Observations can be qualitative or quantitative. Qualitative observations describe properties or occurrences in ways that do not rely on numbers. Examples of qualitative observations include the following: the outside air temperature is cooler during the winter season, table salt is a crystalline solid, sulfur crystals are yellow, and dissolving a penny in dilute nitric acid forms a blue solution and a brown gas. Quantitative observations are measurements, which by definition consist of both a number and a unit. Examples of quantitative observations include the following: the melting point of crystalline sulfur is 115.21 °C, and 35.9 grams of table salt —whose chemical name is sodium chloride—dissolve in 100 grams of water at 20 °C. An example of a quantitative observation was the initial observation leading to the modern theory of the dinosaurs' extinction: iridium concentrations in sediments dating to 66 million years ago were found to be 20–160 times higher than normal. The development of this theory is a good exemplar of the scientific method in action (see Figure 1.2.2 below).

After deciding to learn more about an observation or a set of observations, scientists generally begin an investigation by forming a **hypothesis**, a tentative explanation for the observation(s). The hypothesis may not be correct, but it puts the scientist's understanding of the system being studied into a form that can be tested. For example, the observation that we experience alternating periods of light and darkness corresponding to observed movements of the sun, moon, clouds, and shadows is consistent with either of two hypotheses:

1. Earth rotates on its axis every 24 hours, alternately exposing one side to the sun, or
2. The sun revolves around Earth every 24 hours.

Suitable experiments can be designed to choose between these two alternatives. For the disappearance of the dinosaurs, the hypothesis was that the impact of a large extraterrestrial object caused their extinction. Unfortunately (or perhaps fortunately), this hypothesis does not lend itself to direct testing by any obvious experiment, but scientists collected additional data that either support or refute it.

After a hypothesis has been formed, scientists conduct experiments to test its validity. **Experiments** are systematic observations or measurements, preferably made under controlled conditions—that is, under conditions in which a single variable changes. For example, in the dinosaur extinction scenario, iridium concentrations were measured worldwide and compared. A properly designed

and executed experiment enables a scientist to determine whether the original hypothesis is valid. Experiments often demonstrate that the hypothesis is incorrect or that it must be modified. More experimental data are then collected and analyzed, at which point a scientist may begin to think that the results are sufficiently reproducible (i.e., dependable) to merit being summarized in a **law**, a verbal or mathematical description of a phenomenon that allows for general predictions. A law simply says what happens; it does not address the question of why.

One example of a law, the Law of Definite Proportions, which was discovered by the French scientist Joseph Proust (1754–1826), states that a chemical substance always contains the same proportions of elements by mass. Thus sodium chloride (table salt) always contains the same proportion by mass of sodium to chlorine, in this case 39.34% sodium and 60.66% chlorine by mass, and sucrose (table sugar) is always 42.11% carbon, 6.48% hydrogen, and 51.41% oxygen by mass. Some solid compounds do not strictly obey the law of definite proportions. The law of definite proportions should seem obvious—we would expect the composition of sodium chloride to be consistent—but the head of the US Patent Office did not accept it as a fact until the early 20th century.

Whereas a law states only what happens, a **theory** attempts to explain why nature behaves as it does. Laws are unlikely to change greatly over time unless a major experimental error is discovered. In contrast, a theory, by definition, is incomplete and imperfect, evolving with time to explain new facts as they are discovered. The theory developed to explain the extinction of the dinosaurs, for example, is that Earth occasionally encounters small- to medium-sized asteroids, and these encounters may have unfortunate implications for the continued existence of most species. This theory is by no means proven, but it is consistent with the bulk of evidence amassed to date. Figure 1.2.2 summarizes the application of the scientific method in this case.

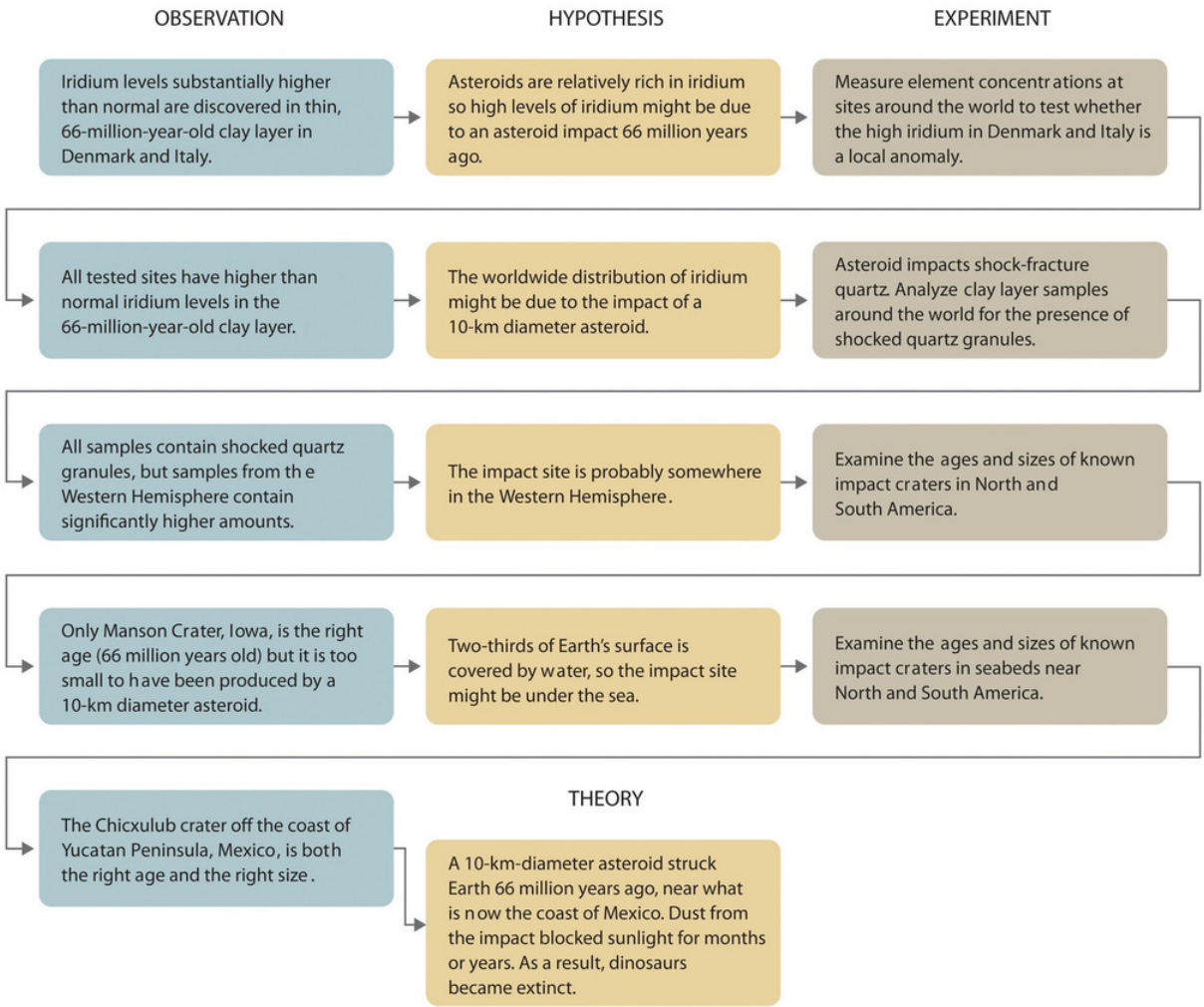


Figure 1.2.2: A Summary of How the Scientific Method Was Used in Developing the Asteroid Impact Theory to Explain the Disappearance of the Dinosaurs from Earth

Example 1.2.1

Classify each statement as a law, a theory, an experiment, a hypothesis, a qualitative observation, or a quantitative observation.

a. Ice always floats on liquid water.
b. Birds evolved from dinosaurs.
c. Hot air is less dense than cold air, probably because the components of hot air are moving more rapidly.
d. When 10 g of ice were added to 100 mL of water at 25 °C, the temperature of the water decreased to 15.5 °C after the ice melted.
e. The ingredients of Ivory soap were analyzed to see whether it really is 99.44% pure, as advertised.

Given: components of the scientific method

Asked for: statement classification

Strategy: Refer to the definitions in this section to determine which category best describes each statement.

Solution

a. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.
b. This is a possible explanation for the origin of birds, so it is a hypothesis.
c. This is a statement that tries to explain the relationship between the temperature and the density of air based on fundamental principles, so it is a theory.
d. The temperature is measured before and after a change is made in a system, so these are quantitative observations.
e. This is an analysis designed to test a hypothesis (in this case, the manufacturer's claim of purity), so it is an experiment.

Exercise 1.2.1

Classify each statement as a law, a theory, an experiment, a hypothesis, a qualitative observation, or a quantitative observation.

a. Measured amounts of acid were added to a Rolaids tablet to see whether it really "consumes 47 times its weight in excess stomach acid."
b. Heat always flows from hot objects to cooler ones, not in the opposite direction.
c. The universe was formed by a massive explosion that propelled matter into a vacuum.
d. Michael Jordan is the greatest pure shooter ever to play professional basketball.
e. Limestone is relatively insoluble in water but dissolves readily in dilute acid with the evolution of a gas.
f. Gas mixtures that contain more than 4% hydrogen in air are potentially explosive.

Answer a

experiment

Answer b

law

Answer c

theory

Answer d

hypothesis

Answer e

qualitative observation

Answer f

quantitative observation

Because scientists can enter the cycle shown in Figure **1.2.1** at any point, the actual application of the scientific method to different topics can take many different forms. For example, a scientist may start with a hypothesis formed by reading about work done by

others in the field, rather than by making direct observations.

It is important to remember that scientists have a tendency to formulate hypotheses in familiar terms simply because it is difficult to propose something that has never been encountered or imagined before. As a result, scientists sometimes discount or overlook unexpected findings that disagree with the basic assumptions behind the hypothesis or theory being tested. Fortunately, truly important findings are immediately subject to independent verification by scientists in other laboratories, so science is a self-correcting discipline. When the Alvarezes originally suggested that an extraterrestrial impact caused the extinction of the dinosaurs, the response was almost universal skepticism and scorn. In only 20 years, however, the persuasive nature of the evidence overcame the skepticism of many scientists, and their initial hypothesis has now evolved into a theory that has revolutionized paleontology and geology.

Summary

Chemists expand their knowledge by making observations, carrying out experiments, and testing hypotheses to develop laws to summarize their results and theories to explain them. In doing so, they are using the scientific method.

Contributors

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1.3: The Classification of Matter

Skills to Develop

- Describe the fundamental properties of each physical state of matter: solid, liquid, and gas
- Define and give examples of atoms and molecules
- Classify matter as an element, or compound; pure substance, or a mixture; a homogenous mixture, or heterogenous mixture
- Classify matter as chemical or physical
- Classify properties of matter as chemical, or physical; an extensive or intensive
- Distinguish between mass and weight

Matter is defined as anything that occupies space and has mass, and it is all around us. Solids and liquids are more obviously matter: We can see that they take up space, and their weight tells us that they have mass. Gases are also matter; if gases did not take up space, a balloon would stay collapsed rather than inflate when filled with gas.

Solids, liquids, and gases are the three states of matter commonly found on earth (Figure **1.3.1**). A solid is rigid and possesses a definite shape. A liquid flows and takes the shape of a container, except that it forms a flat or slightly curved upper surface when acted upon by gravity. (In zero gravity, liquids assume a spherical shape.) Both liquid and solid samples have volumes that are very nearly independent of pressure. A gas takes both the shape and volume of its container.

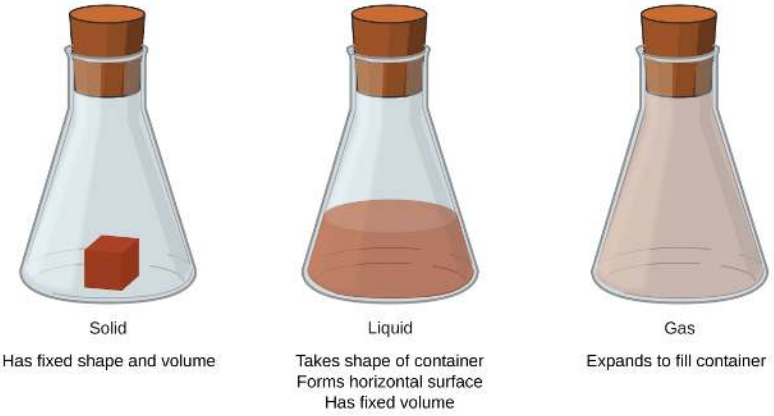


Figure **1.3.1**: The three most common states or phases of matter are solid, liquid, and gas.

A fourth state of matter, plasma, occurs naturally in the interiors of stars. A plasma is a gaseous state of matter that contains appreciable numbers of electrically charged particles (Figure **1.3.2**). The presence of these charged particles imparts unique properties to plasmas that justify their classification as a state of matter distinct from gases. In addition to stars, plasmas are found in some other high-temperature environments (both natural and man-made), such as lightning strikes, certain television screens, and specialized analytical instruments used to detect trace amounts of metals.



Figure **1.3.2**: A plasma torch can be used to cut metal. (credit: “Hypertherm”/Wikimedia Commons)



Video* 1.3.1: In a tiny cell in a plasma television, the plasma emits ultraviolet light, which in turn causes the display at that location to appear a specific color. The composite of these tiny dots of color makes up the image that you see. Watch this video to learn more about plasma and the places you encounter it.*

Some samples of matter appear to have properties of solids, liquids, and/or gases at the same time. This can occur when the sample is composed of many small pieces. For example, we can pour sand as if it were a liquid because it is composed of many small grains of solid sand. Matter can also have properties of more than one state when it is a mixture, such as with clouds. Clouds appear to behave somewhat like gases, but they are actually mixtures of air (gas) and tiny particles of water (liquid or solid).

The **mass** of an object is a measure of the amount of matter in it. One way to measure an object's mass is to measure the force it takes to accelerate the object. It takes much more force to accelerate a car than a bicycle because the car has much more mass. A more common way to determine the mass of an object is to use a balance to compare its mass with a standard mass.

Although weight is related to mass, it is not the same thing. Weight refers to the force that gravity exerts on an object. This force is directly proportional to the mass of the object. The weight of an object changes as the force of gravity changes, but its mass does not. An astronaut's mass does not change just because she goes to the moon. But her weight on the moon is only one-sixth her earth-bound weight because the moon's gravity is only one-sixth that of the earth's. She may feel "weightless" during her trip when she experiences negligible external forces (gravitational or any other), although she is, of course, never "massless."

The **law of conservation of matter** summarizes many scientific observations about matter: It states that *there is no detectable change in the total quantity of matter present when matter converts from one type to another (a chemical change) or changes among solid, liquid, or gaseous states (a physical change)*. Brewing beer and the operation of batteries provide examples of the conservation of matter (Figure **1.3.3**). During the brewing of beer, the ingredients (water, yeast, grains, malt, hops, and sugar) are converted into beer (water, alcohol, carbonation, and flavoring substances) with no actual loss of substance. This is most clearly seen during the bottling process, when glucose turns into ethanol and carbon dioxide, and the total mass of the substances does not change. This can also be seen in a lead-acid car battery: The original substances (lead, lead oxide, and sulfuric acid), which are capable of producing electricity, are changed into other substances (lead sulfate and water) that do not produce electricity, with no change in the actual amount of matter.

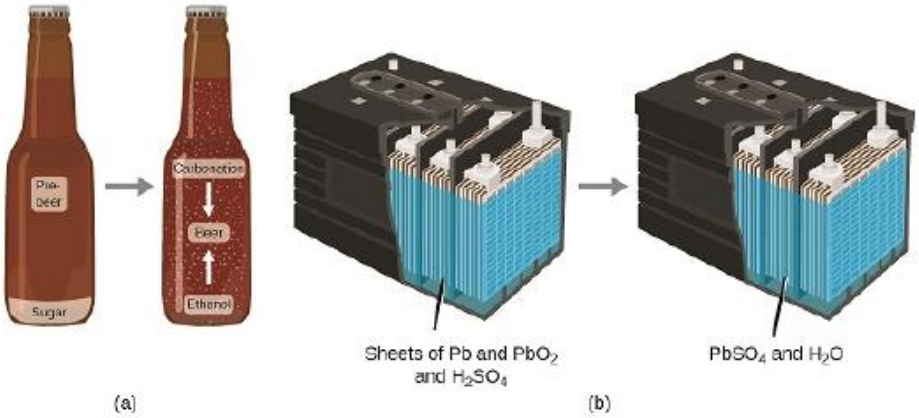


Figure **1.3.3**: (a) The mass of beer precursor materials is the same as the mass of beer produced: Sugar has become alcohol and carbonation. (b) The mass of the lead, lead oxide plates, and sulfuric acid that goes into the production of electricity is exactly equal to the mass of lead sulfate and water that is formed.

Although this conservation law holds true for all conversions of matter, convincing examples are few and far between because, outside of the controlled conditions in a laboratory, we seldom collect all of the material that is produced during a particular conversion. For example, when you eat, digest, and assimilate food, all of the matter in the original food is preserved. But because some of the matter is incorporated into your body, and much is excreted as various types of waste, it is challenging to verify by measurement.

Atoms and Molecules

An **atom** is the smallest particle of an element that has the properties of that element and can enter into a chemical combination. Consider the element gold, for example. Imagine cutting a gold nugget in half, then cutting one of the halves in half, and repeating this process until a piece of gold remained that was so small that it could not be cut in half (regardless of how tiny your knife may be). This minimally sized piece of gold is an atom (from the Greek *atomos*, meaning “indivisible”)

The first suggestion that matter is composed of atoms is attributed to the Greek philosophers Leucippus and Democritus, who developed their ideas in the 5th century BCE. However, it was not until the early nineteenth century that John Dalton (1766–1844), a British schoolteacher with a keen interest in science, supported this hypothesis with quantitative measurements. Since that time, repeated experiments have confirmed many aspects of this hypothesis, and it has become one of the central theories of chemistry. Other aspects of Dalton’s atomic theory are still used but with minor revisions (details of Dalton’s theory are provided in the chapter on atoms and molecules).

An atom is so small that its size is difficult to imagine. One of the smallest things we can see with our unaided eye is a single thread of a spider web: These strands are about 1/10,000 of a centimeter (0.0001 cm) in diameter. Although the cross-section of one strand is almost impossible to see without a microscope, it is huge on an atomic scale. A single carbon atom in the web has a diameter of about 0.000000015 centimeter, and it would take about 7000 carbon atoms to span the diameter of the strand. To put this in perspective, if a carbon atom were the size of a dime, the cross-section of one strand would be larger than a football field, which would require about 150 million carbon atom “dimes” to cover it. (Figure **1.3.5**) shows increasingly close microscopic and atomic-level views of ordinary cotton.

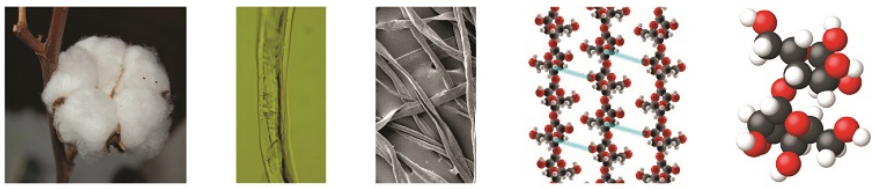


Figure **1.3.5**: These images provide an increasingly closer view: (a) a cotton boll, (b) a single cotton fiber viewed under an optical microscope (magnified 40 times), (c) an image of a cotton fiber obtained with an electron microscope (much higher magnification than with the optical microscope); and (d and e) atomic-level models of the fiber (spheres of different colors represent atoms of different elements). (credit c: modification of work by “Featheredtar”/Wikimedia Commons)

An atom is so light that its mass is also difficult to imagine. A billion lead atoms (1,000,000,000 atoms) weigh about 3×10^{-13} grams, a mass that is far too light to be weighed on even the world's most sensitive balances. It would require over 300,000,000,000,000 lead atoms (300 trillion, or 3×10^{14}) to be weighed, and they would weigh only 0.0000001 gram.

It is rare to find collections of individual atoms. Only a few elements, such as the gases helium, neon, and argon, consist of a collection of individual atoms that move about independently of one another. Other elements, such as the gases hydrogen, nitrogen, oxygen, and chlorine, are composed of units that consist of pairs of atoms (Figure **1.3.6**). One form of the element phosphorus consists of units composed of four phosphorus atoms. The element sulfur exists in various forms, one of which consists of units composed of eight sulfur atoms. These units are called molecules. A molecule consists of two or more atoms joined by strong forces called chemical bonds. The atoms in a molecule move around as a unit, much like the cans of soda in a six-pack or a bunch of keys joined together on a single key ring. A molecule may consist of two or more identical atoms, as in the molecules found in the elements hydrogen, oxygen, and sulfur, or it may consist of two or more different atoms, as in the molecules found in water. Each water molecule is a unit that contains two hydrogen atoms and one oxygen atom. Each glucose molecule is a unit that contains 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. Like atoms, molecules are incredibly small and light. If an ordinary glass of water were enlarged to the size of the earth, the water molecules inside it would be about the size of golf balls.

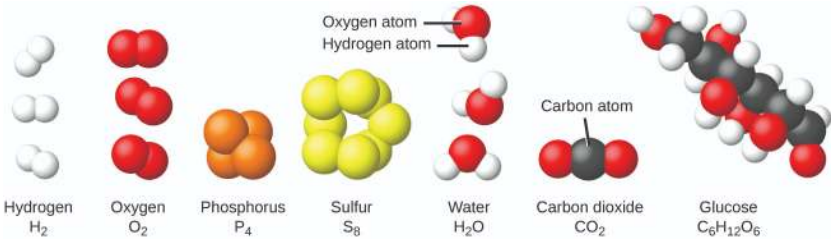


Figure **1.3.6**: The elements hydrogen, oxygen, phosphorus, and sulfur form molecules consisting of two or more atoms of the same element. The compounds water, carbon dioxide, and glucose consist of combinations of atoms of different elements.

Classifying Matter

We can classify matter into several categories. Two broad categories are mixtures and pure substances. A pure substance has a constant composition. All specimens of a pure substance have exactly the same makeup and properties. Any sample of sucrose (table sugar) consists of 42.1% carbon, 6.5% hydrogen, and 51.4% oxygen by mass. Any sample of sucrose also has the same physical properties, such as melting point, color, and sweetness, regardless of the source from which it is isolated.

We can divide pure substances into two classes: elements and compounds. Pure substances that cannot be broken down into simpler substances by chemical changes are called elements. Iron, silver, gold, aluminum, sulfur, oxygen, and copper are familiar examples of the more than 100 known elements, of which about 90 occur naturally on the earth, and two dozen or so have been created in laboratories.

Pure substances that can be broken down by chemical changes are called compounds. This breakdown may produce either elements or other compounds, or both. Mercury(II) oxide, an orange, crystalline solid, can be broken down by heat into the elements mercury and oxygen (Figure **1.3.7**). When heated in the absence of air, the compound sucrose is broken down into the element carbon and the compound water. (The initial stage of this process, when the sugar is turning brown, is known as caramelization—this is what imparts the characteristic sweet and nutty flavor to caramel apples, caramelized onions, and caramel). Silver(I) chloride is a white solid that can be broken down into its elements, silver and chlorine, by absorption of light. This property is the basis for the use of this compound in photographic films and photochromic eyeglasses (those with lenses that darken when exposed to light).



Figure 1.3.7: (a)The compound mercury(II) oxide, (b)when heated, (c) decomposes into silvery droplets of liquid mercury and invisible oxygen gas. (credit: modification of work by Paul Flowers)

The properties of combined elements are different from those in the free, or uncombined, state. For example, white crystalline sugar (sucrose) is a compound resulting from the chemical combination of the element carbon, which is a black solid in one of its uncombined forms, and the two elements hydrogen and oxygen, which are colorless gases when uncombined. Free sodium, an element that is a soft, shiny, metallic solid, and free chlorine, an element that is a yellow-green gas, combine to form sodium chloride (table salt), a compound that is a white, crystalline solid.

A mixture is composed of two or more types of matter that can be present in varying amounts and can be separated by physical changes, such as evaporation (you will learn more about this later). A mixture with a composition that varies from point to point is called a heterogeneous mixture. Italian dressing is an example of a heterogeneous mixture (Figure 1.3.8*a*). Its composition can vary because we can make it from varying amounts of oil, vinegar, and herbs. It is not the same from point to point throughout the mixture—one drop may be mostly vinegar, whereas a different drop may be mostly oil or herbs because the oil and vinegar separate and the herbs settle. Other examples of heterogeneous mixtures are chocolate chip cookies (we can see the separate bits of chocolate, nuts, and cookie dough) and granite (we can see the quartz, mica, feldspar, and more).

A homogeneous mixture, also called a solution, exhibits a uniform composition and appears visually the same throughout. An example of a solution is a sports drink, consisting of water, sugar, coloring, flavoring, and electrolytes mixed together uniformly (Figure 1.3.8*b*). Each drop of a sports drink tastes the same because each drop contains the same amounts of water, sugar, and other components. Note that the composition of a sports drink can vary—it could be made with somewhat more or less sugar, flavoring, or other components, and still be a sports drink. Other examples of homogeneous mixtures include air, maple syrup, gasoline, and a solution of salt in water.



Figure 1.3.8: (a) Oil and vinegar salad dressing is a heterogeneous mixture because its composition is not uniform throughout. (b) A commercial sports drink is a homogeneous mixture because its composition is uniform throughout. (credit a “left”: modification of work by John Mayer; credit a “right”: modification of work by Umberto Salvagnin; credit b “left: modification of work by Jeff Bedford)

Although there are just over 100 elements, tens of millions of chemical compounds result from different combinations of these elements. Each compound has a specific composition and possesses definite chemical and physical properties by which we can distinguish it from all other compounds. And, of course, there are innumerable ways to combine elements and compounds to form different mixtures. The overall organization of matter and the methods used to separate mixtures are summarized in Figure 1.3.6.

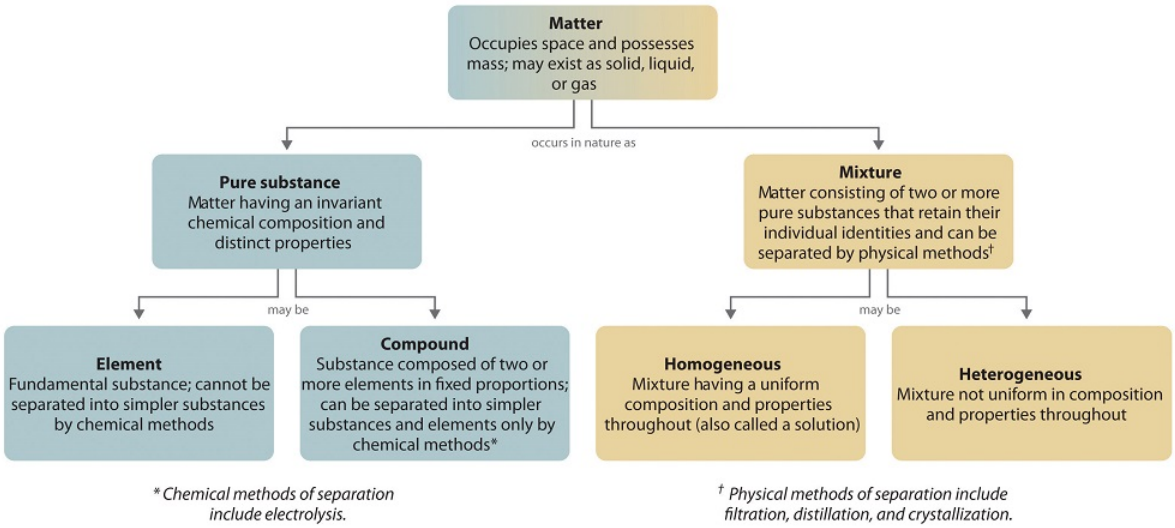


Figure **1.3.6**: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

Example 1.3.1

Identify each substance as a compound, an element, a heterogeneous mixture, or a homogeneous mixture (solution).

a. filtered tea
b. freshly squeezed orange juice
c. a compact disc
d. aluminum oxide, a white powder that contains a 2:3 ratio of aluminum and oxygen atoms
e. selenium

Given: a chemical substance

Asked for: its classification

Strategy:

A. Decide whether a substance is chemically pure. If it is pure, the substance is either an element or a compound. If a substance can be separated into its elements, it is a compound.
B. If a substance is not chemically pure, it is either a heterogeneous mixture or a homogeneous mixture. If its composition is uniform throughout, it is a homogeneous mixture.

Solution

a. **A** Tea is a solution of compounds in water, so it is not chemically pure. It is usually separated from tea leaves by filtration. **B** Because the composition of the solution is uniform throughout, it is a homogeneous mixture.
b. **A** Orange juice contains particles of solid (pulp) as well as liquid; it is not chemically pure. B Because its composition is not uniform throughout, orange juice is a heterogeneous mixture.
c. **A** A compact disc is a solid material that contains more than one element, with regions of different compositions visible along its edge. Hence a compact disc is not chemically pure. **B** The regions of different composition indicate that a compact disc is a heterogeneous mixture.
d. **A** Aluminum oxide is a single, chemically pure compound.
e. **A** Selenium is one of the known elements.

Exercise 1.3.1

Identify each substance as a compound, an element, a heterogeneous mixture, or a homogeneous mixture (solution).

a. white wine
b. mercury

c. ranch-style salad dressing
d. table sugar (sucrose)

Answer A

solution

Answer B

element

Answer C

heterogeneous mixture

Answer D

compound

Physical and Chemical Properties of Matter

The characteristics that enable us to distinguish one substance from another are called properties. A physical property is a characteristic of matter that is not associated with a change in its chemical composition. Familiar examples of physical properties include density, color, hardness, melting and boiling points, and electrical conductivity. We can observe some physical properties, such as density and color, without changing the physical state of the matter observed. Other physical properties, such as the melting temperature of iron or the freezing temperature of water, can only be observed as matter undergoes a physical change. A physical change is a change in the state or properties of matter without any accompanying change in its chemical composition (the identities of the substances contained in the matter). We observe a physical change when wax melts, when sugar dissolves in coffee, and when steam condenses into liquid water (Figure 1.3.1). Other examples of physical changes include magnetizing and demagnetizing metals (as is done with common antitheft security tags) and grinding solids into powders (which can sometimes yield noticeable changes in color). In each of these examples, there is a change in the physical state, form, or properties of the substance, but no change in its chemical composition.



Figure 1.3.1: (a) Wax undergoes a physical change when solid wax is heated and forms liquid wax. (b) Steam condensing inside a cooking pot is a physical change, as water vapor is changed into liquid water. (credit a: modification of work by "95jb14"/Wikimedia Commons; credit b: modification of work by "mjneuby"/Flickr).

The change of one type of matter into another type (or the inability to change) is a chemical property. Examples of chemical properties include flammability, toxicity, acidity, reactivity (many types), and heat of combustion. Iron, for example, combines with oxygen in the presence of water to form rust; chromium does not oxidize (Figure 1.3.2). Nitroglycerin is very dangerous because it explodes easily; neon poses almost no hazard because it is very unreactive.



(a) (b)

Figure 1.3.2: (a) One of the chemical properties of iron is that it rusts; (b) one of the chemical properties of chromium is that it does not. (credit a: modification of work by Tony Hisgett; credit b: modification of work by “Atoma”/Wikimedia Commons)

To identify a chemical property, we look for a chemical change. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance. Other examples of chemical changes include reactions that are performed in a lab (such as copper reacting with nitric acid), all forms of combustion (burning), and food being cooked, digested, or rotting.

Properties of matter fall into one of two categories. If the property depends on the amount of matter present, it is an extensive property. The mass and volume of a substance are examples of extensive properties; for instance, a gallon of milk has a larger mass and volume than a cup of milk. The value of an extensive property is directly proportional to the amount of matter in question. If the property of a sample of matter does not depend on the amount of matter present, it is an intensive property. Temperature is an example of an intensive property. If the gallon and cup of milk are each at 20 °C (room temperature), when they are combined, the temperature remains at 20 °C. As another example, consider the distinct but related properties of heat and temperature. A drop of hot cooking oil spattered on your arm causes brief, minor discomfort, whereas a pot of hot oil yields severe burns. Both the drop and the pot of oil are at the same temperature (an intensive property), but the pot clearly contains much more heat (extensive property).

While many elements differ dramatically in their chemical and physical properties, some elements have similar properties. We can identify sets of elements that exhibit common behaviors. For example, many elements conduct heat and electricity well, whereas others are poor conductors. These properties can be used to sort the elements into three classes: metals (elements that conduct well), nonmetals (elements that conduct poorly), and metalloids (elements that have properties of both metals and nonmetals).

The periodic table is a table of elements that places elements with similar properties close together (Figure 1.3.5). You will learn more about the periodic table as you continue your study of chemistry.

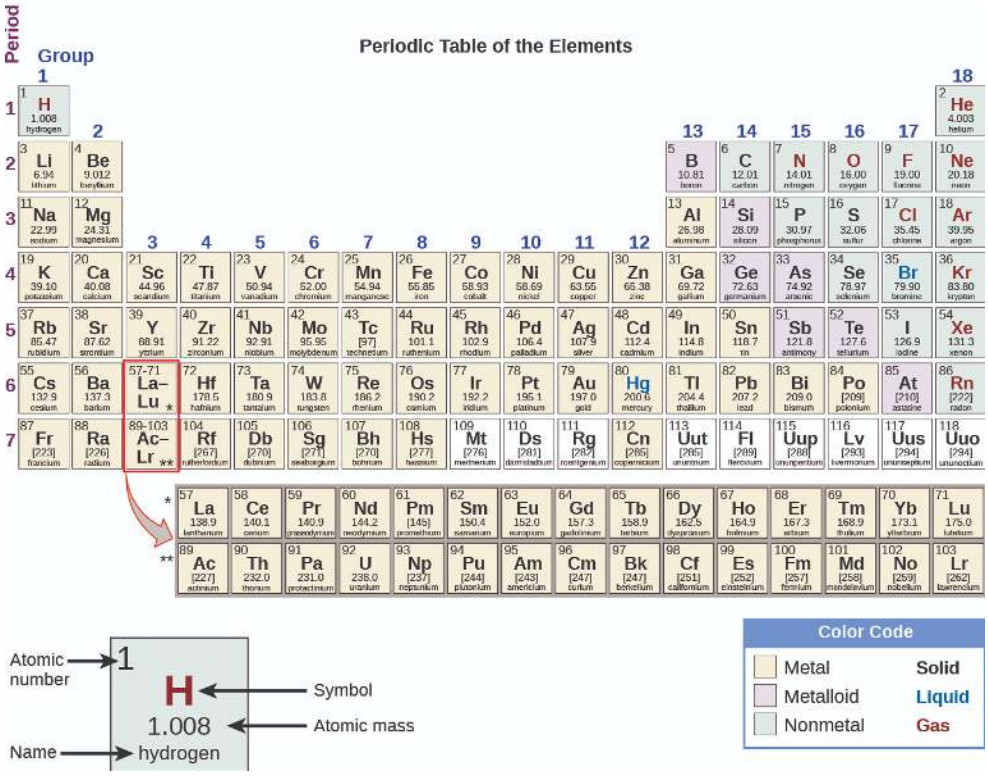


Figure **1.3.5**: The periodic table shows how elements may be grouped according to certain similar properties. Note the background color denotes whether an element is a metal, metalloid, or nonmetal, whereas the element symbol color indicates whether it is a solid, liquid, or gas.

Glossary

atom
: smallest particle of an element that can enter into a chemical combination

chemical change
: change producing a different kind of matter from the original kind of matter

chemical property
: behavior that is related to the change of one kind of matter into another kind of matter

compound
: pure substance that can be decomposed into two or more elements

element
: substance that is composed of a single type of atom; a substance that cannot be decomposed by a chemical change

extensive property
: property of a substance that depends on the amount of the substance

gas
: state in which matter has neither definite volume nor shape

heterogeneous mixture
: combination of substances with a composition that varies from point to point

homogeneous mixture
: (also, solution) combination of substances with a composition that is uniform throughout

intensive property

property of a substance that is independent of the amount of the substance

liquid

state of matter that has a definite volume but indefinite shape

law of conservation of matter

when matter converts from one type to another or changes form, there is no detectable change in the total amount of matter present

mass

fundamental property indicating amount of matter

matter

anything that occupies space and has mass

mixture

matter that can be separated into its components by physical means

molecule

bonded collection of two or more atoms of the same or different elements

physical change

change in the state or properties of matter that does not involve a change in its chemical composition

physical property

characteristic of matter that is not associated with any change in its chemical composition

plasma

gaseous state of matter containing a large number of electrically charged atoms and/or molecules

pure substance

homogeneous substance that has a constant composition

solid

state of matter that is rigid, has a definite shape, and has a fairly constant volume

weight

force that gravity exerts on an object

Summary

Matter is anything that occupies space and has mass. The basic building block of matter is the atom, the smallest unit of an element that can enter into combinations with atoms of the same or other elements. In many substances, atoms are combined into molecules. On earth, matter commonly exists in three states: solids, of fixed shape and volume; liquids, of variable shape but fixed volume; and gases, of variable shape and volume. Under high-temperature conditions, matter also can exist as a plasma. Most matter is a mixture: It is composed of two or more types of matter that can be present in varying amounts and can be separated by physical means. Heterogeneous mixtures vary in composition from point to point; homogeneous mixtures have the same composition from point to point. Pure substances consist of only one type of matter. A pure substance can be an element, which consists of only one type of atom and cannot be broken down by a chemical change, or a compound, which consists of two or more types of atoms.

Matter can be classified according to physical and chemical properties. Matter is anything that occupies space and has mass. The three states of matter are solid, liquid, and gas. A physical change involves the conversion of a substance from one state of matter to another, without changing its chemical composition. Most matter consists of mixtures of pure substances, which can be homogeneous (uniform in composition) or heterogeneous (different regions possess different compositions and properties). Pure substances can be either chemical compounds or elements. Compounds can be broken down into elements by chemical reactions, but elements cannot be separated into simpler substances by chemical means. The properties of substances can be classified as either physical or chemical. Scientists can observe physical properties without changing the composition of the substance, whereas chemical properties describe the tendency of a substance to undergo chemical changes (chemical reactions) that change its

chemical composition. Physical properties can be intensive or extensive. Intensive properties are the same for all samples; do not depend on sample size; and include, for example, color, physical state, and melting and boiling points. Extensive properties depend on the amount of material and include mass and volume. The ratio of two extensive properties, mass and volume, is an important intensive property called density.

Contributors and Attributions

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- Modified by Joshua Halpern (Howard University)

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1.4: Measurements

Skills to Develop

- Explain the process of measurement
- Identify the two basic parts of a quantity
- Describe the properties and units of length, mass, volume, density, and time
- Perform basic unit calculations and conversions in the metric and other unit systems

An Introduction to Units From Crash Course Chemistry



Video* 1.4.1*: Watch this video for an introduction to units for this unit

Measurements provide the macroscopic information that is the basis of most of the hypotheses, theories, and laws that describe the behavior of matter and energy in both the macroscopic and microscopic domains of chemistry. For our purposes, every measurement provides two kinds of information: the size or magnitude of the measurement (a number or quantity); and a standard of comparison for the measurement (a unit). The number and unit are explicitly represented when a quantity is written.

The number in the measurement can be represented in different ways, including decimal form and scientific notation. For example, the maximum takeoff weight of a Boeing 777-200ER airliner is 298,000 kilograms, which can also be written as 2.98×10^5 kg. The mass of the average mosquito is about 0.0000025 kilograms, which can be written as 2.5×10^{-6} kg.

Units, such as liters, pounds, and centimeters, are standards of comparison for measurements. When we buy a 2-liter bottle of a soft drink, we expect that the volume of the drink was measured, so it is two times larger than the volume that everyone agrees to be 1 liter. The meat used to prepare a 0.25-pound hamburger is measured so it weighs one-fourth as much as 1 pound. Without units, a number can be meaningless, confusing, or possibly life threatening. Suppose a doctor prescribes phenobarbital to control a patient's seizures and states a dosage of "100" without specifying units. Not only will this be confusing to the medical professional giving the dose, but the consequences can be dire: 100 mg given three times per day can be effective as an anticonvulsant, but a single dose of 100 g is more than 10 times the lethal amount.

We usually report the results of scientific measurements in SI units, an updated version of the metric system, using the units listed in Table **1.4.1**. Other units can be derived from these base units. The standards for these units are fixed by international agreement, and they are called the International System of Units or SI Units (from the French, *Le Système International d'Unités*). SI units have been used by the United States National Institute of Standards and Technology (NIST) since 1964.

Table* 1.4.1*: *Base Units of the SI System*

Property Measured	Name of Unit	Symbol of Unit
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K
electric current	ampere	A
amount of substance	mole	mol
luminous intensity	candela	cd

Sometimes we use units that are fractions or multiples of a base unit. Ice cream is sold in quarts (a familiar, non-SI base unit), pints (0.5 quart), or gallons (4 quarts). We also use fractions or multiples of units in the SI system, but these fractions or multiples are always powers of 10. Fractional or multiple SI units are named using a prefix and the name of the base unit. For example, a length of 1000 meters is also called a kilometer because the prefix *kilo* means "one thousand," which in scientific notation is 10^3 (1 kilometer = 1000 m = 10^3 m). The prefixes used and the powers to which 10 are raised are listed in Table **1.4.2**.

***Table* 1.4.2:** *Common Unit Prefixes*

Prefix	Symbol	Factor	Example
femto	f	10^{-15}	1 femtosecond (fs) = 1×10^{-15} s (0.000000000000001 s)
pico	p	10^{-12}	1 picometer (pm) = 1×10^{-12} m (0.000000000001 m)
nano	n	10^{-9}	4 nanograms (ng) = 4×10^{-9} g (0.000000004 g)
micro	μ	10^{-6}	1 microliter (μL) = 1×10^{-6} L (0.000001 L)
milli	m	10^{-3}	2 millimoles (mmol) = 2×10^{-3} mol (0.002 mol)
centi	c	10^{-2}	7 centimeters (cm) = 7×10^{-2} m (0.07 m)
deci	d	10^{-1}	1 deciliter (dL) = 1×10^{-1} L (0.1 L)
kilo	k	10^3	1 kilometer (km) = 1×10^3 m (1000 m)
mega	M	10^6	3 megahertz (MHz) = 3×10^6 Hz (3,000,000 Hz)
giga	G	10^9	8 gigayears (Gyr) = 8×10^9 yr (8,000,000,000 Gyr)
tera	T	10^{12}	5 terawatts (TW) = 5×10^{12} W (5,000,000,000,000 W)

SI Base Units

The initial units of the metric system, which eventually evolved into the SI system, were established in France during the French Revolution. The original standards for the meter and the kilogram were adopted there in 1799 and eventually by other countries.

This section introduces four of the SI base units commonly used in chemistry. Other SI units will be introduced in subsequent chapters.

Length

The standard unit of length in both the SI and original metric systems is the meter (m). A meter was originally specified as 1/10,000,000 of the distance from the North Pole to the equator. It is now defined as the distance light in a vacuum travels in 1/299,792,458 of a second. A meter is about 3 inches longer than a yard (Figure **1.4.1**); one meter is about 39.37 inches or 1.094 yards. Longer distances are often reported in kilometers (1 km = 1000 m = 10^3 m), whereas shorter distances can be reported in centimeters (1 cm = 0.01 m = 10^{-2} m) or millimeters (1 mm = 0.001 m = 10^{-3} m).

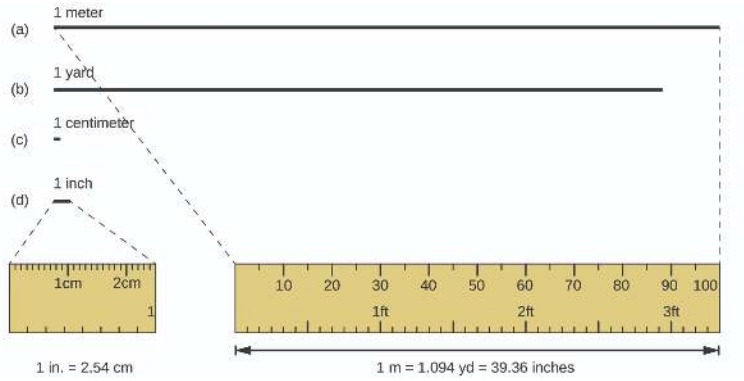


Figure* 1.4.1: The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.*

Mass

The standard unit of mass in the SI system is the kilogram (kg). A kilogram was originally defined as the mass of a liter of water (a cube of water with an edge length of exactly 0.1 meter). In 1889, it was redefined by a certain cylinder of platinum-iridium alloy, which was kept in France (Figure **1.4.2**). Any object with the same mass as this cylinder was said to have a mass of 1 kilogram (which can lead to uncertainties unacceptable to the precision of modern instrumentation). One kilogram is about 2.2 pounds. The gram (g) is exactly equal to 1/1000 of the mass of the kilogram (10^{-3} kg). Over the past 100 years, the IPK has lost 50 millionths of a gram - a seemingly negligible amount, but something that has caused it to be lighter - or all standard replicas to be heavier - and changing the definition of a kilogram in the process. As all balances in the world are standardized to this value, it is important that this value, itself, be standard. On May 20, 2019, a new definition will be used for the kilogram, based on the unchanging Planck's constant.[1]



***Figure* 1.4.2*:** This replica prototype kilogram is housed at the National Institute of Standards and Technology (NIST) in Maryland. (credit: National Institutes of Standards and Technology).*



Video* 1.4.2*: For more information on the new definition of the kilogram, check out this video!

Time

The SI base unit of time is the second (s). Small and large time intervals can be expressed with the appropriate prefixes; for example, 3 microseconds = 0.000003 s = 3×10^{-6} and 5 megaseconds = 5,000,000 s = 5×10^{6} s. Alternatively, hours, days, and years can be used.

Derived SI Units

We can derive many units from the seven SI base units. For example, we can use the base unit of length to define a unit of volume, and the base units of mass and length to define a unit of density.

Volume

Volume is the measure of the amount of space occupied by an object. The standard SI unit of volume is defined by the base unit of length (Figure **1.4.3**). The standard volume is a cubic meter (m^3), a cube with an edge length of exactly one meter. To dispense a cubic meter of water, we could build a cubic box with edge lengths of exactly one meter. This box would hold a cubic meter of water or any other substance.

A more commonly used unit of volume is derived from the decimeter (0.1 m, or 10 cm). A cube with edge lengths of exactly one decimeter contains a volume of one cubic decimeter (dm^3). A liter (L) is the more common name for the cubic decimeter. One liter is about 1.06 quarts. A cubic centimeter (cm^3) is the volume of a cube with an edge length of exactly one centimeter. The abbreviation **cc** (for **c**ubic **c**entimeter) is often used by health professionals. A cubic centimeter is also called a milliliter (mL) and is 1/1000 of a liter.

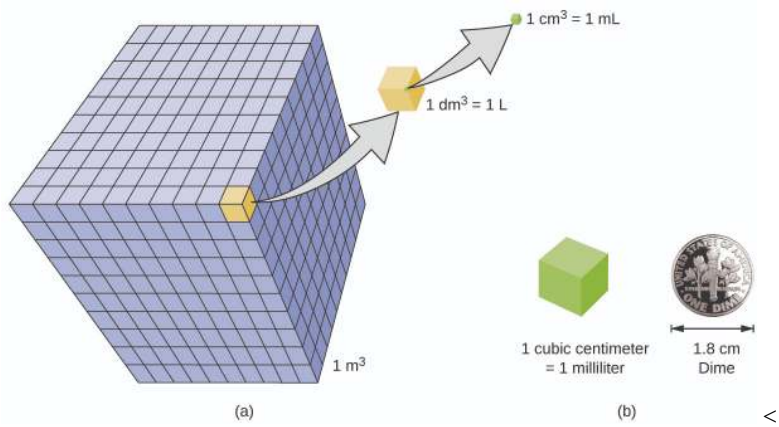


Figure* 1.4.3: (a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube.*

Density

We use the mass and volume of a substance to determine its density. Thus, the units of density are defined by the base units of mass and length.

The density of a substance is the ratio of the mass of a sample of the substance to its volume. The SI unit for density is the kilogram per cubic meter (kg/m^3). For many situations, however, this as an inconvenient unit, and we often use grams per cubic centimeter (g/cm^3) for the densities of solids and liquids, and grams per liter (g/L) for gases. Although there are exceptions, most liquids and solids have densities that range from about 0.7 g/cm^3 (the density of gasoline) to 19 g/cm^3 (the density of gold). The density of air is about 1.2 g/L. Table **1.4.3** shows the densities of some common substances.

Table 1.4.3: Densities of Common Substances

Solids	Liquids	Gases (at 25 °C and 1 atm)
ice (at 0 °C) 0.92 g/cm^3	water 1.0 g/cm^3	dry air 1.20 g/L
oak (wood) 0.60–0.90 g/cm^3	ethanol 0.79 g/cm^3	oxygen 1.31 g/L
iron 7.9 g/cm^3	acetone 0.79 g/cm^3	nitrogen 1.14 g/L
copper 9.0 g/cm^3	glycerin 1.26 g/cm^3	carbon dioxide 1.80 g/L
lead 11.3 g/cm^3	olive oil 0.92 g/cm^3	helium 0.16 g/L
silver 10.5 g/cm^3	gasoline 0.70–0.77 g/cm^3	neon 0.83 g/L
gold 19.3 g/cm^3	mercury 13.6 g/cm^3	radon 9.1 g/L

While there are many ways to determine the density of an object, perhaps the most straightforward method involves separately finding the mass and volume of the object, and then dividing the mass of the sample by its volume. In the following example, the mass is found directly by weighing, but the volume is found indirectly through length measurements.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{1.4.1}$$

Example 1.4.1

Calculation of Density Gold—in bricks, bars, and coins—has been a form of currency for centuries. In order to swindle people into paying for a brick of gold without actually investing in a brick of gold, people have considered filling the centers of hollow gold bricks with lead to fool buyers into thinking that the entire brick is gold. It does not work: Lead is a dense substance, but its density is not as great as that of gold, 19.3 g/cm^3. What is the density of lead if a cube of lead has an edge length of 2.00 cm and a mass of 90.7 g?

Solution

The density of a substance can be calculated by dividing its mass by its volume. The volume of a cube is calculated by cubing the edge length.

$$\text{volume of lead cube} = 2.00\ \text{cm} \times 2.00\ \text{cm} \times 2.00\ \text{cm} = 8.00\ \text{cm}^3$$

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{90.7\ \text{g}}{8.00\ \text{cm}^3} = \frac{11.3\ \text{g}}{1.00\ \text{cm}^3} = 11.3\ \text{g/cm}^3$$

(We will discuss the reason for rounding to the first decimal place in the next section.)

Exercise 1.4.1

a. To three decimal places, what is the volume of a cube (cm^3) with an edge length of 0.843 cm?
b. If the cube in part (a) is copper and has a mass of 5.34 g, what is the density of copper to two decimal places?

Answer a

0.599 cm^3;

Answer b

8.91 g/cm^3

Example 1.4.2: Using Displacement of Water to Determine Density

This PhET simulation illustrates another way to determine density, using displacement of water. Determine the density of the red and yellow blocks.

Solution

When you open the density simulation and select Same Mass, you can choose from several 5.00-kg colored blocks that you can drop into a tank containing 100.00 L water. The yellow block floats (it is less dense than water), and the water level rises to 105.00 L. While floating, the yellow block displaces 5.00 L water, an amount equal to the weight of the block. The red block sinks (it is more dense than water, which has density = 1.00 kg/L), and the water level rises to 101.25 L.

The red block therefore displaces 1.25 L water, an amount equal to the volume of the block. The density of the red block is:

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{5.00\ \text{kg}}{1.25\ \text{L}} = 4.00\ \text{kg/L} \tag{1.4.2}$$

Note that since the yellow block is not completely submerged, you cannot determine its density from this information. But if you hold the yellow block on the bottom of the tank, the water level rises to 110.00 L, which means that it now displaces 10.00 L water, and its density can be found:

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{5.00\ \text{kg}}{10.00\ \text{L}} = 0.500\ \text{kg/L} \tag{1.4.3}$$

Exercise 1.4.1

Remove all of the blocks from the water and add the green block to the tank of water, placing it approximately in the middle of the tank. Determine the density of the green block.

Answer

2.00 kg/L

Summary

Measurements provide quantitative information that is critical in studying and practicing chemistry. Each measurement has an amount, a unit for comparison, and an uncertainty. Measurements can be represented in either decimal or scientific notation. Scientists primarily use the SI (International System) or metric systems. We use base SI units such as meters, seconds, and kilograms, as well as derived units, such as liters (for volume) and g/cm^3 (for density). In many cases, we find it convenient to use unit prefixes that yield fractional and multiple units, such as microseconds (10^{-6} seconds) and megahertz (10^6 hertz), respectively.

Footnotes

1. Read more about the redefinition of SI units including the kilogram here (Laura Howe, *CE&N*, Nov. 16, 2018).

Key Equations

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

Glossary

cubic centimeter (cm^3 or cc)

volume of a cube with an edge length of exactly 1 cm

cubic meter (m^3)

SI unit of volume

density

ratio of mass to volume for a substance or object

kilogram (kg)
standard SI unit of mass; 1 kg = approximately 2.2 pounds

length
measure of one dimension of an object

liter (L)
(also, cubic decimeter) unit of volume; 1 L = 1,000 cm^3

meter (m)
standard metric and SI unit of length; 1 m = approximately 1.094 yards

milliliter (mL)
1/1,000 of a liter; equal to 1 cm^3

second (s)
SI unit of time

SI units (International System of Units)
standards fixed by international agreement in the International System of Units (*Le Système International d'Unités*)

unit
standard of comparison for measurements

volume
amount of space occupied by an object

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1.4.1: Practice Problems on Measurements

PROBLEM 1.4.1.1

Indicate the SI base units or derived units that are appropriate for the following measurements.

(a) the mass of the moon

(b) the distance from Dallas to Oklahoma City

(c) the speed of sound

(d) the density of air

(e) the area of the state of Oregon

(f) the volume of a flu shot

Answer a

kilogram

Answer b

meter

Answer c

meter per second

Answer d

kilogram per cubic meter

Answer e

square meters

Answer f

cubic meters

PROBLEM 1.4.1.2

Indicate the SI base units or derived units that are appropriate for the following measurements.

(a) the length of a marathon race (26.2 miles)

(b) the mass of an automobile

(c) the volume of a swimming pool

(d) the speed of an airplane

(e) the density of gold

(f) the area of a football field

Answer a

meter

Answer b

kilogram

Answer c

cubic meter

Answer d

meter per second

Answer e

kilogram per cubic meter

Answer f

square meter

PROBLEM 1.4.1.3

A large piece of jewelry has a mass of 132.6 g. A graduated cylinder initially contains 48.6 mL of water. When the jewelry is submerged in the graduated cylinder, the total volume increases to 61.2 mL.

(a) Determine the density of this piece of jewelry (in g/mL).

(b) Assuming the jewelry is made of only one substance, what substance is it likely to be (Refer to Table 1.1.3 in the previous section). Explain your answer.

Answer a

10.5 g/mL

Answer b

Silver (Density is 10.5 g/mL in Table 1.1.3)

Click here to see a video of the solution.



PROBLEM 1.4.1.4

Give the name and symbol of the prefixes used with SI units to indicate multiplication by the following exact quantities.

(a) 10^3

(b) 10^{-2}

(c) 0.1

(d) 10^{-3}

(e) 1,000,000

(f) 0.000001

Answer a

kilo- (k)

Answer b

centi- (c)

Answer c

deci- (d)

Answer d

milli- (m)

Answer e

Mega (M)

Answer f

micro (u or μ)

PROBLEM 1.4.1.5

Give the name of the prefix and the quantity indicated by the following symbols that are used with SI base units.

(a) c

(b) d

(c) G

(d) k

(e) m

(f) n

(g) p

(h) T

Answer a

centi (10^{-2})

Answer b

deci (10^{-1})

Answer c

giga (10^{9})

Answer d

kilo (10^{3})

Answer e

milli (10^{-3})

Answer f

nano (10^{-9})

Answer g

pico (10^{-12})

Answer h

Tera (10^{12})

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1.5: The Mole is a Measure of Amount

Skills to Develop

- Define the amount unit mole and the related quantity Avogadro's number

The Mole

The identity of a substance is defined not only by the types of atoms or ions it contains, but by the quantity of each type of atom or ion. For example, water, H_2O, and hydrogen peroxide, H_2O_2, are alike in that their respective molecules are composed of hydrogen and oxygen atoms. However, because a hydrogen peroxide molecule contains two oxygen atoms, as opposed to the water molecule, which has only one, the two substances exhibit very different properties. Today, we possess sophisticated instruments that allow the direct measurement of these defining microscopic traits; however, the same traits were originally derived from the measurement of macroscopic properties (the masses and volumes of bulk quantities of matter) using relatively simple tools (balances and volumetric glassware). This experimental approach required the introduction of a new unit for amount of substances, the *mole*, which remains indispensable in modern chemical science.

The mole is an amount unit similar to familiar units like pair, dozen, gross, etc. It provides a specific measure of *the number* of atoms or molecules in a bulk sample of matter. A mole is defined as the amount of substance containing the same number of discrete entities (such as atoms, molecules, and ions) as the number of atoms in a sample of pure ^{12}C weighing exactly 12 g. One Latin connotation for the word "mole" is "large mass" or "bulk," which is consistent with its use as the name for this unit. The mole provides a link between an easily measured macroscopic property, bulk mass, and an extremely important fundamental property, number of atoms, molecules, and so forth.

The number of entities composing a mole has been experimentally determined to be $\mathbf{6.02214179 \times 10^{23}}$, a fundamental constant named Avogadro's number (N_A) or the Avogadro constant in honor of Italian scientist Amedeo Avogadro. This constant is properly reported with an explicit unit of "per mole," a conveniently rounded version being $\mathbf{6.022 \times 10^{23}/mol}$.



Video* 1.5.1: What is Avogadro's Number?*

Consistent with its definition as an amount unit, 1 mole of any element contains the same number of atoms as 1 mole of any other element. The masses of 1 mole of different elements, however, are different, since the masses of the individual atoms are drastically different. The molar mass of an element (or compound) is the mass in grams of 1 mole of that substance, a property expressed in units of grams per mole (g/mol).

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1.6: Accuracy and Precision



Figure **1.6.1** (Public Domain; Daniel Arizpe via Commons Wikimedia, High School Basketball Game [commons.wikimedia.org])

How do professional basketball players improve their shooting accuracy?

Basketball is one of those sports where you need to hit the target. A football field goal kicker might have room for some deviation from a straight line – for college and pro football there is an 18 foot 6 inch space for the ball to go through. In basketball, the basket is only 18 inches across and the ball is a little less than 10 inches across – not much room for error. The ball has to be on target in order to go into the basket and score.

Accuracy and Precision

In everyday speech, the terms **accuracy** and **precision** are frequently used interchangeably. However, their scientific meanings are quite different. Accuracy is a measure of how close a measurement is to the correct or accepted value of the quantity being measured. Precision is a measure of how close a series of measurements are to one another. Precise measurements are highly reproducible, even if the measurements are not near the correct value. Darts thrown at a dartboard are helpful in illustrating accuracy and precision.

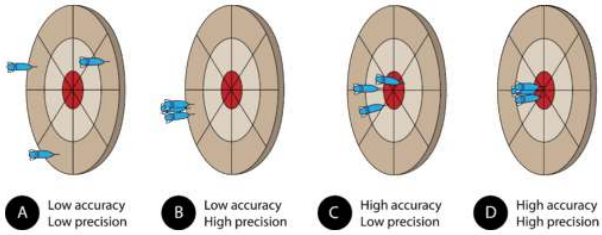


Figure **1.6.2**: The distribution of darts on a dartboard shows the difference between accuracy and precision. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Assume that three darts are thrown at the dartboard, with the bulls-eye representing the true, or accepted, value of what is being measured. A dart that hits the bulls-eye is highly accurate, whereas a dart that lands far away from the bulls-eye displays poor accuracy. The figure above demonstrates four possible outcomes:

a. The darts have landed far from each other and far from the bulls-eye. This grouping demonstrates measurements that are neither accurate nor precise.
b. The darts are close to one another, but far from the bulls-eye. This grouping demonstrates measurements that are precise, but not accurate. In a laboratory situation, high precision with low accuracy often results from a systematic error. Either the measurer makes the same mistake repeatedly, or the measuring tool is somehow flawed. A poorly calibrated balance may give the same mass reading every time, but it will be far from the true mass of the object.
c. The darts are not grouped very near to each other, but are generally centered around the bulls-eye. This demonstrates poor precision, but fairly high accuracy. This situation is not desirable in a lab situation because the "high" accuracy may simply be random chance and not a true indicator of good measuring skill.
d. The darts are grouped together and have hit the bulls-eye. This demonstrates high precision and high accuracy. Scientists always strive to maximize both in their measurements.



Figure **1.6.3**: Students in a chemistry lab are making careful measurements with a series of volumetric flasks. Accuracy and precision are critical in every experiment. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)



Summary

- Accuracy is a measure of how close a measurement is to the correct or accepted value of the quantity being measured.
- Precision is a measure of how close a series of measurements are to one another.

Review

1. Define accuracy.
2. Define precision.
3. What can be said about the reproducibility of precise values?

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1.7: Dimensional Analysis

Skills to Develop

- Explain the dimensional analysis (factor label) approach to mathematical calculations involving quantities
- Use dimensional analysis to carry out unit conversions for a given property and computations involving two or more properties
- Perform dimensional analysis calculations with units raised to a power
- Use density as a conversion factor

An Introduction to Dimensional Analysis From Crash Course Chemistry



Video 1.7.1: Watch this video for an introduction to dimensional analysis

It is often the case that a quantity of interest may not be easy (or even possible) to measure directly but instead must be calculated from other directly measured properties and appropriate mathematical relationships. For example, consider measuring the average speed of an athlete running sprints. This is typically accomplished by measuring the *time* required for the athlete to run from the starting line to the finish line, and the *distance* between these two lines, and then computing *speed* from the equation that relates these three properties:

$$\text{speed} = \frac{\text{distance}}{\text{time}} \tag{1.7.1}$$

An Olympic-quality sprinter can run 100 m in approximately 10 s, corresponding to an average speed of

$$\frac{100\ \text{m}}{10\ \text{s}} = 10\ \text{m/s} \tag{1.7.2}$$

Note that this simple arithmetic involves dividing the numbers of each measured quantity to yield the number of the computed quantity (100/10 = 10) *and likewise* dividing the units of each measured quantity to yield the unit of the computed quantity (m/s = m/s). Now, consider using this same relation to predict the time required for a person running at this speed to travel a distance of 25 m. The same relation between the three properties is used, but in this case, the two quantities provided are a speed (10 m/s) and a distance (25 m). To yield the sought property, time, the equation must be rearranged appropriately:

$$\text{time} = \frac{\text{distance}}{\text{speed}} \tag{1.7.3}$$

The time can then be computed as:

$$\frac{25\ \text{m}}{10\ \text{m/s}} = 2.5\ \text{s} \tag{1.7.4}$$

Again, arithmetic on the numbers (25/10 = 2.5) was accompanied by the same arithmetic on the units (m/m/s = s) to yield the number and unit of the result, 2.5 s. Note that, just as for numbers, when a unit is divided by an identical unit (in this case, m/m), the result is “1”—or, as commonly phrased, the units “cancel.”

These calculations are examples of a versatile mathematical approach known as dimensional analysis (or the factor-label method). Dimensional analysis is based on this premise: *the units of quantities must be subjected to the same mathematical operations as their associated numbers*. This method can be applied to computations ranging from simple unit conversions to more complex, multi-step calculations involving several different quantities.

Conversion Factors and Dimensional Analysis

A ratio of two equivalent quantities expressed with different measurement units can be used as a unit conversion factor. For example, the lengths of 2.54 cm and 1 in. are equivalent (by definition), and so a unit conversion factor may be derived from the ratio,

$$\frac{2.54\ \text{cm}}{1\ \text{in.}}\ (2.54\ \text{cm} = 1\ \text{in.})\ \text{or}\ 2.54\ \frac{\text{cm}}{\text{in.}} \tag{1.7.5}$$

Several other commonly used conversion factors are given in Table 1.7.1.

***Table* 1.7.1:** *Common Conversion Factors*

Length	Volume	Mass
1 m = 1.0936 yd = 3.28 ft	1 L = 1.0567 qt = 10^{-3} m^3 = 1 dm^3 = 10^3 cm^3	1 kg = 2.2046 lb
1 in. = 2.54 cm (exact)	1 qt = 0.94635 L	1 lb = 453.59 g = 16 oz
1 cm = 0.39370 in	1 gal = 4 qt = 3.7854 L	28.35 g = 1 oz
1 km = 0.62137 mi	1 ft^3 = 28.317 L	
1 mi = 1609.3 m = 5280 ft	1 tbsp = 14.787 mL	
	1 cm^3 = 1 mL	
	1 in^3 = 16.4 cm^3	
	1 fl. oz. = 29.5735 mL	

When we multiply a quantity (such as distance given in inches) by an appropriate unit conversion factor, we convert the quantity to an equivalent value with different units (such as distance in centimeters). For example, a basketball player's vertical jump of 34 inches can be converted to centimeters by:

$$34\ \cancel{\text{in.}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in.}}} = 86\ \text{cm} \tag{1.7.6}$$

Since this simple arithmetic involves *quantities*, the premise of dimensional analysis requires that we multiply both *numbers and units*. The numbers of these two quantities are multiplied to yield the number of the product quantity, 86, whereas the units are multiplied to yield

$$\frac{\text{in.} \times \text{cm}}{\text{in.}}. \tag{1.7.7}$$

Just as for numbers, a ratio of identical units is also numerically equal to one,

$$\frac{\text{in.}}{\text{in.}} = 1 \tag{1.7.8}$$

and the unit product thus simplifies to *cm*. (When identical units divide to yield a factor of 1, they are said to "cancel.") Using dimensional analysis, we can determine that a unit conversion factor has been set up correctly by checking to confirm that the original unit will cancel, and the result will contain the sought (converted) unit.

Example 1.7.1: Using a Unit Conversion Factor

The mass of a competition Frisbee is 125 g. Convert its mass to ounces using the unit conversion factor derived from the relationship 1 oz = 28.349 g (Table 1.7.1).

Solution

If we have the conversion factor, we can determine the mass in kilograms using an equation similar the one used for converting length from inches to centimeters.

$$x \text{ oz} = 125 \text{ g} \times \text{unit conversion factor}$$

We write the unit conversion factor in its two forms:

$$\frac{1 \text{ oz}}{28.349 \text{ g}} \text{ and } \frac{28.349 \text{ g}}{1 \text{ oz}}$$

The correct unit conversion factor is the ratio that cancels the units of grams and leaves ounces.

$$\begin{aligned} x \text{ oz} &= 125 \cancel{\text{g}} \times \frac{1 \text{ oz}}{28.349 \cancel{\text{g}}} \\ &= \left(\frac{125}{28.349}\right) \text{ oz} \\ &= 4.41 \text{ oz (three significant figures)} \end{aligned}$$

Exercise 1.7.1

Convert a volume of 9.345 qt to liters.

Answer

8.844 L

Beyond simple unit conversions, the factor-label method can be used to solve more complex problems involving computations. Regardless of the details, the basic approach is the same—all the *factors* involved in the calculation must be appropriately oriented to insure that their *labels* (units) will appropriately cancel and/or combine to yield the desired unit in the result. This is why it is referred to as the factor-label method. As your study of chemistry continues, you will encounter many opportunities to apply this approach.

Example 1.7.2: Computing Quantities from Measurement Results

What is the density of common antifreeze in units of g/mL? A 4.00-qt sample of the antifreeze weighs 9.26 lb.

Solution

Since $\text{density} = \frac{\text{mass}}{\text{volume}}$, we need to divide the mass in grams by the volume in milliliters. In general: the number of units of B = the number of units of A × unit conversion factor. The necessary conversion factors are given in Table 1.7.1: 1 lb = 453.59 g; 1 L = 1.0567 qt; 1 L = 1,000 mL. We can convert mass from pounds to grams in one step:

$$9.26 \cancel{\text{lb}} \times \frac{453.59 \text{ g}}{1 \cancel{\text{lb}}} = 4.20 \times 10^3 \text{ g}$$

We need to use two steps to convert volume from quarts to milliliters.

1. *Convert quarts to liters.*

$$4.00 \cancel{\text{qt}} \times \frac{1 \text{ L}}{1.0567 \cancel{\text{qt}}} = 3.78 \text{ L}$$

2. *Convert liters to milliliters.*

$$3.78 \cancel{\text{L}} \times \frac{1000 \text{ mL}}{1 \cancel{\text{L}}} = 3.78 \times 10^3 \text{ mL}$$

Then,

$$\text{density} = \frac{4.20 \times 10^3 \text{ g}}{3.78 \times 10^3 \text{ mL}} = 1.11 \text{ g/mL}$$

Alternatively, the calculation could be set up in a way that uses three unit conversion factors sequentially as follows:

$$\frac{9.26\ \cancel{\text{lb}}}{4.00\ \cancel{\text{qt}}} \times \frac{453.59\ \text{g}}{1\ \cancel{\text{lb}}} \times \frac{1.0567\ \cancel{\text{qt}}}{1\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{L}}}{1000\ \text{mL}} = 1.11\ \text{g/mL}$$

Exercise 1.7.2

What is the volume in liters of 1.000 oz, given that 1 L = 1.0567 qt and 1 qt = 32 oz (exactly)?

Answer

$2.956 \times 10^{-2}\ \text{L}$

Example 1.7.3: Computing Quantities from Measurement Results

While being driven from Philadelphia to Atlanta, a distance of about 1250 km, a 2014 Lamborghini Aventador Roadster uses 213 L gasoline.

a. What (average) fuel economy, in miles per gallon, did the Roadster get during this trip?
b. If gasoline costs $3.80 per gallon, what was the fuel cost for this trip?

Solution

(a) We first convert distance from kilometers to miles:

$$1250\ \text{km} \times \frac{0.62137\ \text{mi}}{1\ \text{km}} = 777\ \text{mi}$$

and then convert volume from liters to gallons:

$$213\ \cancel{\text{L}} \times \frac{1.0567\ \cancel{\text{qt}}}{1\ \cancel{\text{L}}} \times \frac{1\ \text{gal}}{4\ \cancel{\text{qt}}} = 56.3\ \text{gal}$$

Then,

$$\text{(average) mileage} = \frac{777\ \text{mi}}{56.3\ \text{gal}} = 13.8\ \text{miles/gallon} = 13.8\ \text{mpg}$$

Alternatively, the calculation could be set up in a way that uses all the conversion factors sequentially, as follows:

$$\frac{1250\ \cancel{\text{km}}}{213\ \cancel{\text{L}}} \times \frac{0.62137\ \text{mi}}{1\ \cancel{\text{km}}} \times \frac{1\ \cancel{\text{L}}}{1.0567\ \cancel{\text{qt}}} \times \frac{4\ \cancel{\text{qt}}}{1\ \text{gal}} = 13.8\ \text{mpg}$$

(b) Using the previously calculated volume in gallons, we find:

$$56.3\ \text{gal} \times \frac{\$3.80}{1\ \text{gal}} = \$214$$

Exercise 1.7.3

A Toyota Prius Hybrid uses 59.7 L gasoline to drive from San Francisco to Seattle, a distance of 1300 km (two significant digits).

a. What (average) fuel economy, in miles per gallon, did the Prius get during this trip?
b. If gasoline costs $3.90 per gallon, what was the fuel cost for this trip?

Answer a

51 mpg

Answer b

$62

Units Raised to a Power

When calculating area or volume, you multiply together lengths, widths, and heights. Just like in our dimensional analysis above, our units and our numbers both undergo the mathematical operation, meaning that multiplying the quantity of length by the quantity of width also multiplies the units.

$$42cm * 33cm = 1386cm^2 \tag{1.7.9}$$

This complicates the conversion of units, however, since our GIVEN conversion factors often only account for one dimension, not two or three. When building a conversion factor for units raised to a power, we simply raise the conversion factor to the power we want our final units in. The trick is to remember to raise the QUANTITY and UNITS by the power.

Example 1.7.4: Computing Quantities from Measurement Results

Convert 5.70 L to in^3.

Solution

We must first convert L to mL, which as we saw in Section 1.1, is equivalent to cm^3.

$$5.70L * \frac{1000mL}{1L} * \frac{1cm^3}{1mL} = 5700cm^3 \tag{1.7.10}$$

Now, if we examine the table of conversion factors (*Table* **1.7.1**), we find that there is 16.4 cm^3 in 1 in^3.

$$5700cm^3 * \frac{1in^3}{16.4cm^3} = 347.6cm^3 \tag{1.7.11}$$

We could have just as easily have done this if we hadn't been given the direct conversion factor between cm^3 and in^3. We simply would have had to raise the conversion factor between cm and in to the third power.

$$5700cm^3 * \left(\frac{1in}{2.54cm}\right)^3 = 347.6in^3 \tag{1.7.12}$$

The trick with this way of doing the calculation is you have to remember to apply the power to EVERYTHING:

$$\left(\frac{1in}{2.54cm}\right)^3 = \frac{(1^3in^3)}{2.54^3cm^3} \tag{1.7.13}$$

Note: We are ignoring a concept known as "significant figures" in this example. You will cover the rules for significant figures in next week's lab.

Exercise 1.7.4

Convert 274 m^2 to in^2.

Answer

424701 in^2

Using Density as your Conversion Factor

Density can also be used as a conversion factor. This is useful for determining how much the volume of something weighs without having to mass the object - which is particularly useful if your object is too heavy to actually weigh, like a jet or rocket.

Exercise 1.7.5

A commercial jet is fueled with 156,874 L of jet fuel.

a) If the density of the fuel is 0.768 g/cm^3, what is the mass of the fuel in kilograms?

b) If the jet weights 443.613 Mg without passengers or fuel, what is the mass when the fuel is added?

Answer a

120,479 kg

Answer b

564,092 kg or 564.092 Mg

Summary

Measurements are made using a variety of units. It is often useful or necessary to convert a measured quantity from one unit into another. These conversions are accomplished using unit conversion factors, which are derived by simple applications of a mathematical approach called the factor-label method or dimensional analysis. This strategy is also employed to calculate sought quantities using measured quantities and appropriate mathematical relations.

Glossary

dimensional analysis

(also, factor-label method) versatile mathematical approach that can be applied to computations ranging from simple unit conversions to more complex, multi-step calculations involving several different quantities

unit conversion factor

ratio of equivalent quantities expressed with different units; used to convert from one unit to a different unit

Contributors

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1.7.1: Practice Problems on Dimensional Analysis

PROBLEM 1.7.1.1

Write the conversion factors (as ratios) for the number of:

(a) kilometers in 1 mile

(b) liters in 1 liquid quart

(c) grams in 1 ounce

Answer a

1.6093 km : 1 mi

Answer b

0.94635 L : 1 qt

Answer c

28.35 g : 1 oz

PROBLEM 1.7.1.2

The label on a soft drink bottle gives the volume in two units: 2.0 L and 67.6 fl oz. Use this information to derive a conversion factor between the English and metric units.

Answer

$$\frac{2.0L}{67.6floz.} = \frac{0.030L}{1floz.} \tag{1.7.1.1}$$

Click here to see a video of the solution



PROBLEM 1.7.1.3

The label on a box of cereal gives the mass of cereal in two units: 978 grams and 34.5 oz. Use this information to find a conversion factor between the English and metric units.

Answer

$$\frac{978g}{34.5oz} = \frac{28.35g}{1oz.} \tag{1.7.1.2}$$

PROBLEM 1.7.1.4

Soccer is played with a round ball having a circumference between 27 and 28 inches and a mass between 14 and 16 oz. What are these specifications in cm and g?

Answer

68.6 cm to 71.1 cm
396.9 g to 453.6 g

PROBLEM 1.7.1.5

How many milliliters are in a 12 oz soda can?

Answer

354.9 mL

PROBLEM 1.7.1.6

A barrel of oil is exactly 42 gal. How many liters of oil are in the barrel?

Answer

159 L

Click here to see a video of the solution.



PROBLEM 1.7.1.7

The diameter of a red blood cell is about 3 x 10^{-4} inches. What is the diameter in centimeters?

Answer

7.6 x 10^{-4} cm

PROBLEM 1.7.1.8

The distance between the centers of two oxygen atoms in an oxygen molecule is 1.21×10^{-8} cm. What is this distance in inches?

Answer

4.76×10^{-9} in

PROBLEM 1.7.1.9

Is a 197-lb weight lifter light enough to compete in a class limited to those weighing 90 kg or less?

Answer

Yes (They weigh 89.35 kg)

PROBLEM 1.7.1.10

Complete the following conversions between SI units.

(a) 612 g = ? mg

(b) 8.160 m = ? cm

(c) 3779 µg = ? g

(d) 781 mL = ? L

(e) 4.18 kg = ? g

(f) 27.8 m = ? km

(g) 0.13 mL = ? L

(h) 1738 km = ? m

(i) 1.9 Gg = ? g

Answer a

612,000 mg

Answer b

816.0 cm

Answer c

3.779×10^{-3} g

Answer d

0.781 L

Answer e

4180 g

Answer f

0.0278 km

Answer g

1.3×10^{-4} L

Answer h

1,738,000 m

Answer i

1.9×10^{9} g

PROBLEM 1.7.1.11

Make the conversion indicated in each of the following:

(a) the men's world record long jump, 29 ft 4.5 in, to meters

(b) the greatest depth of the ocean, about 6.5 mi, to kilometers

(c) the area of an 8.5 by 11 inch sheet of paper in cm^2

(d) The displacement volume of an automobile engine, 161 in^3, to L

(e) the estimated mass of the atmosphere, 5.6×10^{15} tons, to kilograms (1 ton = 2000 lbs)

(f) the mass of a bushel of rye, 32.0 lb, to kilograms

(g) the mass of a 5.00 grain aspirin tablet to milligrams (1 grain = 0.00229 oz)

Answer a

8.96 m

Answer b

10.46 km

Answer c

603.22 cm^2

Answer d

2.64 L

Answer e

5.08×10^{18} kg

Answer f

14.52 kg

Answer g

324 mg

Click here to see a video of the solution(s).



PROBLEM 1.7.1.12

Many chemistry conferences have held a 50-Trillion Angstrom (Å) Run. How long is this run in kilometers and in miles? (1 Å = 1 x 10^{-10} m)

Answer

5 kilometers or 3.1 miles

PROBLEM 1.7.1.13

As an instructor is preparing for an experiment, he requires 225 g phosphoric acid. The only container readily available is a 150-mL Erlenmeyer flask. Is it large enough to contain the acid, the density of which is 1.83 g/mL?

Answer

Yes, because the acid's volume will be 122.95 mL

PROBLEM 1.7.1.14

In a recent Grand Prix, the winner completed the race with an average speed of 229.8 km/h. What was the speed in miles per hour, meters per second, and feet per second?

Answer

142.8 mi/h; 63.8 m/s; 209 ft/s

Click here to see a video of the solution.



PROBLEM 1.7.1.15

Calculate these masses.

(a) what is the mass of 6.00 cm^3 of mercury (density = 13.5939 g/cm^3)?

(b) what is the mass of 25.0 mL octane (density = 0.702 g/cm^3)?

(c) what is the mass of 4.00 cm^3 of sodium (density = 0.97 g/cm^3)?

(d) What is the mass of 125 mL gaseous chlorine (density = 3.16 g/L)?

Answer a

81.5634 g

Answer b

17.55 g

Answer c

3.88 g

Answer d

0.395 g

Click here to see a video of the soltion.



PROBLEM 1.7.1.16

Calculate the following volumes.

(a) What is the volume of 25 g of iodine (density = 4.93 g/cm^3)?

(b) What is the volume of 3.28 g gaseous hydrogen (density = 0.089 g/L)?

(c) What is the volume of 11.3 g graphite (density = 2.25 g/cm^3)?

(d) What is the volume of 39.657 g bromine (density = 2.928 g/cm^3)?

Answer a

5.07 mL

Answer b

36.9 L

Answer c

5.02 mL

Answer d

13.54 mL

Click here to see a video of the solution.



PROBLEM 1.7.1.17

Convert 195.7 in^2 to m^2.

Answer

0.126 m^2

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1.8: Significant Digits

Accuracy and precision are very important in chemistry. However, the laboratory equipment and machines used in labs are limited in such a way that they can only determine a certain amount of data. For example, a scale can only mass an object up until a certain decimal place, because no machine is advanced enough to determine an infinite amount of digits. Machines are only able to determine a certain amount of digits precisely. These numbers that are determined precisely are called significant digits. Thus, a scale that could only mass until 99.999 mg, could only measure up to 5 figures of accuracy (5 significant digits). Furthermore, in order to have accurate calculations, the end calculation should not have more significant digits than the original set of data.

Introduction

Significant Digits - Number of digits in a figure that express the precision of a measurement instead of its magnitude. The easiest method to determine significant digits is done by first determining whether or not a number has a decimal point. This rule is known as the *Atlantic-Pacific Rule*. The rule states that if a decimal point is Absent, then the zeroes on the Atlantic/right side are insignificant. If a decimal point is Present, then the zeroes on the Pacific/left side are insignificant.

General Rules for Determining Number of Significant Figures

1. All nonzero digits are significant.
2. Zeros are also significant with two exceptions:
    1. zeros preceding the decimal point.
    2. zeros following the decimal point and preceding the first nonzero digit.
3. Terminal zeros preceding the decimal point in amounts greater than one is an ambiguous case.

Rules for Numbers WITHOUT a Decimal Point

1. START counting for sig. figs. On the FIRST non-zero digit.
2. STOP counting for sig. figs. On the LAST non-zero digit.
3. Non-zero digits are ALWAYS significant
4. Zeroes in between two non-zero digits are significant. All other zeroes are insignificant.

Example 1.8.1:

The first two zeroes in 200500 (four significant digits) are significant because they are between two non-zero digits, and the last two zeroes are insignificant because they are after the last non-zero digit.

It should be noted that both constants and quantities of real world objects have an infinite number of significant figures. For example if you were to count three oranges, a real world object, the value three would be considered to have an infinite number of significant figures in this context.

Example 1.8.1

How many significant digits are in 5010?

Solution

1. Start counting for significant digits On the first non-zero digit (5).
2. Stop counting for significant digits On the last non-zero digit (1).

5 0 1 0 Key: 0 = significant zero. 0 = insignificant zero.

3 significant digits.

Rules for Numbers WITH a Decimal Point

1. START counting for sig. figs. On the FIRST non-zero digit.
2. STOP counting for sig. figs. On the VERY LAST digit (regardless whether or not the last digit is a zero or non-zero number).
3. Non-zero digits are ALWAYS significant.
4. Any zero AFTER the first non-zero digit is STILL significant. The zeroes BEFORE the first non-zero digit are insignificant.

Example 1.8.3

The first two zeroes in 0.058000 (five significant digits) are insignificant because they are before the first non-zero digit, and the last three zeroes are significant because they are after the first non-zero digit.

Example 1.8.4

How many significant digits are in 0.70620?

Solution

1. Start counting for significant digits On the first non-zero digit (7).
2. Stop counting for significant digits On the last digit (0).

0 . 7 0 6 2 0 Key: 0 = significant zero.0 = insignificant zero.

5 significant digits.

Scientific Notation

Scientific notation form: a x 10^b, where “a” and “b” are integers, and "a" has to be between 1 and 10.

Example 1.8.5

The scientific notation for 4548 is 4.548 x 10^3.

Solution

- Disregard the “10^b,” and determine the significant digits in “a.”
- 4.548 x 10^3 has 4 significant digits.

Example 1.8.6

How many significant digits are in 1.52 x 10^6?

NOTE: Only determine the amount of significant digits in the "1.52" part of the scientific notation form.

Answer

3 significant digits.

Rounding Significant Digits

When rounding numbers to a significant digit, keep the amount of significant digits wished to be kept, and replace the other numbers with insignificant zeroes. The reason for rounding a number to a particular amount of significant digits is because in a calculation, some values have less significant digits than other values, and the answer to a calculation is only accurate to the amount of significant digits of the value with the least amount. NOTE: be careful when rounding numbers with a decimal point. Any zeroes added after the first non-zero digit is considered to be a significant zero. TIP: When doing calculations for quizzes/tests/midterms/finals, it would be best to not round in the middle of your calculations, and round to the significant digit only at the end of your calculations.

Example 1.8.7

Round 32445.34 to 2 significant digits.

Answer

32000 (NOT 32000.00, which has 7 significant digits. Due to the decimal point, the zeroes after the first non-zero digit become significant).

Rules for Addition and Subtraction

When adding or subtracting numbers, the end result should have the same amount of decimal places as the number with the least amount of decimal places.

Example 1.8.8

Y = 232.234 + 0.27 Find Y.

Answer

Y = 232.50

NOTE: 232.234 has 3 decimal places and 0.27 has 2 decimal places. The least amount of decimal places is 2. Thus, the answer must be rounded to the 2nd decimal place (thousandth).

Rules for Multiplication and Division

When multiplying or dividing numbers, the end result should have the same amount of significant digits as the number with the least amount of significant digits.

Example 1.8.9

Y = 28 x 47.3 Find Y

Answer

Y = 1300

NOTE: 28 has 2 significant digits and 47.3 has 3 significant digits. The least amount of significant digits is 2. Thus, the answer must me rounded to 2 significant digits (which is done by keeping 2 significant digits and replacing the rest of the digits with insignificant zeroes).

Exact Numbers

Exact numbers can be considered to have an unlimited number of significant figures, as such calculations are not subject to errors in measurement. This may occur:

1. By definition (1 minute = 60 seconds, 1 inch = 2.54 cm, 12 inches = 1 foot, etc.)
2. As a result of counting (6 faces on a cube or dice, two hydrogen atoms in a water molecule, 3 peas in a pod, etc.)

References

1. Brown, Theodore E., H. Eugene LeMay, and Bruce E. Bursten. Chemistry: The Central Science, Tenth Edition. Pearson Education Inc. Upper Saddle River, New Jersey: 2005.
2. Petrucci, Ralph H., William S. Harwood, F. Geoffrey Herring, and Jeffry D. Madura. General Chemistry: Principles and Modern Applications, Ninth Edition. Pearson Education Inc. Upper Saddle River, New Jersey: 2007.
3. Petrucci, Ralph H., William S. Harwood, F. Geoffrey Herring, and Jeffry D. Madura. General Chemistry: Principles and Modern Applications, Tenth Edition. Pearson Education Inc. Upper Saddle River, New Jersey: 2011. Custom Edition for Chem 2, University of California, Davis

Additional Problems

1. a. a) How many significant digits Are in 50?
   b. b) How many significant digits Are in 50.0?
2. How many significant digits Are in 3.670×10^{35}?
3. Round 4279852.243 to 3 significant digits.
4. Round 0.0573000 to 1 significant digit.
5. Y = 45.2 + 16.730 Find Y.
6. Y = 23 – 26.2 Find Y.
7. Y = 16.7 x 33.2 x 16.72 Find Y.
8. Y = 346 ÷ 22 Find Y.
9. Y = (23.2 + 16.723) x 28 Find Y
10. Y = (16.7 x 23) – (23.2 ÷ 2.13) Find Y

Solutions

1. a) 1 significant digit.

b) 2 significant digits.

2. 4 significant digits.

3. 4280000

4. 0.06

5. Y = 61.9

6. Y = -3

7. Y = 9270

8. Y = 16

9. Y = (23.2 + 16.723) x 28

Y = 39.923 x 28 (TIP: Do not round until the end of calculations.)

Y = 1100 (NOTE: 28 has the least amount of significant digits (2 sig. figs.) Thus, answer must be rounded to 2 sig. figs.)

10. Y = (16.7 x 23) – (23.2 ÷ 2.13)

Y = 384.1 – 10.89201878 (TIP: Do not round until the end of calculations.)

Y = 373.2 (NOTE: 384.1 has the least amount of decimal point (tenth). Thus, answer must be rounded to the tenth.)

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CHAPTER OVERVIEW

2: Atoms and Elements

Unit 2 Objectives

By the end of this unit, you will be able to:

- State the postulates of Dalton's atomic theory
- Use postulates of Dalton's atomic theory to explain the laws of definite and multiple proportions
- Outline milestones in the development of modern atomic theory
- Summarize and interpret the results of the experiments of Thomson, Millikan, and Rutherford
- Describe the three subatomic particles that compose atoms
- Define isotopes and give examples for several elements
- Write and interpret symbols that depict the atomic number, mass number, and charge of an atom or ion
- Define the atomic mass unit and average atomic mass
- Calculate average atomic mass and isotopic abundance

2.1: A History of Atomic Theory

2.1.1: Practice Problems- A History of Atomic Theory (Optional)

2.2: The Structure of the Atom and How We Represent It

2.2.1: Practice Problems- The Structure of the Atom and How We Represent It (Optional)

2.3: Calculating Atomic Masses

2.3.1: Practice Problems- Calculating Atomic Masses (Optional)

2.4: Using the Mole in Calculations

2.5: Highlighting Instrumentation- Mass Spectrometry

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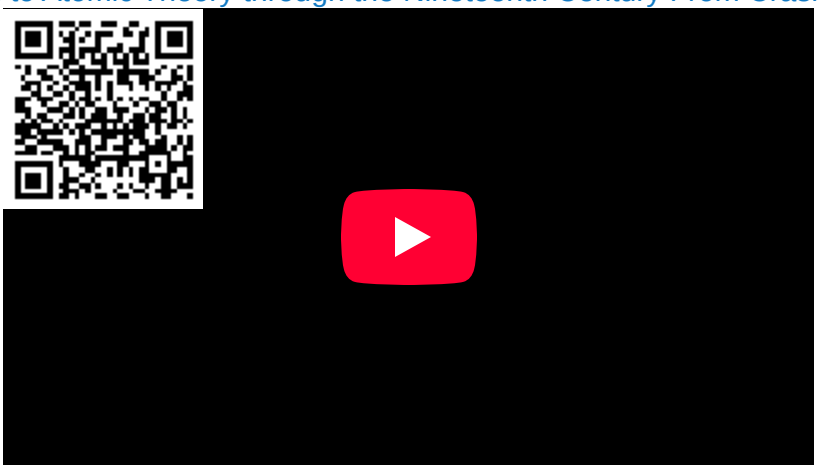
2.1: A History of Atomic Theory

Skills to Develop

By the end of this section, you will be able to:

- State the postulates of Dalton's atomic theory
- Use postulates of Dalton's atomic theory to explain the laws of definite and multiple proportions
- Outline milestones in the development of modern atomic theory
- Summarize and interpret the results of the experiments of Thomson, Millikan, and Rutherford

A Video Introduction to Atomic Theory through the Nineteenth Century From Crash Course Chemistry



***Video* 2.1.1*:* Lavoisier's discovery of The Law of Conservation of Matter led to the Laws of Definite and Multiple Proportions and eventually Dalton's Atomic Theory.**

Atomic Theory through the Nineteenth Century

The earliest recorded discussion of the basic structure of matter comes from ancient Greek philosophers, the scientists of their day. In the fifth century BC, Leucippus and Democritus argued that all matter was composed of small, finite particles that they called *atomos*, a term derived from the Greek word for "indivisible." They thought of atoms as moving particles that differed in shape and size, and which could join together. Later, Aristotle and others came to the conclusion that matter consisted of various combinations of the four "elements"—fire, earth, air, and water—and could be infinitely divided. Interestingly, these philosophers thought about atoms and "elements" as philosophical concepts, but apparently never considered performing experiments to test their ideas.

The Aristotelian view of the composition of matter held sway for over two thousand years, until English schoolteacher John Dalton helped to revolutionize chemistry with his hypothesis that the behavior of matter could be explained using an atomic theory. First published in 1807, many of Dalton's hypotheses about the microscopic features of matter are still valid in modern atomic theory. Here are the postulates of Dalton's atomic theory.

1. Matter is composed of exceedingly small particles called atoms. An atom is the smallest unit of an element that can participate in a chemical change.
2. An element consists of only one type of atom, which has a mass that is characteristic of the element and is the same for all atoms of that element (Figure **2.1.1**). A macroscopic sample of an element contains an incredibly large number of atoms, all of which have identical chemical properties.
3. Atoms of one element differ in properties from atoms of all other elements.
4. A compound consists of atoms of two or more elements combined in a small, whole-number ratio. In a given compound, the numbers of atoms of each of its elements are always present in the same ratio (Figure **2.1.2**).
5. Atoms are neither created nor destroyed during a chemical change, but are instead rearranged to yield substances that are different from those present before the change (Figure **2.1.3**).

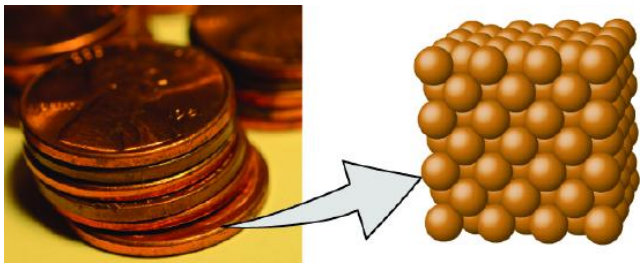


Figure* 2.1.1*: *A pre-1982 copper penny (left) contains approximately 3×10^{22} copper atoms (several dozen are represented as brown spheres at the right), each of which has the same chemical properties. (credit: modification of work by "slgckgc"/Flickr)*

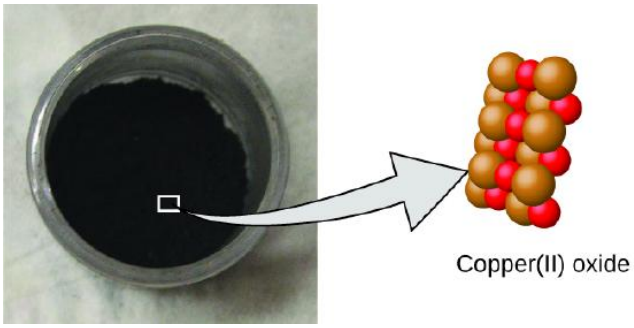


Figure* 2.1.2*: *Copper(II) oxide, a powdery, black compound, results from the combination of two types of atoms—copper (brown spheres) and oxygen (red spheres)—in a 1:1 ratio. (credit: modification of work by "Chemicalinterest"/Wikimedia Commons)*

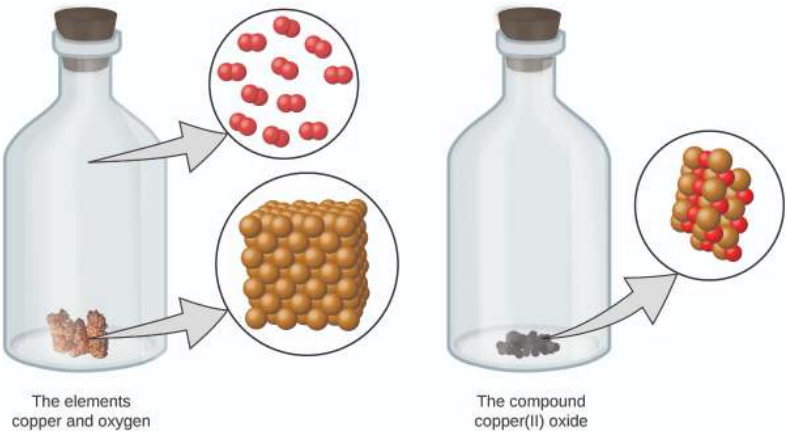


Figure* 2.1.3: When the elements copper (a shiny, red-brown solid, shown here as brown spheres) and oxygen (a clear and colorless gas, shown here as red spheres) react, their atoms rearrange to form a compound containing copper and oxygen (a powdery, black solid). (credit copper: modification of work by http://images-of-elements.com/copper.php).*

Dalton's atomic theory provides a microscopic explanation of the many macroscopic properties of matter that you've learned about. For example, if an element such as copper consists of only one kind of atom, then it cannot be broken down into simpler substances, that is, into substances composed of fewer types of atoms. And if atoms are neither created nor destroyed during a chemical change, then the total mass of matter present when matter changes from one type to another will remain constant (the law of conservation of matter (or mass)).

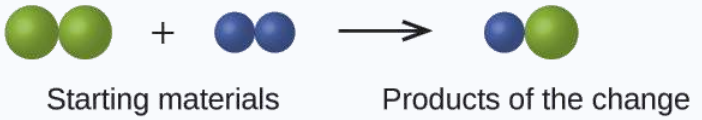
Want to learn more about the Law of Conservation of Mass?



Video 2.1.2: "We are made of star stuff" - Carl Sagan.

Example 2.1.1: Testing Dalton's Atomic Theory

In the following drawing, the green spheres represent atoms of a certain element. The purple spheres represent atoms of another element. If the spheres touch, they are part of a single unit of a compound. Does the following chemical change represented by these symbols violate any of the ideas of Dalton's atomic theory? If so, which one?

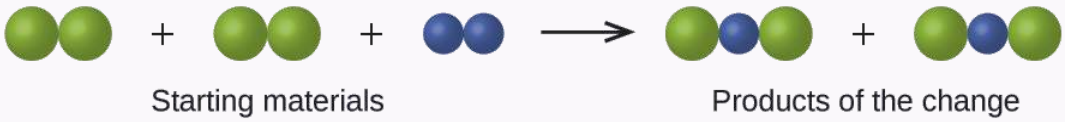


Solution

The starting materials consist of two green spheres and two purple spheres. The products consist of only one green sphere and one purple sphere. This violates Dalton's postulate that atoms are neither created nor destroyed during a chemical change, but are merely redistributed. (In this case, atoms appear to have been destroyed.)

Exercise 2.1.1

In the following drawing, the green spheres represent atoms of a certain element. The purple spheres represent atoms of another element. If the spheres touch, they are part of a single unit of a compound. Does the following chemical change represented by these symbols violate any of the ideas of Dalton's atomic theory? If so, which one



Answer

The starting materials consist of four green spheres and two purple spheres. The products consist of four green spheres and two purple spheres. This does not violate any of Dalton's postulates: Atoms are neither created nor destroyed, but are redistributed in small, whole-number ratios.

Dalton knew of the experiments of French chemist Joseph Proust, who demonstrated that *all samples of a pure compound contain the same elements in the same proportion by mass*. This statement is known as the law of definite proportions or the law of constant composition. The suggestion that the numbers of atoms of the elements in a given compound always exist in the same ratio is consistent with these observations. For example, when different samples of isooctane (a component of gasoline and one of the standards used in the octane rating system) are analyzed, they are found to have a carbon-to-hydrogen mass ratio of 5.33:1, as shown in Table **2.1.1**.

***Table* 2.1.1**: *Constant Composition of Isooctane*

Sample	Carbon	Hydrogen	Mass Ratio
A	14.82 g	2.78 g	$\frac{14.82\text{ g carbon}}{2.78\text{ g hydrogen}} = \frac{5.33\text{ g carbon}}{1.00\text{ g hydrogen}}$
B	22.33 g	4.19 g	$\frac{22.33\text{ g carbon}}{4.19\text{ g hydrogen}} = \frac{5.33\text{ g carbon}}{1.00\text{ g hydrogen}}$
C	19.40 g	3.64 g	$\frac{19.40\text{ g carbon}}{3.63\text{ g hydrogen}} = \frac{5.33\text{ g carbon}}{1.00\text{ g hydrogen}}$

It is worth noting that although all samples of a particular compound have the same mass ratio, the converse is not true in general. That is, samples that have the same mass ratio are not necessarily the same substance. For example, there are many compounds other than isooctane that also have a carbon-to-hydrogen mass ratio of 5.33:1.00.

Dalton also used data from Proust, as well as results from his own experiments, to formulate another interesting law. The law of multiple proportions states that *when two elements react to form more than one compound, a fixed mass of one element will react with masses of the other element in a ratio of small, whole numbers*. For example, copper and chlorine can form a green, crystalline solid with a mass ratio of 0.558 g chlorine to 1 g copper, as well as a brown crystalline solid with a mass ratio of 1.116 g chlorine to 1 g copper. These ratios by themselves may not seem particularly interesting or informative; however, if we take a ratio of these ratios, we obtain a useful and possibly surprising result: a small, whole-number ratio.

$$\frac{\frac{1.116\text{ g Cl}}{1\text{ g Cu}}}{\frac{0.558\text{ g Cl}}{1\text{ g Cu}}} = \frac{2}{1} \tag{2.1.1}$$

This 2-to-1 ratio means that the brown compound has twice the amount of chlorine per amount of copper as the green compound.

This can be explained by atomic theory if the copper-to-chlorine ratio in the brown compound is 1 copper atom to 2 chlorine atoms, and the ratio in the green compound is 1 copper atom to 1 chlorine atom. The ratio of chlorine atoms (and thus the ratio of their masses) is therefore 2 to 1 (Figure **2.1.4**).

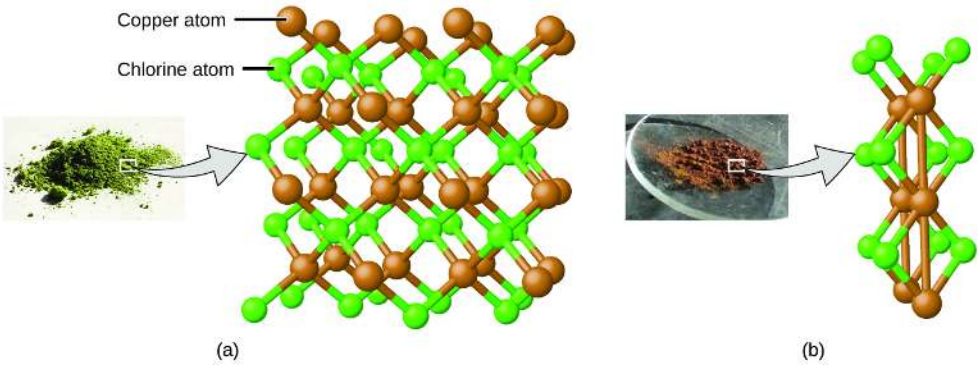


***Figure* 2.1.4**: *Compared to the copper chlorine compound in (a), where copper is represented by brown spheres and chlorine by green spheres, the copper chlorine compound in (b) has twice as many chlorine atoms per copper atom. (credit a: modification of work by "Benjah-bmm27"/Wikimedia Commons; credit b: modification of work by "Walkerma"/Wikimedia Commons)*

Example 2.1.2: Laws of Definite and Multiple Proportions

A sample of compound A (a clear, colorless gas) is analyzed and found to contain 4.27 g carbon and 5.69 g oxygen. A sample of compound B (also a clear, colorless gas) is analyzed and found to contain 5.19 g carbon and 13.84 g oxygen. Are these data an example of the law of definite proportions, the law of multiple proportions, or neither? What do these data tell you about substances A and B?

Solution

In compound A, the mass ratio of carbon to oxygen is:

$$\frac{1.33\ \text{g O}}{1\ \text{g C}}$$

In compound B, the mass ratio of carbon to oxygen is:

$$\frac{2.67\ \text{g O}}{1\ \text{g C}}$$

The ratio of these ratios is:

$$\frac{\frac{1.33\ \text{g O}}{1\ \text{g C}}}{\frac{2.67\ \text{g O}}{1\ \text{g C}}} = \frac{1}{2}$$

This supports the law of multiple proportions. This means that A and B are different compounds, with A having one-half as much carbon per amount of oxygen (or twice as much oxygen per amount of carbon) as B. A possible pair of compounds that would fit this relationship would be A = CO_2 and B = CO.

Exercise 2.1.2

A sample of compound X (a clear, colorless, combustible liquid with a noticeable odor) is analyzed and found to contain 14.13 g carbon and 2.96 g hydrogen. A sample of compound Y (a clear, colorless, combustible liquid with a noticeable odor that is slightly different from X's odor) is analyzed and found to contain 19.91 g carbon and 3.34 g hydrogen. Are these data an example of the law of definite proportions, the law of multiple proportions, or neither? What do these data tell you about substances X and Y?

Answer

In compound X, the mass ratio of carbon to hydrogen is $\frac{14.13\ \text{g C}}{2.96\ \text{g H}}$.

In compound Y, the mass ratio of carbon to oxygen is $\frac{19.91\ \text{g C}}{3.34\ \text{g H}}$.

The ratio of these ratios is

$$\frac{\frac{14.13\ \text{g C}}{2.96\ \text{g H}}}{\frac{19.91\ \text{g C}}{3.34\ \text{g H}}} = \frac{4.77\ \text{g C/g H}}{5.96\ \text{g C/g H}} = 0.800 = \frac{4}{5}.$$

This small, whole-number ratio supports the law of multiple proportions. This means that X and Y are different compounds.

In the two centuries since Dalton developed his ideas, scientists have made significant progress in furthering our understanding of atomic theory. Much of this came from the results of several seminal experiments that revealed the details of the internal structure of atoms. Here, we will discuss some of those key developments, with an emphasis on application of the scientific method, as well as understanding how the experimental evidence was analyzed. While the historical persons and dates behind these experiments can be quite interesting, it is most important to understand the concepts resulting from their work.

Atomic Theory after the Nineteenth Century

If matter were composed of atoms, what were atoms composed of? Were they the smallest particles, or was there something smaller? In the late 1800s, a number of scientists interested in questions like these investigated the electrical discharges that could be produced in low-pressure gases, with the most significant discovery made by English physicist J. J. Thomson using a cathode ray tube. This apparatus consisted of a sealed glass tube from which almost all the air had been removed; the tube contained two metal electrodes. When high voltage was applied across the electrodes, a visible beam called a cathode ray appeared between them. This beam was deflected toward the positive charge and away from the negative charge, and was produced in the same way with identical properties when different metals were used for the electrodes. In similar experiments, the ray was simultaneously deflected by an applied magnetic field, and measurements of the extent of deflection and the magnetic field strength allowed

Thomson to calculate the charge-to-mass ratio of the cathode ray particles. The results of these measurements indicated that these particles were much lighter than atoms (Figure **2.1.1**).

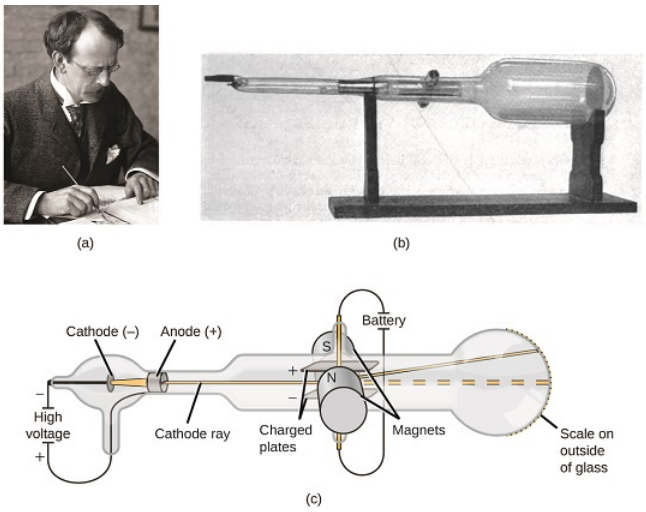
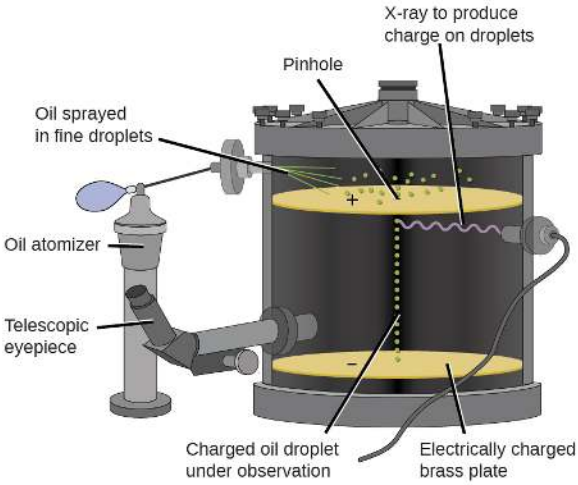


Figure* 2.1.5: (a) J. J. Thomson produced a visible beam in a cathode ray tube. (b) This is an early cathode ray tube, invented in 1897 by Ferdinand Braun. (c) In the cathode ray, the beam (shown in yellow) comes from the cathode and is accelerated past the anode toward a fluorescent scale at the end of the tube. Simultaneous deflections by applied electric and magnetic fields permitted Thomson to calculate the mass-to-charge ratio of the particles composing the cathode ray. (credit a: modification of work by Nobel Foundation; credit b: modification of work by Eugen Nesper; credit c: modification of work by "Kurzon"/Wikimedia Commons).*

Based on his observations, here is what Thomson proposed and why: The particles are attracted by positive (+) charges and repelled by negative (−) charges, so they must be negatively charged (like charges repel and unlike charges attract); they are less massive than atoms and indistinguishable, regardless of the source material, so they must be fundamental, subatomic constituents of all atoms. Although controversial at the time, Thomson's idea was gradually accepted, and his cathode ray particle is what we now call an electron, a negatively charged, subatomic particle with a mass more than one thousand-times less that of an atom. The term "electron" was coined in 1891 by Irish physicist George Stoney, from "*electr*ic i*on*."

In 1909, more information about the electron was uncovered by American physicist Robert A. Millikan via his "oil drop" experiments. Millikan created microscopic oil droplets, which could be electrically charged by friction as they formed or by using X-rays. These droplets initially fell due to gravity, but their downward progress could be slowed or even reversed by an electric field lower in the apparatus. By adjusting the electric field strength and making careful measurements and appropriate calculations, Millikan was able to determine the charge on individual drops (Figure **2.1.2**).



Oil drop	Charge in coulombs (C)
A	4.8×10^{-19} C
B	3.2×10^{-19} C
C	6.4×10^{-19} C
D	1.6×10^{-19} C
E	4.8×10^{-19} C

Figure* 2.1.6*: *Millikan's experiment measured the charge of individual oil drops. The tabulated data are examples of a few possible values.*

Looking at the charge data that Millikan gathered, you may have recognized that the charge of an oil droplet is always a multiple of a specific charge, 1.6×10^{-19} C. Millikan concluded that this value must therefore be a fundamental charge—the charge of a single electron—with his measured charges due to an excess of one electron (1 times 1.6×10^{-19} C), two electrons (2 times 1.6×10^{-19} C), three electrons (3 times 1.6×10^{-19} C), and so on, on a given oil droplet. Since the charge of an electron was now known due to Millikan's research, and the charge-to-mass ratio was already known due to Thomson's research (1.759×10^{11} C/kg), it only required a simple calculation to determine the mass of the electron as well.

$$\text{Mass of electron} = 1.602 \times 10^{-19} \cancel{\text{C}} \times \frac{1 \text{ kg}}{1.759 \times 10^{11} \cancel{\text{C}}} = 9.107 \times 10^{-31} \text{ kg} \tag{2.3.1}$$

Scientists had now established that the atom was not indivisible as Dalton had believed, and due to the work of Thomson, Millikan, and others, the charge and mass of the negative, subatomic particles—the electrons—were known. However, the positively charged part of an atom was not yet well understood. In 1904, Thomson proposed the "plum pudding" model of atoms, which described a positively charged mass with an equal amount of negative charge in the form of electrons embedded in it, since all atoms are electrically neutral. A competing model had been proposed in 1903 by Hantaro Nagaoka, who postulated a Saturn-like atom, consisting of a positively charged sphere surrounded by a halo of electrons (Figure **2.1.3**).

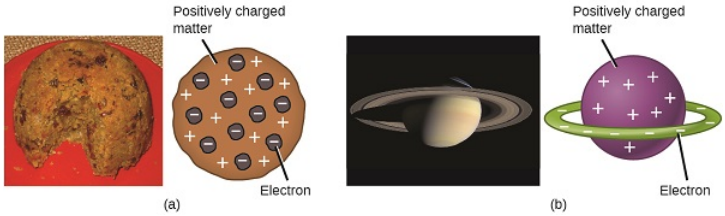


Figure* 2.1.7*: *(a) Thomson suggested that atoms resembled plum pudding, an English dessert consisting of moist cake with embedded raisins ("plums"). (b) Nagaoka proposed that atoms resembled the planet Saturn, with a ring of electrons surrounding a positive "planet." (credit a: modification of work by "Man vyi"/Wikimedia Commons; credit b: modification of work by "NASA"/Wikimedia Commons).*

The next major development in understanding the atom came from Ernest Rutherford, a physicist from New Zealand who largely spent his scientific career in Canada and England. He performed a series of experiments using a beam of high-speed, positively charged alpha particles (α particles) that were produced by the radioactive decay of radium; α particles consist of two protons and two neutrons (you will learn more about radioactive decay in the chapter on nuclear chemistry). Rutherford and his colleagues Hans Geiger (later famous for the Geiger counter) and Ernest Marsden aimed a beam of α particles, the source of which was embedded in a lead block to absorb most of the radiation, at a very thin piece of gold foil and examined the resultant scattering of the α particles using a luminescent screen that glowed briefly where hit by an α particle.

What did they discover? Most particles passed right through the foil without being deflected at all. However, some were diverted slightly, and a very small number were deflected almost straight back toward the source (Figure **2.1.4**). Rutherford described finding these results: "It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you"[1] (p. 68).

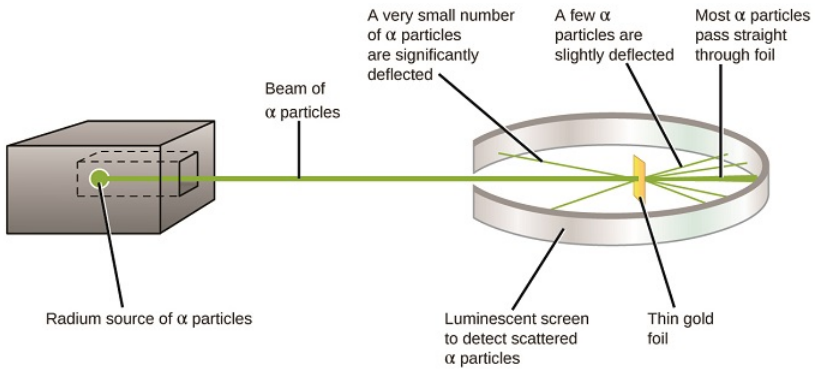


Figure* 2.1.8*: *Geiger and Rutherford fired α particles at a piece of gold foil and detected where those particles went, as shown in this schematic diagram of their experiment. Most of the particles passed straight through the foil, but a few were deflected slightly and a very small number were significantly deflected.*

Here is what Rutherford deduced: Because most of the fast-moving α particles passed through the gold atoms undeflected, they must have traveled through essentially empty space inside the atom. Alpha particles are positively charged, so deflections arose when they encountered another positive charge (like charges repel each other). Since like charges repel one another, the few positively charged α particles that changed paths abruptly must have hit, or closely approached, another body that also had a highly concentrated, positive charge. Since the deflections occurred a small fraction of the time, this charge only occupied a small amount of the space in the gold foil. Analyzing a series of such experiments in detail, Rutherford drew two conclusions:

1. The volume occupied by an atom must consist of a large amount of empty space.
2. A small, relatively heavy, positively charged body, the nucleus, must be at the center of each atom.

This analysis led Rutherford to propose a model in which an atom consists of a very small, positively charged nucleus, in which most of the mass of the atom is concentrated, surrounded by the negatively charged electrons, so that the atom is electrically neutral (Figure 2.1.5).

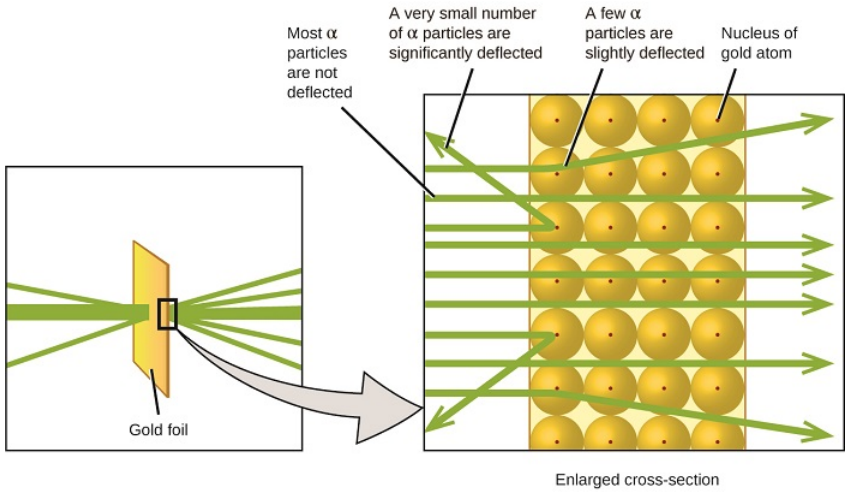
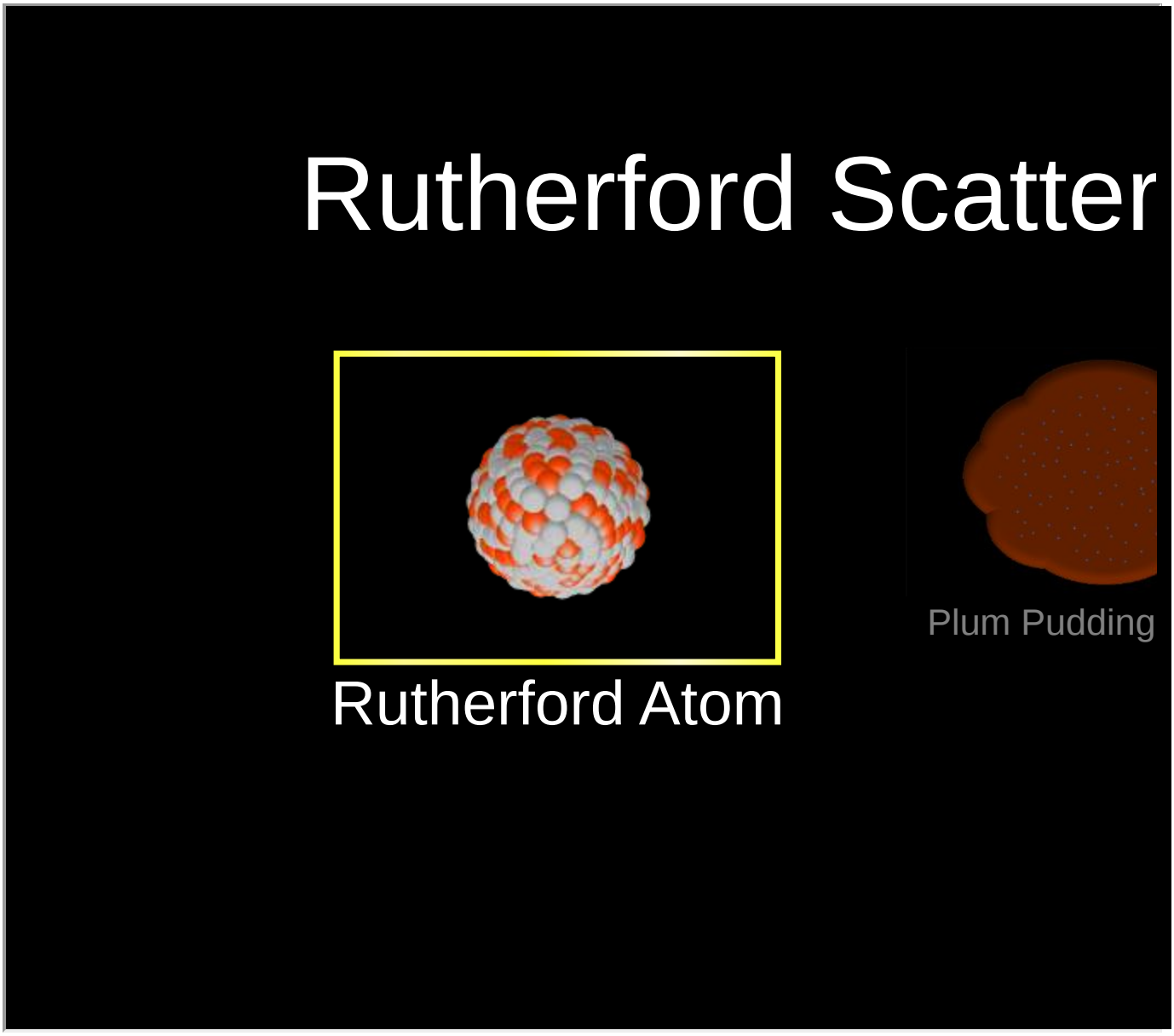


Figure* 2.1.9*: *The α particles are deflected only when they collide with or pass close to the much heavier, positively charged gold nucleus. Because the nucleus is very small compared to the size of an atom, very few α particles are deflected. Most pass through the relatively large region occupied by electrons, which are too light to deflect the rapidly moving particles.*

After many more experiments, Rutherford also discovered that the nuclei of other elements contain the hydrogen nucleus as a "building block," and he named this more fundamental particle the proton, the positively charged, subatomic particle found in the nucleus. With one addition, which you will learn next, this nuclear model of the atom, proposed over a century ago, is still used today.



Another important finding was the discovery of isotopes. During the early 1900s, scientists identified several substances that appeared to be new elements, isolating them from radioactive ores. For example, a “new element” produced by the radioactive decay of thorium was initially given the name mesothorium. However, a more detailed analysis showed that mesothorium was chemically identical to radium (another decay product), despite having a different atomic mass. This result, along with similar findings for other elements, led the English chemist Frederick Soddy to realize that an element could have types of atoms with different masses that were chemically indistinguishable. These different types are called isotopes—atoms of the same element that differ in mass. Soddy was awarded the Nobel Prize in Chemistry in 1921 for this discovery.

One puzzle remained: The nucleus was known to contain almost all of the mass of an atom, with the number of protons only providing half, or less, of that mass. Different proposals were made to explain what constituted the remaining mass, including the existence of neutral particles in the nucleus. As you might expect, detecting uncharged particles is very challenging, and it was not until 1932 that James Chadwick found evidence of neutrons, uncharged, subatomic particles with a mass approximately the same as that of protons. The existence of the neutron also explained isotopes: They differ in mass because they have different numbers of neutrons, but they are chemically identical because they have the same number of protons. This will be explained in more detail later in this unit.



Video 2.1.2: An Introduction to Subatomic Particles

Summary



Video 2.1.3: *A summary of discoveries in atomic theory.*



Video 2.1.4: *A different summary of discoveries in atomic theory.*

The ancient Greeks proposed that matter consists of extremely small particles called atoms. Dalton postulated that each element has a characteristic type of atom that differs in properties from atoms of all other elements, and that atoms of different elements can combine in fixed, small, whole-number ratios to form compounds. Samples of a particular compound all have the same elemental proportions by mass. When two elements form different compounds, a given mass of one element will combine with masses of the other element in a small, whole-number ratio. During any chemical change, atoms are neither created nor destroyed.

Although no one has actually seen the inside of an atom, experiments have demonstrated much about atomic structure. Thomson's cathode ray tube showed that atoms contain small, negatively charged particles called electrons. Millikan discovered that there is a fundamental electric charge—the charge of an electron. Rutherford's gold foil experiment showed that atoms have a small, dense, positively charged nucleus; the positively charged particles within the nucleus are called protons. Chadwick discovered that the nucleus also contains neutral particles called neutrons. Soddy demonstrated that atoms of the same element can differ in mass; these are called isotopes.

Footnotes

1. Ernest Rutherford, "The Development of the Theory of Atomic Structure," ed. J. A. Ratcliffe, in *Background to Modern Science*, eds. Joseph Needham and Walter Pagel, (Cambridge, UK: Cambridge University Press, 1938), 61–74. Accessed September 22, 2014, https://ia600508.us.archive.org/3/it...e032734mbp.pdf.

Glossary

Dalton's atomic theory
set of postulates that established the fundamental properties of atoms

law of constant composition
(also, law of definite proportions) all samples of a pure compound contain the same elements in the same proportions by mass

law of multiple proportions
when two elements react to form more than one compound, a fixed mass of one element will react with masses of the other element in a ratio of small whole numbers

law of definite proportions
(also, law of constant composition) all samples of a pure compound contain the same elements in the same proportions by mass

alpha particle (α particle)
positively charged particle consisting of two protons and two neutrons

electron
negatively charged, subatomic particle of relatively low mass located outside the nucleus

isotopes
atoms that contain the same number of protons but different numbers of neutrons

neutron
uncharged, subatomic particle located in the nucleus

proton
positively charged, subatomic particle located in the nucleus

nucleus
massive, positively charged center of an atom made up of protons and neutrons

Contributors

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- Adelaide Clark, Oregon Institute of Technology
- Crash Course Chemistry: Crash Course is a division of Complexly and videos are free to stream for educational purposes.
- TED-Ed's commitment to creating lessons worth sharing is an extension of TED's mission of spreading great ideas. Within TED-Ed's growing library of TED-Ed animations, you will find carefully curated educational videos, many of which represent collaborations between talented educators and animators nominated through the TED-Ed website.

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2.1.1: Practice Problems- A History of Atomic Theory (Optional)

PROBLEM 2.1.1.1

In the following drawing, the green spheres represent atoms of a certain element. The purple spheres represent atoms of another element. If the spheres of different elements touch, they are part of a single unit of a compound. The following chemical change represented by these spheres may violate one of the ideas of Dalton's atomic theory. Which one?



Answer

The starting materials consist of one green sphere and two purple spheres. The products consist of two green spheres and two purple spheres. This violates Dalton’s postulate that that atoms are not created during a chemical change, but are merely redistributed.

PROBLEM 2.1.1.2

Which postulate of Dalton's theory is consistent with the following observation concerning the weights of reactants and products?

When 100 grams of solid calcium carbonate is heated, 44 g of CO_2 and 56 g of CaO are produced.

Answer

Atoms are neither created nor destroyed during a chemical change, but are instead rearranged to yield substances that are different from those present before the change (Based on the Law of Conservation of Mass).

PROBLEM 2.1.1.3

Samples of compound X, Y, and Z are analyzed, with results shown here. Do these data provide example(s) of the law of definite proportions, the law of multiple proportions, neither, or both? What do these data tell you about compounds X, Y, and Z?

Compound	Description	Mass of Carbon	Mass of Hydrogen
X	clear, colorless, liquid with strong odor	1.776 g	0.148 g
Y	clear, colorless, liquid with strong odor	1.974 g	0.329 g
Z	clear, colorless, liquid with strong odor	7.812 g	0.651 g

Answer

X+Z are similar compounds (same ratios of C and H), aligning with the Law of Definite Proportions

X+Y and Y+Z are different compounds (differing ratios of C and H), aligning with the Law of Multiple Proportions

Click here to see a video of the solution.



PROBLEM 2.1.1.4

How are electrons and protons similar? How are they different?

Answer

Electrons and protons are both ***charged*** subatomic particles.

Protons are much larger than electrons (contributing more mass to the overall atom).

Changing the number of protons changes the *identity* of the atom, which changing the number of electrons changes the *charge*.

PROBLEM 2.1.1.5

How are protons and neutrons similar? How are they different?

Answer

Protons and neutrons are both located in the **nucleus** of the atom.

Protons and neutrons both contribute to the overall *mass* of the atom.

Protons carry a *charge* while neutrons are neutral.

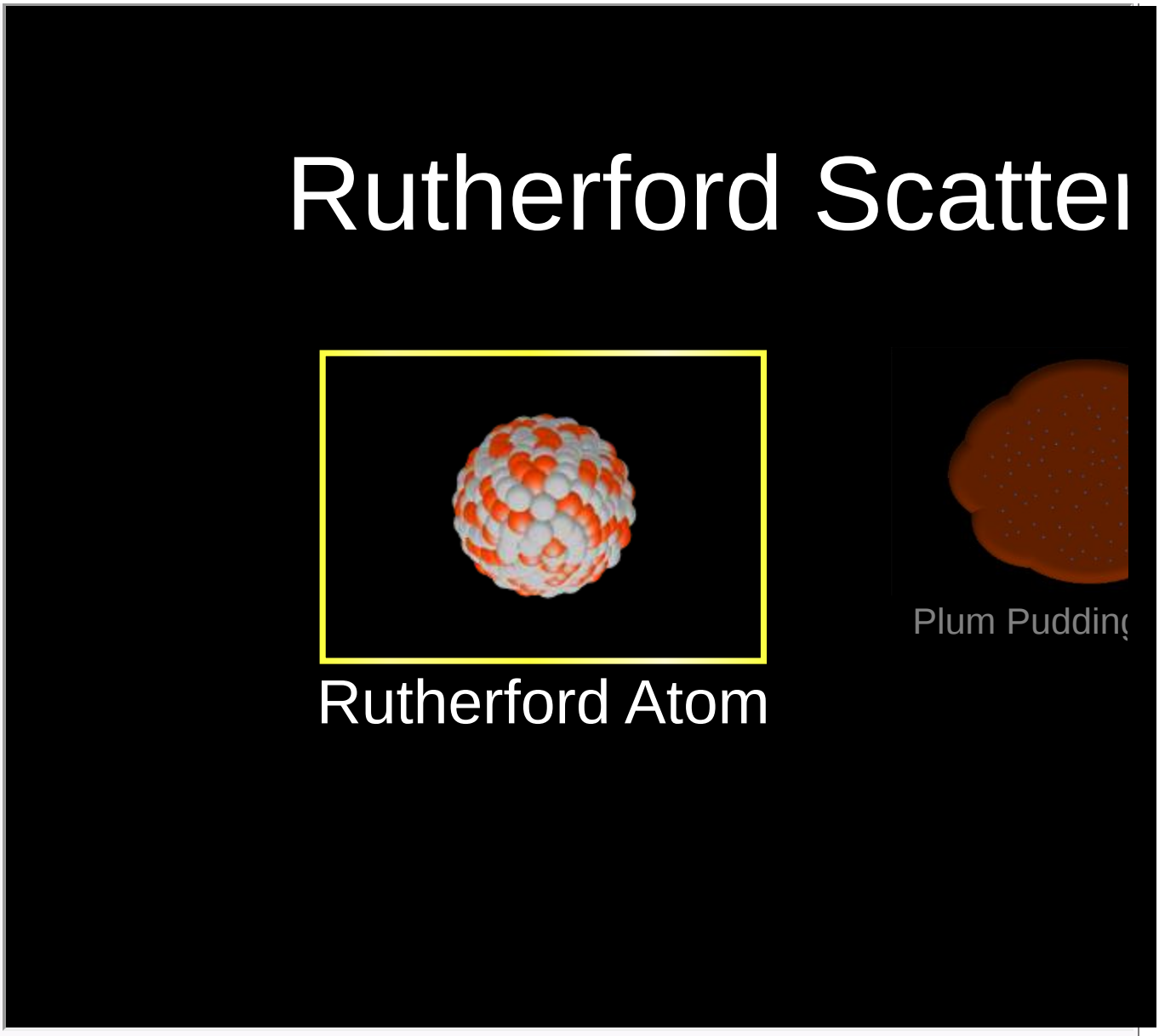
PROBLEM 2.1.1.6

Predict and test the behavior of α particles fired at a “plum pudding” model atom.

(a) Predict the paths taken by α particles that are fired at atoms with a Thomson’s plum pudding model structure. Explain why you expect the α particles to take these paths.

(b) If α particles of higher energy than those in (a) are fired at plum pudding atoms, predict how their paths will differ from the lower-energy α particle paths. Explain your reasoning.

(c) Now test your predictions from (a) and (b).



Select the “Plum Pudding Atom” tab above. Set “Alpha Particles Energy” to “min,” and select “show traces.” Click on the gun to start firing α particles. Does this match your prediction from (a)? If not, explain why the actual path would be that shown in the simulation. Hit the pause button, or “Reset All.” Set “Alpha Particles Energy” to “max,” and start firing α particles. Does this match your prediction from (b)? If not, explain the effect of increased energy on the actual paths as shown in the simulation.

Answer a

The plum pudding model indicates that the positive charge is spread uniformly throughout the atom, so we expect the α particles to (perhaps) be slowed somewhat by the positive-positive repulsion, but to follow straight-line paths (i.e., not to be deflected) as they pass through the atoms.

Answer b

Higher-energy α particles will be traveling faster (and perhaps slowed less) and will also follow straight-line paths through the atoms.

Answer c

The α particles followed straight-line paths through the plum pudding atom. There was no apparent slowing of the α particles as they passed through the atoms.

PROBLEM 2.1.1.7

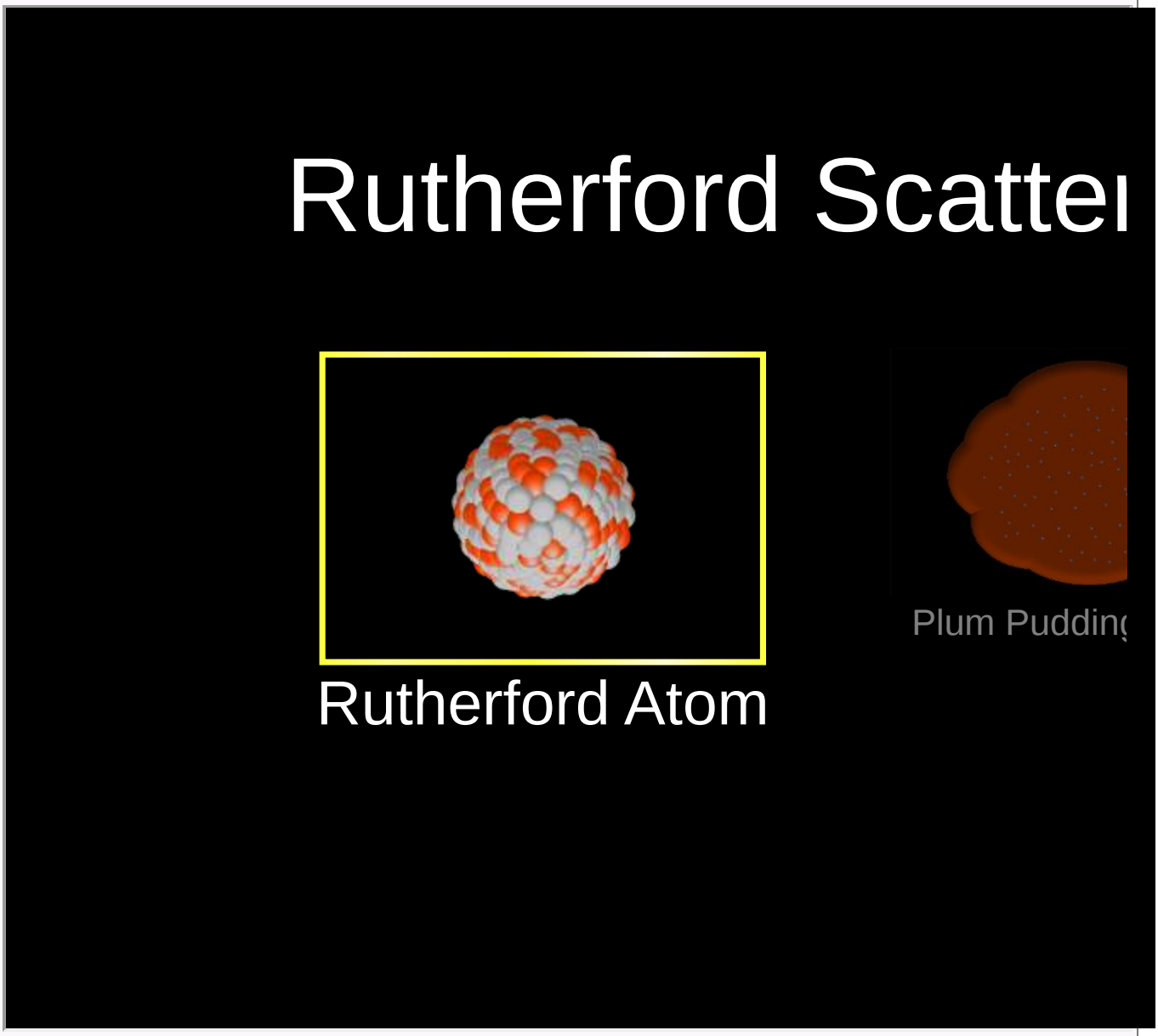
Predict and test the behavior of α particles fired at a Rutherford atom model.

(a) Predict the paths taken by α particles that are fired at atoms with a Rutherford atom model structure. Explain why you expect the α particles to take these paths.

(b) If α particles of higher energy than those in (a) are fired at Rutherford atoms, predict how their paths will differ from the lower-energy α particle paths. Explain your reasoning.

(c) Predict how the paths taken by the α particles will differ if they are fired at Rutherford atoms of elements other than gold. What factor do you expect to cause this difference in paths, and why?

(d) Now test your predictions from (a), (b), and (c).



Select the “Rutherford Atom” tab above. Due to the scale of the simulation, it is best to start with a small nucleus, so select “20” for both protons and neutrons, “min” for energy, show traces, and then start firing α particles. Does this match your prediction from (a)? If not, explain why the actual path would be that shown in the simulation. Pause or reset, set energy to “max,” and start firing α particles. Does this match your prediction from (b)? If not, explain the effect of increased energy on the actual path as shown in the simulation. Pause or reset, select “40” for both protons and neutrons, “min” for energy, show traces, and fire away. Does this match your prediction from (c)? If not, explain why the actual path would be that shown in the simulation. Repeat this with larger numbers of protons and neutrons. What generalization can you make regarding the type of atom and effect on the path of α particles? Be clear and specific.

Answer a

The Rutherford atom has a small, positively charged nucleus, so most α particles will pass through empty space far from the nucleus and be undeflected. Those α particles that pass near the nucleus will be deflected from their paths due to positive-positive repulsion. The more directly toward the nucleus the α particles are headed, the larger the deflection angle will be.

Answer b

Higher-energy α particles that pass near the nucleus will still undergo deflection, but the faster they travel, the less the expected angle of deflection

Answer c

If the nucleus is smaller, the positive charge is smaller and the expected deflections are smaller—both in terms of how closely the α particles pass by the nucleus undeflected and the angle of deflection. If the nucleus is larger, the positive charge is larger and the expected deflections are larger—more α particles will be deflected, and the deflection angles will be larger.

Answer d

The paths followed by the α particles match the predictions from (a), (b), and (c)

Contributors

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- Adelaide Clark, Oregon Institute of Technology

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2.2: The Structure of the Atom and How We Represent It

Skills to Develop

- Describe the three subatomic particles that compose atoms
- Define isotopes and give examples for several elements
- Write and interpret symbols that depict the atomic number, mass number, and charge of an atom or ion

A Primer on Subatomic Particles and Atoms



Video* 2.2.1: A Primer on What an Atom is from Fuse School Chemistry.*

The development of modern atomic theory revealed much about the inner structure of atoms. It was learned that an atom contains a very small nucleus composed of positively charged protons and uncharged neutrons, surrounded by a much larger volume of space containing negatively charged electrons. The nucleus contains the majority of an atom's mass because protons and neutrons are much heavier than electrons, whereas electrons occupy almost all of an atom's volume. The diameter of an atom is on the order of 10^{-10} m, whereas the diameter of the nucleus is roughly 10^{-15} m—about 100,000 times smaller. For a perspective about their relative sizes, consider this: If the nucleus were the size of a blueberry, the atom would be about the size of a football stadium (Figure **2.2.1**).



Figure* 2.2.1: If an atom could be expanded to the size of a football stadium, the nucleus would be the size of a single blueberry. (credit middle: modification of work by "babyknight"/Wikimedia Commons; credit right: modification of work by Paxson Woelber).*

Atoms—and the protons, neutrons, and electrons that compose them—are extremely small. For example, a carbon atom weighs less than 2×10^{-23} g, and an electron has a charge of less than 2×10^{-19} C (coulomb). When describing the properties of tiny objects such as atoms, we use appropriately small units of measure, such as the atomic mass unit (amu) and the fundamental unit of charge (e). The amu was originally defined based on hydrogen, the lightest element, then later in terms of oxygen. Since 1961, it has been defined with regard to the most abundant isotope of carbon, atoms of which are assigned masses of exactly 12 amu. (This isotope is known as "carbon-12" as will be discussed later in this module.) Thus, one amu is exactly $\mathbf{1/12}$ of the mass of one

carbon-12 atom: 1 amu = 1.6605 × 10^{-24} g. (The Dalton (Da) and the unified atomic mass unit (u) are alternative units that are equivalent to the amu.) The fundamental unit of charge (also called the elementary charge) equals the magnitude of the charge of an electron (e) with e = 1.602 × 10^{-19} C.

A proton has a mass of 1.0073 amu and a charge of 1+. A neutron is a slightly heavier particle with a mass 1.0087 amu and a charge of zero; as its name suggests, it is neutral. The electron has a charge of 1− and is a much lighter particle with a mass of about 0.00055 amu (it would take about 1800 electrons to equal the mass of one proton. The properties of these fundamental particles are summarized in Table **2.2.1**. (An observant student might notice that the sum of an atom's subatomic particles does not equal the atom's actual mass: The total mass of six protons, six neutrons, and six electrons is 12.0993 amu, slightly larger than the 12.00 amu of an actual carbon-12 atom. This "missing" mass is known as the mass defect, and you will learn about it in the chapter on nuclear chemistry.)

Table* 2.2.1: Properties of Subatomic Particles*

Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
electron	outside nucleus	-1.602×10^{-19}	1−	0.00055	0.00091×10^{-24}
proton	nucleus	1.602×10^{-19}	1+	1.00727	1.67262×10^{-24}
neutron	nucleus	0	0	1.00866	1.67493×10^{-24}

The number of protons in the nucleus of an atom is its atomic number (Z). This is the defining trait of an element: Its value determines the identity of the atom. For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have. A neutral atom must contain the same number of positive and negative charges, so the number of protons equals the number of electrons. Therefore, the atomic number also indicates the number of electrons in an atom. The total number of protons and neutrons in an atom is called its mass number (A). The number of neutrons is therefore the difference between the mass number and the atomic number: A – Z = number of neutrons.

$$\begin{aligned} \text{atomic number (Z)} &= \text{number of protons} \\ \text{mass number (A)} &= \text{number of protons} + \text{number of neutrons} \\ \text{A} - \text{Z} &= \text{number of neutrons} \end{aligned}$$

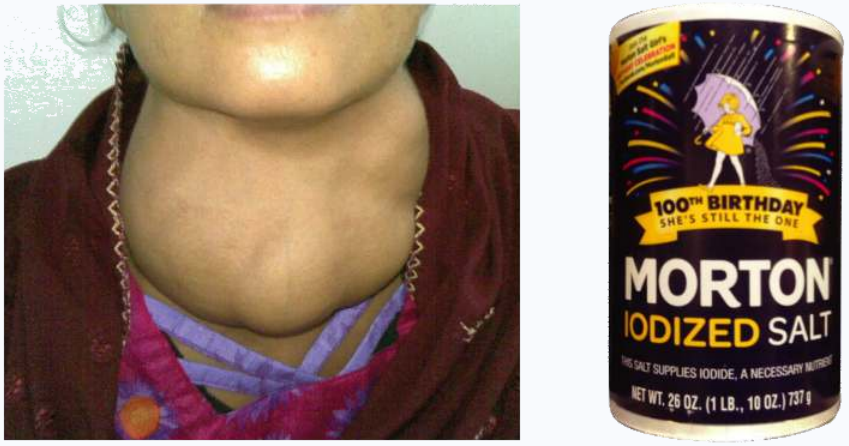
Atoms are electrically neutral if they contain the same number of positively charged protons and negatively charged electrons. When the numbers of these subatomic particles are *not* equal, the atom is electrically charged and is called an ion. The charge of an atom is defined as follows:

$$\text{Atomic charge} = \text{number of protons} - \text{number of electrons}$$

As will be discussed in more detail later in this chapter, atoms (and molecules) typically acquire charge by gaining or losing electrons. An atom that gains one or more electrons will exhibit a negative charge and is called an anion. Positively charged atoms called cations are formed when an atom loses one or more electrons. For example, a neutral sodium atom (Z = 11) has 11 electrons. If this atom loses one electron, it will become a cation with a 1+ charge (11 − 10 = 1+). A neutral oxygen atom (Z = 8) has eight electrons, and if it gains two electrons it will become an anion with a 2− charge (8 − 10 = 2−).

Example 2.2.1: Composition of an Atom

Iodine is an essential trace element in our diet; it is needed to produce thyroid hormone. Insufficient iodine in the diet can lead to the development of a goiter, an enlargement of the thyroid gland (Figure **2.2.2**).



(a) (b)

Figure 2.2.2: (a) Insufficient iodine in the diet can cause an enlargement of the thyroid gland called a goiter. (b) The addition of small amounts of iodine to salt, which prevents the formation of goiters, has helped eliminate this concern in the US where salt consumption is high. (credit a: modification of work by "Almazi"/Wikimedia Commons; credit b: modification of work by Mike Mozart)

The addition of small amounts of iodine to table salt (iodized salt) has essentially eliminated this health concern in the United States, but as much as 40% of the world's population is still at risk of iodine deficiency. The iodine atoms are added as anions, and each has a 1− charge and a mass number of 127. Determine the numbers of protons, neutrons, and electrons in one of these iodine anions.

Solution

The atomic number of iodine (53) tells us that a neutral iodine atom contains 53 protons in its nucleus and 53 electrons outside its nucleus. Because the sum of the numbers of protons and neutrons equals the mass number, 127, the number of neutrons is 74 (127 − 53 = 74). Since the iodine is added as a 1− anion, the number of electrons is 54 [53 – (1–) = 54].

Exercise 2.2.1

An ion of platinum has a mass number of 195 and contains 74 electrons. How many protons and neutrons does it contain, and what is its charge?

Answer

78 protons; 117 neutrons; charge is 4+

Chemical Symbols

A chemical symbol is an abbreviation that we use to indicate an element or an atom of an element. For example, the symbol for mercury is Hg (Figure 2.2.3). We use the same symbol to indicate one atom of mercury (microscopic domain) or to label a container of many atoms of the element mercury (macroscopic domain).



***Figure* 2.2.3**: *The symbol Hg represents the element mercury regardless of the amount; it could represent one atom of mercury or a large amount of mercury. Image used with permission from Wikipedia (user: Materialscientist).*

The symbols for several common elements and their atoms are listed in Table **2.2.2**. Some symbols are derived from the common name of the element; others are abbreviations of the name in another language. Symbols have one or two letters, for example, H for hydrogen and Cl for chlorine. To avoid confusion with other notations, only the first letter of a symbol is capitalized. For example, Co is the symbol for the element cobalt, but CO is the notation for the compound carbon monoxide, which contains atoms of the elements carbon (C) and oxygen (O). All known elements and their symbols are in the periodic table.

***Table* 2.2.2**: *Some Common Elements and Their Symbols*

Element	Symbol	Element	Symbol
aluminum	Al	iron	Fe (from *ferrum*)
bromine	Br	lead	Pb (from *plumbum*)
calcium	Ca	magnesium	Mg
carbon	C	mercury	Hg (from *hydrargyrum*)
chlorine	Cl	nitrogen	N
chromium	Cr	oxygen	O
cobalt	Co	potassium	K (from *kalium*)
copper	Cu (from *cuprum*)	silicon	Si
fluorine	F	silver	Ag (from *argentum*)
gold	Au (from *aurum*)	sodium	Na (from *natrium*)
helium	He	sulfur	S
hydrogen	H	tin	Sn (from *stannum*)
iodine	I	zinc	Zn

Traditionally, the discoverer (or discoverers) of a new element names the element. However, until the name is recognized by the International Union of Pure and Applied Chemistry (IUPAC), the recommended name of the new element is based on the Latin word(s) for its atomic number. For example, element 106 was called unnilhexium (Unh), element 107 was called unnilseptium (Uns), and element 108 was called unniloctium (Uno) for several years. These elements are now named after scientists or locations; for example, element 106 is now known as *seaborgium* (Sg) in honor of Glenn Seaborg, a Nobel Prize winner who was active in the discovery of several heavy elements.

IUPAC

Visit this site to learn more about IUPAC, the International Union of Pure and Applied Chemistry, and explore its periodic table.

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2.2.1: Practice Problems- The Structure of the Atom and How We Represent It (Optional)

PROBLEM 2.2.1.1

In what way are isotopes of a given element always different? In what way(s) are they always the same?

Answer

They **always** have different masses due to ***different numbers of neutrons.***

They **always** have the **same number** of protons (which determines the identity).

PROBLEM 2.2.1.2

Write the symbol for each of the following ions:

(a) the ion with a 1+ charge, atomic number 55, and mass number 133

(b) the ion with 54 electrons, 53 protons, and 74 neutrons

(c) the ion with atomic number 15, mass number 31, and a 3− charge

(d) the ion with 24 electrons, 30 neutrons, and a 3+ charge

Answer a

$^{133}_{55}Cs^{+}$

Answer b

$^{127}_{53}I^{-}$

Answer c

$^{31}_{15}P^{3-}$

Answer d

$^{57}_{27}Co^{3+}$

PROBLEM 2.2.1.3

Write the symbol for each of the following ions:

(a) the ion with a 3+ charge, 28 electrons, and a mass number of 71

(b) the ion with 36 electrons, 35 protons, and 45 neutrons

(c) the ion with 86 electrons, 142 neutrons, and a 4+ charge

(d) the ion with a 2+ charge, atomic number 38, and mass number 87

Answer a

$^{71}_{31}Ga^{3+}$

Answer b

$^{80}_{35}Br^{-}$

Answer c

$^{232}_{90}Th^{4+}$

Answer d

$^{87}_{38}Sr^{2+}$

Click here for a video of the solution.



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2.3: Calculating Atomic Masses

Skills to Develop

- Define the atomic mass unit and average atomic mass
- Calculate average atomic mass and isotopic abundance
- Define the amount unit mole and the related quantity Avogadro's number
- Explain the relation between mass, moles, and numbers of atoms or molecules, and perform calculations deriving these quantities from one another



Video* 2.3.1*: A review of counting subatomic particles and a preview of isotopes and relative atomic mass.

Isotopes

The symbol for a specific isotope of any element is written by placing the mass number as a superscript to the left of the element symbol (Figure **2.3.4**). The atomic number is sometimes written as a subscript preceding the symbol, but since this number defines the element's identity, as does its symbol, it is often omitted. For example, magnesium exists as a mixture of three isotopes, each with an atomic number of 12 and with mass numbers of 24, 25, and 26, respectively. These isotopes can be identified as ^{24}Mg, ^{25}Mg, and ^{26}Mg. These isotope symbols are read as "element, mass number" and can be symbolized consistent with this reading. For instance, ^{24}Mg is read as "magnesium 24," and can be written as "magnesium-24" or "Mg-24." ^{25}Mg is read as "magnesium 25," and can be written as "magnesium-25" or "Mg-25." All magnesium atoms have 12 protons in their nucleus. They differ only because a ^{24}Mg atom has 12 neutrons in its nucleus, a ^{25}Mg atom has 13 neutrons, and a ^{26}Mg has 14 neutrons.

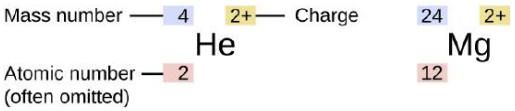
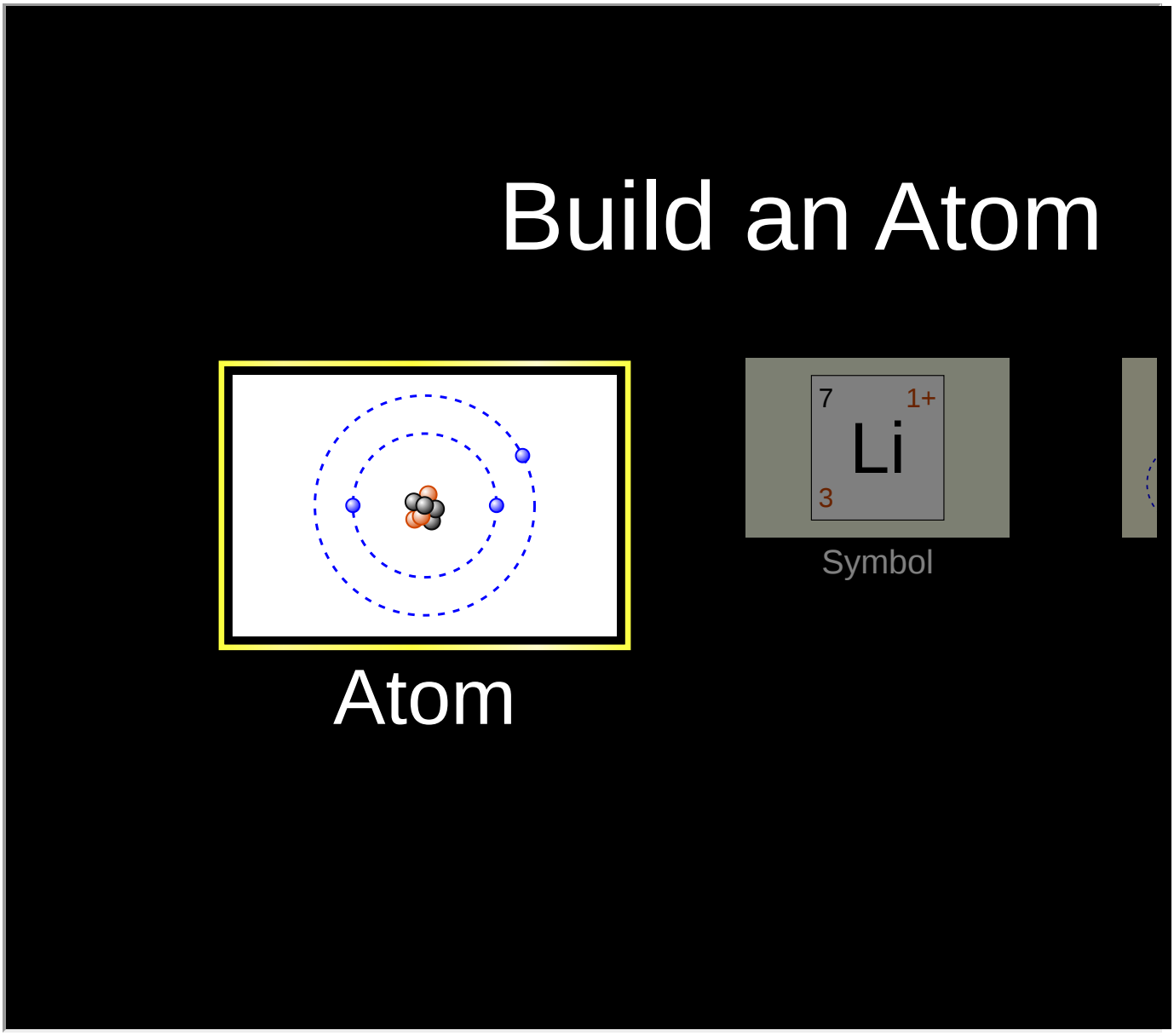


Figure* 2.3.1*: *The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted), and the charge as a right superscript.*

Information about the naturally occurring isotopes of elements with atomic numbers 1 through 10 is given in Table **2.3.2**. Note that in addition to standard names and symbols, the isotopes of hydrogen are often referred to using common names and accompanying symbols. Hydrogen-2, symbolized ^{2}H, is also called deuterium and sometimes symbolized D. Hydrogen-3, symbolized ^{3}H, is also called tritium and sometimes symbolized T.

Table* 2.3.1*: *Nuclear Compositions of Atoms of the Very Light Elements*

Element	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass (amu)	% Natural Abundance
hydrogen	$^{1}_{1}H$ (protium)	1	1	0	1.0078	99.989
	$^{2}_{1}H$ (deuterium)	1	1	1	2.0141	0.0115
	$^{3}_{1}H$ (tritium)	1	1	2	3.01605	— (trace)
helium	$^{3}_{2}He$	2	2	1	3.01603	0.00013
	$^{4}_{2}He$	2	2	2	4.0026	100
lithium	$^{6}_{3}Li$	3	3	3	6.0151	7.59
	$^{7}_{3}Li$	3	3	4	7.0160	92.41
beryllium	$^{9}_{4}Be$	4	4	5	9.0122	100
boron	$^{10}_{5}B$	5	5	5	10.0129	19.9
	$^{11}_{5}B$	5	5	6	11.0093	80.1
carbon	$^{12}_{6}C$	6	6	6	12.0000	98.89
	$^{13}_{6}C$	6	6	7	13.0034	1.11
	$^{14}_{6}C$	6	6	8	14.0032	— (trace)
nitrogen	$^{14}_{7}N$	7	7	7	14.0031	99.63
	$^{15}_{7}N$	7	7	8	15.0001	0.37
oxygen	$^{16}_{8}O$	8	8	8	15.9949	99.757
	$^{17}_{8}O$	8	8	9	16.9991	0.038
	$^{18}_{8}O$	8	8	10	17.9992	0.205
fluorine	$^{19}_{9}F$	9	9	10	18.9984	100
neon	$^{20}_{10}Ne$	10	10	10	19.9924	90.48
	$^{21}_{10}Ne$	10	10	11	20.9938	0.27
	$^{22}_{10}Ne$	10	10	12	21.9914	9.25



Use this Build an Atom simulator to build atoms of the first 10 elements, see which isotopes exist, check nuclear stability, and gain experience with isotope symbols.

Atomic Mass

Because each proton and each neutron contribute approximately one amu to the mass of an atom, and each electron contributes far less, the atomic mass of a single atom is approximately equal to its mass number (a whole number). However, the average masses of atoms of most elements are not whole numbers because most elements exist naturally as mixtures of two or more isotopes.

The mass of an element shown in a periodic table or listed in a table of atomic masses is a weighted, average mass of all the isotopes present in a naturally occurring sample of that element. This is equal to the sum of each individual isotope's mass multiplied by its fractional abundance.

$$\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i \tag{2.3.1}$$

For example, the element boron is composed of two isotopes: About 19.9% of all boron atoms are ^{10}B with a mass of 10.0129 amu, and the remaining 80.1% are ^{11}B with a mass of 11.0093 amu. The average atomic mass for boron is calculated to be:

$$\begin{aligned}\text{boron average mass} &= (0.199 \times 10.0129\ \text{amu}) + (0.801 \times 11.0093\ \text{amu}) \\ &= 1.99\ \text{amu} + 8.82\ \text{amu} \\ &= 10.81\ \text{amu}\end{aligned}$$

It is important to understand that no single boron atom weighs exactly 10.8 amu; 10.8 amu is the average mass of all boron atoms, and individual boron atoms weigh either approximately 10 amu or 11 amu.

Example 2.3.1: Calculation of Average Atomic Mass

A meteorite found in central Indiana contains traces of the noble gas neon picked up from the solar wind during the meteorite's trip through the solar system. Analysis of a sample of the gas showed that it consisted of 91.84% ^{20}Ne (mass 19.9924 amu), 0.47% ^{21}Ne (mass 20.9940 amu), and 7.69% ^{22}Ne (mass 21.9914 amu). What is the average mass of the neon in the solar wind?

Solution

$$\begin{aligned}\text{average mass} &= (0.9184 \times 19.9924\ \text{amu}) + (0.0047 \times 20.9940\ \text{amu}) + (0.0769 \times 21.9914\ \text{amu}) \\ &= (18.36 + 0.099 + 1.69)\ \text{amu} \\ &= 20.15\ \text{amu}\end{aligned}$$

The average mass of a neon atom in the solar wind is 20.15 amu. (The average mass of a terrestrial neon atom is 20.1796 amu. This result demonstrates that we may find slight differences in the natural abundance of isotopes, depending on their origin.)

Exercise 2.3.1

A sample of magnesium is found to contain 78.70% of ^{24}Mg atoms (mass 23.98 amu), 10.13% of ^{25}Mg atoms (mass 24.99 amu), and 11.17% of ^{26}Mg atoms (mass 25.98 amu). Calculate the average mass of a Mg atom.

Answer

24.31 amu

We can also do variations of this type of calculation, as shown in the next example.

Example 2.3.2: Calculation of Percent Abundance

Naturally occurring chlorine consists of ^{35}Cl (mass 34.96885 amu) and ^{37}Cl (mass 36.96590 amu), with an average mass of 35.453 amu. What is the percent composition of Cl in terms of these two isotopes?

Solution

The average mass of chlorine is the fraction that is ^{35}Cl times the mass of ^{35}Cl plus the fraction that is ^{37}Cl times the mass of ^{37}Cl.

$$\text{average mass} = (\text{fraction of } ^{35}\text{Cl} \times \text{mass of } ^{35}\text{Cl}) + (\text{fraction of } ^{37}\text{Cl} \times \text{mass of } ^{37}\text{Cl}) \tag{2.3.2}$$

If we let x represent the fraction that is ^{35}Cl, then the fraction that is ^{37}Cl is represented by $1.00 - x$.

(The fraction that is ^{35}Cl + the fraction that is ^{37}Cl must add up to 1, so the fraction of ^{37}Cl must equal 1.00 − the fraction of ^{35}Cl.)

Substituting this into the average mass equation, we have:

$$\begin{aligned}35.453\ \text{amu} &= (x \times 34.96885\ \text{amu}) + [(1.00 - x) \times 36.96590\ \text{amu}] \\ 35.453 &= 34.96885x + 36.96590 - 36.96590x \\ 1.99705x &= 1.513 \\ x &= \frac{1.513}{1.99705} = 0.7576\end{aligned}$$

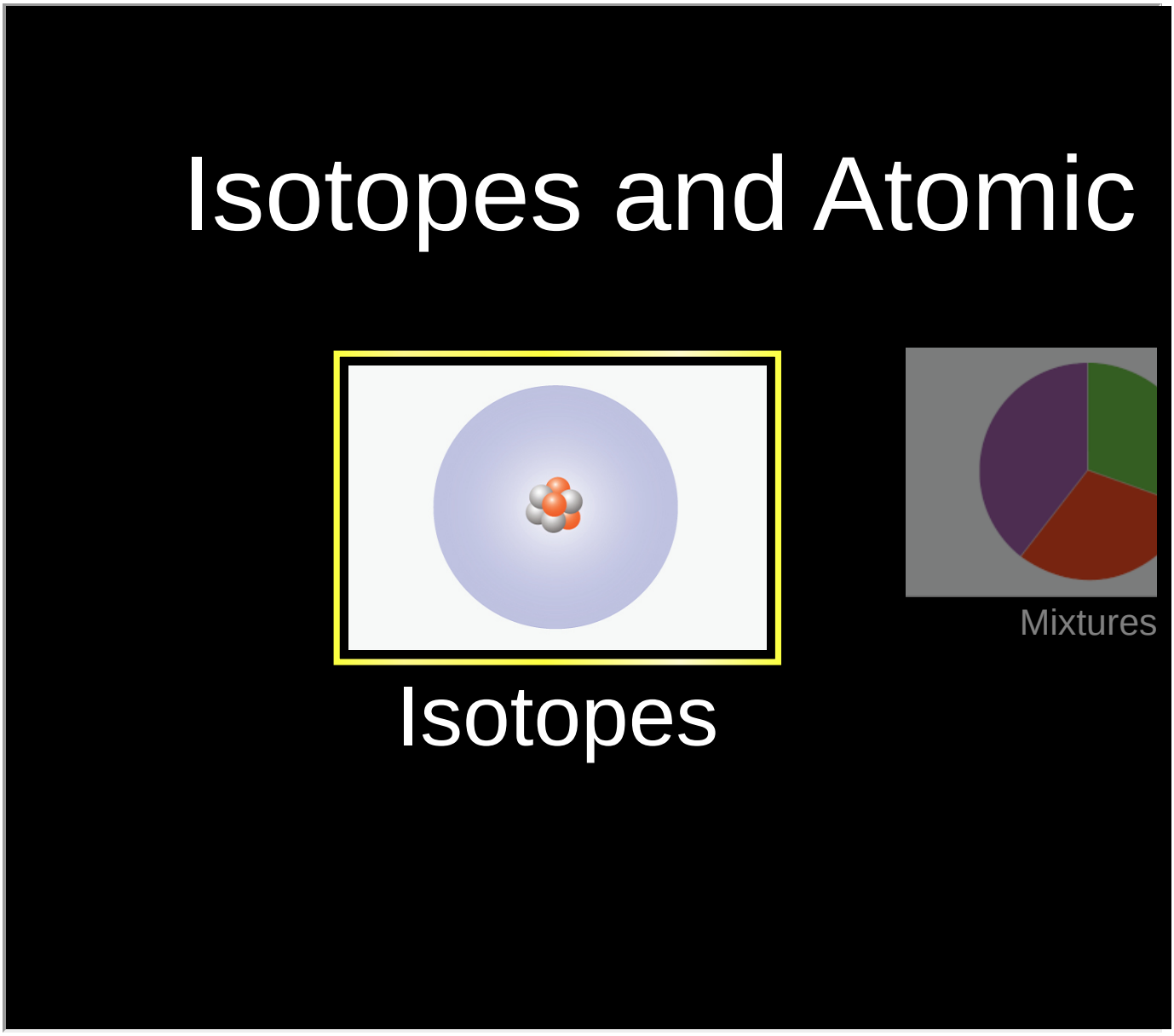
So solving yields: x = 0.7576, which means that 1.00 − 0.7576 = 0.2424. Therefore, chlorine consists of 75.76% ^{35}Cl and 24.24% ^{37}Cl.

Exercise 2.3.2

Naturally occurring copper consists of ^{63}Cu (mass 62.9296 amu) and ^{65}Cu (mass 64.9278 amu), with an average mass of 63.546 amu. What is the percent composition of Cu in terms of these two isotopes?

Answer

69.15% Cu-63 and 30.85% Cu-65



Summary



Video* 2.3.6: Watch this video for a review of relative atomic mass and isotopes.*

An atom consists of a small, positively charged nucleus surrounded by electrons. The nucleus contains protons and neutrons; its diameter is about 100,000 times smaller than that of the atom. The mass of one atom is usually expressed in atomic mass units (amu), which is referred to as the atomic mass. An amu is defined as exactly $\mathbf{1/12}$ of the mass of a carbon-12 atom and is equal to 1.6605×10^{-24} g.

Protons are relatively heavy particles with a charge of 1+ and a mass of 1.0073 amu. Neutrons are relatively heavy particles with no charge and a mass of 1.0087 amu. Electrons are light particles with a charge of 1− and a mass of 0.00055 amu. The number of protons in the nucleus is called the atomic number (Z) and is the property that defines an atom's elemental identity. The sum of the numbers of protons and neutrons in the nucleus is called the mass number and, expressed in amu, is approximately equal to the mass of the atom. An atom is neutral when it contains equal numbers of electrons and protons.

Isotopes of an element are atoms with the same atomic number but different mass numbers; isotopes of an element, therefore, differ from each other only in the number of neutrons within the nucleus. When a naturally occurring element is composed of several isotopes, the atomic mass of the element represents the average of the masses of the isotopes involved. A chemical symbol identifies the atoms in a substance using symbols, which are one-, two-, or three-letter abbreviations for the atoms.

Looking Beyond



Video* 2.3.7: Remember our exploration into the size of an atom last week? This video goes deeper into investigating the size of the subatomic particles we just discussed.*

Footnotes

1. Read more about the redefinition of SI units including the kilogram here (Laura Howe, *CE&N*, Nov. 16, 2018).

Key Equations

- $\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i$

Glossary

anion
negatively charged atom or molecule (contains more electrons than protons)

atomic mass
average mass of atoms of an element, expressed in amu

atomic mass unit (amu)
(also, unified atomic mass unit, u, or Dalton, Da) unit of mass equal to $\frac{1}{12}$ of the mass of a ^{12}C atom

atomic number (Z)
number of protons in the nucleus of an atom

cation
positively charged atom or molecule (contains fewer electrons than protons)

chemical symbol
one-, two-, or three-letter abbreviation used to represent an element or its atoms

Dalton (Da)
alternative unit equivalent to the atomic mass unit

fundamental unit of charge
(also called the elementary charge) equals the magnitude of the charge of an electron (e) with $e = 1.602 \times 10^{-19}$ C

ion
electrically charged atom or molecule (contains unequal numbers of protons and electrons)

mass number (A)
sum of the numbers of neutrons and protons in the nucleus of an atom

mole
amount of substance containing the same number of atoms, molecules, ions, or other entities as the number of atoms in exactly 12 grams of ^{12}C

molar mass
mass in grams of 1 mole of a substance

unified atomic mass unit (u)
alternative unit equivalent to the atomic mass unit

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2.3.1: Practice Problems- Calculating Atomic Masses (Optional)

PROBLEM 2.3.1.1

Determine the number of protons, neutrons, and electrons in the following isotopes that are used in medical diagnoses:

(a) atomic number 9, mass number 18, charge of 1−

(b) atomic number 43, mass number 99, charge of 7+

(c) atomic number 53, atomic mass number 131, charge of 1−

(d) atomic number 81, atomic mass number 201, charge of 1+

(e) Name the elements in parts (a), (b), (c), and (d)

Answer a

p: 9; n: 9; e: 10

Answer b

p: 43; n: 56; e: 36

Answer c

p: 53; n: 78; e: 54

Answer d

p: 81; n: 120; e: 80

Answer e

a - F; b - Tc; c - I; d - Tl

Click here to see a video of the solution.



PROBLEM 2.3.1.2

Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:

(a) $^{10}_{5}\mathrm{B}$

(b) $^{199}_{80}\mathrm{Hg}$

(c) $^{63}_{29}\mathrm{Cu}$

(d) $^{13}_{6}\mathrm{C}$

(e) $^{77}_{34}Se$

Answer a

p&e: 5; n: 5

Answer b

p&e: 80; n: 119

Answer c

p&e: 29; n: 34

Answer d

p&e: 6; n: 7

Answer e

p&e: 34; n: 43

PROBLEM 2.3.1.3

An element has the following natural abundances and isotopic masses: 90.92% abundance with 19.99 amu, 0.26% abundance with 20.99 amu, and 8.82% abundance with 21.99 amu. Calculate the average atomic mass of this element.

Answer

20.16 amu

Click here to see a video of the solution.



PROBLEM 2.3.1.4

Average atomic masses listed by IUPAC are based on a study of experimental results. Bromine has two isotopes, ^{79}Br and ^{81}Br, whose masses (78.9183 and 80.9163 amu) and abundances (50.69% and 49.31%) were determined in earlier experiments. Calculate the average atomic mass of Br based on these experiments. How does this compare to the value given on the periodic table?

Answer

79.90 amu; this matches the value on the periodic table

PROBLEM 2.3.1.5

The ^{18}O:^{16}O abundance ratio in some meteorites is greater than that used to calculate the average atomic mass of oxygen on earth. Is the average atomic mass of an oxygen atom in these meteorites greater than, less than, or equal to a terrestrial oxygen

atom?

Answer

Greater, since the contribution to the average atomic mass of ^{18}O is greater, that will raise the average atomic mass in meteorites compared to on earth.

PROBLEM 2.3.1.6

Compare 1 mole of H_2, 1 mole of O_2, and 1 mole of F_2.

(a) Which has the largest number of molecules? Explain why.

(b) Which has the greatest mass? Explain why.

Answer a

1 mole is always 6.022 x 10^{23} molecules. They have the same number of molecules.

Answer b

F_2; it has the highest molar mass.

PROBLEM 2.3.1.7

Which contains the greatest mass of oxygen: 0.75 mol of ethanol (C_2H_5OH), 0.60 mol of formic acid (HCO_2H), or 1.0 mol of water (H_2O)? Explain why.

Answer

Formic acid. Its formula has twice as many oxygen atoms as the other two compounds (one each). Therefore, 0.60 mol of formic acid would be equivalent to 1.20 mol of a compound containing a single oxygen atom.

PROBLEM 2.3.1.8

Determine the mass of each of the following:

(a) 0.0146 mol KOH
(b) 10.2 mol ethane, C_2H_6
(c) 1.6×10^{-3} mol Na_2SO_4
(d) 6.854×10^3 mol glucose, $C_6H_{12}O_6$
(e) 2.86 mol $Co(NH_3)_6Cl_3$

Answer a

0.819 g

Answer b

307 g

Answer c

0.23 g

Answer d

1.235×10^6 g (1235 kg)

Answer e

765 g

PROBLEM 2.3.1.9

Which of the following represents the least number of molecules?

a. 20.0 g of H_2O (18.02 g/mol)
b. 77.0 g of CH_4 (16.06 g/mol)
c. 68.0 g of CaH_2 (42.09 g/mol)
d. 100.0 g of N_2O (44.02 g/mol)
e. 84.0 g of HF (20.01 g/mol)

Answer

20.0 g of H_2O represents the smallest number of moles, meaning the least number of molecules present. Since 1 mole = 6.022×10^{23} molecules (or atoms) regardless of identity, the least number of moles will equal the least number of molecules.

Click here to see a video of the solution.



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2.4: Using the Mole in Calculations

The Mole

The identity of a substance is defined not only by the types of atoms or ions it contains, but by the quantity of each type of atom or ion. For example, water, H_2O, and hydrogen peroxide, H_2O_2, are alike in that their respective molecules are composed of hydrogen and oxygen atoms. However, because a hydrogen peroxide molecule contains two oxygen atoms, as opposed to the water molecule, which has only one, the two substances exhibit very different properties. Today, we possess sophisticated instruments that allow the direct measurement of these defining microscopic traits; however, the same traits were originally derived from the measurement of macroscopic properties (the masses and volumes of bulk quantities of matter) using relatively simple tools (balances and volumetric glassware). This experimental approach required the introduction of a new unit for amount of substances, the *mole*, which remains indispensable in modern chemical science.

The mole is an amount unit similar to familiar units like pair, dozen, gross, etc. It provides a specific measure of *the number* of atoms or molecules in a bulk sample of matter. A mole is defined as the amount of substance containing the same number of discrete entities (such as atoms, molecules, and ions) as the number of atoms in a sample of pure ^{12}C weighing exactly 12 g. One Latin connotation for the word “mole” is “large mass” or “bulk,” which is consistent with its use as the name for this unit. The mole provides a link between an easily measured macroscopic property, bulk mass, and an extremely important fundamental property, number of atoms, molecules, and so forth.

The number of entities composing a mole has been experimentally determined to be $6.02214179 \times 10^{23}$, a fundamental constant named Avogadro’s number (N_A) or the Avogadro constant in honor of Italian scientist Amedeo Avogadro. This constant is properly reported with an explicit unit of “per mole,” a conveniently rounded version being 6.022 \times 10^{23}\/ce{mol}\).



***Video* 2.4.1**: *What is Avogadro's Number?*

Consistent with its definition as an amount unit, 1 mole of any element contains the same number of atoms as 1 mole of any other element. The masses of 1 mole of different elements, however, are different, since the masses of the individual atoms are drastically different. The molar mass of an element (or compound) is the mass in grams of 1 mole of that substance, a property expressed in units of grams per mole (g/mol) (Figure **2.4.1**).

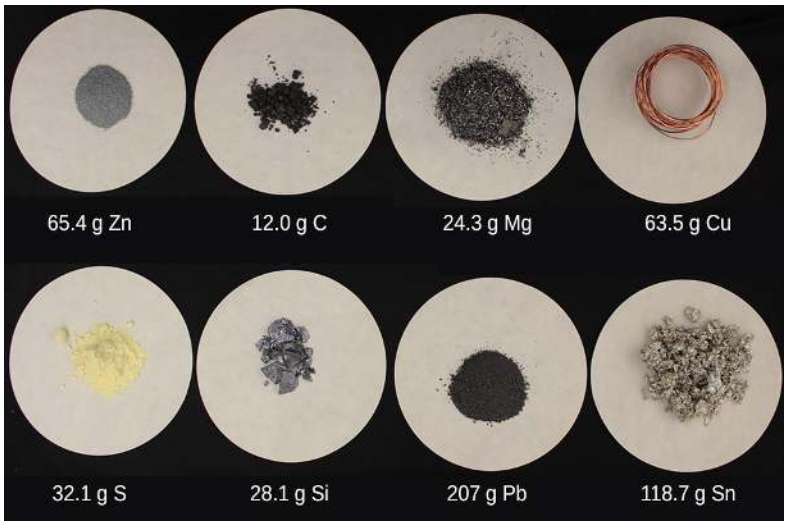


Figure* 2.4.1: Each sample contains* 6.022×10^{23} *atoms —1.00 mol of atoms. From left to right (top row): 65.4 g zinc, 12.0 g carbon, 24.3 g magnesium, and 63.5 g copper. From left to right (bottom row): 32.1 g sulfur, 28.1 g silicon, 207 g lead, and 118.7 g tin. (credit: modification of work by Mark Ott).*

Because the definitions of both the mole and the atomic mass unit are based on the same reference substance, ^{12}C, the molar mass of any substance is numerically equivalent to its atomic or formula weight in amu. Per the amu definition, a single ^{12}C atom weighs 12 amu (its atomic mass is 12 amu). The former definition of the mole was that a mole was 12 g of ^{12}C contains 1 mole of ^{12}C atoms (its molar mass is 12 g/mol). This relationship holds for all elements, since their atomic masses are measured relative to that of the amu-reference substance, ^{12}C. Extending this principle, the molar mass of a compound in grams is likewise numerically equivalent to its formula mass in amu. On May 20, 2019 the definition was permanently changed to Avogadro's number: a mole is $6.02214179 \times 10^{23}$ of any object, from atoms to apples.[1]

Table* 2.4.2*: *Mass of one mole of elements*

Element	Average Atomic Mass (amu)	Molar Mass (g/mol)	Atoms/Mole
C	12.01	12.01	6.022×10^{23}
H	1.008	1.008	6.022×10^{23}
O	16.00	16.00	6.022×10^{23}
Na	22.99	22.99	6.022×10^{23}
Cl	33.45	35.45	6.022×10^{23}

While atomic mass and molar mass are numerically equivalent, keep in mind that they are vastly different in terms of scale, as represented by the vast difference in the magnitudes of their respective units (amu versus g). To appreciate the enormity of the mole, consider a small drop of water after a rainfall. Although this represents just a tiny fraction of 1 mole of water (~18 g), it contains more water molecules than can be clearly imagined. If the molecules were distributed equally among the roughly seven billion people on earth, each person would receive more than 100 billion molecules.



***Video* 2.4.2**: *The mole is used in chemistry to represent* 6.022×10^{23} *of something, but it can be difficult to conceptualize such a large number. Watch this video and then complete the "Think" questions that follow. Explore more about the mole by reviewing the information under "Dig Deeper."*

The relationships between formula mass, the mole, and Avogadro's number can be applied to compute various quantities that describe the composition of substances and compounds. For example, if we know the mass and chemical composition of a substance, we can determine the number of moles and calculate number of atoms or molecules in the sample. Likewise, if we know the number of moles of a substance, we can derive the number of atoms or molecules and calculate the substance's mass.

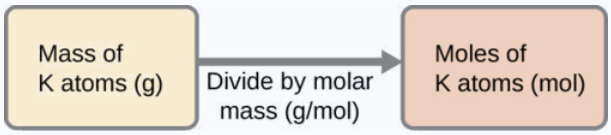
Example 2.4.1: Deriving Moles from Grams for an Element

According to nutritional guidelines from the US Department of Agriculture, the estimated average requirement for dietary potassium is 4.7 g. What is the estimated average requirement of potassium in moles?

Solution

The mass of K is provided, and the corresponding amount of K in moles is requested. Referring to the periodic table, the atomic mass of K is 39.10 amu, and so its molar mass is 39.10 g/mol. The given mass of K (4.7 g) is a bit more than one-tenth the molar mass (39.10 g), so a reasonable "ballpark" estimate of the number of moles would be slightly greater than 0.1 mol.

The molar amount of a substance may be calculated by dividing its mass (g) by its molar mass (g/mol):



The factor-label method supports this mathematical approach since the unit "g" cancels and the answer has units of "mol:"

$$4.7 \cancel{g}\,\text{K}\left(\frac{\text{mol K}}{39.10 \cancel{g}}\right) = 0.12 \text{ mol K}$$

The calculated magnitude (0.12 mol K) is consistent with our ballpark expectation, since it is a bit greater than 0.1 mol.

Exercise 2.4.1: Beryllium

Beryllium is a light metal used to fabricate transparent X-ray windows for medical imaging instruments. How many moles of Be are in a thin-foil window weighing 3.24 g?

Answer

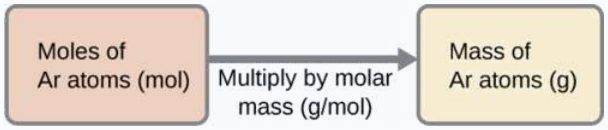
0.360 mol

Example 2.4.2: Deriving Grams from Moles for an Element

A liter of air contains 9.2×10^{-4} mol argon. What is the mass of Ar in a liter of air?

Solution

The molar amount of Ar is provided and must be used to derive the corresponding mass in grams. Since the amount of Ar is less than 1 mole, the mass will be less than the mass of 1 mole of Ar, approximately 40 g. The molar amount in question is approximately one-one thousandth (~10^{-3}) of a mole, and so the corresponding mass should be roughly one-one thousandth of the molar mass (~0.04 g):



In this case, logic dictates (and the factor-label method supports) multiplying the provided amount (mol) by the molar mass (g/mol):

$$9.2 \times 10^{-4} \cancel{\text{mol Ar}} \left(\frac{39.95 \text{ g}}{\cancel{\text{mol Ar}}} \right) = 0.037 \text{ g Ar}$$

The result is in agreement with our expectations, around 0.04 g Ar.

Exercise 2.4.2

What is the mass of 2.561 mol of gold?

Answer

504.4 g

Example 2.4.3: Deriving Number of Atoms from Mass for an Element

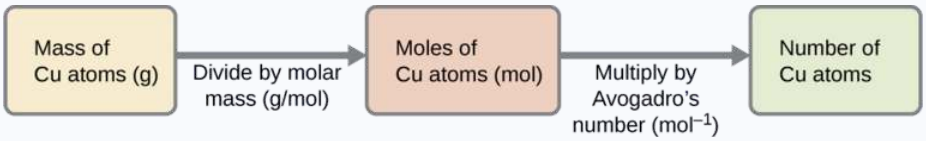
Copper is commonly used to fabricate electrical wire (Figure 2.4.6). How many copper atoms are in 5.00 g of copper wire?



Figure 2.4.6: *Copper wire is composed of many, many atoms of Cu. (credit: Emilian Robert Vicol)*

Solution

The number of Cu atoms in the wire may be conveniently derived from its mass by a two-step computation: first calculating the molar amount of Cu, and then using Avogadro's number (N_A) to convert this molar amount to number of Cu atoms:



Considering that the provided sample mass (5.00 g) is a little less than one-tenth the mass of 1 mole of Cu (~64 g), a reasonable estimate for the number of atoms in the sample would be on the order of one-tenth N_A, or approximately 10^{22} Cu atoms. Carrying out the two-step computation yields:

$$5.00 \cancel{\text{ g Cu}} \left(\frac{\cancel{\text{mol Cu}}}{63.55 \cancel{\text{ g}}} \right) \left(\frac{6.022 \times 10^{23} \text{ atoms}}{\cancel{\text{mol}}} \right) = 4.74 \times 10^{22} \text{ atoms of copper} \quad (2.4.1)$$

The factor-label method yields the desired cancellation of units, and the computed result is on the order of 10^{22} as expected.

Exercise 2.4.3

A prospector panning for gold in a river collects 15.00 g of pure gold. How many Au atoms are in this quantity of gold?

Answer

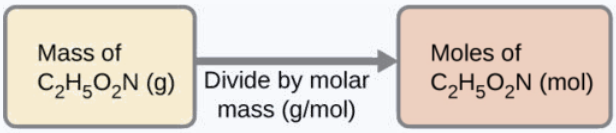
4.586×10^{22} Au atoms

Example 2.4.4: Deriving Moles from Grams for a Compound

Our bodies synthesize protein from amino acids. One of these amino acids is glycine, which has the molecular formula $C_2H_5O_2N$. How many moles of glycine molecules are contained in 28.35 g of glycine?

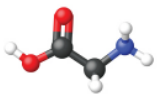
Solution

We can derive the number of moles of a compound from its mass following the same procedure we used for an element in Example **2.4.6**:



The molar mass of glycine is required for this calculation, and it is computed in the same fashion as its molecular mass. One mole of glycine, $C_2H_5O_2N$, contains 2 moles of carbon, 5 moles of hydrogen, 2 moles of oxygen, and 1 mole of nitrogen:

Element	Quantity (mol element/ mol compound)		Molar mass (g/mol element)		Subtotal (g/mol compound)
C	2	×	12.01	=	24.02
H	5	×	1.008	=	5.040
O	2	×	16.00	=	32.00
N	1	×	14.007	=	14.007
Molecular mass (g/mol compound)					75.07



The provided mass of glycine (~28 g) is a bit more than one-third the molar mass (~75 g/mol), so we would expect the computed result to be a bit greater than one-third of a mole (~0.33 mol). Dividing the compound's mass by its molar mass yields:

$$28.35 \cancel{g} \text{ glycine} \left(\frac{\text{mol glycine}}{75.07 \cancel{g}} \right) = 0.378 \text{ mol glycine}$$

This result is consistent with our rough estimate.

Exercise 2.4.4

How many moles of sucrose, $C_{12}H_{22}O_{11}$, are in a 25-g sample of sucrose?

Answer

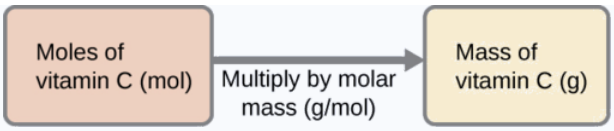
0.073 mol

Example 2.4.5: Deriving Grams from Moles for a Compound

Vitamin C is a covalent compound with the molecular formula $C_6H_8O_6$. The recommended daily dietary allowance of vitamin C for children aged 4–8 years is 1.42×10^{-4} mol. What is the mass of this allowance in grams?

Solution

As for elements, the mass of a compound can be derived from its molar amount as shown:



The molar mass for this compound is computed to be 176.124 g/mol. The given number of moles is a very small fraction of a mole (~10^{-4} or one-ten thousandth); therefore, we would expect the corresponding mass to be about one-ten thousandth of the molar mass (~0.02 g). Performing the calculation, we get:

$$1.42 \times 10^{-4} \cancel{\text{mol}}\ \text{vitamin C} \left(\frac{176.124\ \text{g}}{\cancel{\text{mol}}\ \text{vitamin C}} \right) = 0.0250\ \text{g vitamin C}$$

This is consistent with the anticipated result.

Exercise 2.4.5

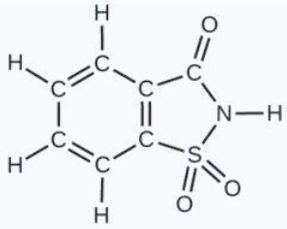
What is the mass of 0.443 mol of hydrazine, N_2H_4?

Answer

14.2 g

Example 2.4.6: Deriving the Number of Molecules from the Compound Mass

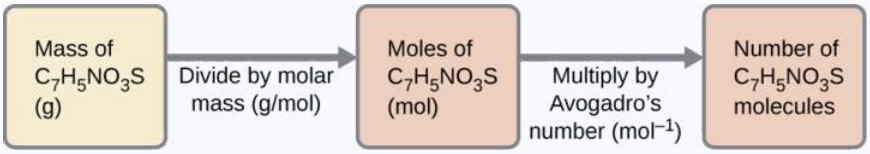
A packet of an artificial sweetener contains 40.0 mg of saccharin ($C_7H_5NO_3S$), which has the structural formula:



Given that saccharin has a molar mass of 183.18 g/mol, how many saccharin molecules are in a 40.0-mg (0.0400-g) sample of saccharin? How many carbon atoms are in the same sample?

Solution

The number of molecules in a given mass of compound is computed by first deriving the number of moles, as demonstrated in Example **2.4.8**, and then multiplying by Avogadro's number:



Using the provided mass and molar mass for saccharin yields:

$$0.0400\ \cancel{\text{g}}\ \text{C}_7\text{H}_5\text{NO}_3\text{S} \left(\frac{\cancel{\text{mol}}\ \text{C}_7\text{H}_5\text{NO}_3\text{S}}{183.18\ \cancel{\text{g}}\ \text{C}_7\text{H}_5\text{NO}_3\text{S}} \right) \left(\frac{6.022 \times 10^{23}\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecules}}{1\ \cancel{\text{mol}}\ \text{C}_7\text{H}_5\text{NO}_3\text{S}} \right) = 1.31 \times 10^{20}\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecules} \tag{2.4.2}$$

The compound's formula shows that each molecule contains seven carbon atoms, and so the number of C atoms in the provided sample is:

$$1.31 \times 10^{20}\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecules} \left(\frac{7\ \text{C atoms}}{1\ \text{C}_7\text{H}_5\text{NO}_3\text{S molecule}} \right) = 9.20 \times 10^{21}\ \text{C atoms}$$

Exercise 2.4.6

How many C_4H_{10} molecules are contained in 9.213 g of this compound? How many hydrogen atoms?

Answer

- 9.545×10^{22} molecules C_4H_{10}
- 9.545×10^{23} atoms H

Summary



Video* 2.4.5: A preview of some of the uses we will have for moles in upcoming units*

Glossary

mole

amount of substance containing the same number of atoms, molecules, ions, or other entities as the number of atoms in exactly 12 grams of ^{12}C

Contributors

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2.5: Highlighting Instrumentation- Mass Spectrometry

Skills to Develop

- Define the atomic mass unit and average atomic mass
- Calculate average atomic mass and isotopic abundance
- Define the amount unit mole and the related quantity Avogadro's number
- Explain the relation between mass, moles, and numbers of atoms or molecules, and perform calculations deriving these quantities from one another

Mass Spectrometry

The occurrence and natural abundances of isotopes can be experimentally determined using an instrument called a mass spectrometer. Mass spectrometry (MS) is widely used in chemistry, forensics, medicine, environmental science, and many other fields to analyze and help identify the substances in a sample of material. In a typical mass spectrometer (Figure **2.5.1**), the sample is vaporized and exposed to a high-energy electron beam that causes the sample's atoms (or molecules) to become electrically charged, typically by losing one or more electrons. These cations then pass through a (variable) electric or magnetic field that deflects each cation's path to an extent that depends on both its mass and charge (similar to how the path of a large steel ball bearing rolling past a magnet is deflected to a lesser extent that that of a small steel BB). The ions are detected, and a plot of the relative number of ions generated versus their mass-to-charge ratios (a *mass spectrum*) is made. The height of each vertical feature or peak in a mass spectrum is proportional to the fraction of cations with the specified mass-to-charge ratio. Since its initial use during the development of modern atomic theory, MS has evolved to become a powerful tool for chemical analysis in a wide range of applications.

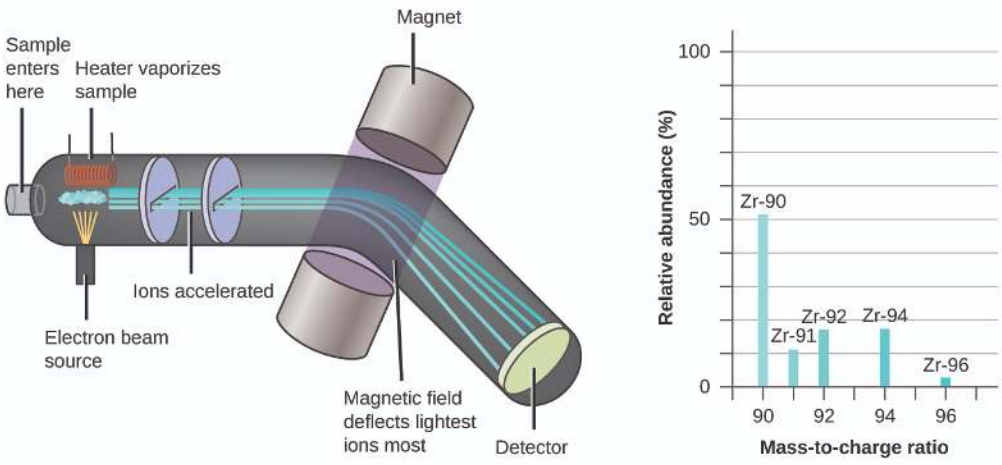


Figure* 2.5.1*: *Analysis of zirconium in a mass spectrometer produces a mass spectrum with peaks showing the different isotopes of Zr.*

T



Video **2.5.1***: Watch this video from the Royal Society for Chemistry for a brief description of the rudiments of mass spectrometry.*

Contributors

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- Adelaide Clark, Oregon Institute of Technology

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CHAPTER OVERVIEW

3: Nuclei, Ions and the Periodic Table

3.1: Nuclear Chemistry and Radioactive Decay

3.1.1: Practice Problems- Nuclear Chemistry and Radioactive Decay (Optional)

3.2: Organization of the Periodic Table

3.2.1: Practice Problems - Organization of the Periodic Table (Optional)

3.3: Predicting Charges of Ions

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3.1: Nuclear Chemistry and Radioactive Decay

Skills to Develop

- Describe nuclear structure in terms of protons, neutrons, and electrons
- Recognize the role of mass defect and binding energy for nuclei in atomic masses
- Explain trends in the relative stability of nuclei
- Identify common particles and energies involved in nuclear reactions
- Write and balance nuclear equations
- Recognize common modes of radioactive decay
- Describe kinetic parameters for decay processes, including half-life
- Describe common radiometric dating techniques

A Review of Isotopes and Introduction to Radioactivity



Video* 3.1.1: A video review of isotopes (Unit 2), with a peek at radioactive isotopes.*

Nuclear Chemistry - An Introduction

Nuclear chemistry is the study of reactions that involve changes in nuclear structure. The chapter on atoms, molecules, and ions introduced the basic idea of nuclear structure, that the nucleus of an atom is composed of protons and, with the exception of $^{1}_{1}\mathrm{H}$, neutrons. Recall that the number of protons in the nucleus is called the atomic number (Z) of the element, and the sum of the number of protons and the number of neutrons is the mass number (A). Atoms with the same atomic number but different mass numbers are isotopes of the same element. When referring to a single type of nucleus, we often use the term nuclide and identify it by the notation:

$$^{A}_{Z}\mathrm{X} \tag{3.1.1}$$

where

- X is the symbol for the element,
- A is the mass number, and
- Z is the atomic number.

Often a nuclide is referenced by the name of the element followed by a hyphen and the mass number. For example, $^{14}_{6}\mathrm{C}$ is called “carbon-14.”

Protons and neutrons, collectively called nucleons, are packed together tightly in a nucleus. With a radius of about 10^{-15} meters, a nucleus is quite small compared to the radius of the entire atom, which is about 10^{-10} meters. Nuclei are extremely dense compared to bulk matter, averaging 1.8×10^{14} grams per cubic centimeter. For example, water has a density of 1 gram per cubic centimeter, and iridium, one of the densest elements known, has a density of 22.6 g/cm^3. If the earth’s density were equal to the average nuclear density, the earth’s radius would be only about 200 meters (earth’s actual radius is approximately 6.4×10^6 meters, 30,000 times larger).

Recall our exploration of atomic size with this video:



***Video* 3.1.2**: *An exploration of atomic size.*

To hold positively charged protons together in the very small volume of a nucleus requires very strong attractive forces because the positively charged protons repel one another strongly at such short distances. The force of attraction that holds the nucleus together is the strong nuclear force. (The strong force is one of the four fundamental forces that are known to exist. The others are the electromagnetic force, the gravitational force, and the nuclear weak force.) This force acts between protons, between neutrons, and between protons and neutrons. It is very different from the electrostatic force that holds negatively charged electrons around a positively charged nucleus (the attraction between opposite charges). Over distances less than 10^{-15} meters and within the nucleus, the strong nuclear force is much stronger than electrostatic repulsions between protons; over larger distances and outside the nucleus, it is essentially nonexistent.

Nuclear Binding Energy

As a simple example of the energy associated with the strong nuclear force, consider the helium atom composed of two protons, two neutrons, and two electrons. The total mass of these six subatomic particles may be calculated as:

$$\underset{\text{protons}}{(2 \times 1.0073 \text{ amu})} + \underset{\text{neutrons}}{(2 \times 1.0087 \text{ amu})} + \underset{\text{electrons}}{(2 \times 0.00055 \text{ amu})} = 4.0331 \text{ amu} \tag{3.1.2}$$

However, mass spectrometric measurements reveal that the mass of an $^{4}_{2}\text{He}$ atom is 4.0026 amu, less than the combined masses of its six constituent subatomic particles. This difference between the calculated and experimentally measured masses is known as the **mass defect** of the atom. In the case of helium, the mass defect indicates a "loss" in mass of 4.0331 amu – 4.0026 amu = 0.0305 amu. *The loss in mass accompanying the formation of an atom from protons, neutrons, and electrons is due to the conversion of that mass into energy that is evolved as the atom forms.* The **nuclear binding energy** is the energy produced when the atoms' nucleons are bound together; this is also the energy needed to break a nucleus into its constituent protons and neutrons. In comparison to chemical bond energies, nuclear binding energies are *vastly* greater, as we will learn in this section. Consequently, the energy changes associated with nuclear reactions are vastly greater than are those for chemical reactions.

The conversion between mass and energy is most identifiably represented by the mass-energy equivalence equation as stated by Albert Einstein:

$$E = mc^2 \tag{3.1.3}$$

where E is energy, m is mass of the matter being converted, and c is the speed of light in a vacuum. This equation can be used to find the amount of energy that results when matter is converted into energy. Using this mass-energy equivalence equation, the nuclear binding energy of a nucleus may be calculated from its **mass defect**, a calculation beyond the scope of our course. A variety of units are commonly used for nuclear binding energies, including electron volts (eV), with 1 eV equaling the amount of energy necessary to the move the charge of an electron across an electric potential difference of 1 volt, making $1 \text{ eV} = 1.602 \times 10^{-19} \text{ J}$.

Because the energy changes for breaking and forming bonds are so small compared to the energy changes for breaking or forming nuclei, the changes in mass during all ordinary chemical reactions are virtually undetectable. As we will discuss later in our unit on

thermochemistry, the most energetic chemical reactions exhibit enthalpies on the order of *thousands* of kJ/mol, which is equivalent to mass differences in the nanogram range (10^{-9} g). On the other hand, nuclear binding energies are typically on the order of *billions* of kJ/mol, corresponding to mass differences in the milligram range (10^{-3} g).

Nuclear Stability

A nucleus is stable if it cannot be transformed into another configuration without adding energy from the outside. Of the thousands of nuclides that exist, about 250 are stable. A plot of the number of neutrons versus the number of protons for stable nuclei reveals that the stable isotopes fall into a narrow band. This region is known as the band of stability (also called the belt, zone, or valley of stability). The straight line in Figure **3.1.1** represents nuclei that have a 1:1 ratio of protons to neutrons (n:p ratio). Note that the lighter stable nuclei, in general, have equal numbers of protons and neutrons. For example, nitrogen-14 has seven protons and seven neutrons. Heavier stable nuclei, however, have increasingly more neutrons than protons. For example: iron-56 has 30 neutrons and 26 protons, an n:p ratio of 1.15, whereas the stable nuclide lead-207 has 125 neutrons and 82 protons, an n:p ratio equal to 1.52. This is because *larger nuclei have more proton-proton repulsions*, and require larger numbers of neutrons to provide compensating strong forces to overcome these electrostatic repulsions and hold the nucleus together.

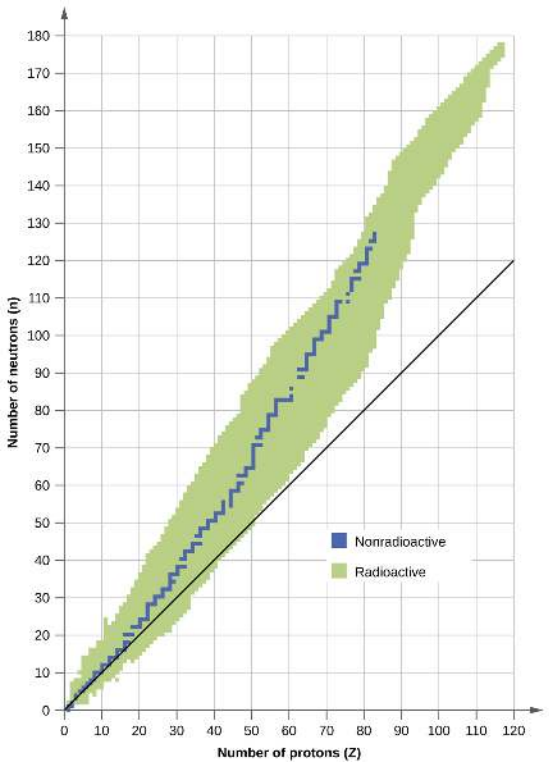


***Figure* 3.1.1**: *This plot shows the nuclides that are known to exist and those that are stable. The stable nuclides are indicated in blue, and the unstable nuclides are indicated in green. Note that all isotopes of elements with atomic numbers greater than 83 are unstable. The solid line is the line where n = Z.*

The nuclei that are to the left or to the right of the band of stability are unstable and exhibit radioactivity. They change spontaneously (decay) into other nuclei that are either in, or closer to, the band of stability. These nuclear decay reactions convert one unstable isotope (or radioisotope) into another, more stable, isotope. We will discuss the nature and products of this radioactive decay in subsequent sections of this unit.

Several observations may be made regarding the relationship between the stability of a nucleus and its structure. Nuclei with even numbers of protons, neutrons, or both are more likely to be stable (Table **3.1.1**). Nuclei with certain numbers of nucleons, known as **magic numbers**, are stable against nuclear decay. These numbers of protons or neutrons (2, 8, 20, 28, 50, 82, and 126) make complete shells in the nucleus. These are similar in concept to the stable electron shells observed for the noble gases. Nuclei that have magic numbers of both protons and neutrons, such as $^{4}_{2}\mathrm{He}$, $^{16}_{8}\mathrm{O}$, $^{40}_{20}\mathrm{Ca}$, and $^{208}_{82}\mathrm{Pb}$ and are particularly stable. These trends in

nuclear stability may be rationalized by considering a quantum mechanical model of nuclear energy states analogous to that used to describe electronic states, which we will discuss later in this course. The details of this model are beyond the scope of this course.

Table* 3.1.1*: *Stable Nuclear Isotopes*

Number of Stable Isotopes	Proton Number	Neutron Number
157	even	even
53	even	odd
50	odd	even
5	odd	odd

The relative stability of a nucleus is correlated with its binding energy per nucleon, the total binding energy for the nucleus divided by the number or nucleons in the nucleus. For instance, the binding energy for a $^4_2\mathrm{He}$ nucleus is therefore:

$$\frac{28.4\ \mathrm{MeV}}{4\ \mathrm{nucleons}} = 7.10\ \mathrm{MeV/nucleon} \tag{3.1.4}$$

The binding energy per nucleon of a nuclide on the curve shown in Figure 3.1.2

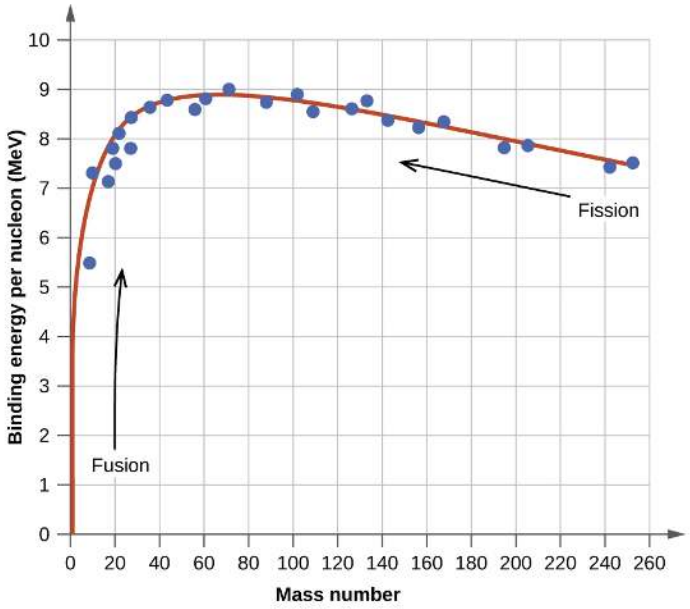


Figure* 3.1.2*: *The binding energy per nucleon is largest for nuclides with mass number of approximately 56.*

Changes of nuclei that result in changes in their atomic numbers, mass numbers, or energy states are **nuclear reactions**. To describe a nuclear reaction, we use an equation that identifies the nuclides involved in the reaction, their mass numbers and atomic numbers, and the other particles involved in the reaction.



Video* 3.1.3: A brief overview of the different types of radioactivity.*

Many entities can be involved in nuclear reactions. The most common are protons, neutrons, alpha particles, beta particles, positrons, and gamma rays, as shown in Figure **3.1.3**. Protons ($_1^1\mathrm{p}$, also represented by the symbol $_1^1\mathrm{H}$) and neutrons ($_0^1\mathrm{n}$) are the constituents of atomic nuclei, and have been described previously. **Alpha particles** ($_2^4\mathrm{He}$, also represented by the symbol $_2^4\alpha$) are *high-energy helium nuclei*. **Beta particles** ($_{-1}^0\beta$, also represented by the symbol $_{-1}^0\mathrm{e}$) are *high-energy electrons*, and **gamma rays** are *photons of very high-energy electromagnetic radiation*. **Positrons** ($_{+1}^0\mathrm{e}$, also represented by the symbol $_{+1}^0\beta$) are *positively charged electrons ("anti-electrons")*. The subscripts and superscripts are necessary for balancing nuclear equations, but are usually optional in other circumstances. For example, an alpha particle is a helium nucleus (He) with a charge of +2 and a mass number of 4, so it is symbolized $_2^4\mathrm{He}$. This works because, in general, the ion charge is not important in the balancing of nuclear equations.

This table has four columns and seven rows. The first row is a header row and it labels each column: "Name," "Symbol(s)," "Representation," and "Description." Under the "Name" column are the following: "Alpha particle," "Beta particle," "Positron," "Proton," "Neutron," and "Gamma ray." Under the "Symbol(s)" column are the following: " superscript 4 stacked over a subscript 2 H e or lowercase alpha," "superscript 0 stacked over a subscript 1 e or lowercase beta," "superscript 0 stacked over a positive subscript 1 e or lowercase beta superscript positive sign," "superscript 1 stacked over a subscript 1 H or lowercase rho superscript 1 stacked over a subscript 1 H," "superscript 1 stacked over a subscript 0 n or lowercase eta superscript 1 stacked over a subscript 0 n," and a lowercase gamma. Under the "Representation column," are the following: two white sphere attached to two blue spheres of about the same size with positive signs in them; a small red sphere with a negative sign in it; a small red sphere with a positive sign in it; a blue spheres with a positive sign in it; a white sphere; and a purple squiggle ling with an arrow pointing right to a lowercase gamma. Under the "Description" column are the following: "(High-energy) helium nuclei consisting of two protons and two neutrons," "(High-energy) elections," "Particles with the same mass as an electron but with 1 unit of positive charge," "Nuclei of hydrogen atoms," "Particles with a mass approximately equal to that of a proton but with no charge," and "Very high-energy electromagnetic radiation."

Figure* 3.1.3: Although many species are encountered in nuclear reactions, this table summarizes the names, symbols, representations, and descriptions of the most common of these.*

Note that p*ositrons are exactly like electrons, **except** they have the opposite charge*. They are the most common example of antimatter, particles with the same mass but the opposite state of another property (for example, charge) than ordinary matter. When antimatter encounters ordinary matter, both are annihilated and their mass is converted into energy in the form of gamma rays (γ)—and other much smaller subnuclear particles, which are beyond the scope of this chapter—according to the mass-energy equivalence equation $E = mc^2$, seen in the preceding section. For example, when a positron and an electron collide, both are annihilated and two gamma ray photons are created:

$$_{-1}^0\mathrm{e} + {}_{+1}^0\mathrm{e} \rightarrow \gamma + \gamma \tag{3.1.5}$$

Gamma rays compose short wavelength, high-energy electromagnetic radiation and are (much) more energetic than better-known X-rays. Gamma rays are produced when a nucleus undergoes a transition from a higher to a lower energy state, similar to how a photon is produced by an electronic transition from a higher to a lower energy level. Due to the much larger energy differences between nuclear energy shells, gamma rays emanating from a nucleus have energies that are typically millions of times larger than electromagnetic radiation emanating from electronic transitions.

Balancing Nuclear Reactions

A balanced chemical reaction equation reflects the fact that during a chemical reaction, bonds break and form, and atoms are rearranged, but the total numbers of atoms of each element are conserved and do not change. A balanced nuclear reaction equation indicates that there is a rearrangement during a nuclear reaction, but of subatomic particles rather than atoms. Nuclear reactions also follow conservation laws, and they are balanced in two ways:

1. The sum of the mass numbers of the reactants equals the sum of the mass numbers of the products.
2. The sum of the charges of the reactants equals the sum of the charges of the products.

If the atomic number and the mass number of all but one of the particles in a nuclear reaction are known, we can identify the particle by balancing the reaction. For instance, we could determine that $^{17}_{8}\mathrm{O}$ is a product of the nuclear reaction of $^{14}_{7}\mathrm{N}$ and $^{4}_{2}\mathrm{He}$ if we knew that a proton, $^{1}_{1}\mathrm{H}$, was one of the two products. Example **3.1.1** shows how we can identify a nuclide by balancing the nuclear reaction.

Example 3.1.1: Balancing Equations for Nuclear Reactions

The reaction of an α particle with magnesium-25 ($^{25}_{12}\mathrm{Mg}$) produces a proton and a nuclide of another element. Identify the new nuclide produced.

Solution

The nuclear reaction can be written as:

$$^{25}_{12}\mathrm{Mg} + {}^{4}_{2}\mathrm{He} \rightarrow {}^{1}_{1}\mathrm{H} + {}^{A}_{Z}\mathrm{X} \tag{3.1.6}$$

where

- **A** is the mass number and
- **Z** is the atomic number of the new nuclide, **X**.

Because the sum of the mass numbers of the reactants must equal the sum of the mass numbers of the products:

$$25 + 4 = A + 1 \tag{3.1.7}$$

so

$$A = 28 \tag{3.1.8}$$

Similarly, the charges must balance, so:

$$12 + 2 = Z + 1 \tag{3.1.9}$$

so

$$Z = 13 \tag{3.1.10}$$

Check the periodic table: The element with nuclear charge = +13 is aluminum. Thus, the product is $^{28}_{13}\mathrm{Al}$.

Exercise 3.1.1

The nuclide $^{125}_{53}\mathrm{I}$ combines with an electron and produces a new nucleus and no other massive particles. What is the equation for this reaction?

Answer:

$$^{125}_{53}\mathrm{I} + {}^{0}_{-1}\mathrm{e} \rightarrow {}^{125}_{52}\mathrm{Te}$$

Following are the equations of several nuclear reactions that have important roles in the history of nuclear chemistry:

- The first naturally occurring unstable element that was isolated, polonium, was discovered by the Polish scientist Marie Curie and her husband Pierre in 1898. It decays, emitting α particles:

$$^{212}_{84}\mathrm{Po} \longrightarrow {}^{208}_{82}\mathrm{Pb} + {}^{4}_{2}\mathrm{He} \tag{3.1.11}$$

- The first nuclide to be prepared by artificial means was an isotope of oxygen, $^{17}\mathrm{O}$. It was made by Ernest Rutherford in 1919 by bombarding nitrogen atoms with α particles:

$$^{14}_{7}\mathrm{N} + {}^{4}_{2}\alpha \longrightarrow {}^{17}_{8}\mathrm{O} + {}^{1}_{1}\mathrm{H} \tag{3.1.12}$$

- James Chadwick discovered the neutron in 1932, as a previously unknown neutral particle produced along with $^{12}\mathrm{C}$ by the nuclear reaction between $^{9}\mathrm{Be}$ and $^{4}\mathrm{He}$:

$$^{9}_{4}\mathrm{Be} + {}^{4}_{2}\mathrm{He} \longrightarrow {}^{12}_{6}\mathrm{C} + {}^{1}_{0}\mathrm{n} \tag{3.1.13}$$

- The first element to be prepared that does not occur naturally on the earth, technetium, was created by bombardment of molybdenum by deuterons (heavy hydrogen, $^{2}_{1}\mathrm{H}$), by Emilio Segre and Carlo Perrier in 1937:

$$^{2}_{1}\mathrm{H} + {}^{97}_{42}\mathrm{Mo} \longrightarrow 2{}^{1}_{0}\mathrm{n} + {}^{97}_{43}\mathrm{Tc} \tag{3.1.14}$$

- The first controlled nuclear chain reaction was carried out in a reactor at the University of Chicago in 1942. One of the many reactions involved was:

$$^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{87}_{35}\mathrm{Br} + {}^{146}_{57}\mathrm{La} + 3{}^{1}_{0}\mathrm{n} \tag{3.1.15}$$

Following the somewhat serendipitous discovery of radioactivity by Becquerel, many prominent scientists began to investigate this new, intriguing phenomenon. Among them were Marie Curie (Video **3.1.4**; the first woman to win a Nobel Prize, and the only person to win two Nobel Prizes in different sciences—chemistry and physics), who was the first to coin the term “radioactivity,” and Ernest Rutherford (of gold foil experiment fame), who investigated and named three of the most common types of radiation. During the beginning of the twentieth century, many radioactive substances were discovered, the properties of radiation were investigated and quantified, and a solid understanding of radiation and nuclear decay was developed.

Learn More about Marie Curie's Life & Work



***Video* 3.1.4**: *A brief overview of the life and work of Marie Curie.*

The spontaneous change of an unstable nuclide into another is radioactive decay. The unstable nuclide is called the parent nuclide; the nuclide that results from the decay is known as the daughter nuclide. The daughter nuclide may be stable, or it may decay itself. The radiation produced during radioactive decay is such that the daughter nuclide lies closer to the band of stability than the parent nuclide, so the location of a nuclide relative to the band of stability can serve as a guide to the kind of decay it will undergo (Figure **3.1.4**).

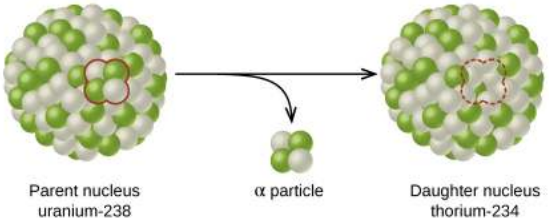


***Figure* 3.1.4**: *A nucleus of uranium-238 (the parent nuclide) undergoes α decay to form thorium-234 (the daughter nuclide). The alpha particle removes two protons (green) and two neutrons (gray) from the uranium-238 nucleus.*

Although the radioactive decay of a nucleus is too small to see with the naked eye, we can indirectly view radioactive decay in an environment called a cloud chamber. Video **3.1.5** is an opportunity to learn about cloud chambers and to view an interesting Cloud Chamber Demonstration from the Jefferson Lab.



Video* 3.1.5*: *How to Build a Cloud Chamber!*

Types of Radioactive Decay

Ernest Rutherford's experiments involving the interaction of radiation with a magnetic or electric field (Figure **3.1.5**) helped him determine that one type of radiation consisted of positively charged and relatively massive α particles; a second type was made up of negatively charged and much less massive β particles; and a third was uncharged electromagnetic waves, γ rays. We now know that α particles are high-energy helium nuclei, β particles are high-energy electrons, and γ radiation compose high-energy electromagnetic radiation. We classify different types of radioactive decay by the radiation produced.

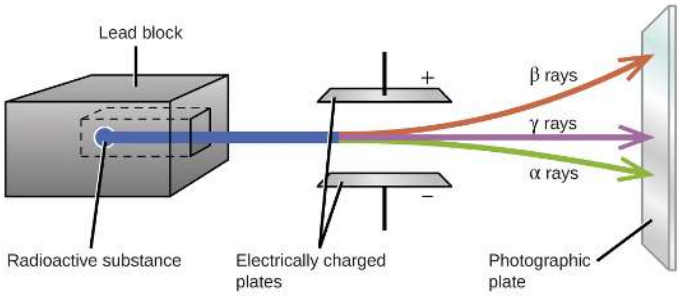


Figure* 3.1.5: Alpha particles, which are attracted to the negative plate and deflected by a relatively small amount, must be positively charged and relatively massive. Beta particles, which are attracted to the positive plate and deflected a relatively large amount, must be negatively charged and relatively light. Gamma rays, which are unaffected by the electric field, must be uncharged.*

Alpha (α) decay is the emission of an α particle from the nucleus. For example, polonium-210 undergoes α decay:

$$^{210}_{84}\text{Po} \longrightarrow {}^{4}_{2}\text{He} + {}^{206}_{82}\text{Pb} \quad \text{or} \quad {}^{210}_{84}\text{Po} \longrightarrow {}^{4}_{2}\alpha + {}^{206}_{82}\text{Pb} \tag{3.1.16}$$

Alpha decay occurs primarily in heavy nuclei (A > 200, Z > 83). Because the loss of an α particle gives a daughter nuclide with a mass number four units smaller and an atomic number two units smaller than those of the parent nuclide, the daughter nuclide has a larger n:p ratio than the parent nuclide. If the parent nuclide undergoing α decay lies below **the band of stability**, the daughter nuclide will lie closer to the band.

Beta (β) decay is the emission of an electron from a nucleus. Iodine-131 is an example of a nuclide that undergoes β decay:

$$^{131}_{53}\text{I} \longrightarrow {}^{0}_{-1}\text{e} + {}^{131}_{54}\text{X} \quad \text{or} \quad {}^{131}_{53}\text{I} \longrightarrow {}^{0}_{-1}\beta + {}^{131}_{54}\text{Xe} \tag{3.1.17}$$

Beta decay, which can be thought of as the conversion of a neutron into a proton and a β particle, is observed in nuclides with a large n:p ratio. The beta particle (electron) emitted is from the atomic nucleus and is not one of the electrons surrounding the nucleus. Such nuclei lie above the band of stability. Emission of an electron does not change the mass number of the nuclide but

does increase the number of its protons and decrease the number of its neutrons. Consequently, the n:p ratio is decreased, and the daughter nuclide lies closer to the band of stability than did the parent nuclide.

Gamma emission (γ emission) is observed when a nuclide is formed in an excited state and then decays to its ground state with the emission of a γ ray, a quantum of high-energy electromagnetic radiation. The presence of a nucleus in an excited state is often indicated by an asterisk (*). Cobalt-60 emits γ radiation and is used in many applications including cancer treatment:

$$^{60}_{27}\text{Co}^{*} \longrightarrow {}^{0}_{0}\gamma + {}^{60}_{27}\text{Co} \tag{3.1.18}$$

There is no change in mass number or atomic number during the emission of a γ ray unless the γ emission accompanies one of the other modes of decay.

Positron emission (β^{+} decay) is the emission of a positron from the nucleus. Oxygen-15 is an example of a nuclide that undergoes positron emission:

$$^{15}_{8}\text{O} \longrightarrow {}^{0}_{+1}\text{e} + {}^{15}_{7}\text{N} \quad \text{or} \quad {}^{15}_{8}\text{O} \longrightarrow {}^{0}_{+1}\beta + {}^{15}_{7}\text{N} \tag{3.1.19}$$

Positron emission is observed for nuclides in which the n:p ratio is low. These nuclides lie below the band of stability. Positron decay is the conversion of a proton into a neutron with the emission of a positron. The n:p ratio increases, and the daughter nuclide lies closer to the band of stability than did the parent nuclide.

Electron capture occurs when one of the inner electrons in an atom is captured by the atom's nucleus. For example, potassium-40 undergoes electron capture:

$$^{40}_{19}\text{K} + {}^{0}_{-1}\text{e} \longrightarrow {}^{40}_{18}\text{Ar} \tag{3.1.20}$$

Electron capture occurs when an inner shell electron combines with a proton and is converted into a neutron. The loss of an inner shell electron leaves a vacancy that will be filled by one of the outer electrons. As the outer electron drops into the vacancy, it will emit energy. In most cases, the energy emitted will be in the form of an X-ray. Like positron emission, electron capture occurs for "proton-rich" nuclei that lie below the band of stability. Electron capture has the same effect on the nucleus as does positron emission: The atomic number is decreased by one and the mass number does not change. This increases the n:p ratio, and the daughter nuclide lies closer to the band of stability than did the parent nuclide. Whether electron capture or positron emission occurs is difficult to predict. The choice is primarily due to kinetic factors, with the one requiring the smaller activation energy being the one more likely to occur. Figure **3.1.3** summarizes these types of decay, along with their equations and changes in atomic and mass numbers.

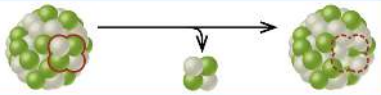
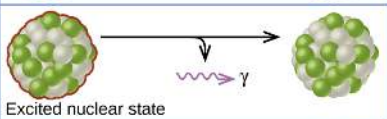
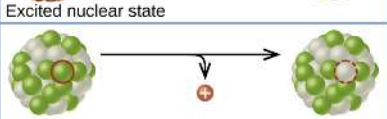
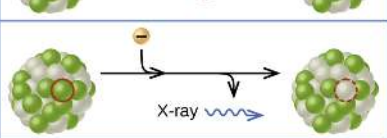
Type	Nuclear equation	Representation	Change in mass/atomic numbers
Alpha decay	$^{A}_{Z}\text{X} \rightarrow {}^{4}_{2}\text{He} + {}^{A-4}_{Z-2}\text{Y}$		A: decrease by 4 Z: decrease by 2
Beta decay	$^{A}_{Z}\text{X} \rightarrow {}^{0}_{-1}\text{e} + {}^{A}_{Z+1}\text{Y}$		A: unchanged Z: increase by 1
Gamma decay	$^{A}_{Z}\text{X} \rightarrow {}^{0}_{0}\gamma + {}^{A}_{Z}\text{Y}$		A: unchanged Z: unchanged
Positron emission	$^{A}_{Z}\text{X} \rightarrow {}^{0}_{+1}\text{e} + {}^{A}_{Y-1}\text{Y}$		A: unchanged Z: decrease by 1
Electron capture	$^{A}_{Z}\text{X} \rightarrow {}^{0}_{-1}\text{e} + {}^{A}_{Y-1}\text{Y}$		A: unchanged Z: decrease by 1

Figure* 3.1.6*: *This table summarizes the type, nuclear equation, representation, and any changes in the mass or atomic numbers for various types of decay.*

PET Scan

Positron emission tomography (PET) scans use radiation to diagnose and track health conditions and monitor medical treatments by revealing how parts of a patient's body function (Figure **3.1.7**). To perform a PET scan, a positron-emitting radioisotope is produced in a cyclotron and then attached to a substance that is used by the part of the body being investigated. This "tagged" compound, or radiotracer, is then put into the patient (injected via IV or breathed in as a gas), and how it is used by the tissue reveals how that organ or other area of the body functions.

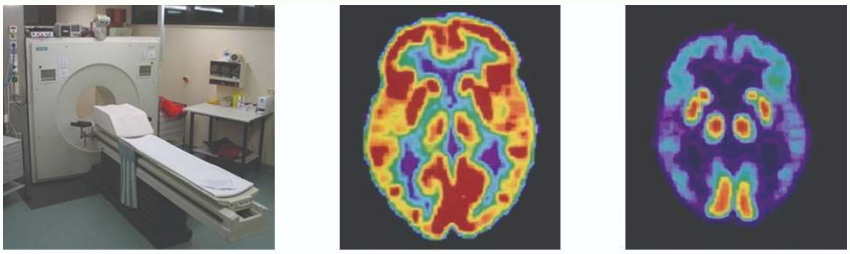


***Figure* 3.1.7:** *A PET scanner (a) uses radiation to provide an image of how part of a patient's body functions. The scans it produces can be used to image a healthy brain (b) or can be used for diagnosing medical conditions such as Alzheimer's disease (c). (credit a: modification of work by Jens Maus)</*

For example, F-18 is produced by proton bombardment of ^{18}O ($^{18}_{8}O + {}^{1}_{1}p \longrightarrow {}^{18}_{9}F + {}^{1}_{0}n$) and incorporated into a glucose analog called fludeoxyglucose (FDG). How FDG is used by the body provides critical diagnostic information; for example, since cancers use glucose differently than normal tissues, FDG can reveal cancers. The ^{18}F emits positrons that interact with nearby electrons, producing a burst of gamma radiation. This energy is detected by the scanner and converted into a detailed, three-dimensional, color image that shows how that part of the patient's body functions. Different levels of gamma radiation produce different amounts of brightness and colors in the image, which can then be interpreted by a radiologist to reveal what is going on. PET scans can detect heart damage and heart disease, help diagnose Alzheimer's disease, indicate the part of a brain that is affected by epilepsy, reveal cancer, show what stage it is, and how much it has spread, and whether treatments are effective. Unlike magnetic resonance imaging and X-rays, which only show how something looks, the big advantage of PET scans is that they show how something functions. PET scans are now usually performed in conjunction with a computed tomography scan.

Radioactive Decay Series

The naturally occurring radioactive isotopes of the heaviest elements fall into chains of successive disintegrations, or decays, and all the species in one chain constitute a radioactive family, or radioactive decay series. Three of these series include most of the naturally radioactive elements of the periodic table. They are the uranium series, the actinide series, and the thorium series. The neptunium series is a fourth series, which is no longer significant on the earth because of the short half-lives of the species involved. Each series is characterized by a parent (first member) that has a long half-life and a series of daughter nuclides that ultimately lead to a stable end-product (Figure **3.1.8**)—that is, a nuclide on the band of stability (Figure 3.1.1). In all three series, the end-product is a stable isotope of lead. The neptunium series, previously thought to terminate with bismuth-209, terminates with thallium-205.

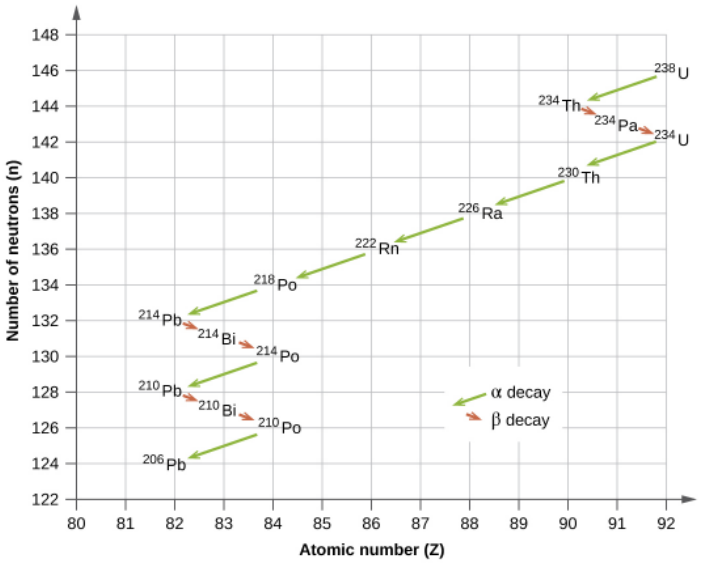


Figure* 3.1.8*: *Uranium-238 undergoes a radioactive decay series consisting of 14 separate steps before producing stable lead-206. This series consists of eight α decays and six β decays.*

Radioactive Half-Lives

Radioactive decay follows first-order kinetics. Since first-order reactions have already been covered in detail in the kinetics chapter, we will now apply those concepts to nuclear decay reactions. Each radioactive nuclide has a characteristic, constant half-life ($t_{1/2}$), the time required for half of the atoms in a sample to decay. An isotope's half-life allows us to determine how long a sample of a useful isotope will be available, and how long a sample of an undesirable or dangerous isotope must be stored before it decays to a low-enough radiation level that is no longer a problem.

For example, cobalt-60, an isotope that emits gamma rays used to treat cancer, has a half-life of 5.27 years (Figure **3.1.6**). In a given cobalt-60 source, since half of the $^{60}_{27}\text{Co}$ nuclei decay every 5.27 years, both the amount of material and the intensity of the radiation emitted is cut in half every 5.27 years. (Note that for a given substance, the intensity of radiation that it produces is directly proportional to the rate of decay of the substance and the amount of the substance.) This is as expected for a process following first-order kinetics. Thus, a cobalt-60 source that is used for cancer treatment must be replaced regularly to continue to be effective.

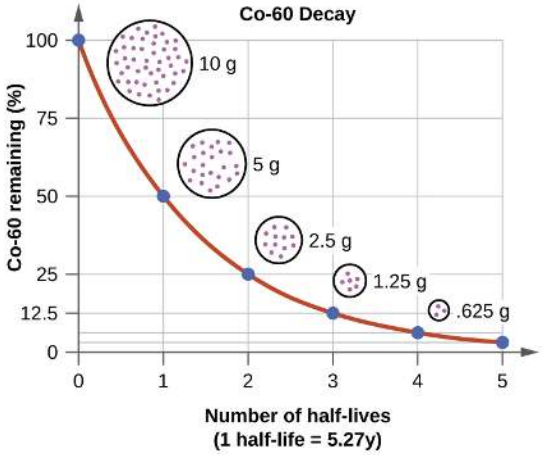


Figure* 3.1.9*: *For cobalt-60, which has a half-life of 5.27 years, 50% remains after 5.27 years (one half-life), 25% remains after 10.54 years (two half-lives), 12.5% remains after 15.81 years (three half-lives), and so on.*

Since nuclear decay follows first-order kinetics, we can adapt the mathematical relationships used for first-order chemical reactions. We generally substitute the number of nuclei, *N*, for the concentration. If the rate is stated in nuclear decays per second, we refer to it as the activity of the radioactive sample. The rate for radioactive decay is:

$$\text{decay rate} = \lambda N \tag{3.1.21}$$

with λ is the **decay constant** for the particular radioisotope.

The decay constant, λ, which is the same as a rate constant discussed in the kinetics chapter. It is possible to express the decay constant in terms of the half-life, $t_{1/2}$:

$$\lambda = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{t_{1/2}} \quad \text{or} \quad t_{1/2} = \frac{\ln 2}{\lambda} = \frac{0.693}{\lambda} \tag{3.1.22}$$

The first-order equations relating amount, N, and time are:

$$N_t = N_0 e^{-kt} \quad \text{or} \quad t = -\frac{1}{\lambda} \ln\left(\frac{N_t}{N_0}\right) \tag{3.1.23}$$

where N_0 is the initial number of nuclei or moles of the isotope, and N_t is the number of nuclei/moles remaining at time t. We will not concern ourselves with the calculation of half-life in this course.

Because each nuclide has a specific number of nucleons, a particular balance of repulsion and attraction, and its own degree of stability, the half-lives of radioactive nuclides vary widely. For example: the half-life of $^{209}_{83}\text{Bi}$ is 1.9×10^{19} years; $^{239}_{94}\text{Ra}$ is 24,000 years; $^{222}_{86}\text{Rn}$ is 3.82 days; and element-111 (Rg for roentgenium) is 1.5×10^{-3} seconds. The half-lives of a number of radioactive isotopes important to medicine are shown in Table **3.1.1**, and others are listed in Appendix N1.

***Table* 3.1.2**: *Half-lives of Radioactive Isotopes Important to Medicine*

Type	Decay Mode	Half-Life	Uses
F-18	β^+ decay	110. minutes	PET scans
Co-60	β decay, γ decay	5.27 years	cancer treatment
Tc-99m[1]	γ decay	8.01 hours	scans of brain, lung, heart, bone
I-131	β decay	8.02 days	thyroid scans and treatment
Tl-201	electron capture	73 hours	heart and arteries scans; cardiac stress tests

The “m” in Tc-99m stands for “metastable,” indicating that this is an unstable, high-energy state of Tc-99. Metastable isotopes emit γ radiation to rid themselves of excess energy and become (more) stable.

Radiometric Dating

Several radioisotopes have half-lives and other properties that make them useful for purposes of “dating” the origin of objects such as archaeological artifacts, formerly living organisms, or geological formations. This process is radiometric dating and has been responsible for many breakthrough scientific discoveries about the geological history of the earth, the evolution of life, and the history of human civilization. We will explore some of the most common types of radioactive dating and how the particular isotopes work for each type.

Radioactive Dating Using Carbon-14

The radioactivity of carbon-14 provides a method for dating objects that were a part of a living organism. This method of radiometric dating, which is also called radiocarbon dating or carbon-14 dating, is accurate for dating carbon-containing substances that are up to about 30,000 years old, and can provide reasonably accurate dates up to a maximum of about 50,000 years old.

Naturally occurring carbon consists of three isotopes: $^{12}_{6}\text{C}$, which constitutes about 99% of the carbon on earth; $^{13}_{6}\text{C}$, about 1% of the total; and trace amounts of $^{14}_{6}\text{C}$. Carbon-14 forms in the upper atmosphere by the reaction of nitrogen atoms with neutrons from cosmic rays in space:

$$^{14}_{7}\text{N} + {}^{1}_{0}\text{n} \longrightarrow {}^{14}_{6}\text{C} + {}^{1}_{1}\text{H} \tag{3.1.24}$$

All isotopes of carbon react with oxygen to produce CO_2 molecules. The ratio of $^{14}_{6}\text{CO}_2$ to $^{12}_{6}\text{CO}_2$ depends on the ratio of $^{14}_{6}\text{CO}$ to $^{12}_{6}\text{CO}$ in the atmosphere. The natural abundance of $^{14}_{6}\text{CO}$ in the atmosphere is approximately 1 part per trillion; until recently, this

has generally been constant over time, as seen is gas samples found trapped in ice. The incorporation of $^{14}_{6}\text{C}$ $^{14}_{6}\text{CO}_2$ and $^{12}_{6}\text{CO}_2$ into plants is a regular part of the photosynthesis process, which means that the $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ ratio found in a living plant is the same as the $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ ratio in the atmosphere. But when the plant dies, it no longer traps carbon through photosynthesis. Because $^{12}_{6}\text{C}$ is a stable isotope and does not undergo radioactive decay, its concentration in the plant does not change. However, carbon-14 decays by β emission with a half-life of 5730 years:

$$^{14}_{6}\text{C} \longrightarrow\ ^{14}_{7}\text{N} + {}^{0}_{-1}\text{e} \tag{3.1.25}$$

Thus, the $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ ratio gradually decreases after the plant dies. The decrease in the ratio with time provides a measure of the time that has elapsed since the death of the plant (or other organism that ate the plant). Figure **3.1.10** visually depicts this process.

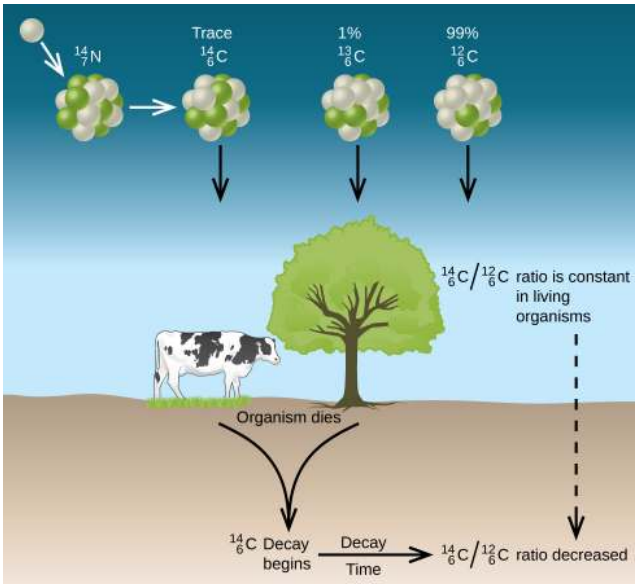


***Figure* 3.1.10:** *Along with stable carbon-12, radioactive carbon-14 is taken in by plants and animals, and remains at a constant level within them while they are alive. After death, the C-14 decays and the C-14:C-12 ratio in the remains decreases. Comparing this ratio to the C-14:C-12 ratio in living organisms allows us to determine how long ago the organism lived (and died).*

For example, with the half-life of $^{14}_{6}\text{C}$ being 5730 years, if the $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ ratio in a wooden object found in an archaeological dig is half what it is in a living tree, this indicates that the wooden object is 5730 years old. Highly accurate determinations of $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ ratios can be obtained from very small samples (as little as a milligram) by the use of a mass spectrometer.

There have been some significant, well-documented changes to the $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ ratio. The accuracy of a straightforward application of this technique depends on the $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ ratio in a living plant being the same now as it was in an earlier era, but this is not always valid. Due to the increasing accumulation of CO_2 molecules (largely $^{12}_{6}\text{CO}_2$) in the atmosphere caused by combustion of fossil fuels (in which essentially all of the $^{14}_{6}\text{C}$ has decayed), the ratio of $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ in the atmosphere may be changing. This manmade increase in $^{12}_{6}\text{CO}_2$ in the atmosphere causes the $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ ratio to decrease, and this in turn affects the ratio in currently living organisms on the earth. Fortunately, however, we can use other data, such as tree dating via examination of annual growth rings, to calculate correction factors. With these correction factors, accurate dates can be determined. In general, radioactive dating only works for about 10 half-lives; therefore, the limit for carbon-14 dating is about 57,000 years.

Radioactive Dating Using Nuclides Other than Carbon-14

Radioactive dating can also use other radioactive nuclides with longer half-lives to date older events. For example, uranium-238 (which decays in a series of steps into lead-206) can be used for establishing the age of rocks (and the approximate age of the oldest rocks on earth). Since U-238 has a half-life of 4.5 billion years, it takes that amount of time for half of the original U-238 to decay into Pb-206. In a sample of rock that does not contain appreciable amounts of Pb-208, the most abundant isotope of lead, we can assume that lead was not present when the rock was formed. Therefore, by measuring and analyzing the ratio of U-238:Pb-206, we can determine the age of the rock. This assumes that all of the lead-206 present came from the decay of uranium-238. If there is additional lead-206 present, which is indicated by the presence of other lead isotopes in the sample, it is necessary to make an

adjustment. Potassium-argon dating uses a similar method. K-40 decays by positron emission and electron capture to form Ar-40 with a half-life of 1.25 billion years. If a rock sample is crushed and the amount of Ar-40 gas that escapes is measured, determination of the Ar-40:K-40 ratio yields the age of the rock. Other methods, such as rubidium-strontium dating (Rb-87 decays into Sr-87 with a half-life of 48.8 billion years), operate on the same principle. To estimate the lower limit for the earth's age, scientists determine the age of various rocks and minerals, making the assumption that the earth is older than the oldest rocks and minerals in its crust. As of 2014, the oldest known rocks on earth are the Jack Hills zircons from Australia, found by uranium-lead dating to be almost 4.4 billion years old.

Summary



***Video* 3.1.6:** *A video summary of some of the nuclear chemistry you just learned about.*

An atomic nucleus consists of protons and neutrons, collectively called nucleons. Although protons repel each other, the nucleus is held tightly together by a short-range, but very strong, force called the strong nuclear force. A nucleus has less mass than the total mass of its constituent nucleons. This "missing" mass is the mass defect, which has been converted into the binding energy that holds the nucleus together according to Einstein's mass-energy equivalence equation, $E = mc^2$. Of the many nuclides that exist, only a small number are stable. Nuclides with even numbers of protons or neutrons, or those with magic numbers of nucleons, are especially likely to be stable. These stable nuclides occupy a narrow band of stability on a graph of number of protons versus number of neutrons. The binding energy per nucleon is largest for the elements with mass numbers near 56; these are the most stable nuclei.

Nuclei can undergo reactions that change their number of protons, number of neutrons, or energy state. Many different particles can be involved in nuclear reactions. The most common are protons, neutrons, positrons (which are positively charged electrons), alpha (α) particles (which are high-energy helium nuclei), beta (β) particles (which are high-energy electrons), and gamma (γ) rays (which compose high-energy electromagnetic radiation). As with chemical reactions, nuclear reactions are always balanced. When a nuclear reaction occurs, the total mass (number) and the total charge remain unchanged.

Nuclei that have unstable n:p ratios undergo spontaneous radioactive decay. The most common types of radioactivity are α decay, β decay, γ emission, positron emission, and electron capture. Nuclear reactions also often involve γ rays, and some nuclei decay by electron capture. Each of these modes of decay leads to the formation of a new nucleus with a more stable n:p ratio. Some substances undergo radioactive decay series, proceeding through multiple decays before ending in a stable isotope. All nuclear decay processes follow first-order kinetics, and each radioisotope has its own characteristic half-life, the time that is required for half of its atoms to decay. Because of the large differences in stability among nuclides, there is a very wide range of half-lives of radioactive substances. Many of these substances have found useful applications in medical diagnosis and treatment, determining the age of archaeological and geological objects, and more.

Key Equations

- $E = mc^2$
- decay rate = λN
- $t_{1/2} = \frac{\ln 2}{\lambda} = \frac{0.693}{\lambda}$

Glossary

alpha (α) decay
loss of an alpha particle during radioactive decay

alpha particle
(α or $^4_2\mathbf{He}$ or $^4_2\boldsymbol{\alpha}$**)** high-energy helium nucleus; a helium atom that has lost two electrons and contains two protons and two neutrons

antimatter
particles with the same mass but opposite properties (such as charge) of ordinary particles

band of stability
(also, belt of stability, zone of stability, or valley of stability) region of graph of number of protons versus number of neutrons containing stable (nonradioactive) nuclides

beta (β) decay
breakdown of a neutron into a proton, which remains in the nucleus, and an electron, which is emitted as a beta particle

beta particle
(β or $^0_{-1}\mathbf{e}$ or $^0_{-1}\boldsymbol{\beta}$**)** high-energy electron

binding energy per nucleon
total binding energy for the nucleus divided by the number of nucleons in the nucleus

daughter nuclide
nuclide produced by the radioactive decay of another nuclide; may be stable or may decay further

electron capture
combination of a core electron with a proton to yield a neutron within the nucleus

electron volt (eV)
measurement unit of nuclear binding energies, with 1 eV equaling the amount energy due to the moving an electron across an electric potential difference of 1 volt

gamma (γ) emission
decay of an excited-state nuclide accompanied by emission of a gamma ray

gamma ray
(γ or $^0_0\boldsymbol{\gamma}$**)** short wavelength, high-energy electromagnetic radiation that exhibits wave-particle duality

half-life ($t_{1/2}$)
time required for half of the atoms in a radioactive sample to decay

magic number
nuclei with specific numbers of nucleons that are within the band of stability

mass defect
difference between the mass of an atom and the summed mass of its constituent subatomic particles (or the mass “lost” when nucleons are brought together to form a nucleus)

mass-energy equivalence equation
Albert Einstein’s relationship showing that mass and energy are equivalent

nuclear binding energy
energy lost when an atom’s nucleons are bound together (or the energy needed to break a nucleus into its constituent protons and neutrons)

nuclear chemistry

study of the structure of atomic nuclei and processes that change nuclear structure

nuclear reaction

change to a nucleus resulting in changes in the atomic number, mass number, or energy state

nucleon

collective term for protons and neutrons in a nucleus

nuclide

nucleus of a particular isotope

parent nuclide

unstable nuclide that changes spontaneously into another (daughter) nuclide

positron ($_{+1}^{0}\beta$ or $_{+1}^{0}e$)

antiparticle to the electron; it has identical properties to an electron, except for having the opposite (positive) charge

positron emission

(also, β^+ decay) conversion of a proton into a neutron, which remains in the nucleus, and a positron, which is emitted

radioactive decay

spontaneous decay of an unstable nuclide into another nuclide

radioactive decay series

chains of successive disintegrations (radioactive decays) that ultimately lead to a stable end-product

radioactivity

phenomenon exhibited by an unstable nucleon that spontaneously undergoes change into a nucleon that is more stable; an unstable nucleon is said to be radioactive

radiocarbon dating

highly accurate means of dating objects 30,000–50,000 years old that were derived from once-living matter; achieved by calculating the ratio of $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ in the object vs. the ratio of $^{14}_{6}\text{C}:\ ^{12}_{6}\text{C}$ in the present-day atmosphere

radioisotope

isotope that is unstable and undergoes conversion into a different, more stable isotope

radiometric dating

use of radioisotopes and their properties to date the formation of objects such as archeological artifacts, formerly living organisms, or geological formations

strong nuclear force

force of attraction between nucleons that holds a nucleus together

Contributors

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3.1.1: Practice Problems- Nuclear Chemistry and Radioactive Decay (Optional)

PROBLEM 3.1.1.1

Write the following isotopes in nuclide notation (e.g., " $^{14}_{6}C$ ")

a. oxygen-14
b. copper-70
c. tantalum-175
d. francium-217

Answer a

$^{14}_{8}O$

Answer b

$^{70}_{29}Cu$

Answer c

$^{175}_{73}Ta$

Answer d

$^{217}_{87}Fr$

Click here for a video of the solution



PROBLEM 3.1.1.2

For the following isotopes that have missing information, fill in the missing information to complete the notation

a. $^{34}_{14}X$
b. $_{P}{}^{36}$
c. $^{57}_{X}Mn$
d. $^{121}_{56}X$

Answer a

$^{34}_{14}Si$

Answer b

$^{36}_{15}P$

Answer c

Answer c

$^{57}_{25}\mathrm{Mn}$

Answer d

$^{121}_{56}\mathrm{Ba}$

PROBLEM 3.1.1.3

Write the nuclide notation, including charge if applicable, for atoms with the following characteristics:

a. 25 protons, 20 neutrons, 24 electrons
b. 45 protons, 24 neutrons, 43 electrons
c. 53 protons, 89 neutrons, 54 electrons
d. 97 protons, 146 neutrons, 97 electrons

Answer a

$^{45}_{25}\mathrm{Mn}^{+1}$

Answer b

$^{69}_{45}\mathrm{Rh}^{+2}$

Answer c

$^{142}_{53}\mathrm{I}^{-1}$

Answer d

$^{243}_{97}\mathrm{Bk}$

Click here for a video of the solution



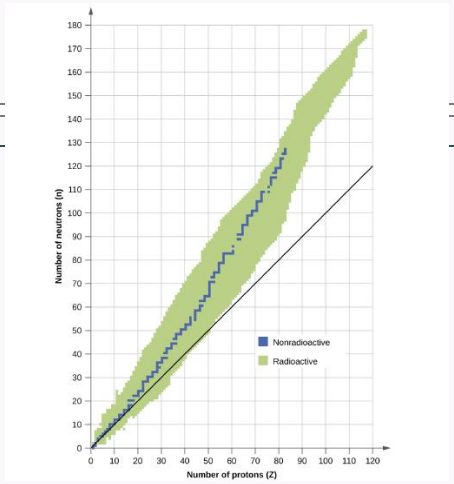
PROBLEM 3.1.1.4

Which of the following nuclei lie within the band of stability?

a. chlorine-37
b. calcium-40
c. $^{204}\mathrm{Bi}$
d. $^{56}\mathrm{Fe}$
e. $^{206}\mathrm{Pb}$
f. $^{211}\mathrm{Pb}$
g. $^{222}\mathrm{Rn}$
h. carbon-14

Answer

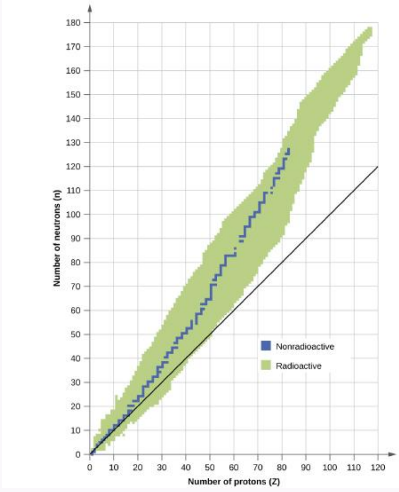
(a), (b), (c), (d), and (e)



PROBLEM 3.1.1.5

Which of the following nuclei lie within the band of stability?

a. argon-40
b. oxygen-16
c. ^{122}Ba
d. ^{58}Ni
e. ^{205}Tl
f. ^{210}Tl
g. ^{226}Ra
h. magnesium-24



Answer

(b), (e - very close), and (h)

Click here for a video of the solution



PROBLEM 3.1.1.6

Write a brief description or definition of each of the following:

a. nucleon
b. α particle
c. β particle
d. positron
e. γ ray
f. nuclide
g. mass number
h. atomic number

Answer a

collective term for protons and neutrons in a nucleus

Answer b

(α or ^4_2He or $^4_2\alpha$) high-energy helium nucleus; a helium atom that has lost two electrons and contains two protons and two neutrons

Answer c

(β or ${}_{-1}^{0}\mathrm{e}$ or ${}_{-1}^{0}\beta$) high-energy electron

Answer d

antiparticle to the electron; it has identical properties to an electron, except for having the opposite (positive) charge

Answer e

(γ or ${}_{0}^{0}\gamma$) short wavelength, high-energy electromagnetic radiation that exhibits wave-particle duality

Answer f

nucleus of a particular isotope

Answer g

sum of the numbers of neutrons and protons in the nucleus of an atom

Answer h

number of protons in the nucleus of an atom

PROBLEM 3.1.1.7

Complete each of the following equations by adding the missing species:

a. ${}_{13}^{27}\mathrm{Al} + {}_{2}^{4}\mathrm{He} \longrightarrow \ ? + {}_{0}^{1}\mathrm{n}$
b. ${}_{94}^{239}\mathrm{Pu} + \ ? \longrightarrow {}_{96}^{242}\mathrm{Cm} + {}_{0}^{1}\mathrm{n}$
c. ${}_{7}^{14}\mathrm{N} + {}_{2}^{4}\mathrm{He} \longrightarrow \ ? + {}_{1}^{1}\mathrm{H}$
d. ${}_{92}^{235}\mathrm{U} \longrightarrow \ ? + {}_{55}^{135}\mathrm{Cs} + 4{}_{0}^{1}\mathrm{n}$

Answer a

${}_{13}^{27}\mathrm{Al} + {}_{2}^{4}\mathrm{He} \longrightarrow {}_{15}^{30}\mathrm{P} + {}_{0}^{1}\mathrm{n}$

Answer b

${}_{94}^{239}\mathrm{Pu} + {}_{2}^{4}\mathrm{He} \longrightarrow {}_{96}^{242}\mathrm{Cm} + {}_{0}^{1}\mathrm{n}$

Answer c

${}_{7}^{14}\mathrm{N} + {}_{2}^{4}\mathrm{He} \longrightarrow {}_{8}^{17}\mathrm{O} + {}_{1}^{1}\mathrm{H}$

Answer d

${}_{92}^{235}\mathrm{U} \longrightarrow {}_{37}^{96}\mathrm{Rb} + {}_{55}^{135}\mathrm{Cs} + 4{}_{0}^{1}\mathrm{n}$

Click here for a video of the solution



PROBLEM 3.1.1.8

Complete each of the following equations:

a. $^{7}_{3}\text{Li} + \ ? \longrightarrow 2^{4}_{2}\text{He}$
b. $^{14}_{6}\text{C} \longrightarrow {}^{14}_{7}\text{N} + \ ?$
c. $^{27}_{13}\text{Al} + {}^{4}_{2}\text{He} \longrightarrow \ ? + {}^{1}_{0}\text{n}$
d. $^{250}_{96}\text{Cm} \longrightarrow \ ? + {}^{98}_{38}\text{Sr} + 4^{1}_{0}\text{n}$

Answer a

$$^{7}_{3}\text{Li} + {}^{1}_{1}\text{H} \longrightarrow 2^{4}_{2}\text{He}$$

Answer b

$$^{14}_{6}\text{C} \longrightarrow {}^{14}_{7}\text{N} + {}^{0}_{-1}\text{e}$$

Answer c

$$^{27}_{13}\text{Al} + {}^{4}_{2}\text{He} \longrightarrow {}^{30}_{15}\text{P} + {}^{1}_{0}\text{n}$$

Answer d

$$^{250}_{96}\text{Cm} \longrightarrow {}^{148}_{58}\text{Ce} + {}^{98}_{38}\text{Sr} + 4^{1}_{0}\text{n}$$

PROBLEM 3.1.1.9

Write a balanced equation for each of the following nuclear reactions:

a. the production of ^{17}O from ^{14}N by α particle bombardment
b. the production of ^{14}C from ^{14}N by neutron bombardment
c. the production of ^{233}Th from ^{232}Th by neutron bombardment
d. the production of ^{239}U from ^{238}U by $^{2}_{1}\text{H}$ bombardment

Answer a

$$^{14}_{7}\text{N} + \text{He}^{2} \longrightarrow {}^{17}_{8}\text{O} + {}^{1}_{1}\text{H}$$

Answer b

$$^{14}_{7}\text{N} + {}^{1}_{0}\text{n} \longrightarrow {}^{14}_{6}\text{N} + {}^{1}_{1}\text{H}$$

Answer c

$$^{232}_{90}\text{Th} + {}^{1}_{0}\text{n} \longrightarrow {}^{233}_{90}\text{Th}$$

Answer d

$$^{238}_{92}\text{U} + {}^{2}_{1}\text{H} \longrightarrow {}^{239}_{92}\text{U} + {}^{1}_{1}\text{H}$$

Click here for a video of the solution



PROBLEM 3.1.1.10

Technetium-99 is prepared f[redacted]n-99, an unstable isotope that emits a β particle to yiel[redacted]eus relaxes to the ground state, represented as ^{99}Tc, by emitting a γ ray. The ground state of ^{99}Tc then emits a β particle. Write the equations for each of these nuclear reactions.

Answer

$$^{98}_{42}\text{Mo} + ^{1}_{0}\text{n} \rightarrow ^{99}_{42}\text{Mo}$$

$$^{99}_{42}\text{Mo} \rightarrow ^{0}_{-1}e + ^{99}_{43}Tc^{*}$$

$$^{99}_{43}\text{Tc}^{*} \rightarrow ^{0}_{0}\gamma + ^{99}_{43}Tc$$

$$^{99}_{43}\text{Tc} \rightarrow ^{0}_{-1}\text{e} + ^{99}_{44}\text{Ru}$$

PROBLEM 3.1.1.11

What changes occur to the atomic number and mass of a nucleus during each of the following decay scenarios?

a. an α particle is emitted
b. a β particle is emitted
c. γ radiation is emitted
d. a positron is emitted
e. an electron is captured

Answer a

Since an α particle is the same as a $^{4}_{2}\text{He}$ nucleus, the mass number will decrease by 4 and the atomic number will decrease by 2.

Answer b

Since a β particle is the same as $^{0}_{-1}\text{e}$, the mass number will not change but the atomic number will ***increase*** by 1.

Answer c

Since a γ ray has no mass (it is energy) the mass number and atomic number do not change.

Answer d

A positron is the opposite of a β particle, it is $^{0}_{+1}\text{e}$, the mass number will not change but the atomic number will decrease by 1

Answer e

Electron capture has the same effect on the nucleus as positron emission: The atomic number is decreased by one and the mass number does not change.

PROBLEM 3.1.1.12

What is the change in the nucleus that results from the following decay scenarios?

a. emission of a β particle
b. emission of a β^{+} particle
c. capture of an electron

Answer a

conversion of a neutron to a proton: ${}^{1}_{0}\text{n} \longrightarrow {}^{1}_{1}\text{p} + {}^{0}_{+1}\text{e}$

Answer b

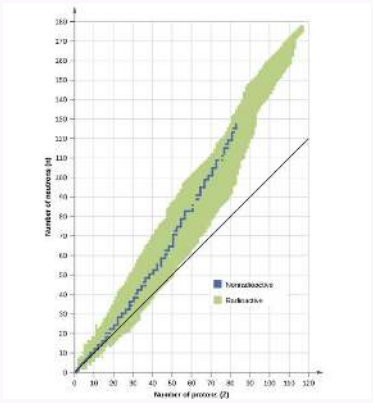
conversion of a proton to a neutron; the positron has the same mass as an electron and the same magnitude of positive charge as the electron has negative charge

when the n:p ratio of a nucleus is too low, a proton is converted into a neutron with the emission of a positron: ${}^{1}_{1}\text{p} \longrightarrow {}^{1}_{0}\text{n} + {}^{0}_{+1}\text{e}$

Answer c

In a proton-rich nucleus, an inner atomic electron can be absorbed. In simplest form, this changes a proton into a neutron: ${}^{1}_{1}\text{p} + {}^{0}_{-1}\text{e} \longrightarrow {}^{1}_{0}\text{p}$

PROBLEM 3.1.1.13



Explain how unstable heavy nuclides (atomic number > 83) may decompose to form nuclides of greater stability if

(a) they are below the band of stability and

(b) they are above the band of stability

Answer

Nuclei below the band of stability will undergo positron decay, while those above the band of stability will undergo beta decay. Heavy nuclei past the band of stability will undergo alpha decay

PROBLEM 3.1.1.14

Which of the following nuclei is most likely to decay by positron emission? Explain your choice.

a. chromium-53
b. manganese-51
c. iron-59

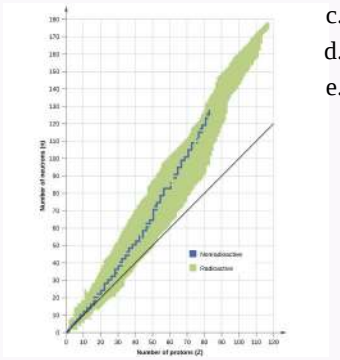
Answer

Manganese-51 is most likely to decay by positron emission. The n:p ratio for Cr-53 is $\frac{29}{24}$ = 1.21; for Mn-51, it is $\frac{26}{25}$ = 1.04; for Fe-59, it is $\frac{33}{26}$ = 1.27. Positron decay occurs when the n:p ratio is low. Mn-51 has the lowest n:p ratio and therefore is most likely to decay by positron emission. Besides, ${}^{53}_{24}\text{Cr}$ is a stable isotope, and ${}^{59}_{26}\text{Fe}$ decays by beta emission.

PROBLEM 3.1.1.15

The following nuclei do not lie in the band of stability. How would they be expected to decay? Explain your answer.

a. ${}^{34}_{15}\text{P}$
b. ${}^{239}_{92}\text{U}$



c. $^{38}_{20}\text{Ca}$
d. $^{3}_{1}\text{H}$
e. $^{245}_{94}\text{Pu}$

Answer a

Above the band of stability, beta decay is expected

Answer b

Beyond the band of stability, heavy nuclei undergo alpha decay

Answer c

Below the band of stability, positron decay is expected

Answer d

Above the band of stability, beta decay is expected

Answer e

Beyond the band of stability, heavy nuclei undergo alpha decay

PROBLEM 3.1.1.16

Write a nuclear reaction for each step in the formation of $^{218}_{84}\text{Po}$ from $^{238}_{92}\text{U}$, which proceeds by a series of decay reactions involving the step-wise emission of α, β, β, α, α, α, α particles, in that order.

Answer

$$^{238}_{92}\text{U} \longrightarrow {}^{234}_{90}\text{Th} + {}^{4}_{2}\text{He}$$

$$^{234}_{90}\text{Th} \longrightarrow {}^{234}_{91}\text{Pa} + {}^{0}_{-1}\text{e}$$

$$^{234}_{91}\text{Pa} \longrightarrow {}^{234}_{92}\text{U} + {}^{0}_{-1}\text{e}$$

$$^{234}_{92}\text{U} \longrightarrow {}^{230}_{90}\text{Th} + {}^{4}_{2}\text{He}$$

$$^{230}_{90}\text{Th} \longrightarrow {}^{226}_{88}\text{Ra} + {}^{4}_{2}\text{He}$$

$$^{226}_{88}\text{Ra} \longrightarrow {}^{222}_{86}\text{Rn} + {}^{4}_{2}\text{He}$$

$$^{222}_{86}\text{Rn} \longrightarrow {}^{218}_{84}\text{Po} + {}^{4}_{2}\text{He}$$

PROBLEM 3.1.1.17

Write a nuclear reaction for each step in the formation of $^{208}_{82}\text{Pb}$ from $^{228}_{90}\text{Th}$, which proceeds by a series of decay reactions involving the step-wise emission of α, α, α, α, β, β, α particles, in that order.

Answer

$$^{228}_{90}\text{Th} \rightarrow {}^{4}_{2}\text{He} + {}^{224}_{88}\text{Ra}$$

$$^{224}_{88}\text{Ra} \rightarrow {}^{4}_{2}\text{He} + {}^{220}_{86}\text{Rn}$$

$$^{220}_{86}\text{Th} \rightarrow {}^{4}_{2}\text{He} + {}^{216}_{84}\text{Po}$$

$$^{216}_{84}\text{Po} \rightarrow {}^{4}_{2}\text{He} + {}^{212}_{82}\text{Pb}$$

$^{212}_{82}\text{Pb} \rightarrow {}^{0}_{-1}\text{e} + {}^{212}_{83}\text{Bi}$

$^{212}_{83}\text{Bi} \rightarrow {}^{0}_{-1}\text{e} + {}^{212}_{84}\text{Po}$

$^{212}_{84}\text{Po} \rightarrow {}^{4}_{2}\text{He} + {}^{208}_{82}\text{Pb}$

Click here to see a video of the solution



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3.2: Organization of the Periodic Table

Skills to Develop

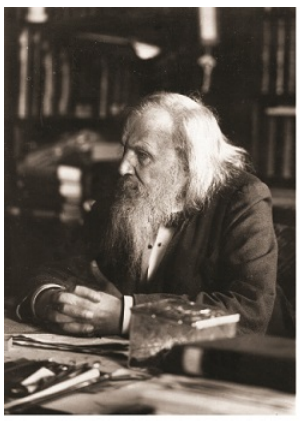
- State the periodic law and explain the organization of elements in the periodic table
- Predict the general properties of elements based on their location within the periodic table
- Identify metals, nonmetals, and metalloids by their properties and/or location on the periodic table



Video* 3.2.1: A brief introduction to the genius of Mendeleev.*

As early chemists worked to purify ores and discovered more elements, they realized that various elements could be grouped together by their similar chemical behaviors. One such grouping includes lithium (Li), sodium (Na), and potassium (K): These elements all are shiny, conduct heat and electricity well, and have similar chemical properties. A second grouping includes calcium (Ca), strontium (Sr), and barium (Ba), which also are shiny, good conductors of heat and electricity, and have chemical properties in common. However, the specific properties of these two groupings are notably different from each other. For example: Li, Na, and K are much more reactive than are Ca, Sr, and Ba; Li, Na, and K form compounds with oxygen in a ratio of two of their atoms to one oxygen atom, whereas Ca, Sr, and Ba form compounds with one of their atoms to one oxygen atom. Fluorine (F), chlorine (Cl), bromine (Br), and iodine (I) also exhibit similar properties to each other, but these properties are drastically different from those of any of the elements above.

Dimitri Mendeleev in Russia (1869) and Lothar Meyer in Germany (1870) independently recognized that there was a periodic relationship among the properties of the elements known at that time. Both published tables with the elements arranged according to increasing atomic mass. But Mendeleev went one step further than Meyer: He used his table to predict the existence of elements that would have the properties similar to aluminum and silicon, but were yet unknown. The discoveries of gallium (1875) and germanium (1886) provided great support for Mendeleev's work. Although Mendeleev and Meyer had a long dispute over priority, Mendeleev's contributions to the development of the periodic table are now more widely recognized (Figure **3.2.1**).



(a)

Reihen	Gruppe I. — R^1O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

(b)

***Figure* 3.2.1**: *(a) Dimitri Mendeleev is widely credited with creating (b) the first periodic table of the elements. (credit a: modification of work by Serge Lachinov; credit b: modification of work by "Den fjättrade ankan"/Wikimedia Commons)*



***Video* 3.2.2**: *An introduction to the organization of the periodic table into periods and groups. This organization will be important as we continue building on the principles of chemistry.*

By the twentieth century, it became apparent that the periodic relationship involved atomic numbers rather than atomic masses. The modern statement of this relationship, the periodic law, is as follows: *the properties of the elements are periodic functions of their atomic numbers*. A modern periodic table arranges the elements in increasing order of their atomic numbers and groups atoms with similar properties in the same vertical column (Figure **3.2.2**). Each box represents an element and contains its atomic number, symbol, average atomic mass, and (sometimes) name. The elements are arranged in seven horizontal rows, called periods or series, and 18 vertical columns, called groups. Groups are labeled at the top of each column. In the United States, the labels traditionally were numerals with capital letters. However, IUPAC recommends that the numbers 1 through 18 be used, and these labels are more common. For the table to fit on a single page, parts of two of the rows, a total of 14 columns, are usually written below the main body of the table.

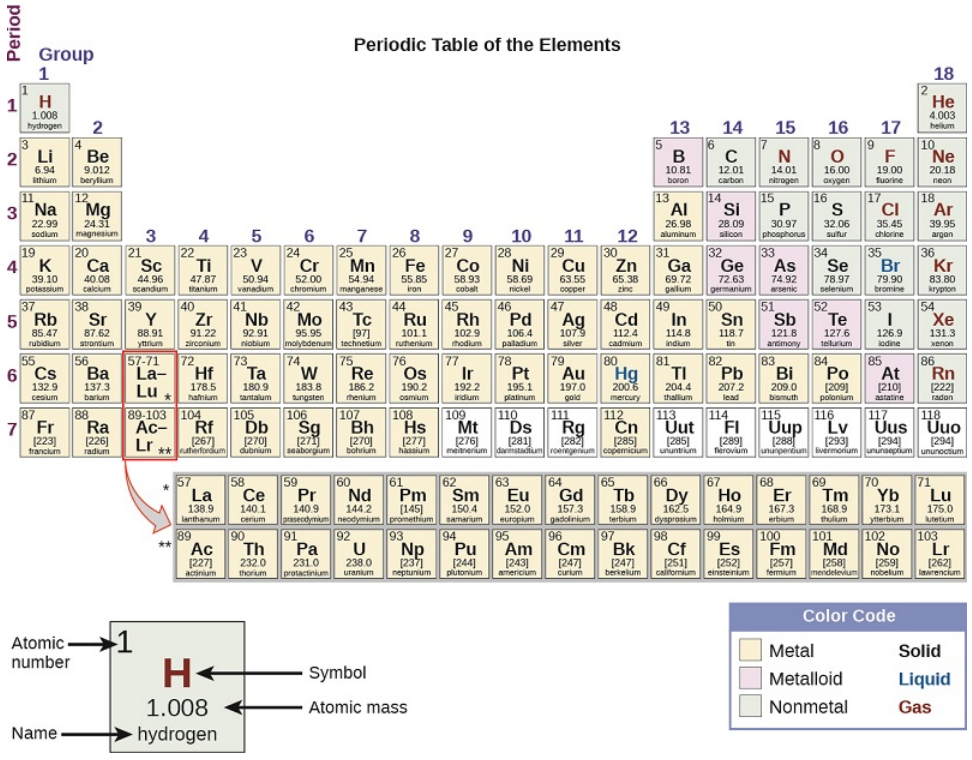


***Figure* 3.2.2**: *Elements in the periodic table are organized according to their properties.*

Many elements differ dramatically in their chemical and physical properties, but some elements are similar in their behaviors. For example, many elements appear shiny, are malleable (able to be deformed without breaking) and ductile (can be drawn into wires), and conduct heat and electricity well. Other elements are not shiny, malleable, or ductile, and are poor conductors of heat and electricity. We can sort the elements into large classes with common properties: metals (elements that are shiny, malleable, good conductors of heat and electricity—shaded yellow); nonmetals (elements that appear dull, poor conductors of heat and electricity—shaded green); and metalloids (elements that conduct heat and electricity moderately well, and possess some properties of metals and some properties of nonmetals—shaded purple).

The elements can also be classified into the main-group elements (or representative elements) in the columns labeled 1, 2, and 13–18; the transition metals in the columns labeled 3–12; and inner transition metals in the two rows at the bottom of the table (the top-row elements are called lanthanides and the bottom-row elements are actinides; Figure **3.2.3**). The elements can be subdivided further by more specific properties, such as the composition of the compounds they form. For example, the elements in group 1 (the first column) form compounds that consist of one atom of the element and one atom of hydrogen. These elements (except hydrogen) are known as alkali metals, and they all have similar chemical properties. The elements in group 2 (the second column) form compounds consisting of one atom of the element and two atoms of hydrogen: These are called alkaline earth metals, with similar properties among members of that group. Other groups with specific names are the pnictogens (group 15), chalcogens (group 16), halogens (group 17), and the noble gases (group 18, also known as inert gases). The groups can also be referred to by the first element of the group: For example, the chalcogens can be called the oxygen group or oxygen family. Hydrogen is a unique, nonmetallic element with properties similar to both group 1 and group 17 elements. For that reason, hydrogen may be shown at the top of both groups, or by itself.

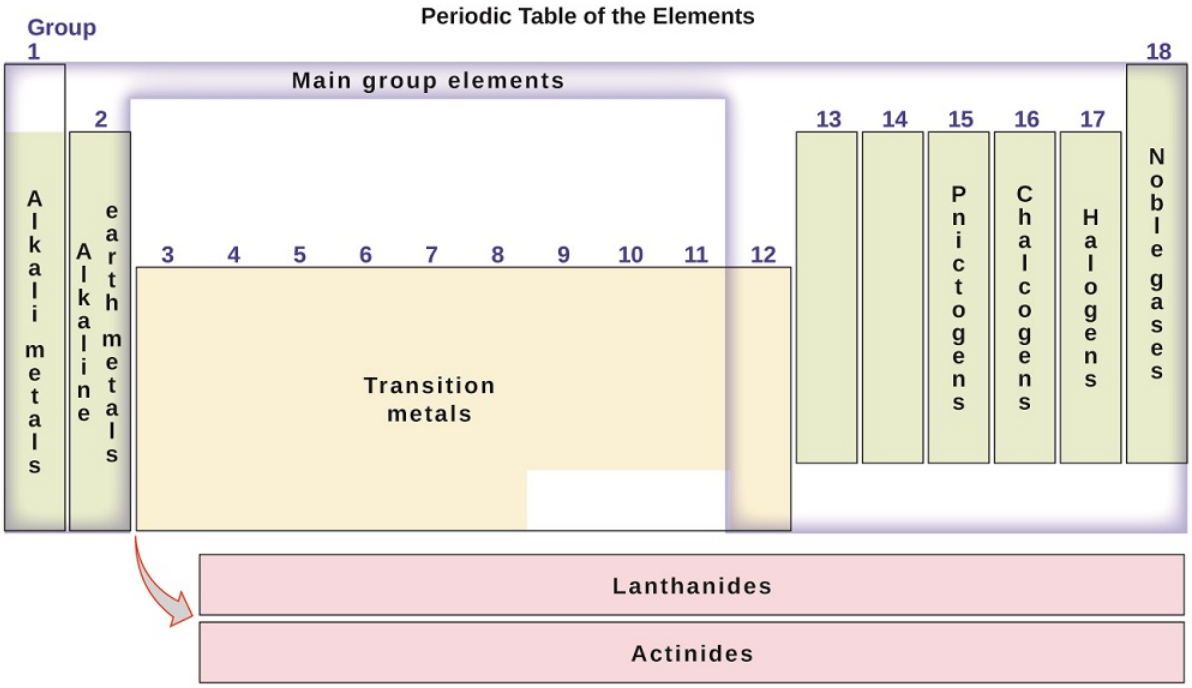


***Figure* 3.2.3:** *The periodic table organizes elements with similar properties into groups.*

Example 3.2.1: Naming Groups of Elements

Atoms of each of the following elements are essential for life. Give the group name for the following elements:

a. chlorine
b. calcium
c. sodium
d. sulfur

Solution

The family names are as follows:

a. halogen
b. alkaline earth metal
c. alkali metal
d. chalcogen

Exercise 3.2.1

Give the group name for each of the following elements:

a. krypton
b. selenium
c. barium
d. lithium

Answer a

noble gas

Answer b

chalcogen

Answer c

alkaline earth metal

Answer d

alkali metal

In studying the periodic table, you might have noticed something about the atomic masses of some of the elements. Element 43 (technetium), element 61 (promethium), and most of the elements with atomic number 84 (polonium) and higher have their atomic mass given in square brackets. This is done for elements that consist entirely of unstable, radioactive isotopes (you will learn more about radioactivity in the nuclear chemistry chapter). An average atomic weight cannot be determined for these elements because their radioisotopes may vary significantly in relative abundance, depending on the source, or may not even exist in nature. The number in square brackets is the atomic mass number (and approximate atomic mass) of the most stable isotope of that element.

Summary



Video* 3.2.3: A summary of the discovery and properties of the periodic table.*

The discovery of the periodic recurrence of similar properties among the elements led to the formulation of the periodic table, in which the elements are arranged in order of increasing atomic number in rows known as periods and columns known as groups. Elements in the same group of the periodic table have similar chemical properties. Elements can be classified as metals, metalloids, and nonmetals, or as a main-group elements, transition metals, and inner transition metals. Groups are numbered 1–18 from left to

right. The elements in group 1 are known as the alkali metals; those in group 2 are the alkaline earth metals; those in 15 are the pnictogens; those in 16 are the chalcogens; those in 17 are the halogens; and those in 18 are the noble gases.

Glossary

actinide
inner transition metal in the bottom of the bottom two rows of the periodic table

alkali metal
element in group 1

alkaline earth metal
element in group 2

chalcogen
element in group 16

group
vertical column of the periodic table

halogen
element in group 17

inert gas
(also, noble gas) element in group 18

inner transition metal
(also, lanthanide or actinide) element in the bottom two rows; if in the first row, also called lanthanide, or if in the second row, also called actinide

lanthanide
inner transition metal in the top of the bottom two rows of the periodic table

main-group element
(also, representative element) element in columns 1, 2, and 12–18

metal
element that is shiny, malleable, good conductor of heat and electricity

metalloid
element that conducts heat and electricity moderately well, and possesses some properties of metals and some properties of nonmetals

noble gas
(also, inert gas) element in group 18

nonmetal
element that appears dull, poor conductor of heat and electricity

period
(also, series) horizontal row of the periodic table

periodic law
properties of the elements are periodic function of their atomic numbers.

periodic table

table of the elements that places elements with similar chemical properties close together

pnictogen

element in group 15

representative element

(also, main-group element) element in columns 1, 2, and 12–18

transition metal

element in columns 3–11

series

(also, period) horizontal row of the period table

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3.2.1: Practice Problems - Organization of the Periodic Table (Optional)

PROBLEM 3.2.1.1

Using the periodic table, classify each of the following elements as a metal or a nonmetal, and then further classify each as a main-group element, transition metal, or inner transition metal:

a. uranium
b. bromine
c. strontium
d. neon
e. gold
f. americium
g. rhodium
h. sulfur
i. carbon
j. potassium

Answer a

metal; inner transition metal

Answer b

nonmetal, main group element

Answer c

metal, main group element

Answer d

nonmetal, main group element

Answer e

metal, transition metal

Answer f

metal, inner transition metal

Answer g

metal, transition metal

Answer h

nonmetal, main group element

Answer i

nonmetal, main group element

Answer j

metal, main group element

PROBLEM 3.2.1.2

Using the periodic table, classify each of the following elements as a metal or a nonmetal, and then further classify each as a main-group (representative) element, transition metal, or inner transition metal:

a. cobalt
b. europium
c. iodine
d. indium
e. lithium
f. oxygen
g. cadmium
h. terbium
i. rhenium

Answer a

metal, transition metal

Answer b

metal, inner transition metal

Answer c

nonmetal, main group element

Answer d

metal, main group element

Answer e

metal, main group element

Answer f

nonmetal, main group element

Answer g

metal, transition metal

Answer h

metal, inner transition metal

Answer i

metal, transition metal

PROBLEM 3.2.1.3

Using the periodic table, identify the lightest member of each of the following groups:

a. noble gases
b. alkaline earth metals
c. alkali metals
d. chalcogens

Answer a

He

Answer b

Be

Answer c

Li

Answer d

O

PROBLEM 3.2.1.4

Using the periodic table, identify the heaviest member of each of the following groups:

a. alkali metals
b. chalcogens
c. noble gases
d. alkaline earth metals

Answer a

Fr

Answer b

Po

Answer c

Rn

Answer d

Ra

PROBLEM 3.2.1.5

Use the periodic table to give the name and symbol for each of the following elements:

(a) the noble gas in the same period as germanium
(b) the alkaline earth metal in the same period as selenium
(c) the halogen in the same period as lithium
(d) the chalcogen in the same period as cadmium

Answer a

Kr; Krypton

Answer b

Ca; Calcium

Answer c

F; Fluorine

Answer d

Te; Tellurium

PROBLEM 3.2.1.6

Use the periodic table to give the name and symbol for each of the following elements:

(a) the halogen in the same period as the alkali metal with 11 protons
(b) the alkaline earth metal in the same period with the neutral noble gas with 18 electrons

(c) the noble gas in the same row as an isotope with 30 neutrons and 25 protons
(d) the noble gas in the same period as gold

Answer a

Cl; Chlorine

Answer b

Mg; Magnesium

Answer c

Kr; Krypton

Answer d

Rn; Radon

PROBLEM 3.2.1.7

Write a symbol for each of the following neutral isotopes. Include the atomic number and mass number for each.

(a) the alkali metal with 11 protons and a mass number of 23
(b) the noble gas element with and 75 neutrons in its nucleus and 54 electrons in the neutral atom
(c) the isotope with 33 protons and 40 neutrons in its nucleus
(d) the alkaline earth metal with 88 electrons and 138 neutrons

Answer a

$^{23}_{11}\mathrm{Na}$

Answer b

$^{129}_{54}\mathrm{Xe}$

Answer c

$^{73}_{33}\mathrm{As}$

Answer d

$^{226}_{88}\mathrm{Ra}$

PROBLEM 3.2.1.8

Write a symbol for each of the following neutral isotopes. Include the atomic number and mass number for each.

a. the chalcogen with a mass number of 125
b. the halogen whose longest-lived isotope is radioactive
c. the noble gas, used in lighting, with 10 electrons and 10 neutrons
d. the lightest alkali metal with three neutrons

Answer a

$^{125}_{52}\mathrm{Te}$

Answer b

$^{210}_{85}\mathrm{At}$

Answer c

$^{20}_{10}\mathrm{Ne}$

Answer d

$^{6}_{3}Li$

Contributors

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3.3: Predicting Charges of Ions

Skills to Develop

- Define ions, cations and anions
- Predict the charge of an anion or cation based upon the location of the element in the periodic table

Ions of Elements

In ordinary chemical reactions, the nucleus of each atom remains unchanged. Electrons, however, can be added to atoms by transfer form other atoms, lost by transfer to other atoms, or shared with other atoms. The transfer and sharing of electrons among atoms govern the chemistry of the elements. During the formation of some compounds, atoms gain or lose electrons, and form electrically charged particles called ions (Figure **3.3.1**; Video **3.3.1**)

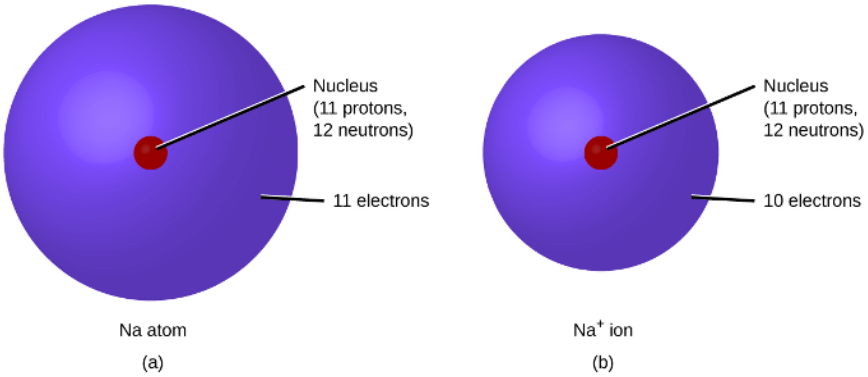


Figure* 3.3.1*: *(a) A sodium atom (Na) has equal numbers of protons and electrons (11) and is uncharged. (b) A sodium cation (Na^+) has lost an electron, so it has one more proton (11) than electrons (10), giving it an overall positive charge, signified by a superscripted plus sign.*

One can use the periodic table to predict whether an atom will form an anion or a cation, and you can often predict the charge of the resulting ion. Atoms of many main-group metals lose enough electrons to leave them with the same number of electrons as an atom of the preceding noble gas. To illustrate, an atom of an alkali metal (group 1) loses one electron and forms a cation with a 1+ charge; an alkaline earth metal (group 2) loses two electrons and forms a cation with a 2+ charge, and so on. For example, a neutral calcium atom, with 20 protons and 20 electrons, readily loses two electrons. This results in a cation with 20 protons, 18 electrons, and a 2+ charge. It has the same number of electrons as atoms of the preceding noble gas, argon, and is symbolized Ca^{2+}. The name of a metal ion is the same as the name of the metal atom from which it forms, so Ca^{2+} is called a calcium ion.

When atoms of nonmetal elements form ions, they generally gain enough electrons to give them the same number of electrons as an atom of the next noble gas in the periodic table. Atoms of group 17 gain one electron and form anions with a 1− charge; atoms of group 16 gain two electrons and form ions with a 2− charge, and so on. For example, the neutral bromine atom, with 35 protons and 35 electrons, can gain one electron to provide it with 36 electrons. This results in an anion with 35 protons, 36 electrons, and a 1− charge. It has the same number of electrons as atoms of the next noble gas, krypton, and is symbolized Br^-. (A discussion of the theory supporting the favored status of noble gas electron numbers reflected in these predictive rules for ion formation is provided in a later chapter of this text.)

Note the usefulness of the periodic table in predicting likely ion formation and charge (Figure **3.3.2**). Moving from the far left to the right on the periodic table, main-group elements tend to form cations with a charge equal to the group number. That is, group 1 elements form 1+ ions; group 2 elements form 2+ ions, and so on. Moving from the far right to the left on the periodic table, elements often form anions with a negative charge equal to the number of groups moved left from the noble gases. For example, group 17 elements (one group left of the noble gases) form 1− ions; group 16 elements (two groups left) form 2− ions, and so on. This trend can be used as a guide in many cases, but its predictive value decreases when moving toward the center of the periodic

table. In fact, transition metals and some other metals often exhibit variable charges that are not predictable by their location in the table. For example, copper can form ions with a 1+ or 2+ charge, and iron can form ions with a 2+ or 3+ charge.

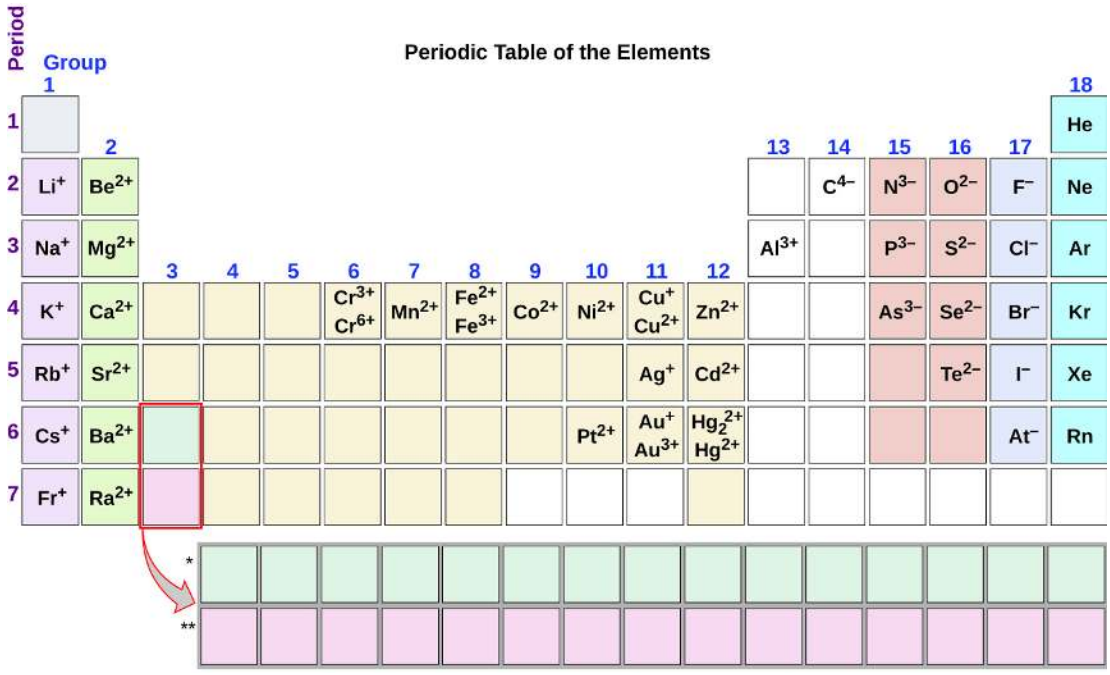


***Figure* 3.3.2:** *Some elements exhibit a regular pattern of ionic charge when they form ions.*

Example 3.3.1: Composition of Ions

An ion found in some compounds used as antiperspirants contains 13 protons and 10 electrons. What is its symbol?

Solution

Because the number of protons remains unchanged when an atom forms an ion, the atomic number of the element must be 13. Knowing this lets us use the periodic table to identify the element as Al (aluminum). The Al atom has lost three electrons and thus has three more positive charges (13) than it has electrons (10). This is the aluminum cation, Al^{3+}.

Exercise 3.3.1

Give the symbol and name for the ion with 34 protons and 36 electrons.

Answer

Se^{2-}, the selenide ion

Example 3.3.2: Formation of Ions

Magnesium and nitrogen react to form an ionic compound. Predict which forms an anion, which forms a cation, and the charges of each ion. Write the symbol for each ion and name them.

Solution

Magnesium's position in the periodic table (group 2) tells us that it is a metal. Metals form positive ions (cations). A magnesium atom must lose two electrons to have the same number electrons as an atom of the previous noble gas, neon. Thus, a magnesium atom will form a cation with two fewer electrons than protons and a charge of 2+. The symbol for the ion is Mg^{2+}, and it is called a magnesium ion.

Nitrogen's position in the periodic table (group 15) reveals that it is a nonmetal. Nonmetals form negative ions (anions). A nitrogen atom must gain three electrons to have the same number of electrons as an atom of the following noble gas, neon. Thus,

a nitrogen atom will form an anion with three more electrons than protons and a charge of 3−. The symbol for the ion is N^{3-}, and it is called a nitride ion.

Exercise 3.3.2

Aluminum and carbon react to form an ionic compound. Predict which forms an anion, which forms a cation, and the charges of each ion. Write the symbol for each ion and name them.

Answer

Al will form a cation with a charge of 3+: Al^{3+}, an aluminum ion. Carbon will form an anion with a charge of 4−: C^{4-}, a carbide ion.

Summary



Video* 3.3.1*: *What are ions?*

Glossary

anion
negative ions that are formed when a nonmental atom gains one or more electrons.

cation
positive ions that are formed when an atom loses electrons

ion
atom or molecule that has gained or lost one or more of its valence electrons, giving it a net electrical charge.

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CHAPTER OVERVIEW

4: Molecules, Compounds and Quantifying Chemicals

4.1: Chemical Bonding

4.2: Representing ionic and molecular compounds and molecules

4.2.1: Practice Problems- Writing Chemical Formulas

4.3: Chemical Nomenclature

4.3.1: Practice Problems- Molecular and Ionic Compounds

4.4: Formula Mass, Percent Composition, and the Mole

4.4.1: Practice Problems- Formula Mass, Percent Composition, and the Mole

4.5: Empirical and Molecular Formulas

4.5.1: Practice Problems- Empirical and Molecular Formulas

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4.1: Chemical Bonding

Skills to Develop

- Define ionic and molecular (covalent) compounds
- Predict the type of compound formed from elements based on their location within the periodic table
- Determine formulas for simple ionic compounds
- Identify polyatomic ions

Atomic Bonding

Chemical bonds form when electrons can be simultaneously close to two or more nuclei, but beyond this, there is no simple, easily understood theory that would not only explain why atoms bind together to form molecules, but would also predict the three-dimensional structures of the resulting compounds as well as the energies and other properties of the bonds themselves. Unfortunately, no one theory exists that accomplishes these goals in a satisfactory way for all of the many categories of compounds that are known. Moreover, it seems likely that if such a theory does ever come into being, it will be far from simple.

When we are faced with a scientific problem of this complexity, experience has shown that it is often more useful to concentrate instead on developing **models**. A scientific model is something like a theory in that it should be able to explain observed phenomena and to make useful predictions. But whereas a theory can be discredited by a single contradictory case, a model can be useful even if it does not encompass all instances of the phenomena it attempts to explain.

Given the extraordinary variety of ways in which atoms combine into aggregates, it should come as no surprise that a number of useful bonding models have been developed. Most of them apply only to certain classes of compounds, or attempt to explain only a restricted range of phenomena. In this section we will provide brief descriptions of some of the bonding models.



Video* 4.1.1*: *A brief overview of the bonds we will discuss below.*

Ionic Bonding and Ionic Compounds

When an element composed of atoms that readily lose electrons (a metal) reacts with an element composed of atoms that readily gain electrons (a nonmetal), a transfer of electrons usually occurs, producing ions. The compound formed by this transfer is stabilized by the electrostatic attractions (ionic bonds) between the ions of opposite charge present in the compound. For example, when each sodium atom in a sample of sodium metal (group 1) gives up one electron to form a sodium cation, Na^+, and each chlorine atom in a sample of chlorine gas (group 17) accepts one electron to form a chloride anion, Cl^-, the resulting compound, NaCl, is composed of sodium ions and chloride ions in the ratio of one Na^+ ion for each Cl^- ion. The solid NaCl, therefore consists of positive and negative ions arranged in a three - dimensional structure, called a crystal lattice. Each ion is attracted to neighboring

ions of opposite charge, and is repelled by ions of like charge; this combination of attractions and repulsions, acting in all directions, causes the ion to be tightly fixed in its own location in the crystal lattice.

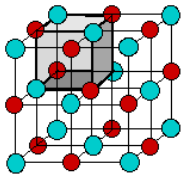


Figure 4.1.1: A model of a crystal lattice structure

Similarly, each calcium atom (group 2) can give up two electrons and transfer one to each of two chlorine atoms to form $CaCl_2$, which is composed of Ca^{2+} and Cl^- ions in the ratio of one Ca^{2+} ion to two Cl^- ions.

A compound that contains ions and is held together by ionic bonds is called an ionic compound. The periodic table can help us recognize many of the compounds that are ionic: When a metal is combined with one or more nonmetals, the compound is usually ionic. This guideline works well for predicting ionic compound formation for most of the compounds typically encountered in an introductory chemistry course. However, it is not always true (for example, aluminum chloride, $AlCl_3$, is not ionic).



Video 4.1.2: *What is an ionic bond?*

Ionic compounds can often be recognized by their properties. Ionic compounds are solids that typically melt at high temperatures and boil at even higher temperatures. For example, sodium chloride melts at 801 °C and boils at 1413 °C. (As a comparison, the molecular compound water melts at 0 °C and boils at 100 °C.) In solid form, an ionic compound is not electrically conductive because its ions are unable to flow (“electricity” is the flow of charged particles). When molten, however, it can conduct electricity because its ions are able to move freely through the liquid (Figure **4.1.3**; Video **4.1.3**).

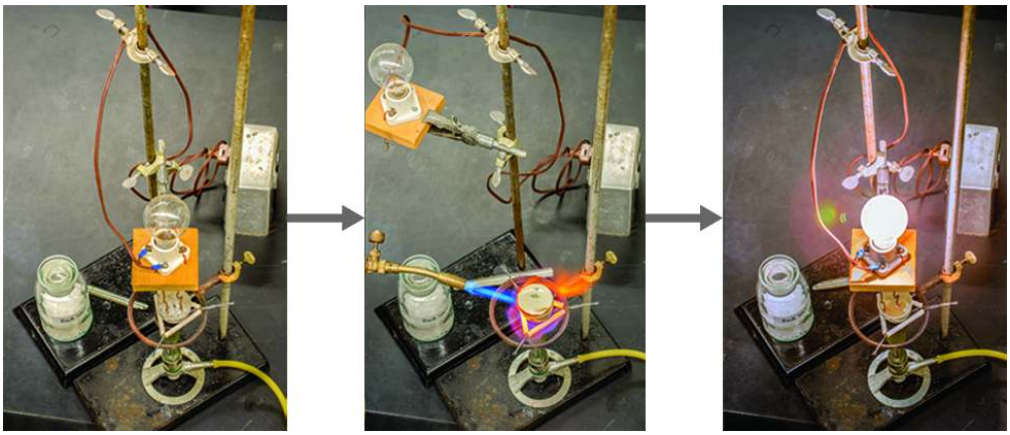


Figure* 4.1.2*: *Sodium chloride melts at 801 °C and conducts electricity when molten. (credit: modification of work by Mark Blaser and Matt Evans)*



Video* 4.1.3*: *Watch this video to see a mixture of salts melt and conduct electricity.*

In every ionic compound, the total number of positive charges of the cations equals the total number of negative charges of the anions. Thus, ionic compounds are electrically neutral overall, even though they contain positive and negative ions. We can use this observation to help us write the formula of an ionic compound. The formula of an ionic compound must have a ratio of ions such that the numbers of positive and negative charges are equal.

Example 4.1.3: Predicting the Formula of an Ionic Compound

The gemstone sapphire (Figure **4.1.4**) is mostly a compound of aluminum and oxygen that contains aluminum cations, Al^{3+}, and oxygen anions, O^{2-}. What is the formula of this compound?



Figure* 4.1.4*: *Although pure aluminum oxide is colorless, trace amounts of iron and titanium give blue sapphire its characteristic color. (credit: modification of work by Stanislav Doronenko)*

Solution Because the ionic compound must be electrically neutral, it must have the same number of positive and negative charges. Two aluminum ions, each with a charge of 3+, would give us six positive charges, and three oxide ions, each with a charge of 2−, would give us six negative charges. The formula would be Al_2O_3.

Exercise 4.1.3

Predict the formula of the ionic compound formed between the sodium cation, Na^+, and the sulfide anion, S^{2-}.

Answer

Na_2S

Molecular (Covalent) Compounds



Video* 4.1.4*: *What are covalent bonds?*

Many compounds do not contain ions but instead consist solely of discrete, neutral molecules. These molecular compounds (covalent compounds) result when atoms share, rather than transfer (gain or lose), electrons. Covalent bonding is an important and extensive concept in chemistry, and it will be treated in considerable detail in a later chapter of this text. We can often identify molecular compounds on the basis of their physical properties. Under normal conditions, molecular compounds often exist as gases, low-boiling liquids, and low-melting solids, although many important exceptions exist.

Whereas ionic compounds are usually formed when a metal and a nonmetal combine, covalent compounds are usually formed by a combination of nonmetals. Thus, the periodic table can help us recognize many of the compounds that are covalent. While we can use the positions of a compound's elements in the periodic table to predict whether it is ionic or covalent at this point in our study of chemistry, you should be aware that this is a very simplistic approach that does not account for a number of interesting exceptions. Shades of gray exist between ionic and molecular compounds, and you'll learn more about those later.

Example 4.1.2: Predicting the Type of Bonding in Compounds

Predict whether the following compounds are ionic or molecular:

a. KI, the compound used as a source of iodine in table salt
b. H_2O_2, the bleach and disinfectant hydrogen peroxide
c. $CHCl_3$, the anesthetic chloroform
d. Li_2CO_3, a source of lithium in antidepressants

Solution

a. Potassium (group 1) is a metal, and iodine (group 17) is a nonmetal; KI is predicted to be ionic.
b. Hydrogen (group 1) is a nonmetal, and oxygen (group 16) is a nonmetal; H_2O_2 is predicted to be molecular.
c. Carbon (group 14) is a nonmetal, hydrogen (group 1) is a nonmetal, and chlorine (group 17) is a nonmetal; $CHCl_3$ is predicted to be molecular.
d. Lithium (group 1) is a metal, and carbonate is a polyatomic ion; Li_2CO_3 is predicted to be ionic.

Exercise 4.1.2

Using the periodic table, predict whether the following compounds are ionic or covalent:

a. SO_2
b. CaF_2
c. N_2H_4
d. $Al_2(SO_4)_3$

Answer a

molecular

Answer b

ionic

Answer c

molecular

Answer d

ionic

Polyatomic Ions

The ions that we have discussed so far are called monatomic ions, that is, they are ions formed from only one atom. Polyatomic ions are are composed of more than one atom, united chemically. For example, the ammonium ion consists of one nitrogen atoms and four hydrogen atoms linked together via covalent bonds, and together they comprise a single ion with a $1+$ charge and a formula of NH_4^+. The carbonate ion consists of one carbon atom and three oxygen atoms and carries an overall charge of $2-$. The formula of the carbonate ion is CO_3^{2-}. The atoms of a polyatomic ion are covalently bonded together and so the entire ion behaves as a single unit.

Some of the more important polyatomic ions are listed in Table 4.1.1. Because you will use them repeatedly, they will soon become familiar.

Note that there is a system for naming some polyatomic ions: Oxyanions are polyatomic ions that contain one or more oxygen atoms; *-ate* and *-ite* are suffixes designating polyatomic ions containing more or fewer oxygen atoms. *Per-* (short for "hyper") and *hypo-* (meaning "under") are prefixes meaning more oxygen atoms than *-ate* and fewer oxygen atoms than *-ite*, respectively. For example, perchlorate is ClO_4^-, chlorate is ClO_3^-, chlorite is ClO_2^- and hypochlorite is ClO^-. Unfortunately, the number of oxygen atoms corresponding to a given suffix or prefix is not consistent; for example, nitrate is NO_3^- while sulfate is SO_4^{2-}.

Table* 4.1.1*: *Common Polyatomic Ions*

Name	Formula	Related Acid	Formula
ammonium	NH_4^+		
hydronium	H_3O^+		
oxide	O^{2-}		
peroxide	O_2^{2-}		
hydroxide	OH^-		
acetate	CH_3COO^-	acetic acid	CH_3COOH
cyanide	CN^-	hydrocyanic acid	HCN
azide	N_3^-	hydrazoic acid	HN_3
carbonate	CO_3^{2-}	carbonic acid	H_2CO_3
bicarbonate	HCO_3^-		
nitrate	NO_3^-	nitric acid	HNO_3
nitrite	NO_2^-	nitrous acid	HNO_2
sulfate	SO_4^{2-}	sulfuric acid	H_2SO_4
hydrogen sulfate	HSO_4^-		
sulfite	SO_3^{2-}	sulfurous acid	H_2SO_3
hydrogen sulfite	HSO_3^-		
phosphate	PO_4^{3-}	phosphoric acid	H_3PO_4
hydrogen phosphate	HPO_4^{2-}		
dihydrogen phosphate	$H_2PO_4^-$		
perchlorate	ClO_4^-	perchloric acid	$HClO_4$
chlorate	ClO_3^-	chloric acid	$HClO_3$
chlorite	ClO_2^-	chlorous acid	$HClO_2$
hypochlorite	ClO^-	hypochlorous acid	$HClO$
chromate	CrO_4^{2-}	chromic acid	H_2CrO_4
dichromate	$Cr_2O_7^{2-}$	dichromic acid	$H_2Cr_2O_7$
permanganate	MnO_4^-	permanganic acid	$HMnO_4$

Compounds Containing Polyatomic Ions

Polyatomic ions will form ionic bonds with monoatomic or polyatomic ions to make ionic compounds. For example, ammonium nitrate is an ionic compound composed of the polyatomic cation ammonium NH_4^+ and the polyatomic anion nitrate NO_3^-. Because ammonium carries a $1+$ charge and nitrate carries a $1-$ charge, the two ions will be in a 1:1 ratio and the ionic compound will be electrically neutral; the formula for ammonium nitrate is NH_4NO_3.

An example of an ionic compound formed between a monoatomic and a polyatomic ion is lithium dichromate: $Li_2Cr_2O_7$. Note that the lithium cation, Li_2^+, carries a $1+$ charge and the dichromate anion, $Cr_2O_7^{2}$, carries a $2-$ charge, therefore in order for the

formula to be neutral, there must be two lithium ions for every one dichromate ion. Use the table of polyatomic ioncs (Table 4.1.1) when predicting formulas for, or naming ionic compounds that include a polyatomic ion.

Example 4.1.3: Predicting the Formula of a Compound with a Polyatomic Anion

Baking powder contains calcium dihydrogen phosphate, an ionic compound composed of the ions Ca^{2+} and $\mathbf{H_2PO_4^-}$. What is the formula of this compound?

Solution

The positive and negative charges must balance, and this ionic compound must be electrically neutral. Thus, we must have two negative charges to balance the 2+ charge of the calcium ion. This requires a ratio of one Ca^{2+} ion to two $\mathbf{H_2PO_4^-}$ ions. We designate this by enclosing the formula for the dihydrogen phosphate ion in parentheses and adding a subscript 2. The formula is $Ca(H_2PO_4)_2$.

Exercise 4.1.3

Predict the formula of the ionic compound formed between the lithium ion and the peroxide ion, $\mathbf{O_2^{2-}}$ (Hint: Use the periodic table to predict the sign and the charge on the lithium ion.)

Answer

Li_2O_2

Summary

The nature of the attractive forces that hold atoms or ions together within a compound is the basis for classifying chemical bonding. When electrons are transferred and ions form, ionic bonds result. Ionic bonds are electrostatic forces of attraction, that is, the attractive forces experienced between objects of opposite electrical charge (in this case, cations and anions). When electrons are “shared” and molecules form, covalent bonds result. Covalent bonds are the attractive forces between the positively charged nuclei of the bonded atoms and one or more pairs of electrons that are located between the atoms. Compounds are classified as ionic or molecular (covalent) on the basis of the bonds present in them.

Metals (particularly those in groups 1 and 2) tend to lose the number of electrons that would leave them with the same number of electrons as in the preceding noble gas in the periodic table. By this means, a positively charged ion is formed. Similarly, nonmetals (especially those in groups 16 and 17, and, to a lesser extent, those in Group 15) can gain the number of electrons needed to provide atoms with the same number of electrons as in the next noble gas in the periodic table. Thus, nonmetals tend to form negative ions. Positively charged ions are called cations, and negatively charged ions are called anions. Ions can be either monatomic (containing only one atom) or polyatomic (containing more than one atom).

Compounds that contain ions are called ionic compounds. Ionic compounds generally form from metals and nonmetals. Compounds that do not contain ions, but instead consist of atoms bonded tightly together in molecules (uncharged groups of atoms that behave as a single unit), are called covalent compounds. Covalent compounds usually form from two or more nonmetals.



Video* 4.1.6*: *A summary of chemical bonds.*

Glossary

binary compound
compound containing two different elements.

covalent bond
attractive force between the nuclei of a molecule's atoms and pairs of electrons between the atoms

covalent compound
(also, molecular compound) composed of molecules formed by atoms of two or more different elements

ionic bond
electrostatic forces of attraction between the oppositely charged ions of an ionic compound

ionic compound
compound composed of cations and anions combined in ratios, yielding an electrically neutral substance

molecular compound
(also, covalent compound) composed of molecules formed by atoms of two or more different elements

monatomic ion
ion composed of a single atom

polyatomic ion
ion composed of more than one atom

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-

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4.2: Representing ionic and molecular compounds and molecules

Skills to Develop

- Symbolize the composition of molecules using molecular formulas and empirical formulas
- Represent the bonding arrangement of atoms within molecules using structural formulas

Communicating the identity and composition of chemicals is an imperative skill for scientists and there are a variety of methods that can be used. Each method has a set of rules, and communicates different types of information about the compound or molecules. This section explains the use of chemical formulas for covalent molecules. The next section will discuss naming, or nomenclature of molecular and ionic compounds. The use of chemical formulas and nomenclature are used in academic and industrial settings and can be important for engineers, biologists, geologists, medical professionals and many others.

Chemical Formulas for Covalent Compounds and Molecules

A **molecular formula** is a representation of a molecule that uses chemical symbols to indicate the types of atoms followed by subscripts to show the number of atoms of each type in the molecule. (A subscript is used only when more than one atom of a given type is present.) Molecular formulas are also used as abbreviations for the names of compounds.

An **empirical formula** gives shows the composition of a compound given as the simplest whole-number ratio of atoms.

The **structural formula** for a compound gives the same information as its molecular formula (the types and numbers of atoms in the molecule), and also shows how the atoms are connected in the molecule.

For example, the molecule cyclohexane is comprised of six carbon atoms and twelve hydrogen atoms; the carbon atoms are covalently bonded in a circular chain and two hydrogen atoms are covalently bonded to each carbon atom. The molecular formula for cyclohexane is C_6H_{12} and the empirical formula of cyclohexane is C_2H_6. Notice how the molecular formula provides more information than the empirical formula - the actual number of atoms compared to the ratio of the atoms. The structural formula also shows the connectivity of those atoms with a straight line indicating a single covalent bond (sharing of two electrons). Figure 1 shows the structural formula for cyclohexane

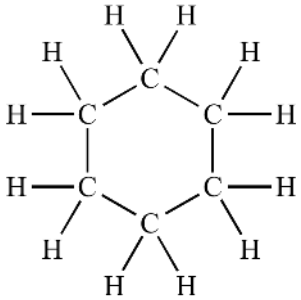


***Figure* 4.2.1**: *The structural formula of cyclohexane*

Two other common representations of molecular (covalent) molecules and compounds are the ball-and-stick model and the space-filling model. A ball-and-stick model shows the geometric arrangement of the atoms with atomic sizes not to scale, and a space-filling model shows the relative sizes of the atoms. Figure two shows the molecular and structural formulas for methane and the ball and stick and space filling models for methane.

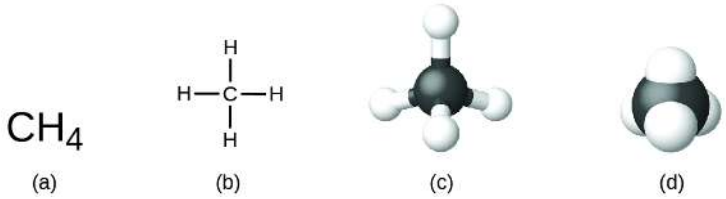


***Figure* 4.2.2**: *A methane molecule can be represented as (a) a molecular formula, (b) a structural formula, (c) a ball-and-stick model, and (d) a space-filling model. Carbon and hydrogen atoms are represented by black and white spheres, respectively.*

In many cases, the molecular formula of a substance is derived from experimental determination of both its empirical formula and its molecular mass (the sum of atomic masses for all atoms composing the molecule). For example, it can be determined experimentally that benzene contains two elements, carbon (C) and hydrogen (H), and that for every carbon atom in benzene, there is one hydrogen atom. Thus, the empirical formula is CH. An experimental determination of the molecular mass reveals that a molecule of benzene contains six carbon atoms and six hydrogen atoms, so the molecular formula for benzene is C_6H_6 (Figure **4.2.3**).

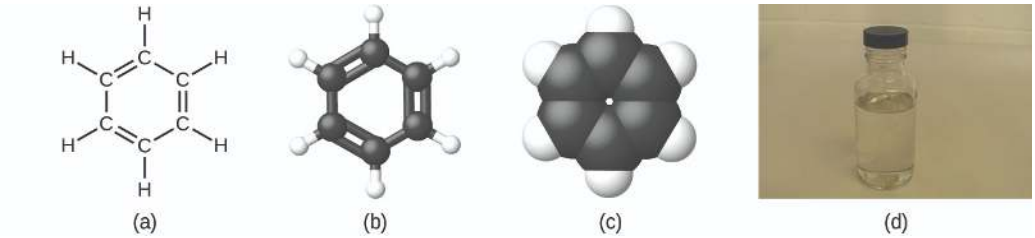


***Figure* 4.2.3**: *Benzene, C_6H_6, is produced during oil refining and has many industrial uses. A benzene molecule can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. (d) Benzene is a clear liquid. (credit d: modification of work by Sahar Atwa).*

If we know a compound's formula, we can easily determine the empirical formula. (This is somewhat of an academic exercise; the reverse chronology is generally followed in actual practice.) For example, the molecular formula for acetic acid, the component that gives vinegar its sharp taste, is $C_2H_4O_2$. This formula indicates that a molecule of acetic acid (Figure **4.2.4**) contains two carbon atoms, four hydrogen atoms, and two oxygen atoms. The ratio of atoms is 2:4:2. Dividing by the lowest common denominator (2) gives the simplest, whole-number ratio of atoms, 1:2:1, so the empirical formula is CH_2O. Note that a molecular formula is always a whole-number multiple of an empirical formula.

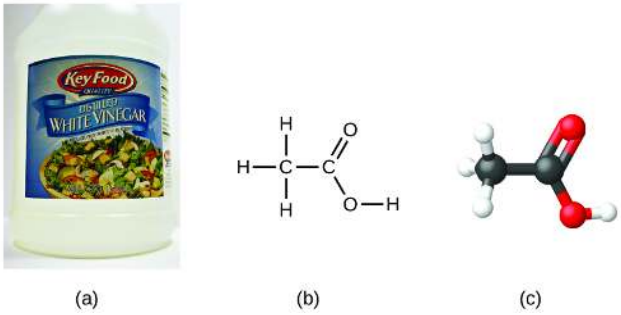


Figure 4.2.4: (a) Vinegar contains acetic acid, $C_2H_4O_2$, which has an empirical formula of CH_2O. It can be represented as (b) a structural formula and (c) as a ball-and-stick model. (credit a: modification of work by "HomeSpot HQ"/Flickr)

Example 4.2.1: Empirical and Molecular Formulas

Molecules of glucose (blood sugar) contain 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. What are the molecular and empirical formulas of glucose?

Solution

The molecular formula is $C_6H_{12}O_6$ because one molecule actually contains 6 C, 12 H, and 6 O atoms. The simplest whole-number ratio of C to H to O atoms in glucose is 1:2:1, so the empirical formula is CH_2O.

Exercise 4.2.1

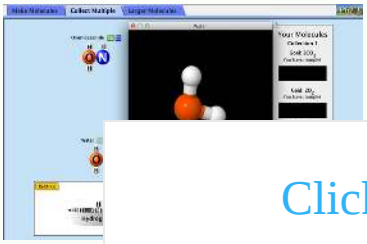
A molecule of metaldehyde (a pesticide used for snails and slugs) contains 8 carbon atoms, 16 hydrogen atoms, and 4 oxygen atoms. What are the molecular and empirical formulas of metaldehyde?

Answer

Molecular formula, $C_8H_{16}O_4$; empirical formula, C_2H_4O

Learn More about Atoms, Molecules and Molecular Formulas

Click the link to use a simulator.



Isomers: A preview of organic chemistry

It is important to be aware that it may be possible for the same atoms to be arranged in different ways: Compounds with the same molecular formula may have different atom-to-atom bonding and therefore different structures. For example, could there be another compound with the same formula as acetic acid, $C_2H_4O_2$? And if so, what would be the structure of its molecules?

If you predict that another compound with the formula $C_2H_4O_2$ could exist, then you demonstrated good chemical insight and are correct. Two C atoms, four H atoms, and two O atoms can also be arranged to form methyl formate, which is used in manufacturing, as an insecticide, and for quick-drying finishes. Methyl formate molecules have one of the oxygen atoms between the two carbon atoms, differing from the arrangement in acetic acid molecules. Acetic acid and methyl formate are examples of isomers—compounds with the same chemical formula but different molecular structures (Figure **4.2.5**). Note that this small difference in the arrangement of the atoms has a major effect on their respective chemical properties. You would certainly not want to use a solution of methyl formate as a substitute for a solution of acetic acid (vinegar) when you make salad dressing.

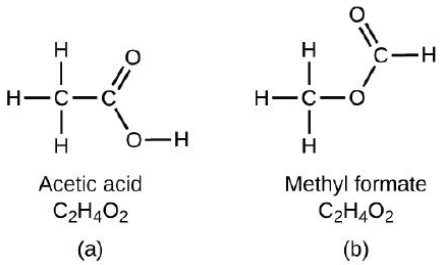


Figure* 4.2.5*: *Molecules of (a) acetic acid and methyl formate (b) are structural isomers; they have the same formula ($C_2H_4O_2$) but different structures (and therefore different chemical properties).*

Many other types of isomers exist (Figure 4.2.6). Acetic acid and methyl formate are structural isomers, compounds in which the molecules differ in how the atoms are connected to each other. There are also various types of spatial isomers, in which the relative orientations of the atoms in space can be different. For example, the compound carvone (found in caraway seeds, spearmint, and mandarin orange peels) consists of two isomers that are mirror images of each other. *S*-(+)-carvone smells like caraway, and *R*-(−)-carvone smells like spearmint. To learn more about these types of isomers, students can take an organic chemistry course!

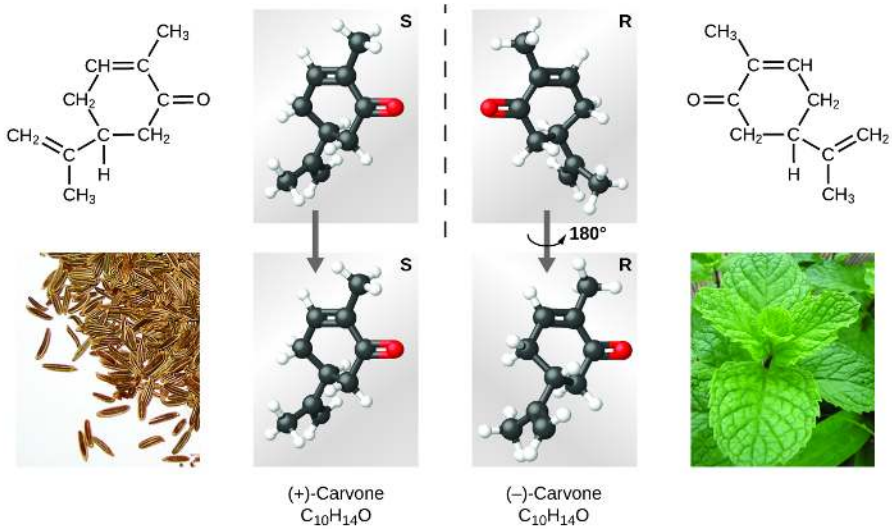


***Figure* 4.2.6**: *Molecules of carvone are spatial isomers; they only differ in the relative orientations of the atoms in space. (credit bottom left: modification of work by "Miansari66"/Wikimedia Commons; credit bottom right: modification of work by Forest & Kim Starr)*

Molecules made up of a single element

Although many elements consist of discrete, individual atoms, some exist as molecules made up of two or more atoms of the element chemically bonded together. For example, most samples of the elements hydrogen, oxygen, and nitrogen are composed of molecules that contain two atoms each (called diatomic molecules) and thus have the molecular formulas H_2, O_2, and N_2, respectively. Other elements commonly found as diatomic molecules are fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). The most common form of the element sulfur is composed of molecules that consist of eight atoms of sulfur; its molecular formula is S_8 (Figure 4.2.3).

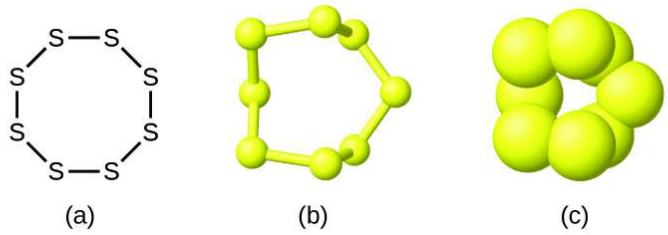


***Figure* 4.2.3**: *A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8. It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.*

It is important to note that a subscript following a symbol and a number in front of a symbol do not represent the same thing; for example, H_2 and 2H represent distinctly different species. H_2 is a molecular formula; it represents a diatomic molecule of hydrogen, consisting of two atoms of the element that are chemically bonded together. The expression 2H, on the other hand, indicates two separate hydrogen atoms that are not combined as a unit. The expression $2H_2$ represents two molecules of diatomic hydrogen (Figure 4.2.3).

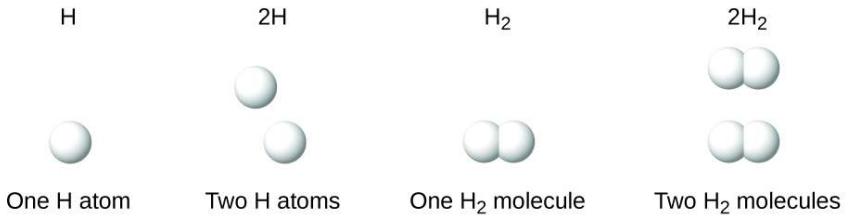


Figure* 4.2.3: The symbols H, 2H, H_2, and $2H_2$ represent very different entities.*

Summary

A molecular formula uses chemical symbols and subscripts to indicate the exact numbers of different atoms in a molecule or compound. An empirical formula gives the simplest, whole-number ratio of atoms in a compound. A structural formula indicates the bonding arrangement of the atoms in the molecule. Ball-and-stick and space-filling models show the geometric arrangement of atoms in a molecule. Isomers are compounds with the same molecular formula but different arrangements of atoms.

Going Beyond

What is it that chemists do? According to Lee Cronin, chemists make very complicated molecules by "chopping up" small molecules and "reverse engineering" them. He wonders if we could "make a really cool universal chemistry set" by what he calls "app-ing" chemistry. Could we "app" chemistry?



Video* 4.2.2*: *2012 TED talk about Cronin's idea to "print" medicine.*

Glossary

empirical formula
formula showing the composition of a compound given as the simplest whole-number ratio of atoms

isomers
compounds with the same chemical formula but different structures

molecular formula
formula indicating the composition of a molecule of a compound and giving the actual number of atoms of each element in a molecule of the compound.

spatial isomers
compounds in which the relative orientations of the atoms in space differ

structural isomer
one of two substances that have the same molecular formula but different physical and chemical properties because their atoms are bonded differently

structural formula
shows the atoms in a molecule and how they are connected

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- Adelaide Clark, Oregon Institute of Technology

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4.2.1: Practice Problems- Writing Chemical Formulas

PROBLEM 4.2.1.1

Explain why the symbol for an atom of the element oxygen and the formula for a molecule of oxygen differ.

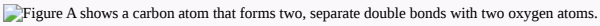
Answer

The symbol for the element oxygen, O, represents both the element and one atom of oxygen. A molecule of oxygen, O_2, contains two oxygen atoms; the subscript 2 in the formula must be used to distinguish the diatomic molecule from two single oxygen atoms.

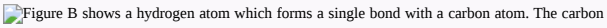
PROBLEM 4.2.1.2

Write the molecular and empirical formulas of the following compounds:

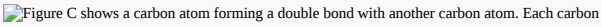
(a)


Figure A shows a carbon atom that forms two, separate double bonds with two oxygen atoms.

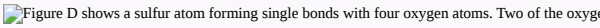
(b)


Figure B shows a hydrogen atom which forms a single bond with a carbon atom. The carbon atom forms a triple bond with another carbon atom. The second carbon atom forms a single bond with a hydrogen atom.

(c)


Figure C shows a carbon atom forming a double bond with another carbon atom. Each carbon atom forms a single bond with two hydrogen atoms.

(d)


Figure D shows a sulfur atom forming single bonds with four oxygen atoms. Two of the oxygen atoms form a single bond with a hydrogen atom.

Answer a

molecular: CO_2

empirical: CO_2

Answer b

molecular: C_2H_2

empirical: CH

Answer c

molecular: C_2H_4

empirical: CH_2

Answer d

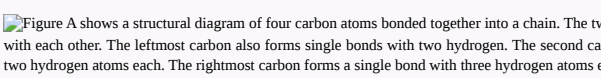
molecular: H_2SO_4

empirical: H_2SO4

PROBLEM 4.2.1.3

Write the molecular and empirical formulas of the following compounds:

(a)


Figure A shows a structural diagram of four carbon atoms bonded together into a chain. The two carbon atoms on the left form a double bond with each other. All of the remaining carbon atoms form single bonds with each other. The leftmost carbon also forms single bonds with two hydrogen. The second carbon in the chain forms a single bond with a hydrogen atom. The third carbon in the chain forms a single bond with two hydrogen atoms each. The rightmost carbon forms a single bond with three hydrogen atoms each.

(b)


Figure B shows a structural diagram of a molecule that has a chain of four carbon atoms. The leftmost carbon atom forms a single bond with three hydrogen atoms each and single bond with the second carbon atom. The second carbon atom forms a triple bond with the third carbon atom. The third carbon atom forms a single bond to the fourth carbon atom. The fourth carbon atom forms a single bond to three hydrogen atoms each.

(c)


Figure C shows a structural diagram of two silicon atoms are bonded together with a single bond. Each of the silicon atoms form single bonds to two chlorine atoms each and one hydrogen atom.

(d)


Figure D shows a structural diagram of a phosphorus atom that forms a single bond to four oxygen atoms each. Three of the oxygen atoms each have a single bond to a hydrogen atom.

Answer a

molecular: C_4H_8

empirical: CH_2

Answer b

molecular: C_4H_6

empirical: C_2H_3

Answer c

molecular: $H_2Si_2Cl_4$

empirical: $HSiCl_2$

Answer d

molecular: H_3PO_4

empirical: H_3PO4

Click here for a video of the solution



PROBLEM 4.2.1.4

Determine the empirical formulas for the following compounds:

(a) caffeine, $C_8H_{10}N_4O_2$
(b) fructose, $C_{12}H_{22}O_{11}$
(c) hydrogen peroxide, H_2O_2
(d) glucose, $C_6H_{12}O_6$
(e) ascorbic acid (vitamin C), $C_6H_8O_6$

Answer a

$C_4H_5N_2O$

Answer b

$C_{12}H_{22}O_{11}$

Answer c

HO

Answer d

CH_2O

Answer e

C_3H_4O3

PROBLEM 4.2.1.5

Determine the empirical formulas for the following compounds:

(a) acetic acid, $C_2H_4O_2$
(b) citric acid, $C_6H_8O_7$
(c) hydrazine, N_2H_4
(d) nicotine, $C_{10}H_{14}N_2$
(e) butane, C_4H_{10}

Answer a

CH_2O

Answer b

$C_6H_8O_7$

Answer c

NH_2

Answer d

C_5H_{7N}

Answer e

C_2H_5

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- Adelaide Clark, Oregon Institute of Technology

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4.3: Chemical Nomenclature

Skills to Develop

- Derive names for common types of inorganic compounds using a systematic approach

Naming Chemicals Compounds

Nomenclature, a collection of rules for naming things, is important in science and in many other situations. This module describes an approach that is used to name simple ionic and molecular compounds, such as NaCl, $CaCO_3$, and N_2O_4. The simplest of these are binary compounds, those containing only two elements, but we will also consider how to name ionic compounds containing polyatomic ions, and one specific, very important class of compounds known as acids (subsequent chapters in this text will focus on these compounds in great detail). We will limit our attention here to inorganic compounds, compounds that are composed principally of elements other than carbon, and will follow the nomenclature guidelines proposed by IUPAC. The rules for organic compounds, in which carbon is the principle element, will be treated in a later chapter on organic chemistry.

To name compounds, we need to consider the answers to several questions. First, is the compound ionic or molecular? If the compound is ionic, does the metal form ions of only one type (fixed charge) or more than one type (variable charge)? Are the ions monatomic or polyatomic? If the compound is molecular, does it contain hydrogen? If so, does it also contain oxygen? From the answers we derive, we place the compound in an appropriate category and then name it accordingly. We will begin with the nomenclature rules for ionic compounds.

Naming Ionic Compounds

Ionic compounds are not made up of single discrete molecules, they are instead, comprised of many atoms in a crystalline lattice structure. Therefore, ionic compounds cannot be symbolized using molecular formulas like we did in the previous section. Instead, ionic compounds are represented using **formula units**, which is the smallest, electrically neutral collection of ions. For many ionic compounds, formula units may look like empirical formulas, however they are different! We will explore the reason for this when we learn about ionic compounds containing polyatomic ions.

As we saw in the section on chemical bonding, the formulas of ionic compounds must be electrically neutral. We determined that an ionic compound comprised of aluminum cations and oxygen anions would have a formula of Al_2O_3. This was determined by studying the location of aluminum and oxygen on the periodic table; the aluminum cation (Al^{3+}) has a charge of +3 and the oxygen anion (O^{2-}) has a charge of -2. The formula unit must have an overall charge of zero; combining two aluminum cations with three oxygen anions, will give an overall charge of zero.

Ionic Compounds Containing Only Monatomic Ions

The name of a binary compound containing monatomic ions consists of the name of the cation (the name of the metal) followed by the name of the anion (the name of the nonmetallic element with its ending replaced by the suffix *–ide*). Some examples are given in Table **4.3.1**.

Table* 4.3.1*: *Names of Some Ionic Compounds*

NaCl, sodium chloride	Na_2O, sodium oxide
KBr, potassium bromide	CdS, cadmium sulfide
CaI_2, calcium iodide	Mg_3N_2, magnesium nitride
CsF, cesium fluoride	Ca_3P_2, calcium phosphide
LiCl, lithium chloride	Al_4C_3, aluminum carbide

Compounds Containing Polyatomic Ions

Compounds containing polyatomic ions are named similarly to those containing only monatomic ions, except there is no need to change to an *–ide* ending, since the suffix is already present in the name of the anion. Examples are shown in Table **4.3.2**.

Table* 4.3.2*: *Names of Some Polyatomic Ionic Compounds*

$KC_2H_3O_2$, potassium acetate	$(NH_4)Cl$, ammonium chloride
$NaHCO_3$, sodium bicarbonate	$CaSO_4$, calcium sulfate
$Al_2(CO_3)_3$, aluminum carbonate	$Mg_3(PO_4)_2$, magnesium phosphate

Ionic Compounds in Your Cabinets

Every day you encounter and use a large number of ionic compounds. Some of these compounds, where they are found, and what they are used for are listed in Table **4.3.3**. Look at the label or ingredients list on the various products that you use during the next few days, and see if you run into any of those in this table, or find other ionic compounds that you could now name or write as a formula.

Table* 4.3.3*: *Everyday Ionic Compounds*

Ionic Compound	Name	Use
NaCl	sodium chloride	ordinary table salt
KI	potassium iodide	added to “iodized” salt for thyroid health
NaF	sodium fluoride	ingredient in toothpaste
$NaHCO_3$	sodium bicarbonate	baking soda; used in cooking (and in antacids)
Na_2CO_3	sodium carbonate	washing soda; used in cleaning agents
NaOCl	sodium hypochlorite	active ingredient in household bleach
$CaCO_3$	calcium carbonate	ingredient in antacids
$Mg(OH)_2$	magnesium hydroxide	ingredient in antacids
$Al(OH)_3$	aluminum hydroxide	ingredient in antacids
NaOH	sodium hydroxide	lye; used as drain cleaner
K_3PO_4	potassium phosphate	food additive (many purposes)
$MgSO_4$	magnesium sulfate	added to purified water
Na_2HPO_4	sodium hydrogen phosphate	anti-caking agent; used in powdered products
Na_2SO_3	sodium sulfite	preservative

Compounds Containing a Metal Ion with a Variable Charge

Most of the transition metals can form two or more cations with different charges. Compounds of these metals with nonmetals are named with the same method as compounds in the first category, except the charge of the metal ion is specified by a Roman numeral in parentheses after the name of the metal. The charge of the metal ion is determined from the formula of the compound and the charge of the anion. For example, consider binary ionic compounds of iron and chlorine. Iron typically exhibits a charge of either 2+ or 3+, and the two corresponding compound formulas are $FeCl_2$ and $FeCl_3$. The simplest name, “iron chloride,” will, in this case, be ambiguous, as it does not distinguish between these two compounds. In cases like this, the charge of the metal ion is included as a Roman numeral in parentheses immediately following the metal name. These two compounds are then unambiguously named iron(II) chloride and iron(III) chloride, respectively. Other examples are provided in Table **4.3.4**.

Table* 4.3.4*: *Names of Some Transition Metal Ionic Compounds*

Transition Metal Ionic Compound	Name
$FeCl_3$	iron(III) chloride
Hg_2O	mercury(I) oxide
HgO	mercury(II) oxide
$Cu_3(PO_4)_2$	copper(II) phosphate

Out-of-date nomenclature used the suffixes *–ic* and *–ous* to designate metals with higher and lower charges, respectively: Iron(III) chloride, $FeCl_3$, was previously called ferric chloride, and iron(II) chloride, $FeCl_2$, was known as ferrous chloride. Though this naming convention has been largely abandoned by the scientific community, it remains in use by some segments of industry. For example, you may see the words *stannous fluoride* on a tube of toothpaste. This represents the formula SnF_2, which is more

properly named tin(II) fluoride. The other fluoride of tin is SnF_4, which was previously called stannic fluoride but is now named tin(IV) fluoride.

Example 4.3.1: Naming Ionic Compounds

Name the following ionic compounds, which contain a metal that can have more than one ionic charge:

a. Fe_2S_3
b. CuSe
c. GaN
d. $CrCl_3$
e. $Ti_2(SO_4)_3$

Solution

The anions in these compounds have a fixed negative charge (S^{2-}, Se^{2-} , N^{3-}, Cl^-, and SO_4^{2-}), and the compounds must be neutral. Because the total number of positive charges in each compound must equal the total number of negative charges, the positive ions must be Fe^{3+}, Cu^{2+}, Ga^{3+}, Cr^{3+}, and Ti^{3+}. These charges are used in the names of the metal ions:

a. iron(III) sulfide
b. copper(II) selenide
c. gallium(III) nitride
d. chromium(III) chloride
e. titanium(III) sulfate

Exercise 4.3.1

Write the formulas of the following ionic compounds:

a. chromium(III) phosphide
b. mercury(II) sulfide
c. manganese(II) phosphate
d. copper(I) oxide
e. chromium(VI) fluoride

Answer a

CrP

Answer b

HgS

Answer c

$Mn_3(PO_4)_2$

Answer d

Cu_2O

Answer e

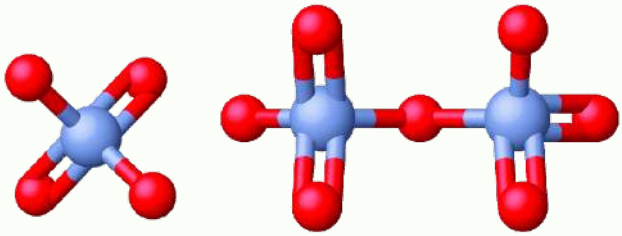
CrF_6

Erin Brokovich and Chromium Contamination

In the early 1990s, legal file clerk Erin Brockovich (Figure **4.3.2**) discovered a high rate of serious illnesses in the small town of Hinckley, California. Her investigation eventually linked the illnesses to groundwater contaminated by Cr(VI) used by Pacific Gas & Electric (PG&E) to fight corrosion in a nearby natural gas pipeline. As dramatized in the film *Erin Brokovich* (for which Julia Roberts won an Oscar), Erin and lawyer Edward Masry sued PG&E for contaminating the water near Hinckley in 1993. The settlement they won in 1996—$333 million—was the largest amount ever awarded for a direct-action lawsuit in the US at that time.



(a)



(b)

Figure* 4.3.5: (a) Erin Brockovich found that Cr(VI), used by PG&E, had contaminated the Hinckley, California, water supply. (b) The Cr(VI) ion is often present in water as the polyatomic ions chromate, CrO_4^{2-} (left), and dichromate, $Cr_2O_7^{2-}$ (right).*

Chromium compounds are widely used in industry, such as for chrome plating, in dye-making, as preservatives, and to prevent corrosion in cooling tower water, as occurred near Hinckley. In the environment, chromium exists primarily in either the Cr(III) or Cr(VI) forms. Cr(III), an ingredient of many vitamin and nutritional supplements, forms compounds that are not very soluble in water, and it has low toxicity. Cr(VI), on the other hand, is much more toxic and forms compounds that are reasonably soluble in water. Exposure to small amounts of Cr(VI) can lead to damage of the respiratory, gastrointestinal, and immune systems, as well as the kidneys, liver, blood, and skin.

Despite cleanup efforts, Cr(VI) groundwater contamination remains a problem in Hinckley and other locations across the globe. A 2010 study by the Environmental Working Group found that of 35 US cities tested, 31 had higher levels of Cr(VI) in their tap water than the public health goal of 0.02 parts per billion set by the California Environmental Protection Agency.

Naming Molecular (Covalent) Compounds

The bonding characteristics of inorganic molecular compounds are different from ionic compounds, and they are named using a different system as well. The charges of cations and anions dictate their ratios in ionic compounds, so specifying the names of the ions provides sufficient information to determine chemical formulas. However, because covalent bonding allows for significant variation in the combination ratios of the atoms in a molecule, the names for molecular compounds must explicitly identify these ratios.

Compounds Composed of Two Elements

When two nonmetallic elements form a molecular compound, several combination ratios are often possible. For example, carbon and oxygen can form the compounds CO and CO_2. Since these are different substances with different properties, they cannot both have the same name (they cannot both be called carbon oxide). To deal with this situation, we use a naming method that is somewhat similar to that used for ionic compounds, but with added prefixes to specify the numbers of atoms of each element. The name of the more metallic element (the one farther to the left and/or bottom of the periodic table) is first, followed by the name of the more nonmetallic element (the one farther to the right and/or top) with its ending changed to the suffix *–ide*. The numbers of atoms of each element are designated by the Greek prefixes shown in Table **4.3.5**.

Table* 4.3.5: Nomenclature Prefixes*

Number	Prefix		Number	Prefix
1 (sometimes omitted)	mono-		6	hexa-
2	di-		7	hepta-
3	tri-		8	octa-
4	tetra-		9	nona-
5	penta-		10	deca-

When only one atom of the first element is present, the prefix *mono-* is usually deleted from that part. Thus, CO is named carbon monoxide, and CO_2 is called carbon dioxide. When two vowels are adjacent, the *a* in the Greek prefix is usually dropped. Some other examples are shown in Table **4.3.6**.

Table* 4.3.6*: *Names of Some Molecular Compounds Composed of Two Elements*

Compound	Name		Compound	Name
SO_2	sulfur dioxide		BCl_3	boron trichloride
SO_3	sulfur trioxide		SF_6	sulfur hexafluoride
NO_2	nitrogen dioxide		PF_5	phosphoru

Compound	Name	Compound	Name
			sentafluoride
N_2O_4	dinitrogen tetroxide	P_4O_{10}	tetraphosphorus decaoxide
N_2O_5	dinitrogen pentoxide	IF_7	iodin

Compound	Name	Compound	Name
			heptafluoride

There are a few common names that you will encounter as you continue your study of chemistry. For example, although NO is often called nitric oxide, its proper name is nitrogen monoxide. Similarly, N_2O is known as nitrous oxide even though our rules would specify the name dinitrogen monoxide. (And H_2O is usually called water, not dihydrogen monoxide.) You should commit to memory the common names of compounds as you encounter them.

Example 4.3.2: Naming Covalent Compounds

Name the following covalent compounds:

a. SF_6
b. N_2O_3
c. Cl_2O_7
d. P_4O_6

Solution

Because these compounds consist solely of nonmetals, we use prefixes to designate the number of atoms of each element:

a. sulfur hexafluoride
b. dinitrogen trioxide
c. dichlorine heptoxide
d. tetraphosphorus hexoxide

Exercise 4.3.2

Write the formulas for the following compounds:

a. phosphorus pentachloride
b. dinitrogen monoxide
c. iodine heptafluoride
d. carbon tetrachloride

Answer a

PCl_5

Answer b

N_2O

Answer c

IF_7

Answer d

CCl_4

Binary Acids

Some compounds containing hydrogen are members of an important class of substances known as acids. The chemistry of these compounds is explored in more detail in later chapters of this text, but for now, it will suffice to note that many acids release hydrogen ions, H^+, when dissolved in water. To denote this distinct chemical property, a mixture of water with an acid is given a name derived from the compound's name. If the compound is a binary acid (comprised of hydrogen and one other nonmetallic element):

1. The word "hydrogen" is changed to the prefix *hydro-*
2. The other nonmetallic element name is modified by adding the suffix *-ic*
3. The word "acid" is added as a second word

For example, when the gas **HCl** (hydrogen chloride) is dissolved in water, the solution is called *hydrochloric acid*. Several other examples of this nomenclature are shown in Table **4.3.7**.

Table* 4.3.7: Names of Some Simple Acids*

Name of Gas	Name of Acid
HF(*g*), hydrogen fluoride	HF(*aq*), hydrofluoric acid
HCl(*g*), hydrogen chloride	HCl(*aq*), hydrochloric acid
HBr(*g*), hydrogen bromide	HBr(*aq*), hydrobromic acid
HI(*g*), hydrogen iodide	HI(*aq*), hydroiodic acid
H_2S(*g*), hydrogen sulfide	H_2S(*aq*), hydrosulfuric acid

Oxyacids

Many compounds containing three or more elements (such as organic compounds or coordination compounds) are subject to specialized nomenclature rules that you will learn later. However, we will briefly discuss the important compounds known as oxyacids, compounds that contain hydrogen, oxygen, and at least one other element, and are bonded in such a way as to impart acidic properties to the compound (you will learn the details of this in a later chapter). Typical oxyacids consist of hydrogen combined with a polyatomic, oxygen-containing ion. To name oxyacids:

1. Omit "hydrogen"
2. Start with the root name of the anion
3. Replace *–ate* with *–ic*, or *–ite* with *–ous*
4. Add "acid"

For example, consider H_2CO_3 (which you might be tempted to call "hydrogen carbonate"). To name this correctly, "hydrogen" is omitted; the *–ate* of carbonate is replace with *–ic*; and acid is added—so its name is carbonic acid. Other examples are given in Table **4.3.8**. There are some exceptions to the general naming method (e.g., H_2SO_4 is called sulfuric acid, not sulfic acid, and H_2SO_3 is sulfurous, not sulfous, acid).

Table* 4.3.8: Names of Common Oxyacids*

Formula	Anion Name	Acid Name
$HC_2H_3O_2$	acetate	acetic acid
HNO_3	nitrate	nitric acid
HNO_2	nitrite	nitrous acid
$HClO_4$	perchlorate	perchloric acid
H_2CO_3	carbonate	carbonic acid
H_2SO_4	sulfate	sulfuric acid
H_2SO_3	sulfite	sulfurous acid
H_3PO_4	phosphate	phosphoric acid

Summary



Video* 4.3.1*: *"You know what's an awful lot like unexpectedly waking up in Belguim with a bunch of people speaking several different languages you don't understand? Chemistry."*

Chemists use nomenclature rules to clearly name compounds. Ionic and molecular compounds are named using somewhat-different methods. Binary ionic compounds typically consist of a metal and a nonmetal. The name of the metal is written first, followed by the name of the nonmetal with its ending changed to *–ide*. For example, K_2O is called potassium oxide. If the metal can form ions with different charges, a Roman numeral in parentheses follows the name of the metal to specify its charge. Thus, $FeCl_2$ is iron(II) chloride and $FeCl_3$ is iron(III) chloride. Some compounds contain polyatomic ions; the names of common polyatomic ions should be memorized. Molecular compounds can form compounds with different ratios of their elements, so prefixes are used to specify the numbers of atoms of each element in a molecule of the compound. Examples include SF_6, sulfur hexafluoride, and N_2O_4, dinitrogen tetroxide. Acids are an important class of compounds containing hydrogen and having special nomenclature rules. Binary acids are named using the prefix *hydro-*, changing the *–ide* suffix to *–ic*, and adding "acid;" HCl is hydrochloric acid. Oxyacids are named by changing the ending of the anion (*-ate* to *–ic*, and *-ite* to *-ous*) and adding "acid;" H_2CO_3 is carbonic acid.

Glossary

binary acid

compound that contains hydrogen and one other element, bonded in a way that imparts acidic properties to the compound (ability to release H^+ ions when dissolved in water)

binary compound

compound containing two different elements.

covalent bond

attractive force between the nuclei of a molecule's atoms and pairs of electrons between the atoms

covalent compound

(also, molecular compound) composed of molecules formed by atoms of two or more different elements

ionic bond

electrostatic forces of attraction between the oppositely charged ions of an ionic compound

ionic compound

compound composed of cations and anions combined in ratios, yielding an electrically neutral substance

molecular compound

(also, covalent compound) composed of molecules formed by atoms of two or more different elements

monatomic ion

ion composed of a single atom

nomenclature

system of rules for naming objects of interest

oxyacid

compound that contains hydrogen, oxygen, and one other element, bonded in a way that imparts acidic properties to the compound (ability to release H^+ ions when dissolved in water)

oxyanion

polyatomic anion composed of a central atom bonded to oxygen atoms

polyatomic ion

ion composed of more than one atom

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- Adelaide Clark, Oregon Institute of Technology
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- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook

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4.3.1: Practice Problems- Molecular and Ionic Compounds

PROBLEM 4.3.1.1

Using the periodic table, predict whether the following chlorides are ionic or covalent: KCl, NCl_3, ICl, $MgCl_2$, PCl_5, and CCl_4.

Answer

Ionic: KCl, $MgCl_2$; Covalent: NCl_3, ICl, PCl_5, CCl_4

PROBLEM 4.3.1.2

Using the periodic table, predict whether the following chlorides are ionic or covalent: $SiCl_4$, PCl_3, $CaCl_2$, CsCl, $CuCl_2$, and $CrCl_3$.

Answer

Ionic: $CaCl_2$, $CuCl_2$, CsCl, $CrCl_3$; Covalent: $SiCl_4$, PCl_3

PROBLEM 4.3.1.3

For each of the following compounds, state whether it is ionic or covalent. If it is ionic, write the symbols for the ions involved:

(a) NF_3
(b) BaO
(c) $(NH_4)_2CO_3$
(d) $Sr(H_2PO_4)_2$
(e) IBr
(f) Na_2O

Answer a

covalent

Answer b

ionic, Ba^{2+}, O^{2-}

Answer c

ionic, $\mathbf{NH_4^+}$, $\mathbf{CO_3^{2-}}$

Answer d

ionic, Sr^{2+}, $\mathbf{H_2PO_4^-}$

Answer e

covalent

Answer f

ionic, Na^+, O^{2-}

PROBLEM 4.3.1.4

For each of the following pairs of ions, write the symbol for the formula of the compound they will form:

(a) Ca^{2+}, S^{2-}
(b) $\mathbf{NH_4^+}$, $\mathbf{SO_4^{2-}}$
(c) Al^{3+}, Br^-

(d) Na^+, HPO_4^{2-}
(e) Mg^{2+}, PO_4^{3-}

Answer a

CaS

Answer b

$(NH_4)_2CO_3$

Answer c

$AlBr_3$

Answer d

Na_2HPO_4

Answer e

$Mg_3(PO_4)_2$

PROBLEM 4.3.1.5

For each of the following pairs of ions, write the symbol for the formula of the compound they will form:

(a) K^+, O^{2-}
(b) NH_4^+, PO_4^{3-}
(c) Al^{3+}, O^{2-}
(d) Na^+, CO_3^{2-}
(e) Ba^{2+}, PO_4^{3-}

Answer a

K_2O

Answer b

$(NH_4)_3PO_4$

Answer c

Al_2O_3

Answer d

Na_2CO_3

Answer e

$Ba_3(PO_4)_2$

Click here for a video of the solution



[redacted]

PROBLEM 4.3.1.6

Name the following compounds:

(a) CsCl
(b) BaO
(c) K_2S
(d) $BeCl_2$
(e) HBr
(f) AlF_3

Answer a

cesium chloride

Answer b

barium oxide

Answer c

potassium sulfide

Answer d

beryllium chloride

Answer e

hydrogen bromide

Answer f

aluminum fluoride

PROBLEM 4.3.1.7

Name the following compounds:

(a) NaF
(b) Rb_2O
(c) BCl_3
(d) H_2Se
(e) P_4O_6
(f) ICl_3

Answer a

sodium fluoride

Answer b

Rubidium oxide

Answer c

boron trichloride

Answer d

hydrogen selenide

Answer e

tetraphosphorous hexaoxide

(if you are googling the answers to your homework, google may disagree with you. But we are naming based on the rules we just learned, which is why you shouldn't trust google for the answers to your homework)

Answer f

Iodine trichloride

PROBLEM 4.3.1.8

Write the formulas of the following compounds:

(a) rubidium bromide
(b) magnesium selenide
(c) sodium oxide
(d) calcium chloride
(e) hydrogen fluoride
(f) gallium phosphide
(g) aluminum bromide
(h) ammonium sulfate

Answer a

RbBr

Answer b

MgSe

Answer c

Na_2O

Answer d

$CaCl_2$

Answer e

HF

Answer f

GaP

Answer g

$AlBr_3$

Answer h

$(NH_4)_2SO_4$

PROBLEM 4.3.1.9

Write the formulas of the following compounds:

(a) lithium carbonate
(b) sodium perchlorate
(c) barium hydroxide

(d) ammonium carbonate
(e) sulfuric acid
(f) calcium acetate
(g) magnesium phosphate
(h) sodium sulfite

Answer a

Li_2CO_3

Answer b

$NaClO_4$

Answer c

$Ba(OH)_2$

Answer d

$(NH_4)_2CO_3$

Answer e

H_2SO_4

Answer f

$Ca(C_2H_3O_2)_2$

Answer g

$Mg_3(PO_4)_2$

Answer h

Na_2SO_3

PROBLEM 4.3.1.10

Write the formulas of the following compounds:

(a) chlorine dioxide
(b) dinitrogen tetraoxide
(c) potassium phosphide
(d) silver(I) sulfide
(e) aluminum nitride
(f) silicon dioxide

Answer a

ClO_2

Answer b

N_2O_4

Answer c

K_3P

Answer d

Ag_2S

Answer e

AlN

Answer f

SiO_2

PROBLEM 4.3.1.11

Write the formulas of the following compounds:

(a) barium chloride
(b) magnesium nitride
(c) sulfur dioxide
(d) nitrogen trichloride
(e) dinitrogen trioxide
(f) tin(IV) chloride

Answer a

$BaCl_2$

Answer b

Mg_3N_2

Answer c

SO_2

Answer d

NCl_3

Answer e

N_2O_3

Answer f

$SnCl_4$

PROBLEM 4.3.1.12

Each of the following compounds contains a metal that can exhibit more than one ionic charge. Name these compounds:

(a) Cr_2O_3
(b) $FeCl_2$
(c) CrO_3
(d) $TiCl_4$
(e) CoO
(f) MoS_2

Answer a

chromium(III) oxide

Answer b

iron(II) chloride

Answer c

chromium(VI) oxide

Answer d

titanium(IV) chloride

Answer e

cobalt(II) oxide

Answer f

molybdenum(IV) sulfide

PROBLEM 4.3.1.13

Each of the following compounds contains a metal that can exhibit more than one ionic charge. Name these compounds:

(a) $NiCO_3$
(b) MoO_3
(c) $Co(NO_3)_2$
(d) V_2O_5
(e) MnO_2
(f) Fe_2O_3

Answer a

nickel (II) carbonate

Answer b

Molybdenum (VI) oxide

Answer c

cobalt (II) nitrate

Answer d

vanadium (V) oxide

Answer e

manganese (IV) oxide

Answer f

iron (III) oxide

PROBLEM 4.3.1.14

The following ionic compounds are found in common household products. Write the formulas for each compound:

(a) potassium phosphate
(b) copper(II) sulfate
(c) calcium chloride
(d) titanium dioxide
(e) ammonium nitrate
(f) sodium bisulfate (the common name for sodium hydrogen sulfate)

Answer a

K_3PO_4

Answer b

$CuSO_4$

Answer c

$CaCl_2$

Answer d

TiO_2

Answer e

NH_4NO_3

Answer f

$NaHSO_4$

PROBLEM 4.3.1.15

The following ionic compounds are found in common household products. Name each of the compounds:

(a) $Ca(H_2PO_4)_2$
(b) $FeSO_4$
(c) $CaCO_3$
(d) MgO
(e) $NaNO_2$
(f) KI

Answer a

calcium dihydrogen phosphate

Answer b

iron (II) sulfate

Answer c

calcium carbonate

Answer d

magnesium oxide

Answer e

sodium nitrite

Answer f

potassium iodide

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4.4: Formula Mass, Percent Composition, and the Mole

Skills to Develop

- Calculate formula masses for covalent and ionic compounds
- Define the amount unit mole and the related quantity Avogadro's number
- Explain the relation between mass, moles, and numbers of atoms or molecules, and perform calculations deriving these quantities from one another
- Compute the percent composition of a compound

We can argue that modern chemical science began when scientists started exploring the quantitative as well as the qualitative aspects of chemistry. For example, Dalton's atomic theory was an attempt to explain the results of measurements that allowed him to calculate the relative masses of elements combined in various compounds. Understanding the relationship between the masses of atoms and the chemical formulas of compounds allows us to quantitatively describe the composition of substances.

Formula Mass

In an earlier chapter, we described the development of the atomic mass unit, the concept of average atomic masses, and the use of chemical formulas to represent the elemental makeup of substances. These ideas can be extended to calculate the formula mass of a substance by summing the average atomic masses of all the atoms represented in the substance's formula.

Formula Mass for Covalent Substances

For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a molecular mass. Consider chloroform ($CHCl_3$), a covalent compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the "anti-stick" polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms. Figure **4.4.1** outlines the calculations used to derive the molecular mass of chloroform, which is 119.37 amu.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
				Molecular mass	119.37

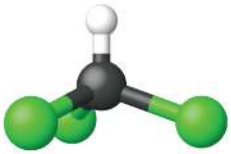


***Figure* 4.4.1:** *The average mass of a chloroform molecule, $CHCl_3$, is 119.37 amu, which is the sum of the average atomic masses of each of its constituent atoms. The model shows the molecular structure of chloroform.*

Likewise, the molecular mass of an aspirin molecule, $C_9H_8O_4$, is the sum of the atomic masses of nine carbon atoms, eight hydrogen atoms, and four oxygen atoms, which amounts to 180.15 amu (Figure **4.4.2**).

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	9	×	12.01	=	108.09
H	8	×	1.008	=	8.064
O	4	×	16.00	=	64.00
				Molecular mass	180.15



***Figure* 4.4.2**: *The average mass of an aspirin molecule is 180.15 amu. The model shows the molecular structure of aspirin, $C_9H_8O_4$.*

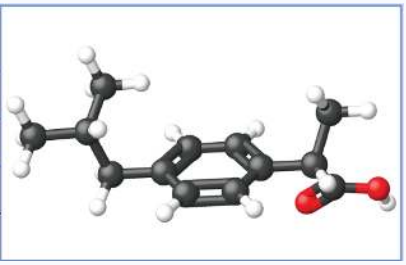
Example 4.4.1: Computing Molecular Mass for a Covalent Compound

Ibuprofen, $C_{13}H_{18}O_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Advil and Motrin. What is the molecular mass (amu) for this compound?

Solution

Molecules of this compound are comprised of 13 carbon atoms, 18 hydrogen atoms, and 2 oxygen atoms. Following the approach described above, the average molecular mass for this compound is therefore:

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	13	×	12.01	=	156.13
H	18	×	1.008	=	18.114
O	2	×	16.00	=	32.00
				Molecular mass	206.27



Exercise 4.4.1

Acetaminophen, $C_8H_9NO_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Tylenol. What is the molecular mass (amu) for this compound?

Answer

151.16 amu

Formula Mass for Ionic Compounds

Ionic compounds are composed of discrete cations and anions combined in ratios to yield electrically neutral bulk matter. The formula mass for an ionic compound is calculated in the same way as the formula mass for covalent compounds: by summing the average atomic masses of all the atoms in the compound's formula. Keep in mind, however, that the formula for an ionic compound does not represent the composition of a discrete molecule, so it may not correctly be referred to as the "molecular mass."

As an example, consider sodium chloride, NaCl, the chemical name for common table salt. Sodium chloride is an ionic compound composed of sodium cations, Na^+, and chloride anions, Cl^-, combined in a 1:1 ratio. The formula mass for this compound is computed as 58.44 amu (Figure **4.4.3**).

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
				Formula mass	58.44

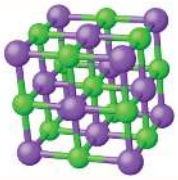


Figure* 4.4.3*: *Table salt, NaCl, contains an array of sodium and chloride ions combined in a 1:1 ratio. Its formula mass is 58.44 amu.*

Note that the average masses of neutral sodium and chlorine atoms were used in this computation, rather than the masses for sodium cations and chlorine anions. This approach is perfectly acceptable when computing the formula mass of an ionic compound. Even though a sodium cation has a slightly smaller mass than a sodium atom (since it is missing an electron), this difference will be offset by the fact that a chloride anion is slightly more massive than a chloride atom (due to the extra electron). Moreover, the mass of an electron is negligibly small with respect to the mass of a typical atom. Even when calculating the mass of an isolated ion, the missing or additional electrons can generally be ignored, since their contribution to the overall mass is negligible, reflected only in the nonsignificant digits that will be lost when the computed mass is properly rounded. The few exceptions to this guideline are very light ions derived from elements with precisely known atomic masses.

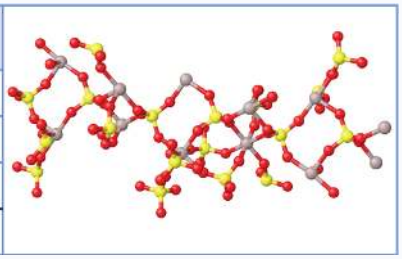
Example 4.4.2: Computing Formula Mass for an Ionic Compound

Aluminum sulfate, $Al_2(SO_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $Al_2S_3O_{12}$. Following the approach outlined above, the formula mass for this compound is calculated as follows:

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
Al	2	×	26.98	=	53.96
S	3	×	32.06	=	96.18
O	12	×	16.00	=	192.00
				Molecular mass	342.14



Exercise 4.4.2

Calcium phosphate, $Ca_3(PO_4)_2$, is an ionic compound and a common anti-caking agent added to food products. What is the formula mass (amu) of calcium phosphate?

Answer

310.18 amu

The Mole

In units 1 and 2, we learned about measuring quanity of atoms using the mole. As a reminder, the mole is an amount unit similar to familiar units like pair, dozen, gross, etc. The number of entities composing a mole has been experimentally determined to be $6.02214179 \times 10^{23}$, a fundamental constant named Avogadro's number (N_A) or the Avogadro constant in honor of Italian scientist Amedeo Avogadro. This constant is properly reported with an explicit unit of "per mole," a conveniently rounded version being $6.022 \times 10^{23}/\text{mol}$.

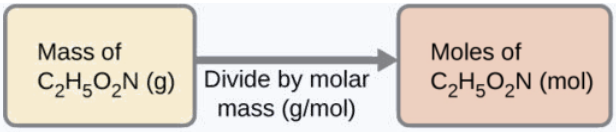
The relationships between formula mass, the mole, and Avogadro's number can be applied to compute various quantities that describe the composition of substances and compounds. In unit 2, we used Avogadro's number to convert between mass of an element, moles of an element, and number of atoms in the sample. Now, we can use Avogadro's number and formula mass (also called molecular weight) to convent between mass of a sample, moles, and number of molecules. For example, if we know the mass and chemical composition of a substance, we can determine the number of moles and calculate number of atoms or molecules in the sample. Likewise, if we know the number of moles of a substance, we can derive the number of atoms or molecules and calculate the substance's mass.s

Example 4.4.7: Deriving Moles from Grams for a Compound

Our bodies synthesize protein from amino acids. One of these amino acids is glycine, which has the molecular formula $C_2H_5O_2N$. How many moles of glycine molecules are contained in 28.35 g of glycine?

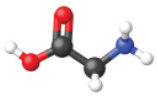
Solution

We can derive the number of moles of a compound from its mass following the same procedure we used for an element in Example 4.4.6:



The molar mass of glycine is required for this calculation, and it is computed in the same fashion as its molecular mass. One mole of glycine, $C_2H_5O_2N$, contains 2 moles of carbon, 5 moles of hydrogen, 2 moles of oxygen, and 1 mole of nitrogen:

Element	Quantity (mol element/ mol compound)		Molar mass (g/mol element)		Subtotal (g/mol compound)
C	2	×	12.01	=	24.02
H	5	×	1.008	=	5.040
O	2	×	16.00	=	32.00
N	1	×	14.007	=	14.007
Molecular mass (g/mol compound)					75.07



The provided mass of glycine (~28 g) is a bit more than one-third the molar mass (~75 g/mol), so we would expect the computed result to be a bit greater than one-third of a mole (~0.33 mol). Dividing the compound's mass by its molar mass yields:

$$28.35 \cancel{\text{ g}} \text{ glycine} \left(\frac{\text{mol glycine}}{75.07 \cancel{\text{ g}}} \right) = 0.378 \text{ mol glycine}$$

This result is consistent with our rough estimate.

Exercise 4.4.7

How many moles of sucrose, $C_{12}H_{22}O_{11}$, are in a 25-g sample of sucrose?

Answer

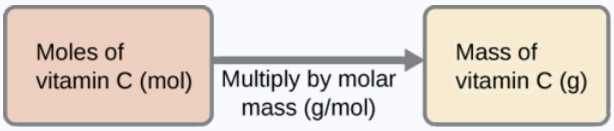
0.073 mol

Example 4.4.8: Deriving Grams from Moles for a Compound

Vitamin C is a covalent compound with the molecular formula $C_6H_8O_6$. The recommended daily dietary allowance of vitamin C for children aged 4–8 years is 1.42×10^{-4} mol. What is the mass of this allowance in grams?

Solution

As for elements, the mass of a compound can be derived from its molar amount as shown:



The molar mass for this compound is computed to be 176.124 g/mol. The given number of moles is a very small fraction of a mole (~10^{-4} or one-ten thousandth); therefore, we would expect the corresponding mass to be about one-ten thousandth of the molar mass (~0.02 g). Performing the calculation, we get:

$$1.42 \times 10^{-4} \cancel{\text{ mol}} \text{ vitamin C} \left(\frac{176.124 \text{ g}}{\cancel{\text{mol}} \text{ vitamin C}} \right) = 0.0250 \text{ g vitamin C}$$

This is consistent with the anticipated result.

Exercise 4.4.8

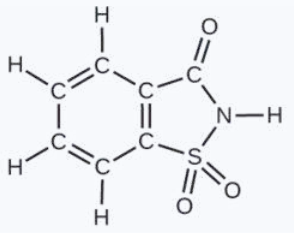
What is the mass of 0.443 mol of hydrazine, N_2H_4?

Answer

14.2 g

Example 4.4.9: Deriving the Number of Molecules from the Compound Mass

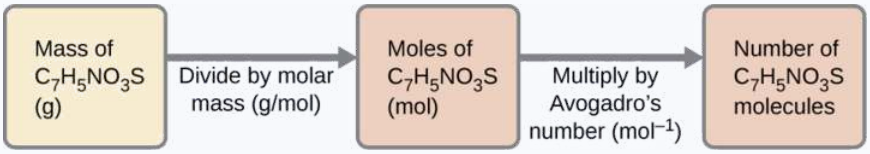
A packet of an artificial sweetener contains 40.0 mg of saccharin ($C_7H_5NO_3S$), which has the structural formula:



Given that saccharin has a molar mass of 183.18 g/mol, how many saccharin molecules are in a 40.0-mg (0.0400-g) sample of saccharin? How many carbon atoms are in the same sample?

Solution

The number of molecules in a given mass of compound is computed by first deriving the number of moles, as demonstrated in Example **4.4.8**, and then multiplying by Avogadro's number:



Using the provided mass and molar mass for saccharin yields:

$$0.0400\ \cancel{g}\ C_7H_5NO_3S \left(\frac{\cancel{mol}\ C_7H_5NO_3S}{183.18\ \cancel{g}\ C_7H_5NO_3S}\right) \left(\frac{6.022 \times 10^{23}\ C_7H_5NO_3S\ molecules}{1\ \cancel{mol}\ C_7H_5NO_3S}\right) = 1.31 \times 10^{20}\ C_7H_5NO_3S\ molecules \tag{4.4.1}$$

The compound's formula shows that each molecule contains seven carbon atoms, and so the number of C atoms in the provided sample is:

$$1.31 \times 10^{20}\ C_7H_5NO_3S\ molecules \left(\frac{7\ C\ atoms}{1\ C_7H_5NO_3S\ molecule}\right) = 9.20 \times 10^{21}\ C\ atoms$$

Exercise 4.4.9

How many C_4H_{10} molecules are contained in 9.213 g of this compound? How many hydrogen atoms?

Answer

- 9.545×10^{22} molecules C_4H_{10}
- 9.545×10^{23} atoms H



Video* 4.4.3: A preview of some of the uses we will have for moles in upcoming units*

Percent Composition

In the previous section, we discussed the relationship between the bulk mass of a substance and the number of atoms or molecules it contains (moles). Given the chemical formula of the substance, we were able to determine the amount of the substance (moles) from its mass, and vice versa. But what if the chemical formula of a substance is unknown? In this section, we will explore how to apply these very same principles in order to derive the chemical formulas of unknown substances from experimental mass measurements.

The elemental makeup of a compound defines its chemical identity, and chemical formulas are the most succinct way of representing this elemental makeup. When a compound's formula is unknown, measuring the mass of each of its constituent elements is often the first step in the process of determining the formula experimentally. The results of these measurements permit the calculation of the compound's percent composition, defined as the percentage by mass of each element in the compound. For example, consider a gaseous compound composed solely of carbon and hydrogen. The percent composition of this compound could be represented as follows:

$$\%\text{H} = \frac{\text{mass H}}{\text{mass compound}} \times 100\% \tag{4.4.2}$$

$$\%\text{C} = \frac{\text{mass C}}{\text{mass compound}} \times 100\% \tag{4.4.3}$$

If analysis of a 10.0-g sample of this gas showed it to contain 2.5 g H and 7.5 g C, the percent composition would be calculated to be 25% H and 75% C:

$$\%\text{H} = \frac{2.5\ \text{g H}}{10.0\ \text{g compound}} \times 100\% = 25\% \tag{4.4.4}$$

$$\%\text{C} = \frac{7.5\ \text{g C}}{10.0\ \text{g compound}} \times 100\% = 75\% \tag{4.4.5}$$

Example 4.4.10: Calculation of Percent Composition

Analysis of a 12.04-g sample of a liquid compound composed of carbon, hydrogen, and nitrogen showed it to contain 7.34 g C, 1.85 g H, and 2.85 g N. What is the percent composition of this compound?

Solution

To calculate percent composition, we divide the experimentally derived mass of each element by the overall mass of the compound, and then convert to a percentage:

$$\%\text{C} = \frac{7.34\ \text{g C}}{12.04\ \text{g compound}} \times 100\% = 61.0\%$$

$$\%\text{H} = \frac{1.85\ \text{g H}}{12.04\ \text{g compound}} \times 100\% = 15.4\%$$

$$\%\text{N} = \frac{2.85\ \text{g N}}{12.04\ \text{g compound}} \times 100\% = 23.7\%$$

The analysis results indicate that the compound is 61.0% C, 15.4% H, and 23.7% N by mass.

Exercise 4.4.10

A 24.81-g sample of a gaseous compound containing only carbon, oxygen, and chlorine is determined to contain 3.01 g C, 4.00 g O, and 17.81 g Cl. What is this compound's percent composition?

Answer

12.1% C, 16.1% O, 71.8% Cl

Determining Percent Composition from Formula Mass



Video* 4.4.4*: *A video overview of how to calculate percent composition of a compound based on its chemical formula.*

Percent composition is also useful for evaluating the relative abundance of a given element in different compounds of known formulas. As one example, consider the common nitrogen-containing fertilizers ammonia (NH_3), ammonium nitrate (NH_4NO_3), and urea (CH_4N_2O). The element nitrogen is the active ingredient for agricultural purposes, so the mass percentage of nitrogen in the compound is a practical and economic concern for consumers choosing among these fertilizers. For these sorts of applications, the percent composition of a compound is easily derived from its formula mass and the atomic masses of its constituent elements. A molecule of NH_3 contains one N atom weighing 14.01 amu and three H atoms weighing a total of (3 × 1.008 amu) = 3.024 amu. The formula mass of ammonia is therefore (14.01 amu + 3.024 amu) = 17.03 amu, and its percent composition is:

$$\%\text{N} = \frac{14.01\ \text{amu N}}{17.03\ \text{amu NH}_3} \times 100\% = 82.27\% \tag{4.4.6}$$

$$\%\text{H} = \frac{3.024\ \text{amu N}}{17.03\ \text{amu NH}_3} \times 100\% = 17.76\% \tag{4.4.7}$$

This same approach may be taken considering a pair of molecules, a dozen molecules, or a mole of molecules, etc. The latter amount is most convenient and would simply involve the use of molar masses instead of atomic and formula masses, as demonstrated Example **4.4.2**. As long as we know the chemical formula of the substance in question, we can easily derive percent composition from the formula mass or molar mass.

Example 4.4.11: Determining Percent Composition from a Molecular Formula

Aspirin is a compound with the molecular formula $C_9H_8O_4$. What is its percent composition?

Solution

To calculate the percent composition, we need to know the masses of C, H, and O in a known mass of $C_9H_8O_4$. It is convenient to consider 1 mol of $C_9H_8O_4$ and use its molar mass (180.159 g/mole, determined from the chemical formula) to calculate the percentages of each of its elements:

$$\%C = \frac{9\ \text{mol C} \times \text{molar mass C}}{\text{molar mass } C_9H_{18}O_4} \times 100$$
$$= \frac{9 \times 12.01\ \text{g/mol}}{180.159\ \text{g/mol}} \times 100 = \frac{108.09\ \text{g/mol}}{180.159\ \text{g/mol}} \times 100$$
$$\%C = 60.00\,\%\,C$$

$$\%H = \frac{8\ \text{mol H} \times \text{molar mass H}}{\text{molar mass } C_9H_{18}O_4} \times 100$$
$$= \frac{8 \times 1.008\ \text{g/mol}}{180.159\ \text{g/mol}} \times 100 = \frac{8.064\ \text{g/mol}}{180.159\ \text{g/mol}} \times 100$$
$$\%H = 4.476\,\%\,H$$

$$\%O = \frac{4\ \text{mol O} \times \text{molar mass O}}{\text{molar mass } C_9H_{18}O_4} \times 100$$
$$= \frac{4 \times 16.00\ \text{g/mol}}{180.159\ \text{g/mol}} \times 100 = \frac{64.00\ \text{g/mol}}{180.159\ \text{g/mol}} \times 100$$
$$\%O = 35.52\%$$

Note that these percentages sum to equal 100.00% when appropriately rounded.

Exercise 4.4.11

To three significant digits, what is the mass percentage of iron in the compound Fe_2O_3?

Answer

69.9% Fe

Summary

The formula mass of a substance is the sum of the average atomic masses of each atom represented in the chemical formula and is expressed in atomic mass units. The formula mass of a covalent compound is also called the molecular mass. A convenient amount unit for expressing very large numbers of atoms or molecules is the mole. Experimental measurements have determined the number of entities composing 1 mole of substance to be 6.022×10^{23}, a quantity called Avogadro's number. The mass in grams of 1 mole of substance is its molar mass. Due to the use of the same reference substance in defining the atomic mass unit and the mole, the formula mass (amu) and molar mass (g/mol) for any substance are numerically equivalent (for example, one H_2O molecule weighs approximately 18 amu and 1 mole of H_2O molecules weighs approximately 18 g).

Footnotes

1. 1 Omiatek, Donna M., Amanda J. Bressler, Ann-Sofie Cans, Anne M. Andrews, Michael L. Heien, and Andrew G. Ewing. "The Real Catecholamine Content of Secretory Vesicles in the CNS Revealed by Electrochemical Cytometry." *Scientific Report* 3 (2013): 1447, accessed January 14, 2015, doi:10.1038/srep01447.
2. Read more about the redefinition of SI units including the kilogram here (Laura Howe, *CE&N*, Nov. 16, 2018).

Glossary

Avogadro's number (N_A)

experimentally determined value of the number of entities comprising 1 mole of substance, equal to $6.022 \times 10^{23}\ \text{mol}^{-1}$

formula mass

sum of the average masses for all atoms represented in a chemical formula; for covalent compounds, this is also the molecular mass

mole

amount of substance containing the same number of atoms, molecules, ions, or other entities as the number of atoms in exactly 12 grams of ^{12}C

molar mass

mass in grams of 1 mole of a substance

percent composition

percentage by mass of the various elements in a compound

Contributors

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4.4.1: Practice Problems- Formula Mass, Percent Composition, and the Mole

PROBLEM 4.4.1.1

What is the total mass (amu) of ***carbon*** in each of the following molecules?

(a) CH_4
(b) $CHCl_3$
(c) $C_{12}H_{10}O_6$
(d) $CH_3CH_2CH_2CH_2CH_3$

Answer a

12.01 amu

Answer b

12.01 amu

Answer c

144.12 amu

Answer d

60.05 amu

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PROBLEM 4.4.1.2

What is the total mass of ***hydrogen*** in each of the molecules?

(a) CH_4
(b) $CHCl_3$
(c) $C_{12}H_{10}O_6$
(d) $CH_3CH_2CH_2CH_2CH_3$

Answer a

4.032 amu

Answer b

1.008 amu

Answer c

10.08 amu

Answer d

12.096 amu

PROBLEM 4.4.1.3

Calculate the molecular or formula mass of each of the following:

(a) P_4
(b) H_2O
(c) $Ca(NO_3)_2$
(d) CH_3CO_2H (acetic acid)
(e) $C_{12}H_{22}O_{11}$ (sucrose, cane sugar).

Answer a

123.896 amu

Answer b

18.015 amu

Answer c

164.086 amu

Answer d

60.052 amu

Answer e

342.297 amu

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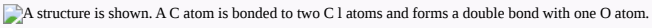


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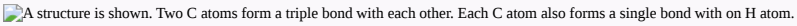
PROBLEM 4.4.1.4

Determine the molecular mass of the following compounds:

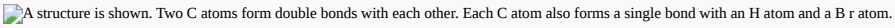
(a)


A structure is shown. A C atom is bonded to two C l atoms and forms a double bond with one O atom.

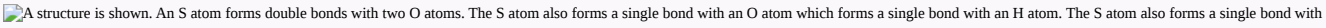
(b)


A structure is shown. Two C atoms form a triple bond with each other. Each C atom also forms a single bond with on H atom.

(c)


A structure is shown. Two C atoms form double bonds with each other. Each C atom also forms a single bond with an H atom and a B r atom.

(d)


A structure is shown. An S atom forms double bonds with two O atoms. The S atom also forms a single bond with an H atom. The S atom also forms a single bond with another O atom which forms a single bond with another H atom.

Answer a

98.906 amu

Answer b

26.018 amu

Answer c

185.818 amu

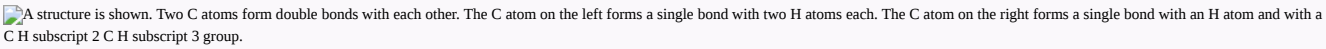
Answer d

98.072 amu

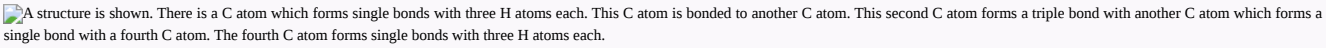
PROBLEM 4.4.1.5

Determine the molecular mass of the following compounds:

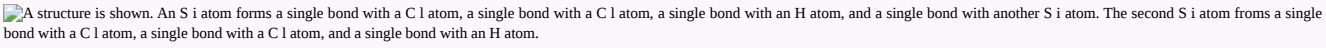
(a)


A structure is shown. Two C atoms form double bonds with each other. The C atom on the left forms a single bond with two H atoms each. The C atom on the right forms a single bond with an H atom and with a C H subscript 2 C H subscript 3 group.

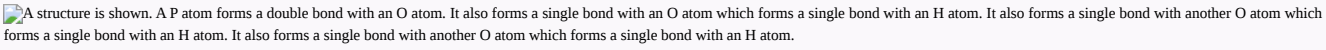
(b)


A structure is shown. There is a C atom which forms single bonds with three H atoms each. This C atom is bonded to another C atom. This second C atom forms a triple bond with another C atom which forms a single bond with a fourth C atom. The fourth C atom forms single bonds with three H atoms each.

(c)


A structure is shown. An S i atom forms a single bond with a C l atom, a single bond with a C l atom, a single bond with an H atom, and a single bond with another S i atom. The second S i atom froms a single bond with a C l atom, a single bond with a C l atom, and a single bond with an H atom.

(d)


A structure is shown. A P atom forms a double bond with an O atom. It also forms a single bond with an O atom which forms a single bond with an H atom. It also forms a single bond with another O atom which forms a single bond with an H atom. It also forms a single bond with another O atom which forms a single bond with an H atom.

Answer a

56.107 amu

Answer b

54.091 amu

Answer c

199.9976 amu

Answer d

97.9950 amu

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PROBLEM 4.4.1.6

Compare 1 mole of H_2, 1 mole of O_2, and 1 mole of F_2.

(a) Which has the largest number of molecules? Explain why.

(b) Which has the greatest mass? Explain why.

Answer a

1 mole is always 6.022 x 10^{23} molecules. They have the same number of molecules.

Answer b

F_2; it has the highest molar mass

PROBLEM 4.4.1.7

Which contains the greatest mass of oxygen: 0.75 mol of ethanol (C_2H_5OH), 0.60 mol of formic acid (HCO_2H), or 1.0 mol of water (H_2O)? Explain why.

Answer

Formic acid. Its formula has twice as many oxygen atoms as the other two compounds (one each). Therefore, 0.60 mol of formic acid would be equivalent to 1.20 mol of a compound containing a single oxygen atom.

PROBLEM 4.4.1.8

Calculate the molar mass of each of the following:

(a) S_8
(b) C_5H_{12}
(c) $Sc_2(SO_4)_3$
(d) CH_3COCH_3 (acetone)
(e) $C_6H_{12}O_6$ (glucose)

Answer a

256.528 g/mol

Answer b

72.150 g/mol

Answer c

378.103 g/mol

Answer d

58.080 g/mol

Answer e

180.158 g/mol

PROBLEM 4.4.1.9

Calculate the molar mass of each of the following:

(a) the anesthetic halothane, $C_2HBrClF_3$
(b) the herbicide paraquat, $C_{12}H_{14}N_2Cl_2$
(c) caffeine, $C_8H_{10}N_4O_2$
(d) urea, $CO(NH_2)_2$
(e) a typical soap, $C_{17}H_{35}CO_2Na$

Answer a

197.382 g/mol

Answer b

257.163 g/mol

Answer c

194.193 g/mol

Answer d

60.056 g/mol

Answer e

306.464 g/mol

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PROBLEM 4.4.1.10

Determine the number of mo he following:

(a) 25.0 g of propylene
(b) 3.06×10^{-3} g of th
(c) 25 lb of the herbici
(d) 0.125 kg of the ins
(e) 325 mg of aspirin,

Answer a

0.595 mol C_3H_6

1.78 mol C

3.57 mol H

Answer b

4.08×10^{-5} mol $C_2H_5NO_2$

8.16×10^{-5} mol C

2.04×10^{-4} mol H

4.08×10^{-5} mol N

8.16×10^{-5} mol O

Answer c

40.106 mol $C_{13}H_{16}N_2O_4F$

521.37 mol C

641.70 mol H

80.212 mol N

160.42 mol O

40.106 mol F

Answer d

0.202 mol $Cu_4(AsO_3)_2(CH_3CO_2)_2$

0.808 mol Cu

0.404 mol As

2.02 mol O

0.808 mol C

1.212 mol H

Answer e

0.00181 mol $C_6H_4(CO_2H)(CO_2CH_3)$

0.01625 mol C

0.0144 mol H

0.0072 mol O

PROBLEM 4.4.1.11

Determine the mass of each of the following:

(a) 0.0146 mol KOH
(b) 10.2 mol ethane, C_2H_6
(c) 1.6×10^{-3} mol $Na_2 SO_4$
(d) 6.854×10^3 mol glucose, $C_6 H_{12} O_6$
(e) 2.86 mol $Co(NH_3)_6Cl_3$

Answer a

0.819 g

Answer b

307 g

Answer c

0.23 g

Answer d

1.235×10^6 g (1235 kg)

Answer e

765 g

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PROBLEM 4.4.1.12

Determine the number of moles of the compound and determine the number of moles of each type of atom in each of the following:

(a) 2.12 g of potassium bromide, KBr
(b) 0.1488 g of phosphoric acid, H_3PO_4
(c) 23 kg of calcium carbonate, $CaCO_3$
(d) 78.452 g of aluminum sulfate, $Al_2(SO_4)_3$
(e) 0.1250 mg of caffeine, $C_8H_{10}N_4O_2$

Answer a

0.0178 mol KBr

0.0178 mol K

0.0178 mol Br

Answer b

0.00152 mol H_3PO_4

0.00455 mol H

0.00152 mol P

0.00607 mol O

Answer c

230 mol $CaCO_3$

230 mol Ca

230 mol C

690 mol O

Answer d

0.229 mol $Al_2(SO_4)_3$

0.459 mol Al

0.688 mol S

2.75 mol O

Answer e

6.44×10^{-7} mol $C_8H_{10}N_4O_2$

5.15×10^{-6} mol C

6.44×10^{-6} mol H

2.58×10^{-6} mol N

1.29×10^{-6} mol O

PROBLEM 4.4.1.13

The approximate minimum daily dietary requirement of the amino acid leucine, $C_6H_{13}NO_2$, is 1.1 g. What is this requirement in moles?

Answer

0.0084 mol $C_6H_{13}NO_2$

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PROBLEM 4.4.1.14

Determine the mass in grams

(a) 0.600 mol of oxyge
(b) 0.600 mol of oxyg
(c) 0.600 mol of ozone

Answer a

9.60 g

Answer b

19.2 g

Answer c

28.2 g

PROBLEM 4.4.1.15

A 55-kg woman has 7.5×10^{-3} mol of hemoglobin (molar mass = 64,456 g/mol) in her blood. How many hemoglobin molecules is this? What is this quantity in grams?

Answer

4.52×10^{21} molecules of hemoglobin

483.42 g of hemoglobin

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PROBLEM 4.4.1.16

One 55-gram serving of a particular cereal supplies 270 mg of sodium, 11% of the recommended daily allowance. How many moles and atoms of sodium are in the **recommended daily allowance**?

Answer

0.107 mol Na; 6.45×10^{22} atoms Na

PROBLEM 4.4.1.17

A tube of toothpaste contains 0.76 g of sodium monofluorophosphate (Na_2PO_3F) in 100 mL.

a. What mass of fluorine atoms in mg was present?
b. How many fluorine atoms were present?

Answer a

100.27 mg F

Answer b

3.18×10^{21} atoms F

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PROBLEM 4.4.1.18

Which of the following represents the least number of molecules?

a. 20.0 g of H_2O (18.02 g/mol)
b. 77.0 g of CH_4 (16.06 g/mol)
c. 68.0 g of CaH_2 (42.09 g/mol)
d. 100.0 g of N_2O (44.02 g/mol)
e. 84.0 g of HF (20.01 g/mol)

Answer

20.0 g of H_2O represents the smallest number of moles, meaning the least number of molecules present. Since 1 mole = 6.022×10^{23} molecules (or atoms) regardless of identity, the least number of moles will equal the least number of molecules

PROBLEM 4.4.1.19

Calculate the following:

(a) the percent composition of ammonia, NH_3
(b) the percent composition of photographic "hypo," $Na_2S_2O_3$
(c) the percent of calcium ion in $Ca_3(PO_4)_2$

Answer a

% N = 82.24%

% H = 17.76%

Answer b

% Na = 29.08%

% S = 40.56%

% O = 30.36%

Answer c

% Ca^{2+} = 38.76%

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PROBLEM 4.4.1.20

A compound of carbon and hydrogen contains 92.3% C and has a molar mass of 78.1 g/mol. What is its molecular formula?

Answer

C_6H6

PROBLEM 4.4.1.21

Dichloroethane, a compound that is often used for dry cleaning, contains carbon, hydrogen, and chlorine. It has a molar mass of 99 g/mol. Analysis of a sample shows that it contains 24.3% carbon and 4.1% hydrogen. What is its molecular formula?

Answer

$C_2H_4Cl_2$

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In the second edition, th[redacted]rrect, but the solution is correct.

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4.5: Empirical and Molecular Formulas

Skills to Develop

- Determine the empirical formula of a compound
- Determine the molecular formula of a compound

Determination of Empirical Formulas

As previously mentioned, the most common approach to determining a compound's chemical formula is to first measure the masses of its constituent elements. However, we must keep in mind that chemical formulas represent the relative *numbers*, not masses, of atoms in the substance. Therefore, any experimentally derived data involving mass must be used to derive the corresponding numbers of atoms in the compound. To accomplish this, we can use molar masses to convert the mass of each element to a number of moles. We then consider the moles of each element relative to each other, converting these numbers into a whole-number ratio that can be used to derive the empirical formula of the substance. Consider a sample of compound determined to contain 1.71 g C and 0.287 g H. The corresponding numbers of atoms (in moles) are:

$$1.71\ \text{g C} \times \frac{1\ \text{mol C}}{12.01\ \text{g C}} = 0.142\ \text{mol C} \tag{4.5.1}$$

$$0.287\ \text{g H} \times \frac{1\ \text{mol H}}{1.008\ \text{g H}} = 0.284\ \text{mol H} \tag{4.5.2}$$

Thus, we can accurately represent this compound with the formula $C_{0.142}H_{0.284}$. Of course, per accepted convention, formulas contain whole-number subscripts, which can be achieved by dividing each subscript by the smaller subscript:

$$\text{C}_{\frac{0.142}{0.142}}\ \text{H}_{\frac{0.284}{0.142}}\ \text{or}\ \text{CH}_2 \tag{4.5.3}$$

(Recall that subscripts of "1" are not written, but rather assumed if no other number is present.)

The empirical formula for this compound is thus CH_2. This may or not be the compound's *molecular formula* as well; however, we would need additional information to make that determination (as discussed later in this section).

Consider as another example a sample of compound determined to contain 5.31 g Cl and 8.40 g O. Following the same approach yields a tentative empirical formula of:

$$\text{Cl}_{0.150}\text{O}_{0.525} = \text{Cl}_{\frac{0.150}{0.150}}\ \text{O}_{\frac{0.525}{0.150}} = \text{ClO}_{3.5} \tag{4.5.4}$$

In this case, dividing by the smallest subscript still leaves us with a decimal subscript in the empirical formula. To convert this into a whole number, we must multiply each of the subscripts by two, retaining the same atom ratio and yielding Cl_2O_7 as the final empirical formula.

Procedure

In summary, empirical formulas are derived from experimentally measured element masses by:

1. Deriving the number of moles of each element from its mass
2. Dividing each element's molar amount by the smallest molar amount to yield subscripts for a tentative empirical formula
3. Multiplying all coefficients by an integer, if necessary, to ensure that the smallest whole-number ratio of subscripts is obtained

Figure 4.5.1 outlines this procedure in flow chart fashion for a substance containing elements A and X.

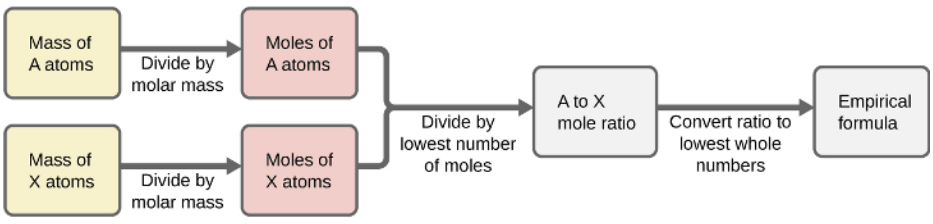


***Figure* 4.5.1**: *The empirical formula of a compound can be derived from the masses of all elements in the sample.*

Example 4.5.3: Determining an Empirical Formula from Masses of Elements

A sample of the black mineral hematite (Figure **4.5.2**), an oxide of iron found in many iron ores, contains 34.97 g of iron and 15.03 g of oxygen. What is the empirical formula of hematite?



***Figure* 4.5.2**: *Hematite is an iron oxide that is used in jewelry. (credit: Mauro Cateb)*

Solution

For this problem, we are given the mass in grams of each element. Begin by finding the moles of each:

$$34.97\ \text{g Fe}\left(\frac{\text{mol Fe}}{55.85\ \text{g}}\right) = 0.6261\ \text{mol Fe}$$

$$15.03\ \text{g O}\left(\frac{\text{mol O}}{16.00\ \text{g}}\right) = 0.9394\ \text{mol O}$$

Next, derive the iron-to-oxygen molar ratio by dividing by the lesser number of moles:

$$\frac{0.6261}{0.6261} = 1.000\ \text{mol Fe}$$

$$\frac{0.9394}{0.6261} = 1.500\ \text{mol O}$$

The ratio is 1.000 mol of iron to 1.500 mol of oxygen ($Fe_1O_{1.5}$). Finally, multiply the ratio by two to get the smallest possible whole number subscripts while still maintaining the correct iron-to-oxygen ratio:

$$2(Fe_1O_{1.5}) = Fe_2O_3$$

The empirical formula is Fe_2O_3.

Exercise 4.5.3

What is the empirical formula of a compound if a sample contains 0.130 g of nitrogen and 0.370 g of oxygen?

Answer

N_2O_5



***Video* 4.5.2**: *Additional worked examples illustrating the derivation of empirical formulas are presented in the brief video clip.*

Deriving Empirical Formulas from Percent Composition

Finally, with regard to deriving empirical formulas, consider instances in which a compound's percent composition is available rather than the absolute masses of the compound's constituent elements. In such cases, the percent composition can be used to calculate the masses of elements present in any convenient mass of compound; these masses can then be used to derive the empirical formula in the usual fashion.

Example 4.5.4: Determining an Empirical Formula from Percent Composition

The bacterial fermentation of grain to produce ethanol forms a gas with a percent composition of 27.29% C and 72.71% O (Figure **4.5.3**). What is the empirical formula for this gas?



***Figure* 4.5.3**: *An oxide of carbon is removed from these fermentation tanks through the large copper pipes at the top. (credit: "Dual Freq"/Wikimedia Commons)*

Solution

Since the scale for percentages is 100, it is most convenient to calculate the mass of elements present in a sample weighing 100 g. The calculation is "most convenient" because, per the definition for percent composition, the mass of a given element in grams is numerically equivalent to the element's mass percentage. This numerical equivalence results from the definition of the "percentage" unit, whose name is derived from the Latin phrase *per centum* meaning "by the hundred." Considering this definition, the mass percentages provided may be more conveniently expressed as fractions:

$$27.29\,\%\,\mathrm{C} = \frac{27.29\ \mathrm{g\ C}}{100\ \mathrm{g\ compound}}$$

$$72.71\,\%\,\mathrm{O} = \frac{72.71\ \mathrm{g\ O}}{100\ \mathrm{g\ compound}}$$

The molar amounts of carbon and hydrogen in a 100-g sample are calculated by dividing each element's mass by its molar mass:

$$27.29\ \text{g C}\left(\frac{\text{mol C}}{12.01\ \text{g}}\right) = 2.272\ \text{mol C}$$

$$72.71\ \text{g O}\left(\frac{\text{mol O}}{16.00\ \text{g}}\right) = 4.544\ \text{mol O}$$

Coefficients for the tentative empirical formula are derived by dividing each molar amount by the lesser of the two:

$$\frac{2.272\ \text{mol C}}{2.272} = 1$$

$$\frac{4.544\ \text{mol O}}{2.272} = 2$$

Since the resulting ratio is one carbon to two oxygen atoms, the empirical formula is CO_2.

Exercise 4.5.4

What is the empirical formula of a compound containing 40.0% C, 6.71% H, and 53.28% O?

Answer

CH_2O

Derivation of Molecular Formulas



***Video* 4.5.3**: *A review of calculating empirical formula from percent composition and an explanation of deriving molecular formula.*

Recall that empirical formulas are symbols representing the *relative* numbers of a compound's elements. Determining the *absolute* numbers of atoms that compose a single molecule of a covalent compound requires knowledge of both its empirical formula and its molecular mass or molar mass. These quantities may be determined experimentally by various measurement techniques. Molecular mass, for example, is often derived from the mass spectrum of the compound (see discussion of this technique in the previous chapter on atoms and molecules). Molar mass can be measured by a number of experimental methods, many of which will be introduced in later chapters of this text.

Molecular formulas are derived by comparing the compound's molecular or molar mass to its empirical formula mass. As the name suggests, an empirical formula mass is the sum of the average atomic masses of all the atoms represented in an empirical formula. If we know the molecular (or molar) mass of the substance, we can divide this by the empirical formula mass in order to identify the number of empirical formula units per molecule, which we designate as *n*:

$$\frac{\text{molecular or molar mass}\left(\text{amu or }\frac{\text{g}}{\text{mol}}\right)}{\text{empirical formula mass}\left(\text{amu or }\frac{\text{g}}{\text{mol}}\right)} = n\ \text{formula units/molecule} \tag{4.5.5}$$

The molecular formula is then obtained by multiplying each subscript in the empirical formula by *n*, as shown by the generic empirical formula A_xB_y:

Double subscripts: use braces to clarify

For example, consider a covalent compound whose empirical formula is determined to be CH_2O. The empirical formula mass for this compound is approximately 30 amu (the sum of 12 amu for one C atom, 2 amu for two H atoms, and 16 amu for one O atom). If the compound's molecular mass is determined to be 180 amu, this indicates that molecules of this compound contain six times the number of atoms represented in the empirical formula:

$$\frac{180\ \text{amu/molecule}}{30\ \frac{\text{amu}}{\text{formula unit}}} = 6\ \text{formula units/molecule} \tag{4.5.6}$$

Molecules of this compound are then represented by molecular formulas whose subscripts are six times greater than those in the empirical formula:

$$(CH_2O)_6 = C_6H_{12}O_6 \tag{4.5.7}$$

Note that this same approach may be used when the molar mass (g/mol) instead of the molecular mass (amu) is used. In this case, we are merely considering one mole of empirical formula units and molecules, as opposed to single units and molecules.

Example 4.5.5: Determination of the Molecular Formula for Nicotine

Nicotine, an alkaloid in the nightshade family of plants that is mainly responsible for the addictive nature of cigarettes, contains 74.02% C, 8.710% H, and 17.27% N. If 40.57 g of nicotine contains 0.2500 mol nicotine, what is the molecular formula?

Solution

Determining the molecular formula from the provided data will require comparison of the compound's empirical formula mass to its molar mass. As the first step, use the percent composition to derive the compound's empirical formula. Assuming a convenient, a 100-g sample of nicotine yields the following molar amounts of its elements:

$$(74.02\ \text{g C})\left(\frac{1\ \text{mol C}}{12.01\ \text{g C}}\right) = 6.163\ \text{mol C} \tag{4.5.8}$$

$$(8.710\ \text{g H})\left(\frac{1\ \text{mol H}}{1.01\ \text{g H}}\right) = 8.624\ \text{mol H} \tag{4.5.9}$$

$$(17.27\ \text{g N})\left(\frac{1\ \text{mol N}}{14.01\ \text{g N}}\right) = 1.233\ \text{mol N} \tag{4.5.10}$$

Next, we calculate the molar ratios of these elements relative to the least abundant element, **N**.

6.163 mol C / **1.233 mol N** = 5

8.264 mol H / **1.233 mol N** = 7

1.233 mol N / **1.233 mol N** = 1

1.233/1.233 = **1.000 mol N**

6.163/1.233 = **4.998 mol C**

8.624/1.233 = **6.994 mol H**

The C-to-N and H-to-N molar ratios are adequately close to whole numbers, and so the empirical formula is C_5H_7N. The empirical formula mass for this compound is therefore 81.13 amu/formula unit, or 81.13 g/mol formula unit.

We calculate the molar mass for nicotine from the given mass and molar amount of compound:

$$\frac{40.57\ \text{g nicotine}}{0.2500\ \text{mol nicotine}} = \frac{162.3\ \text{g}}{\text{mol}}$$

Comparing the molar mass and empirical formula mass indicates that each nicotine molecule contains two formula units:

$$\frac{162.3\ \text{g/mol}}{81.13\ \frac{\text{g}}{\text{formula unit}}} = 2\ \text{formula units/molecule}$$

Thus, we can derive the molecular formula for nicotine from the empirical formula by multiplying each subscript by two:

$$(C_5H_7N)_2 = C_{10}H_{14}N_2$$

Exercise 4.5.5

What is the molecular formula of a compound with a percent composition of 49.47% C, 5.201% H, 28.84% N, and 16.48% O, and a molecular mass of 194.2 amu?

Answer

$C_8H_{10}N_4O_2$

Summary

The chemical identity of a substance is defined by the types and relative numbers of atoms composing its fundamental entities (molecules in the case of covalent compounds, ions in the case of ionic compounds). A compound's percent composition provides the mass percentage of each element in the compound, and it is often experimentally determined and used to derive the compound's empirical formula. The empirical formula mass of a covalent compound may be compared to the compound's molecular or molar mass to derive a molecular formula.

Key Equations

- $\%X = \dfrac{\text{mass X}}{\text{mass compound}} \times 100\%$
- $\dfrac{\text{molecular or molar mass}\left(\text{amu or } \dfrac{\text{g}}{\text{mol}}\right)}{\text{empirical formula mass}\left(\text{amu or } \dfrac{\text{g}}{\text{mol}}\right)} = n \text{ formula units/molecule}$
- $(A_xB_y)_n = A_{nx}B_{ny}$

Glossary

percent composition

percentage by mass of the various elements in a compound

empirical formula mass

sum of average atomic masses for all atoms represented in an empirical formula

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4.5.1: Practice Problems- Empirical and Molecular Formulas

PROBLEM 4.5.1.1

Determine the empirical formulas for compounds with the following percent compositions:

(a) 15.8% carbon and 84.2% sulfur
(b) 40.0% carbon, 6.7% hydrogen, and 53.3% oxygen

Answer a

CS_2

Answer b

CH_2O

Click here to see a video of the solution



PROBLEM 4.5.1.2

Determine the empirical and molecular formula for chrysotile asbestos. Chrysotile has the following percent composition: 28.03% Mg, 21.60% Si, 1.16% H, and 49.21% O. The molar mass for chrysotile is 520.8 g/mol.

Answer

$Mg_3Si_2H_3O_8$ (empirical formula), $Mg_6Si_4H_6O_{16}$ (molecular formula)

PROBLEM 4.5.1.3

Polymers are large molecules composed of simple units repeated many times. Thus, they often have relatively simple empirical formulas. Calculate the empirical formulas of the following polymers:

(a) Lucite (Plexiglas); 59.9% C, 8.06% H, 32.0% O
(b) Saran; 24.8% C, 2.0% H, 73.1% Cl
(c) polyethylene; 86% C, 14% H
(d) polystyrene; 92.3% C, 7.7% H
(e) Orlon; 67.9% C, 5.70% H, 26.4% N

Answer a

$C_5H_8O_2$

Answer b

CHCl

Answer c

CH_2

Answer d

CH

Answer e

C_3H_3N

Click here to see a video of the solution



PROBLEM 4.5.1.4

A major textile dye manufacturer developed a new yellow dye. The dye has a percent composition of 75.95% C, 17.72% N, and 6.33% H by mass with a molar mass of about 240 g/mol. Determine the molecular formula of the dye.

Answer

$C_{15}H_{15}N_3$

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- Adelaide Clark, Oregon Institute of Technology

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CHAPTER OVERVIEW

5: Transformations of Matter

Unit 5 Objectives

By the end of this unit, you will be able to:

- Derive chemical equations from narrative descriptions of chemical reactions.
- Write and balance chemical equations
- Explain the concept of stoichiometry as it pertains to chemical reactions
- Use balanced chemical equations to derive stoichiometric factors relating amounts of reactants and products
- Perform stoichiometric calculations involving mass and moles
- Explain the concepts of theoretical yield and limiting reactants/reagents.
- Derive the theoretical yield for a reaction under specified conditions.
- Calculate the percent yield for a reaction.

5.1: Writing and Balancing Chemical Equations

5.1.1: Practice Problems- Writing and Balancing Chemical Equations

5.2: Reaction Stoichiometry

5.2.1: Practice Problems- Reaction Stoichiometry

5.3: Calculating Reaction Yields

5.3.1: Practice Problems Calculating Reaction Yields

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5.1: Writing and Balancing Chemical Equations

Skills to Develop

- Derive chemical equations from narrative descriptions of chemical reactions.
- Write and balance chemical equations



Video* 5.1.1: An introduction to writing chemical equations.*

The preceding unit introduced the use of element symbols to represent individual atoms. When atoms gain or lose electrons to yield ions, or combine with other atoms to form molecules, their symbols are modified or combined to generate chemical formulas that appropriately represent these species. Extending this symbolism to represent both the identities and the relative quantities of substances undergoing a chemical (or physical) change involves writing and balancing a chemical equation. Consider as an example the reaction between one methane molecule (CH_4) and two diatomic oxygen molecules (O_2) to produce one carbon dioxide molecule (CO_2) and two water molecules (H_2O). The chemical equation representing this process is provided in the upper half of Figure **5.1.1**, with space-filling molecular models shown in the lower half of the figure.

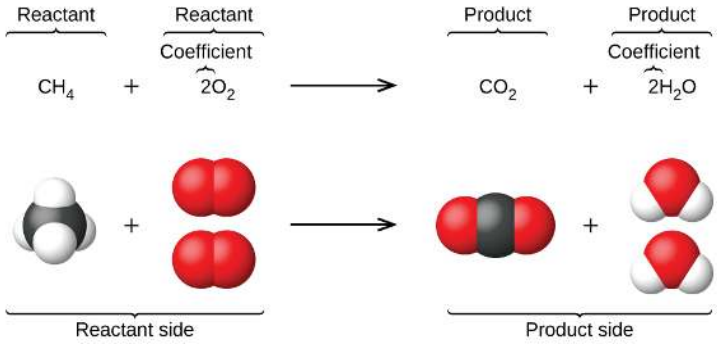


Figure* 5.1.1: The reaction between methane and oxygen to yield carbon dioxide and water (shown at bottom) may be represented by a chemical equation using formulas (top).*

This example illustrates the fundamental aspects of any chemical equation:

1. The substances undergoing reaction are called reactants, and their formulas are placed on the left side of the equation.
2. The substances generated by the reaction are called products, and their formulas are placed on the right sight of the equation.
3. Plus signs (+) separate individual reactant and product formulas, and an arrow (⟶) separates the reactant and product (left and right) sides of the equation.
4. The relative numbers of reactant and product species are represented by coefficients (numbers placed immediately to the left of each formula). A coefficient of 1 is typically omitted.

It is common practice to use the smallest possible whole-number coefficients in a chemical equation, as is done in this example. Realize, however, that these coefficients represent the *relative* numbers of reactants and products, and, therefore, they may be

correctly interpreted as ratios. Methane and oxygen react to yield carbon dioxide and water in a 1:2:1:2 ratio. This ratio is satisfied if the numbers of these molecules are, respectively, 1-2-1-2, or 2-4-2-4, or 3-6-3-6, and so on (Figure **5.1.2**). Likewise, these coefficients may be interpreted with regard to any amount (number) unit, and so this equation may be correctly read in many ways, including:

- *One* methane molecule and *two* oxygen molecules react to yield *one* carbon dioxide molecule and *two* water molecules.
- *One dozen* methane molecules and *two dozen* oxygen molecules react to yield *one dozen* carbon dioxide molecules and *two dozen* water molecules.
- *One mole* of methane molecules and *2 moles* of oxygen molecules react to yield *1 mole* of carbon dioxide molecules and *2 moles* of water molecules.

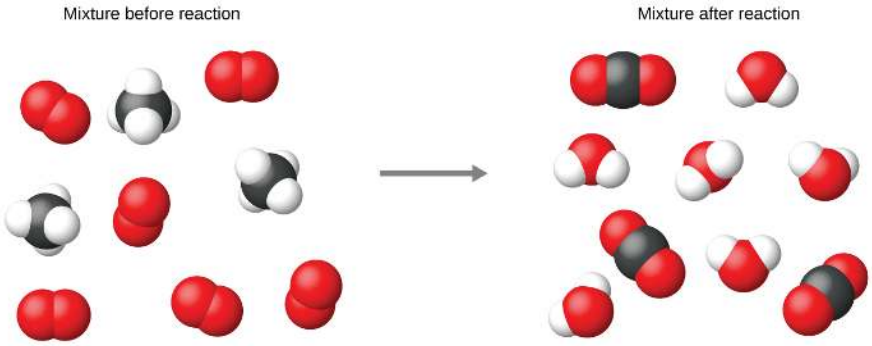


***Figure* 5.1.2**: *Regardless of the absolute number of molecules involved, the ratios between numbers of molecules of each species that react (the reactants) and molecules of each species that form (the products) are the same and are given in the chemical equation.*

Balancing Equations



***Video* 5.1.2**: *An introduction to counting atoms in a chemical equation.*

When a chemical equation is balanced it means that equal numbers of atoms for each element involved in the reaction are represented on the reactant and product sides. This is a requirement the equation must satisfy to be consistent with the law of conservation of matter. It may be confirmed by simply summing the numbers of atoms on either side of the arrow and comparing these sums to ensure they are equal. Note that the number of atoms for a given element is calculated by multiplying the coefficient of any formula containing that element by the element's subscript in the formula. If an element appears in more than one formula on a given side of the equation, the number of atoms represented in each must be computed and then added together. For example, both product species in the example reaction, $\mathbf{CO_2}$ and $\mathbf{H_2O}$, contain the element oxygen, and so the number of oxygen atoms on the product side of the equation is

$$\left(1\ \cancel{CO_2\ \text{molecule}} \times \frac{2\ \text{O atoms}}{\cancel{CO_2\ \text{molecule}}}\right) + \left(\cancel{2\,H_2O\ \text{molecule}} \times \frac{1\ \text{O atom}}{\cancel{H_2O\ \text{molecule}}}\right) = 4\ \text{O atoms} \tag{5.1.1}$$

The equation for the reaction between methane and oxygen to yield carbon dioxide and water is confirmed to be balanced per this approach, as shown here:

$$CH_4 + 2\,O_2 \rightarrow CO_2 + 2\,H_2O \tag{5.1.2}$$

Element	Reactants	Products	Balanced?
C	1 × 1 = 1	1 × 1 = 1	1 = 1, yes
H	4 × 1 = 4	2 × 2 = 4	4 = 4, yes
O	2 × 2 = 4	(1 × 2) + (2 × 1) = 4	4 = 4, yes

A balanced chemical equation often may be derived from a qualitative description of some chemical reaction by a fairly simple approach known as balancing by inspection. Consider as an example the decomposition of water to yield molecular hydrogen and oxygen. This process is represented qualitatively by an *unbalanced* chemical equation:

$$H_2O \rightarrow H_2 + O_2 \quad \text{(unbalanced)}$$

Comparing the number of H and O atoms on either side of this equation confirms its imbalance:

Element	Reactants	Products	Balanced?
H	1 × 2 = 2	1 × 2 = 2	2 = 2, yes
O	1 × 1 = 1	1 × 2 = 2	1 ≠ 2, no

The numbers of H atoms on the reactant and product sides of the equation are equal, but the numbers of O atoms are not. To achieve balance, the *coefficients* of the equation may be changed as needed. Keep in mind, of course, that the *formula subscripts* define, in part, the identity of the substance, and so these cannot be changed without altering the qualitative meaning of the equation. For example, changing the reactant formula from H_2O to H_2O_2 would yield balance in the number of atoms, but doing so also changes the reactant's identity (it's now hydrogen peroxide and not water). The O atom balance may be achieved by changing the coefficient for H_2O to 2.

$$2\,H_2O \rightarrow H_2 + O_2 \quad \text{(unbalanced)}$$

Element	Reactants	Products	Balanced?
H	**2** × 2 = 4	1 × 2 = 2	4 ≠ 2, no
O	2 × 1 = 2	1 × 2 = 2	2 = 2, yes

The H atom balance was upset by this change, but it is easily reestablished by changing the coefficient for the H_2 product to 2.

$$2\,H_2O \rightarrow 2\,H_2 + O_2 \quad \text{(balanced)}$$

Element	Reactants	Products	Balanced?
H	2 × 2 = 4	**2** × 2 = 2	4 = 4, yes
O	2 × 1 = 2	1 × 2 = 2	2 = 2, yes

These coefficients yield equal numbers of both H and O atoms on the reactant and product sides, and the balanced equation is, therefore:

$$2\,H_2O \rightarrow 2\,H_2 + O_2 \tag{5.1.3}$$



***Video* 5.1.3**: *A review of balancing chemical equations, with extra practice problems.*

Example 5.1.1: Balancing Chemical Equations

Write a balanced equation for the reaction of molecular nitrogen (N_2) and oxygen (O_2) to form dinitrogen pentoxide.

Solution

First, write the unbalanced equation.

$$N_2 + O_2 \rightarrow N_2O_5 \qquad \text{(unbalanced)}$$

Next, count the number of each type of atom present in the unbalanced equation.

Element	Reactants	Products	Balanced?
N	1 × 2 = 2	1 × 2 = 2	2 = 2, yes
O	1 × 2 = 2	1 × 5 = 5	2 ≠ 5, no

Though nitrogen is balanced, changes in coefficients are needed to balance the number of oxygen atoms. To balance the number of oxygen atoms, a reasonable first attempt would be to change the coefficients for the O_2 and N_2O_5 to integers that will yield 10 O atoms (the least common multiple for the O atom subscripts in these two formulas).

$$N_2 + 5\,O_2 \rightarrow 2\,N_2O_5 \qquad \text{(unbalanced)}$$

Element	Reactants	Products	Balanced?
N	1 × 2 = 2	**2** × 2 = 4	2 ≠ 4, no
O	**5** × 2 = 10	**2** × 5 = 10	10 = 10, yes

The N atom balance has been upset by this change; it is restored by changing the coefficient for the reactant N_2 to 2.

$$2\,N_2 + 5\,O_2 \rightarrow 2\,N_2O_5 \tag{5.1.4}$$

Element	Reactants	Products	Balanced?
N	**2** × 2 = 4	2 × 2 = 4	4 = 4, yes
O	5 × 2 = 10	2 × 5 = 10	10 = 10, yes

The numbers of N and O atoms on either side of the equation are now equal, and so the equation is balanced.

Exercise 5.1.1

Write a balanced equation for the decomposition of ammonium nitrate to form molecular nitrogen, molecular oxygen, and water. (Hint: Balance oxygen last, since it is present in more than one molecule on the right side of the equation.)

Answer

$$2\,NH_4NO_3 \rightarrow 2\,N_2 + O_2 + 4\,H_2O$$

It is sometimes convenient to use fractions instead of integers as intermediate coefficients in the process of balancing a chemical equation. When balance is achieved, all the equation's coefficients may then be multiplied by a whole number to convert the fractional coefficients to integers without upsetting the atom balance. For example, consider the reaction of ethane (C_2H_6) with oxygen to yield H_2O and CO_2, represented by the unbalanced equation:

$$C_2H_6 + O_2 \rightarrow H_2O + CO_2 \quad \text{(unbalanced)}$$

Following the usual inspection approach, one might first balance C and H atoms by changing the coefficients for the two product species, as shown:

$$C_2H_6 + O_2 \rightarrow 3\,H_2O + 2\,CO_2 \quad \text{(unbalanced)}$$

This results in seven O atoms on the product side of the equation, an odd number—no integer coefficient can be used with the O_2 reactant to yield an odd number, so a fractional coefficient, $\frac{7}{2}$, is used instead to yield a provisional balanced equation:

$$C_2H_6 + \tfrac{7}{2}\,O_2 \rightarrow 3\,H_2O + 2\,CO_2$$

A conventional balanced equation with integer-only coefficients is derived by multiplying each coefficient by 2:

$$2\,C_2H_6 + 7\,O_2 \rightarrow 6\,H_2O + 4\,CO_2 \quad (5.1.5)$$

Finally with regard to balanced equations, recall that convention dictates use of the *smallest whole-number coefficients*. Although the equation for the reaction between molecular nitrogen and molecular hydrogen to produce ammonia is, indeed, balanced,

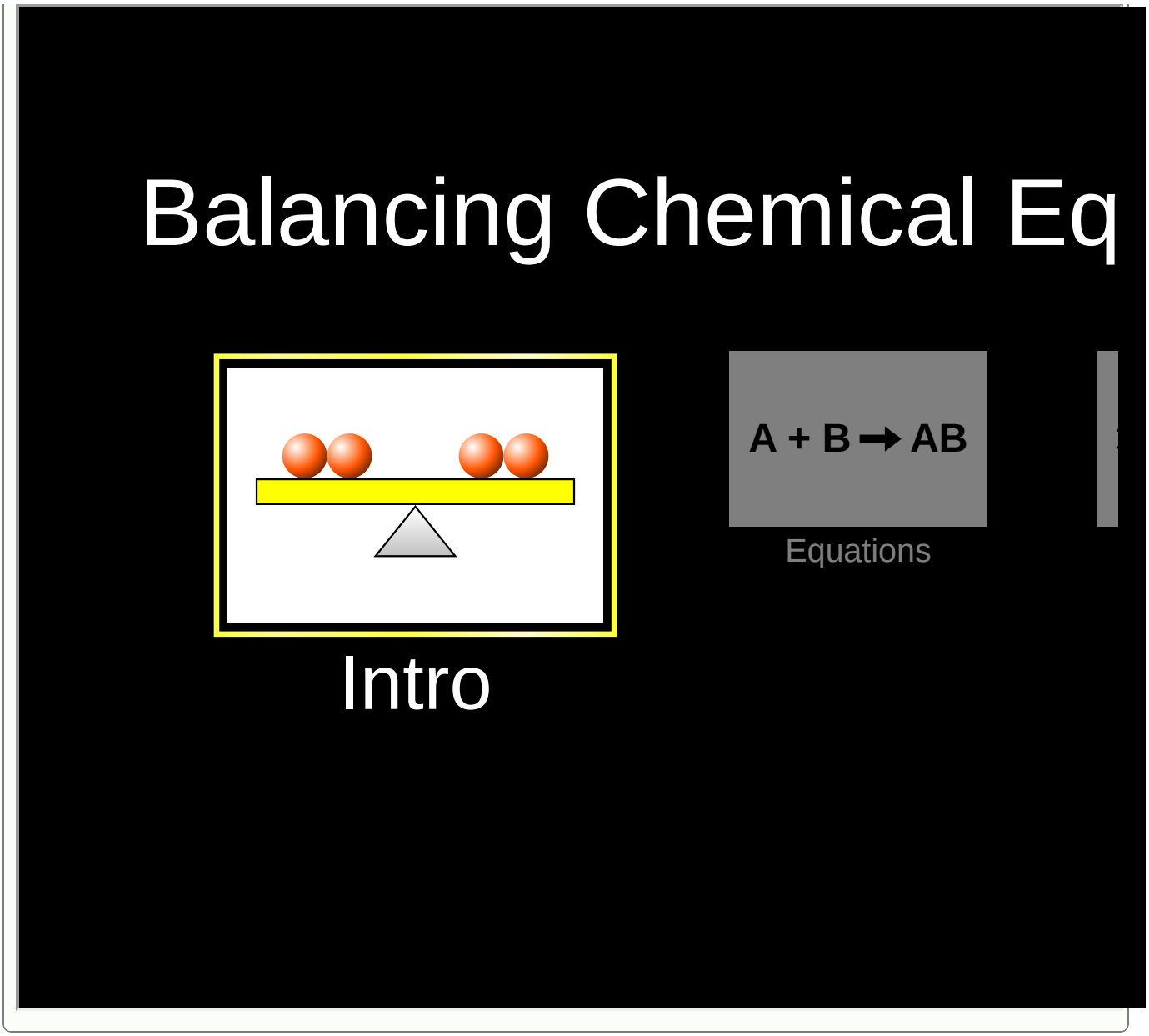
$$3\,N_2 + 9\,H_2 \rightarrow 6\,NH_3 \quad (5.1.6)$$

the coefficients are not the smallest possible integers representing the relative numbers of reactant and product molecules. Dividing each coefficient by the greatest common factor, 3, gives the preferred equation:

$$N_2 + 3\,H_2 \rightarrow 2\,NH_3 \quad (5.1.7)$$

Phet Simulation

Use this interactive tutorial for additional practice balancing equations.



State Symbols in Chemical Equations



***Video* 5.1.4**: *How to use state symbol to give more information in your chemical equations.*

The physical states of reactants and products in chemical equations very often are indicated with a parenthetical abbreviation following the formulas. Common abbreviations include *s* for solids, *l* for liquids, *g* for gases, and *aq* for substances dissolved in water (*aqueous solutions*, as introduced in the preceding chapter). These notations are illustrated in the example equation here:

$$2\,Na\,(s) + 2\,H_2O\,(l) \rightarrow 2\,NaOH\,(aq) + H_2\,(g) \tag{5.1.8}$$

This equation represents the reaction that takes place when sodium metal is placed in water. The solid sodium reacts with liquid water to produce molecular hydrogen gas and the ionic compound sodium hydroxide (a solid in pure form, but readily dissolved in water).

Special conditions necessary for a reaction are sometimes designated by writing a word or symbol above or below the equation's arrow. For example, a reaction carried out by heating may be indicated by the uppercase Greek letter delta (Δ) over the arrow.

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g) \tag{5.1.9}$$

Other examples of these special conditions will be encountered in more depth in later units.

Key Concepts and Summary



***Video* 5.1.5**: *An overview of balancing chemical equations in the context of stoichiometry.*

Chemical equations are symbolic representations of chemical and physical changes. Formulas for the substances undergoing the change (reactants) and substances generated by the change (products) are separated by an arrow and preceded by integer coefficients indicating their relative numbers. Balanced equations are those whose coefficients result in equal numbers of atoms for each element in the reactants and products. Chemical reactions in aqueous solution that involve ionic reactants or products may be represented more realistically by complete ionic equations and, more succinctly, by net ionic equations.

Glossary

balanced equation
chemical equation with equal numbers of atoms for each element in the reactant and product

chemical equation
symbolic representation of a chemical reaction

coefficient
number placed in front of symbols or formulas in a chemical equation to indicate their relative amount

molecular equation
chemical equation in which all reactants and products are represented as neutral substances

product

substance formed by a chemical or physical change; shown on the right side of the arrow in a chemical equation

reactant

substance undergoing a chemical or physical change; shown on the left side of the arrow in a chemical equation

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5.1.1: Practice Problems- Writing and Balancing Chemical Equations

PROBLEM 5.1.1.1

Balance the following equations:

a. $PCl_5(s) + H_2O(l) \rightarrow POCl_3(l) + HCl(aq)$
b. $Cu(s) + HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + H_2O(l) + NO(g)$
c. $H_2(g) + I_2(s) \rightarrow HI(s)$
d. $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$
e. $Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$
f. $(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + H_2O(g)$
g. $P_4(s) + Cl_2(g) \rightarrow PCl_3(l)$
h. $PtCl_4(s) \rightarrow Pt(s) + Cl_2(g)$

Answer a

$$PCl_5(s) + H_2O(l) \rightarrow POCl_3(l) + 2\,HCl(aq)$$

Answer b

$$3\,Cu(s) + 8\,HNO_3(aq) \rightarrow 3\,Cu(NO_3)_2(aq) + 4\,H_2O(l) + 2\,NO(g)$$

Answer c

$$H_2(g) + I_2(s) \rightarrow 2\,HI(s)$$

Answer d

$$4\,Fe(s) + 3\,O_2(g) \rightarrow 2\,Fe_2O_3(s)$$

Answer e

$$2\,Na(s) + 2\,H_2O(l) \rightarrow 2\,NaOH(aq) + H_2(g)$$

Answer f

$$(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + 4\,H_2O(g)$$

Answer g

$$P_4(s) + 6\,Cl_2(g) \rightarrow 4\,PCl_3(l)$$

Answer h

$$PtCl_4(s) \rightarrow Pt(s) + 2\,Cl_2(g)$$

Click here to see a video of the solution



PROBLEM 5.1.1.2

Balance the following equations:

a. $\mathrm{Ag}(s) + \mathrm{H_2S}(g) + \mathrm{O_2}(g) \rightarrow \mathrm{Ag_2S}(s) + \mathrm{H_2O}(l)$
b. $\mathrm{P_4}(s) + \mathrm{O_2}(g) \rightarrow \mathrm{P_4O_{10}}(s)$
c. $\mathrm{Pb}(s) + \mathrm{H_2O}(l) + \mathrm{O_2}(g) \rightarrow \mathrm{Pb(OH)_2}(s)$
d. $\mathrm{Fe}(s) + \mathrm{H_2O}(l) \rightarrow \mathrm{Fe_3O_4}(s) + \mathrm{H_2}(g)$
e. $\mathrm{Sc_2O_3}(s) + \mathrm{SO_3}(l) \rightarrow \mathrm{Sc_2(SO_4)_3}(s)$
f. $\mathrm{Ca_3(PO_4)_2}(aq) + \mathrm{H_3PO_4}(aq) \rightarrow \mathrm{Ca(H_2PO_4)_2}(aq)$
g. $\mathrm{Al}(s) + \mathrm{H_2SO_4}(aq) \rightarrow \mathrm{Al_2(SO_4)_3}(s) + \mathrm{H_2}(g)$
h. $\mathrm{TiCl_4}(s) + \mathrm{H_2O}(g) \rightarrow \mathrm{TiO_2}(s) + \mathrm{HCl}(g)$

Answer a

$$4\,\mathrm{Ag}(s) + 2\,\mathrm{H_2S}(g) + \mathrm{O_2}(g) \rightarrow 2\,\mathrm{Ag_2S}(s) + 2\,\mathrm{H_2O}(l)$$

Answer b

$$\mathrm{P_4}(s) + 5\,\mathrm{O_2}(g) \rightarrow \mathrm{P_4O_{10}}(s)$$

Answer c

$$2\,\mathrm{Pb}(s) + 2\,\mathrm{H_2O}(l) + \mathrm{O_2}(g) \rightarrow 2\,\mathrm{Pb(OH)_2}(s)$$

Answer d

$$3\,\mathrm{Fe}(s) + 4\,\mathrm{H_2O}(l) \rightarrow \mathrm{Fe_3O_4}(s) + 4\,\mathrm{H_2}(g)$$

Answer e

$$\mathrm{Sc_2O_3}(s) + 3\,\mathrm{SO_3}(l) \rightarrow \mathrm{Sc_2(SO_4)_3}(s)$$

Answer f

$$\mathrm{Ca_3(PO_4)_2}(aq) + 4\,\mathrm{H_3PO_4}(aq) \rightarrow 3\,\mathrm{Ca(H_2PO_4)_2}(aq)$$

Answer g

$$2\,\mathrm{Al}(s) + 3\,\mathrm{H_2SO_4}(aq) \rightarrow \mathrm{Al_2(SO_4)_3}(s) + 3\,\mathrm{H_2}(g)$$

Answer h

$$\mathrm{TiCl_4}(s) + 2\,\mathrm{H_2O}(g) \rightarrow \mathrm{TiO_2}(s) + 4\,\mathrm{HCl}(g)$$

PROBLEM 5.1.1.3

Write a balanced molecular equation describing each of the following chemical reactions.

a. Solid calcium carbonate is heated and decomposes to solid calcium oxide and carbon dioxide gas.
b. Gaseous butane, C_4H_{10}, reacts with diatomic oxygen gas to yield gaseous carbon dioxide and water vapor.
c. Aqueous solutions of magnesium chloride and sodium hydroxide react to produce solid magnesium hydroxide and aqueous sodium chloride.
d. Water vapor reacts with sodium metal to produce solid sodium hydroxide and hydrogen gas.

Answer a

$$\mathrm{CaCO_3}(s) \rightarrow \mathrm{CaO}(s) + \mathrm{CO_2}(g)$$

Answer b

$$2\,C_4H_{10}(g) + 13\,O_2(g) \rightarrow 8\,CO_2(g) + 10\,H_2O(g)$$

Answer c

$$MgCl_2(aq) + 2\,NaOH(aq) \rightarrow Mg(OH)_2(s) + 2\,NaCl(aq)$$

Answer d

$$2\,H_2O(g) + 2\,Na(s) \rightarrow 2\,NaOH(s) + H_2(g)$$

Click here to see a video of the solution



PROBLEM 5.1.1.4

Write a balanced equation describing each of the following chemical reactions.

a. Solid potassium chlorate, $KClO_3$, decomposes to form solid potassium chloride and diatomic oxygen gas.
b. Solid aluminum metal reacts with solid diatomic iodine to form solid Al_2I_6.
c. When solid sodium chloride is added to aqueous sulfuric acid, hydrogen chloride gas and aqueous sodium sulfate are produced.
d. Aqueous solutions of phosphoric acid and potassium hydroxide react to produce aqueous potassium dihydrogen phosphate and liquid water.

Answer a

$$2\,KClO_3(s) \rightarrow 2\,KCl(s) + 3\,O_2(g)$$

Answer b

$$2\,Al(s) + 3\,I_2(s) \rightarrow Al_2I_6(s)$$

Answer c

$$2\,NaCl(s) + H_2SO_4(aq) \rightarrow 2\,HCl(g) + Na_2SO_4(aq)$$

Answer d

$$H_3PO_4(aq) + KOH(aq) \rightarrow KH_2PO_4(aq) + H_2O(l)$$

PROBLEM 5.1.1.5

Colorful fireworks often involve the decomposition of barium nitrate and potassium chlorate and the reaction of the metals magnesium, aluminum, and iron with oxygen.

a. Write the formulas of barium nitrate and potassium chlorate.
b. The decomposition of solid potassium chlorate leads to the formation of solid potassium chloride and diatomic oxygen gas. Write an equation for the reaction.
c. The decomposition of solid barium nitrate leads to the formation of solid barium oxide, diatomic nitrogen gas, and diatomic oxygen gas. Write an equation for the reaction.
d. Write separate equations for the reactions of the solid metals magnesium, aluminum, and iron with diatomic oxygen gas to yield the corresponding metal oxides. (Assume the iron oxide contains Fe^{3+} ions.)

Answer a

$Ba(NO_3)_2$, $KClO_3$

Answer b

$$2\,KClO_3(s) \rightarrow 2\,KCl(s) + 3\,O_2(g)$$

Answer c

$$2\,Ba(NO_3)_2(s) \rightarrow 2\,BaO(s) + 2\,N_2(g) + 5\,O_2(g)$$

Answer d

$$2\,Mg(s) + O_2(g) \rightarrow 2\,MgO(s)\ ;\ 4\,Al(s) + 3\,O_2(g) \rightarrow 2\,Al_2O_3(g);\ 4\,Fe(s) + 3\,O_2(g) \rightarrow 2\,Fe_2O_3(s)$$

Click here to see a video of the solution



PROBLEM 5.1.1.6

Aqueous hydrogen fluoride (hydrofluoric acid) is used to etch glass and to analyze minerals for their silicon content. Hydrogen fluoride will also react with sand (silicon dioxide).

a. Write an equation for the reaction of solid silicon dioxide with hydrofluoric acid to yield gaseous silicon tetrafluoride and liquid water.
b. The mineral fluorite (calcium fluoride) occurs extensively in Illinois. Solid calcium fluoride can also be prepared by the reaction of aqueous solutions of calcium chloride and sodium fluoride, yielding aqueous sodium chloride as the other product. Write the equation for this reaction.

Answer a

$$4\,HF(aq) + SiO_2(s) \rightarrow SiF_4(g) + 2\,H_2O(l)$$

Answer b

$$CaCl_2(aq) + 2\,NaF(aq) \rightarrow 2\,NaCl(aq) + CaF_2(s)$$

PROBLEM 5.1.1.7

A novel process for obtaining magnesium from sea water involves several reactions. Write a balanced chemical equation for each step of the process.

a. The first step is the decomposition of solid calcium carbonate from seashells to form solid calcium oxide and gaseous carbon dioxide.
b. The second step is the formation of solid calcium hydroxide as the only product from the reaction of the solid calcium oxide with liquid water.
c. Solid calcium hydroxide is then added to the seawater, reacting with dissolved magnesium chloride to yield solid magnesium hydroxide and aqueous calcium chloride.
d. The solid magnesium hydroxide is added to a hydrochloric acid solution, producing dissolved magnesium chloride and liquid water.
e. Finally, the magnesium chloride is melted and electrolyzed to yield liquid magnesium metal and diatomic chlorine gas.

Answer a

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Answer b

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

Answer c

$$Ca(OH)_2(s) + MgCl_2(aq) \rightarrow Mg(OH)_2(s) + CaCl_2(aq)$$

Answer d

$$Mg(OH)_2(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + 2\,H_2O(l)$$

Answer e

$$MgCl_2(s) \rightarrow Mg(s) + Cl_2(g)$$

Click here to see a video of the solution



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- Adelaide Clark, Oregon Institute of Technology

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5.2: Reaction Stoichiometry

Skills to Develop

- Explain the concept of stoichiometry as it pertains to chemical reactions
- Use balanced chemical equations to derive stoichiometric factors relating amounts of reactants and products
- Perform stoichiometric calculations involving mass and moles



Video* 5.2.1: A preview of the reaction stoichiometry process.*

A balanced chemical equation provides a great deal of information in a very succinct format. Chemical formulas provide the identities of the reactants and products involved in the chemical change, allowing classification of the reaction. Coefficients provide the relative numbers of these chemical species, allowing a quantitative assessment of the relationships between the amounts of substances consumed and produced by the reaction. These quantitative relationships are known as the reaction's stoichiometry, a term derived from the Greek words *stoicheion* (meaning "element") and *metron* (meaning "measure"). In this module, the use of balanced chemical equations for various stoichiometric applications is explored.

The general approach to using stoichiometric relationships is similar in concept to the way people go about many common activities. Cooking, for example, offers an appropriate comparison. Suppose a recipe for making eight pancakes calls for 1 cup pancake mix, $\frac{3}{4}$ cup milk, and one egg. The "equation" representing the preparation of pancakes per this recipe is

$$\text{1 cup mix} + \frac{3}{4}\ \text{cup milk} + \text{1 egg} \rightarrow \text{8 pancakes} \tag{5.2.1}$$

If two dozen pancakes are needed for a big family breakfast, the ingredient amounts must be increased proportionally according to the amounts given in the recipe. For example, the number of eggs required to make 24 pancakes is

$$24\ \cancel{\text{pancakes}} \times \frac{\text{1 egg}}{8\ \cancel{\text{pancakes}}} = \text{3 eggs} \tag{5.2.2}$$

Balanced chemical equations are used in much the same fashion to determine the amount of one reactant required to react with a given amount of another reactant, or to yield a given amount of product, and so forth. The coefficients in the balanced equation are used to derive stoichiometric factors that permit computation of the desired quantity. To illustrate this idea, consider the production of ammonia by reaction of hydrogen and nitrogen:

$$\mathrm{N_2}(g) + 3\,\mathrm{H_2}(g) \rightarrow 2\,\mathrm{NH_3}(g) \tag{5.2.3}$$

This equation shows that ammonia molecules are produced from hydrogen molecules in a 2:3 ratio, and stoichiometric factors may be derived using any amount (number) unit:

$$\frac{2\,NH3\ molecules}{3H2\ molecules} \text{ or } \frac{2\ doz\ NH3\ molecules}{3\ doz\ H2\ molecules} \text{ or } \frac{2\ mol\ NH3\ molecules}{3\ mol\ H2\ molecules} \tag{5.2.4}$$

These stoichiometric factors can be used to compute the number of ammonia molecules produced from a given number of hydrogen molecules, or the number of hydrogen molecules required to produce a given number of ammonia molecules. Similar factors may be derived for any pair of substances in any chemical equation.

Example 5.2.1: Moles of Reactant Required in a Reaction

How many moles of I_2 are required to react with 0.429 mol of Al according to the following equation (see Figure **5.2.2**)?

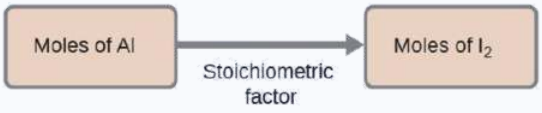
$$2\,Al + 3\,I_2 \rightarrow 2\,AlI_3 \tag{5.2.5}$$



***Figure* 5.2.1**: *Aluminum and iodine react to produce aluminum iodide. The heat of the reaction vaporizes some of the solid iodine as a purple vapor. (credit: modification of work by Mark Ott)*

Solution

Referring to the balanced chemical equation, the stoichiometric factor relating the two substances of interest is $\frac{3\ mol\ I2}{2\ mol\ Al}$. The molar amount of iodine is derived by multiplying the provided molar amount of aluminum by this factor:



$$\begin{aligned} \text{mol } I_2 &= 0.429\ \cancel{\text{mol Al}} \times \frac{3\ \text{mol } I_2}{2\ \cancel{\text{mol Al}}} \\ &= 0.644\ \text{mol } I_2 \end{aligned}$$

Exercise 5.2.1

How many moles of $Ca(OH)_2$ are required to react with 1.36 mol of H_3PO_4 to produce $Ca_3(PO_4)_2$ according to the equation $3\,Ca(OH)_2 + 2\,H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6\,H_2O$?

Answer

2.04 mol

Example 5.2.2: Number of Product Molecules Generated by a Reaction

How many carbon dioxide molecules are produced when 0.75 mol of propane is combusted according to this equation?

$$C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O \tag{5.2.6}$$

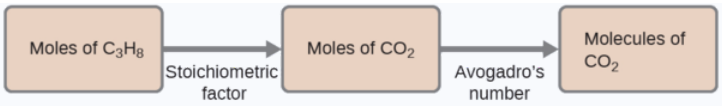
Solution

The approach here is the same as for Example **5.2.1**, though the absolute number of molecules is requested, not the number of moles of molecules. This will simply require use of the moles-to-numbers conversion factor, Avogadro's number.

The balanced equation shows that carbon dioxide is produced from propane in a 3:1 ratio:

$$\frac{3\ mol\ CO2}{1\ mol\ C3H8} \tag{5.2.7}$$

Using this stoichiometric factor, the provided molar amount of propane, and Avogadro's number,



$$0.75 \cancel{\text{ mol } C_3H_8} \times \frac{3 \cancel{\text{ mol } CO_2}}{1 \cancel{\text{ mol } C_3H_8}} \times \frac{6.022 \times 10^{23} \text{ CO}_2 \text{ molecules}}{\cancel{\text{mol } CO_2}} = 1.4 \times 10^{24} \text{ CO}_2 \text{ molecules} \tag{5.2.8}$$

Exercise 5.2.1

How many NH_3 molecules are produced by the reaction of 4.0 mol of $Ca(OH)_2$ according to the following equation:

$$(NH_4)_2SO_4 + Ca(OH)_2 \rightarrow 2\,NH_3 + CaSO_4 + 2\,H_2O$$

Answer

4.8×10^{24} NH_3 molecules

These examples illustrate the ease with which the amounts of substances involved in a chemical reaction of known stoichiometry may be related. Directly measuring numbers of atoms and molecules is, however, not an easy task, and the practical application of stoichiometry requires that we use the more readily measured property of mass.

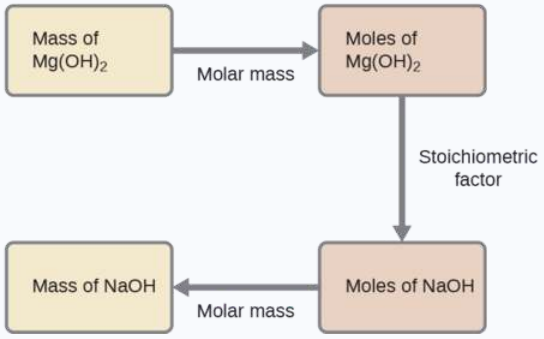
Example 5.2.3: Relating Masses of Reactants and Products

What mass of sodium hydroxide, NaOH, would be required to produce 16 g of the antacid milk of magnesia [magnesium hydroxide, $Mg(OH)_2$] by the following reaction?

$$MgCl_2(aq) + 2\,NaOH(aq) \rightarrow Mg(OH)_2(s) + 2\,NaCl(aq)$$

Solution

The approach used previously in Examples **5.2.1** and **5.2.2** is likewise used here; that is, we must derive an appropriate stoichiometric factor from the balanced chemical equation and use it to relate the amounts of the two substances of interest. In this case, however, masses (not molar amounts) are provided and requested, so additional steps of the sort learned in the previous chapter are required. The calculations required are outlined in this flowchart:



$$16 \cancel{\text{ g } Mg(OH)_2} \times \frac{1 \cancel{\text{ mol } Mg(OH)_2}}{58.3 \cancel{\text{ g } Mg(OH)_2}} \times \frac{2 \cancel{\text{ mol NaOH}}}{1 \cancel{\text{ mol } Mg(OH)_2}} \times \frac{40.0 \text{ g NaOH}}{\cancel{\text{mol NaOH}}} = 22 \text{ g NaOH} \tag{5.2.9}$$

Exercise 5.2.3

What mass of gallium oxide, Ga_2O_3, can be prepared from 29.0 g of gallium metal? The equation for the reaction is $4\,Ga + 3\,O_2 \rightarrow 2\,Ga_2O_3$.

Answer

39.0 g

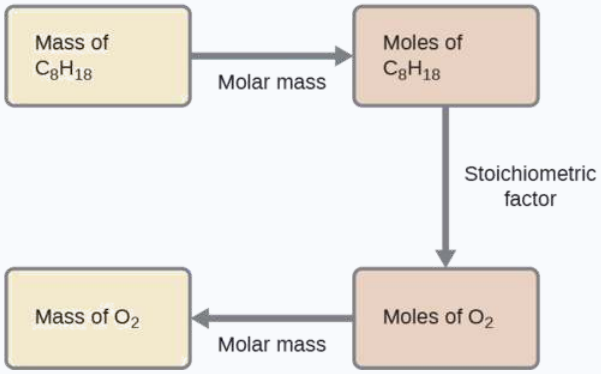
Example 5.2.4: Relating Masses of Reactants

What mass of oxygen gas, O_2, from the air is consumed in the combustion of 702 g of octane, C_8H_{18}, one of the principal components of gasoline?

$$2\,C_8H_{18} + 25\,O_2 \rightarrow 16\,CO_2 + 18\,H_2O$$

Solution

The approach required here is the same as for the Example **5.2.3**, differing only in that the provided and requested masses are both for reactant species.



$$702\ \cancel{\text{g } C_8H_{18}} \times \frac{1\ \cancel{\text{mol } C_8H_{18}}}{114.23\ \cancel{\text{g } C_8H_{18}}} \times \frac{25\ \cancel{\text{mol } O_2}}{2\ \cancel{\text{mol } C_8H_{18}}} \times \frac{32.00\ \text{g } O_2}{\cancel{\text{mol } O_2}} = 2.46 \times 10^3\ \text{g } O_2$$

Exercise 5.2.4

What mass of CO is required to react with 25.13 g of Fe_2O_3 according to the equation $Fe_2O_3 + 3\,CO \rightarrow 2\,Fe + 3\,CO_2$?

Answer

13.22 g

These examples illustrate just a few instances of reaction stoichiometry calculations. Numerous variations on the beginning and ending computational steps are possible depending upon what particular quantities are provided and sought (volumes, solution concentrations, and so forth). Regardless of the details, all these calculations share a common essential component: the use of stoichiometric factors derived from balanced chemical equations. Figure **5.2.2** provides a general outline of the various computational steps associated with many reaction stoichiometry calculations.

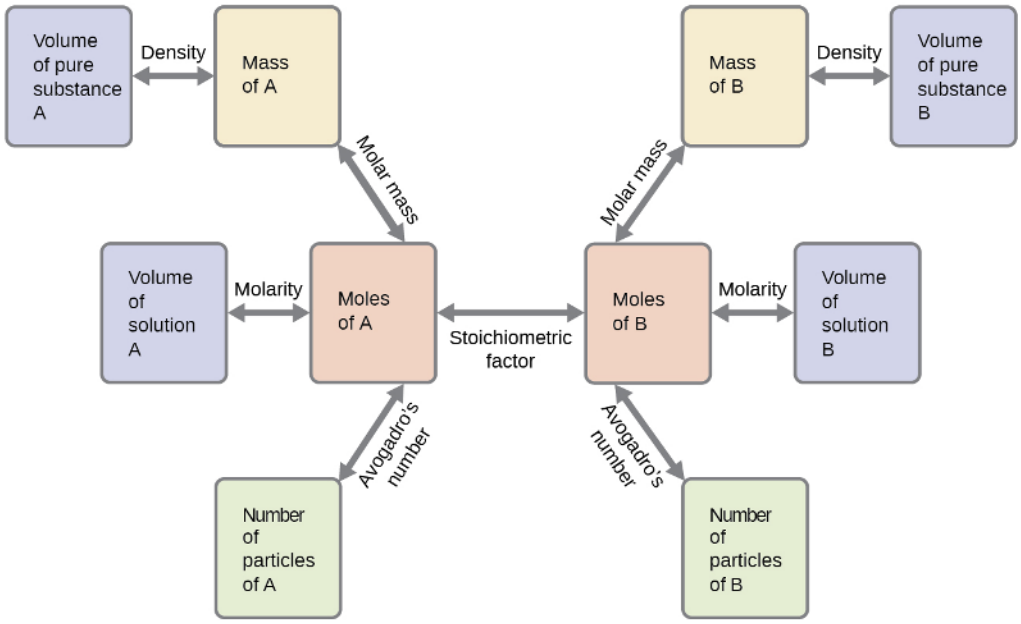


***Figure* 5.2.2**: *The flowchart depicts the various computational steps involved in most reaction stoichiometry calculations. We will discuss Molarity in Unit 6.*

Airbags

Airbags (Figure **5.2.3**) are a safety feature provided in most automobiles since the 1990s. The effective operation of an airbag requires that it be rapidly inflated with an appropriate amount (volume) of gas when the vehicle is involved in a collision. This requirement is satisfied in many automotive airbag systems through use of explosive chemical reactions, one common choice being the decomposition of sodium azide, NaN_3. When sensors in the vehicle detect a collision, an electrical current is passed through a carefully measured amount of NaN_3 to initiate its decomposition:

$$2\,NaN_3(s) \rightarrow 3\,N_2(g) + 2\,Na(s) \tag{5.2.10}$$

This reaction is very rapid, generating gaseous nitrogen that can deploy and fully inflate a typical airbag in a fraction of a second (~0.03–0.1 s). Among many engineering considerations, the amount of sodium azide used must be appropriate for generating enough nitrogen gas to fully inflate the air bag and ensure its proper function. For example, a small mass (~100 g) of NaN_3 will generate approximately 50 L of N_2.



Figure* 5.2.3:* Airbags deploy upon impact to minimize serious injuries to passengers. (credit: Jon Seidman)

Summary



Video* 5.2.2: An overview of the process covered in this section with a preview of the next section.*

A balanced chemical equation may be used to describe a reaction's stoichiometry (the relationships between amounts of reactants and products). Coefficients from the equation are used to derive stoichiometric factors that subsequently may be used for computations relating reactant and product masses, molar amounts, and other quantitative properties.

Glossary

stoichiometric factor
ratio of coefficients in a balanced chemical equation, used in computations relating amounts of reactants and products

stoichiometry
relationships between the amounts of reactants and products of a chemical reaction

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5.2.1: Practice Problems- Reaction Stoichiometry

PROBLEM 5.2.1.1

Write the balanced equation and determine the information requested. Don't worry about state symbols in these reactions.

a. The number of moles and the mass (in grams) of chlorine, Cl_2, required to react with 10.0 g of sodium metal, Na, to produce sodium chloride, NaCl.
b. The number of moles and the mass (in milligrams) of diatomic oxygen formed by the decomposition of 1.252 g of mercury(II) oxide.
c. The number of moles and the mass (in g) of sodium nitrate, $NaNO_3$, required to decompose and produce 128 g of diatomic oxygen, where $NaNO_2$ is the other product.
d. The number of moles and the mass (in kg) of carbon dioxide formed by the combustion of 20.0 kg of carbon in an excess of diatomic oxygen.
e. The number of moles and the mass (in kg) of copper(II) carbonate needed to decompose in order to produce 1.500 kg of copper(II) oxide, where CO_2 is the other product.
f. The number of moles and mass (in grams) of C_2H_4 required to react with water to produce 9.55 g C_2H_6O.

Answer a

$$\mathbf{2\,Na + Cl_2 \rightarrow 2\,NaCl}$$

0.217 mol Cl_2

15.43 g Cl_2

Answer b

$$\mathbf{2\,HgO \rightarrow 2\,Hg + O_2}$$

0.00289 mol O_2

92 mg O_2

Answer c

$$\mathbf{2\,NaNO_3 \rightarrow 2\,NaNO_3 + O_2}$$

8 mol $NaNO_3$

680 g $NaNO_3$

Answer d

$$\mathbf{C + O_2 \rightarrow CO_2}$$

1666.67 mol CO_2

73.3 kg CO_2

Answer e

$$\mathbf{CuCO_3 \rightarrow CuO + CO_2}$$

18.87 mol $CuCO_3$

2.330 kg $CuCO_3$

Answer f

$$\mathbf{C_2H_4 + H_2O \rightarrow C_2H_6O}$$

0.207 mol C_2H_4

5.81 g C_2H_4

Click here to see a video of the solution

Apologies for the brief phone ringing



PROBLEM 5.2.1.2

I_2 is produced by the reaction of 0.4235 mol of $CuCl_2$ according to the following equation: $2\,CuCl_2 + 4\,KI \rightarrow 2\,CuI + 4\,KCl + I_2$.

a. How many molecules of I_2 are produced?
b. What mass of I_2 is produced?

Answer a

1.28×10^{23} molecules I_2

Answer b

53.8 g I2

PROBLEM 5.2.1.3

Silver is often extracted from ores as $K[Ag(CN)_2]$ and then recovered by the reaction

$$2\,K[Ag(CN)_2](aq) + Zn(s) \rightarrow 2\,Ag(s) + Zn(CN)_2(aq) + 2\,KCN(aq)$$

a. How many molecules of $Zn(CN)_2$ are produced by the reaction of 35.27 g of $K[Ag(CN)_2]$?
b. What mass of $Zn(CN)_2$ is produced?

Answer a

5.337×10^{22} molecules

Answer b

10.41 g $Zn(CN)_2$

Click here to see a video of the solution





PROBLEM 5.2.1.4

What mass of silver oxide, Ag_2O, is required to produce 25.0 g of silver sulfadiazine, $AgC_{10}H_9N_4SO_2$, from the reaction of silver oxide and sulfadiazine?

$$2\,C_{10}H_{10}N_4SO_2 + Ag_2O \rightarrow 2\,AgC_{10}H_9N_4SO_2 + H_2O$$

Answer

8.12 g Ag_2O

PROBLEM 5.2.1.5

Carborundum is silicon carbide, SiC, a very hard material used as an abrasive on sandpaper and in other applications. It is prepared by the reaction of pure sand, SiO_2, with carbon at high temperature. Carbon monoxide, CO, is the other product of this reaction. Write the balanced equation for the reaction, and calculate how much SiO_2 is required to produce 3.00 kg of SiC.

Answer

$$SiO_2 + 3\,C \rightarrow SiC + 2\,CO$$

4.50 kg SiO_2

Click here to see a video of the solution



PROBLEM 5.2.1.6

Automotive air bags inflate when a sample of sodium azide, NaN_3, is very rapidly decomposed.

$$2\,NaN_3(s) \rightarrow 2\,Na(s) + 3\,N_2(g)$$

What mass of sodium azide is required to produce 2.6 ft^3 (73.6 L) of nitrogen gas with a density of 1.25 g/L?

Answer

142g NaN_3

PROBLEM 5.2.1.7

Urea, $CO(NH_2)_2$, is manufactured on a large scale for use in producing urea-formaldehyde plastics and as a fertilizer. What is the maximum mass of urea that can be manufactured from the CO_2 produced by combustion of 1.00×10^3 kg of carbon followed

by the reaction?

$$CO_2(g) + 2\,NH_3(g) \rightarrow CO(NH_2)_2(s) + H_2O(l) \tag{5.2.1.1}$$

Answer

5.00 kg Urea

Click here to see a video of the solution



PROBLEM 5.2.1.8

In an accident, a solution containing 2.5 kg of nitric acid was spilled. Two kilograms of Na_2CO_3 was quickly spread on the area and CO_2 was released by the reaction. Was sufficient Na_2CO_3 used to neutralize all of the acid? (in this reaction, water and sodium nitrate are the other two products)

Answer

No, you will need 2.1 kg of sodium carbonate to neutralize 2.5 kg of nitric acid.

PROBLEM 5.2.1.9

A compact car gets 37.5 miles per gallon on the highway. If gasoline contains 84.2% carbon by mass and has a density of 0.8205 g/mL, determine the mass of carbon dioxide produced during a 500-mile trip (3.785 liters per gallon).

Answer

1.28×10^5 g CO_2

Click here to see a video of the solution





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5.3: Calculating Reaction Yields

Skills to Develop

- Explain the concepts of theoretical yield and limiting reactants/reagents.
- Derive the theoretical yield for a reaction under specified conditions.
- Calculate the percent yield for a reaction.

The relative amounts of reactants and products represented in a balanced chemical equation are often referred to as *stoichiometric amounts*. All the exercises of the preceding module involved stoichiometric amounts of reactants. For example, when calculating the amount of product generated from a given amount of reactant, it was assumed that any other reactants required were available in stoichiometric amounts (or greater). In this module, more realistic situations are considered, in which reactants are not present in stoichiometric amounts.

Limiting Reactant

Consider another food analogy, making grilled cheese sandwiches (Figure 5.3.1):

$$\text{1 slice of cheese} + \text{2 slices of bread} \rightarrow \text{1 sandwich} \tag{5.3.1}$$

Stoichiometric amounts of sandwich ingredients for this recipe are bread and cheese slices in a 2:1 ratio. Provided with 28 slices of bread and 11 slices of cheese, one may prepare 11 sandwiches per the provided recipe, using all the provided cheese and having six slices of bread left over. In this scenario, the number of sandwiches prepared has been *limited* by the number of cheese slices, and the bread slices have been provided in *excess*.

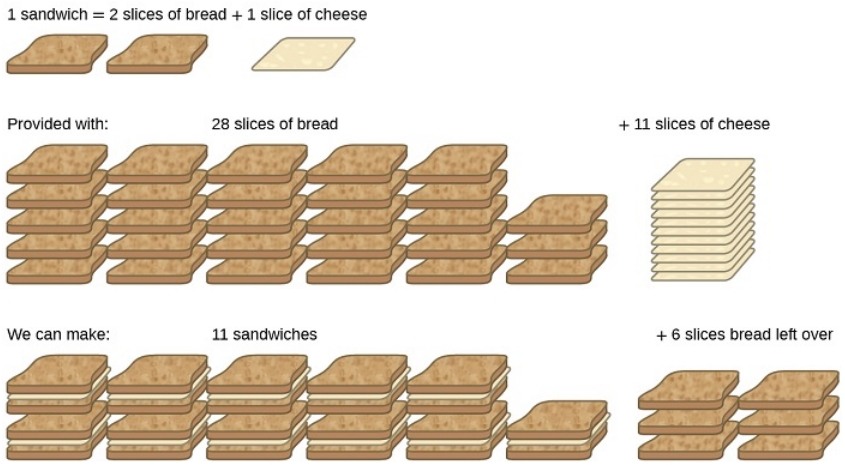


Figure* 5.3.1*: *Sandwich making can illustrate the concepts of limiting and excess reactants.*

Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$\mathrm{Cl_2}(g) \rightarrow 2\,\mathrm{HCl}(g) \tag{5.3.2}$$

The balanced equation shows the hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the limiting reactant, and the other substance is the excess reactant. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation. For example, imagine combining 3 moles of H_2 and 2 moles of Cl_2. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all the provided chlorine (2 mol) will consume 2 mol of the 3 mol of hydrogen provided, leaving 1 mol of hydrogen nonreacted.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the

reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant. For the example in the previous paragraph, complete reaction of the hydrogen would yield

$$\text{mol HCl produced} = 3\ \text{mol } H_2 \times \frac{2\ \text{mol HCl}}{1\ \text{mol } H_2} = 6\ \text{mol HCl} \tag{5.3.3}$$

Complete reaction of the provided chlorine would produce

$$\text{mol HCl produced} = 2\ \text{mol } Cl_2 \times \frac{2\ \text{mol HCl}}{1\ \text{mol } Cl_2} = 4\ \text{mol HCl} \tag{5.3.4}$$

The chlorine will be completely consumed once 4 moles of HCl have been produced. Since enough hydrogen was provided to yield 6 moles of HCl, there will be non-reacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (Figure **5.3.2**).

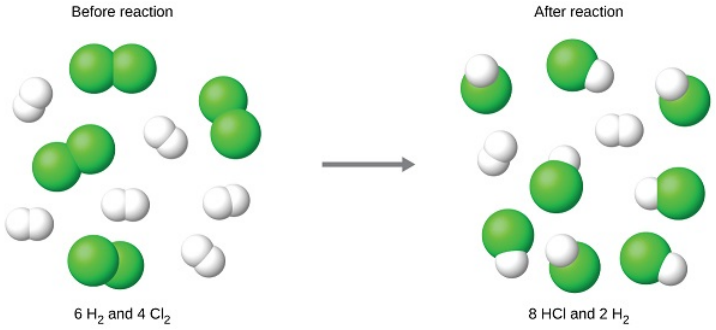
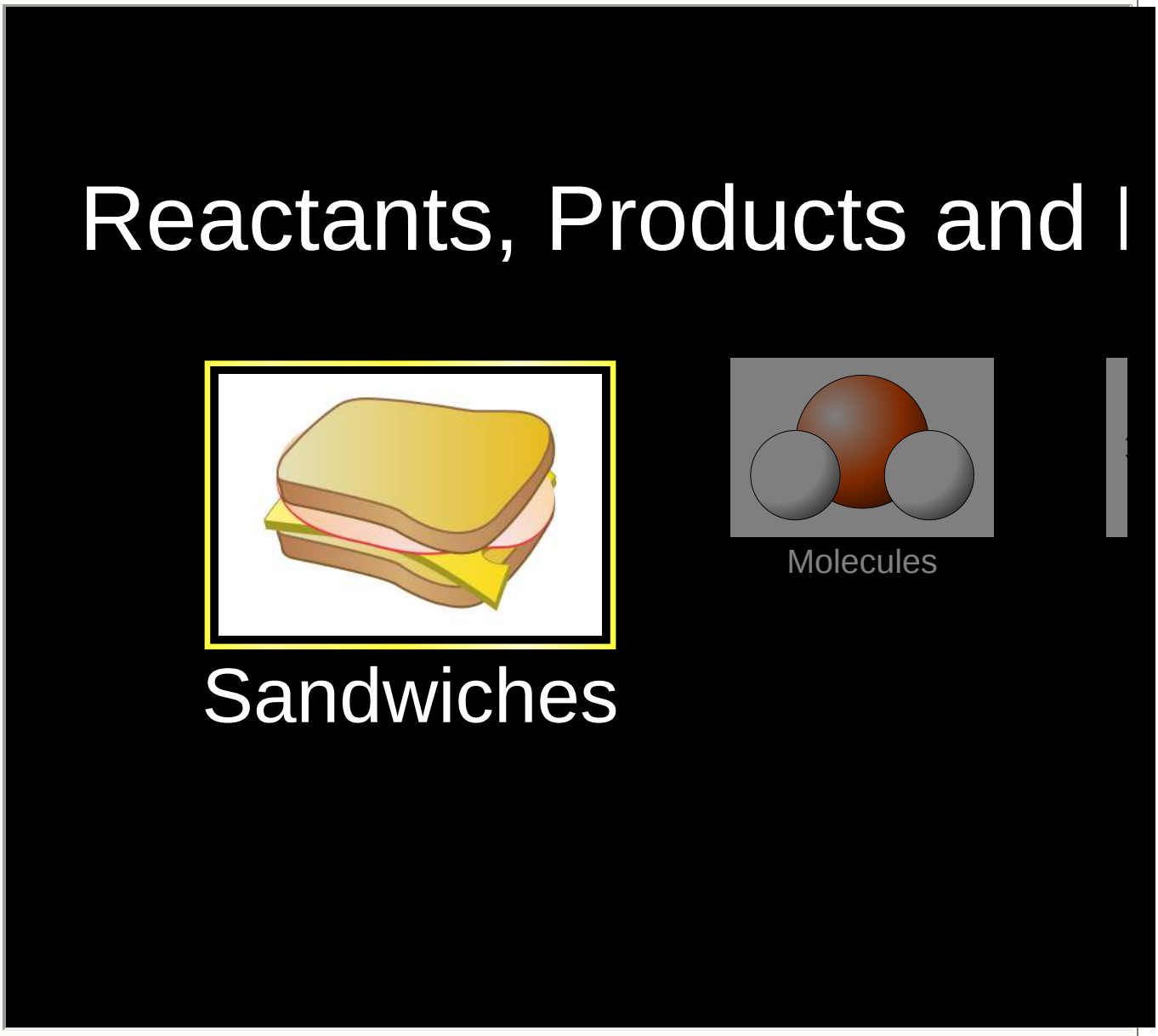


***Figure* 5.3.2:** *When H_2 and Cl_2 are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant.*

Interactive simulation of limiting and excess reactants



Example 5.3.1: Identifying the Limiting Reactant

Silicon nitride is a very hard, high-temperature-resistant ceramic used as a component of turbine blades in jet engines. It is prepared according to the following equation:

$$3\,\mathrm{Si}(s) + 2\,\mathrm{N_2}(g) \rightarrow \mathrm{Si_3N_4}(s)$$

Which is the limiting reactant when 2.00 g of Si and 1.50 g of N_2 react?

Solution

Compute the provided molar amounts of reactants, and then compare these amounts to the balanced equation to identify the limiting reactant.

$$\text{mol Si} = 2.00\ \cancel{\text{g Si}} \times \frac{1\ \text{mol Si}}{28.09\ \cancel{\text{g Si}}} = 0.0712\ \text{mol Si}$$

$$\text{mol N}_2 = 1.50\ \cancel{\text{g N}_2} \times \frac{1\ \text{mol N}_2}{28.02\ \cancel{\text{g N}_2}} = 0.0535\ \text{mol N}_2$$

The provided Si:N_2 molar ratio is:

$$\frac{0.0712\ \text{mol Si}}{0.0535\ \text{mol N}_2} = \frac{1.33\ \text{mol Si}}{1\ \text{mol N}_2}$$

The stoichiometric Si:N_2 ratio is:

$$\frac{3\ \text{mol Si}}{2\ \text{mol N}_2} = \frac{1.5\ \text{mol Si}}{1\ \text{mol N}_2}$$

Comparing these ratios shows that Si is provided in a less-than-stoichiometric amount, and so is the limiting reactant.

Alternatively, compute the amount of product expected for complete reaction of each of the provided reactants. The 0.0712 moles of silicon would yield

$$\text{mol Si}_3\text{N}_4\ \text{produced} = 0.0712\ \text{mol Si} \times \frac{1\ \text{mol Si}_3\text{N}_4}{3\ \text{mol Si}} = 0.0237\ \text{mol Si}_3\text{N}_4$$

while the 0.0535 moles of nitrogen would produce

$$\text{mol Si}_3\text{N}_4\ \text{produced} = 0.0535\ \text{mol N}_2 \times \frac{1\ \text{mol Si}_3\text{N}_4}{2\ \text{mol N}_2} = 0.0268\ \text{mol Si}_3\text{N}_4$$

Since silicon yields the lesser amount of product, it is the limiting reactant.

Exercise 5.3.1

Which is the limiting reactant when 5.00 g of H_2 and 10.0 g of O_2 react and form water?

Answer

O_2

Percent Yield



Video* 5.3.1*: *An introduction to Theoretical Yield.*

The amount of product that *may be* produced by a reaction under specified conditions, as calculated per the stoichiometry of an appropriate balanced chemical equation, is called the theoretical yield of the reaction. In practice, the amount of product obtained is called the actual yield, and it is often less than the theoretical yield for a number of reasons. Some reactions are inherently inefficient, being accompanied by *side reactions* that generate other products. Others are, by nature, incomplete (consider the partial reactions of weak acids and bases discussed earlier in this text). Some products are difficult to collect without some loss, and so less than perfect recovery will reduce the actual yield. The extent to which a reaction's theoretical yield is achieved is commonly expressed as its percent yield:

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \tag{5.3.5}$$

Actual and theoretical yields may be expressed as masses or molar amounts (or any other appropriate property; e.g., volume, if the product is a gas). As long as both yields are expressed using the same units, these units will cancel when percent yield is calculated.



***Video* 5.3.2:** *A video example of calculating percent yield.*

Example 5.3.2: Calculation of Percent Yield

Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

$$\ce{CuSO4}(aq) + \ce{Zn}(s) \rightarrow \ce{Cu}(s) + \ce{ZnSO4}(aq)$$

What is the percent yield?

Solution

The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield is found by the approach illustrated in the previous module, as shown here:

Double subscripts: use braces to clarify

Using this theoretical yield and the provided value for actual yield, the percent yield is calculated to be

$$\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100$$

$$\begin{aligned}\text{percent yield} &= \left(\frac{0.392\ \text{g Cu}}{0.5072\ \text{g Cu}}\right) \times 100 \\ &= 77.3\%\end{aligned}$$

Exercise 5.3.2

What is the percent yield of a reaction that produces 12.5 g of the gas Freon CF_2Cl_2 from 32.9 g of CCl_4 and excess HF?

$$\ce{CCl4} + 2\,\ce{HF} \rightarrow \ce{CF2Cl2} + 2\,\ce{HCl}$$

Answer

48.3%

Green Chemistry and Atom Economy

The purposeful design of chemical products and processes that minimize the use of environmentally hazardous substances and the generation of waste is known as *green chemistry*. Green chemistry is a philosophical approach that is being applied to many areas of science and technology, and its practice is summarized by guidelines known as the “Twelve Principles of Green Chemistry”. One of the 12 principles is aimed specifically at maximizing the efficiency of processes for synthesizing chemical

products. The *atom economy* of a process is a measure of this efficiency, defined as the percentage by mass of the final product of a synthesis relative to the masses of *all* the reactants used:

$$\text{atom economy} = \frac{\text{mass of product}}{\text{mass of reactants}} \times 100\% \tag{5.3.6}$$

Though the definition of atom economy at first glance appears very similar to that for percent yield, be aware that this property represents a difference in the *theoretical* efficiencies of *different* chemical processes. The percent yield of a given chemical process, on the other hand, evaluates the efficiency of a process by comparing the yield of product actually obtained to the maximum yield predicted by stoichiometry.

The synthesis of the common nonprescription pain medication, ibuprofen, nicely illustrates the success of a green chemistry approach (Figure **5.3.3**). First marketed in the early 1960s, ibuprofen was produced using a six-step synthesis that required 514 g of reactants to generate each mole (206 g) of ibuprofen, an atom economy of 40%. In the 1990s, an alternative process was developed by the BHC Company (now BASF Corporation) that requires only three steps and has an atom economy of ~80%, nearly twice that of the original process. The BHC process generates significantly less chemical waste; uses less-hazardous and recyclable materials; and provides significant cost-savings to the manufacturer (and, subsequently, the consumer). In recognition of the positive environmental impact of the BHC process, the company received the Environmental Protection Agency's Greener Synthetic Pathways Award in 1997.



***Figure* 5.3.3**: *(a) Ibuprofen is a popular nonprescription pain medication commonly sold as 200 mg tablets. (b) The BHC process for synthesizing ibuprofen requires only three steps and exhibits an impressive atom economy. (credit a: modification of work by Derrick Coetzee)*

Summary

When reactions are carried out using less-than-stoichiometric quantities of reactants, the amount of product generated will be determined by the limiting reactant. The amount of product generated by a chemical reaction is its actual yield. This yield is often less than the amount of product predicted by the stoichiometry of the balanced chemical equation representing the reaction (its theoretical yield). The extent to which a reaction generates the theoretical amount of product is expressed as its percent yield.

Key Equations

- $\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100$

Glossary

actual yield

amount of product formed in a reaction

excess reactant

reactant present in an amount greater than required by the reaction stoichiometry

limiting reactant

reactant present in an amount lower than required by the reaction stoichiometry, thus limiting the amount of product generated

percent yield

measure of the efficiency of a reaction, expressed as a percentage of the theoretical yield

theoretical yield

amount of product that may be produced from a given amount of reactant(s) according to the reaction stoichiometry

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5.3.1: Practice Problems Calculating Reaction Yields

PROBLEM 5.3.1.1

The following quantities are placed in a container: 1.5×10^{24} molecules of diatomic hydrogen, 1.0 mol of sulfur, and 88.0 g of diatomic oxygen.

a. What is the total mass in grams for the collection of all three elements?
b. What is the total number of moles of atoms for the three elements?
c. If the mixture of the three elements formed a compound with molecules that contain two hydrogen atoms, one sulfur atom, and four oxygen atoms, which substance is consumed first?
d. How many atoms or molecules of each remaining element would remain unreacted in the change described above ?

Answer a

4.98 g H_2, 32 g S; 124.98 g total

Answer b

2.49 mol H_2, 2.75 mol O_2; 6.24 mol total

Answer c

S is the limiting reactant (what would be consumed first)

Answer d

8.978×10^{23} molecules of H_2 and 4.6×10^{23} molecules of O_2 remain

Click here to see a video of the solution



PROBLEM 5.3.1.2

What is the limiting reactant in a reaction that produces sodium chloride from 8 g of sodium and 8 g of diatomic chlorine?

Answer

The limiting reactant is Cl_2

PROBLEM 5.3.1.3

Which of the postulates of Dalton's atomic theory explains why we can calculate a theoretical yield for a chemical reaction?

Answer

Postulate 4 (A compound consists of atoms of two or more elements combined in a small, whole-number ratio. In a given compound, the numbers of atoms of each of its elements are always present in the same ratio).

PROBLEM 5.3.1.4

A student isolated 25 g of a compound following a procedure that would theoretically yield 81 g. What was his percent yield?

Answer

$\textbf{Percent yield} = 31\%$

PROBLEM 5.3.1.5

A sample of 0.53 g of carbon dioxide was obtained by heating 1.31 g of calcium carbonate. What is the percent yield for this reaction?

$$\mathrm{CaCO_3}(s) \rightarrow \mathrm{CaO}(s) + \mathrm{CO_2}(s) \tag{5.3.1.1}$$

Answer

$\textbf{Percent yield} = 91.9\%$

Click here to see a video of the solution



PROBLEM 5.3.1.6

Freon-12, CCl_2F_2, is prepared from CCl_4 by reaction with HF. The other product of this reaction is HCl. Outline the steps needed to determine the percent yield of a reaction that produces 12.5 g of CCl_2F_2 from 32.9 g of CCl_4. Freon-12 has been banned and is no longer used as a refrigerant because it catalyzes the decomposition of ozone and has a very long lifetime in the atmosphere. Determine the percent yield.

Answer

$\textbf{g } \mathrm{CCl_4} \rightarrow \textbf{mol } \mathrm{CCl_4} \rightarrow \textbf{mol } \mathrm{CCl_2F_2} \rightarrow \textbf{g } \mathrm{CCl_2F_2}$

$\textbf{percent yield} = 48.3\%$

PROBLEM 5.3.1.7

Citric acid, $C_6H_8O_7$, a component of jams, jellies, and fruity soft drinks, is prepared industrially via fermentation of sucrose by the mold *Aspergillus niger*. The equation representing this reaction is

$$\mathrm{C_{12}H_{22}O_{11} + H_2O + 3\,O_2 \rightarrow 2\,C_6H_8O_7 + 4\,H_2O} \tag{5.3.1.2}$$

What mass of citric acid is produced from exactly 1 metric ton (1.000×10^3 kg) of sucrose ($C_{12}H_{22}O_{11}$) if the yield is 92.30%?

Answer

1036 kg citric acid

Click here to see a video of the solution



PROBLEM 5.3.1.8

Toluene, $C_6H_5CH_3$, is oxidized by air under carefully controlled conditions to benzoic acid, $C_6H_5CO_2H$, which is used to prepare the food preservative sodium benzoate, $C_6H_5CO_2Na$. What is the percent yield of a reaction that converts 1.000 kg of toluene to 1.21 kg of benzoic acid?

$$\mathbf{2\,C_6H_5CH_3 + 3\,O_2 \rightarrow 2\,C_6H_5CO_2H + 2\,H_2O} \qquad \mathbf{(5.3.1.3)}$$

Answer

$\mathbf{\text{percent yield} = 91.3\%}$

PROBLEM 5.3.1.9

In a laboratory experiment, the reaction of 3.0 mol of H_2 with 2.0 mol of I_2 produced 1.0 mol of HI. Determine the theoretical yield in grams and the percent yield for this reaction.

Answer

512 g (theoretical yield); $\mathbf{\text{percent yield} = 25\%}$

Click here to see a video of the solution



PROBLEM 5.3.1.10

What is the limiting reactant when 1.50 g of lithium and 1.50 g of nitrogen combine to form lithium nitride, a component of advanced batteries, according to the following unbalanced equation?

$$\mathrm{Li} + \mathrm{N_2} \rightarrow \mathrm{Li_3N} \tag{5.3.1.4}$$

Answer

$$6\,\mathrm{Li} + \mathrm{N_2} \rightarrow 2\,\mathrm{Li_3N} \tag{5.3.1.5}$$

$$1.50g\,\mathrm{Li} \times \frac{1\,mole\,\mathrm{Li}}{6.94g\,\mathrm{Li}} \times \frac{2\,mole\,\mathrm{Li_3N}}{6\,mole\,\mathrm{Li}} = 0.0720\,moles\,\mathrm{Li_3N} \tag{5.3.1.6}$$

$$1.50g\,\mathrm{N_2} \times \frac{1\,mole\,\mathrm{N_2}}{28.02g\,\mathrm{N_2}} \times \frac{2\,mole\,\mathrm{Li_3N}}{1\,mole\,\mathrm{N_2}} = 0.107\,moles\,\mathrm{Li_3N} \tag{5.3.1.7}$$

Li is the limiting reactant

PROBLEM 5.3.1.11

Uranium can be isolated from its ores by dissolving it as $UO_2(NO_3)_2$, then separating it as solid $UO_2(C_2O_4)\cdot 3H_2O$. Addition of 0.4031 g of sodium oxalate, $Na_2C_2O_4$, to a solution containing 1.481 g of uranyl nitrate, $UO_2(NO_3)_2$, yields 1.073 g of solid $UO_2(C_2O_4)\cdot 3H_2O$.

$$\mathrm{Na_2C_2O_4 + UO_2(NO_3)_2 + 3\,H_2O \longrightarrow UO_2(C_2O_4)\cdot 3\,H_2O + 2\,NaNO_3} \tag{5.3.1.8}$$

Determine the limiting reactant and the percent yield of this reaction.

Answer

$Na_2C_2O_4$ is the limiting reactant. percent yield = 86.6%

Click here to see a video of the solution



PROBLEM 5.3.1.12

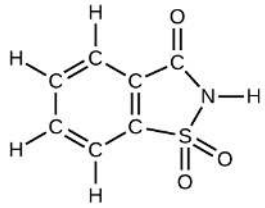
How many molecules of $C_2H_4Cl_2$ can be prepared from 15 C_2H_4 molecules and 8 Cl_2 molecules?

Answer

Only 8 molecules can be formed

PROBLEM 5.3.1.13

How many molecules of the sweetener saccharin can be prepared from 30 C atoms, 25 H atoms, 12 O atoms, 8 S atoms, and 14 N atoms?



Answer

Only four molecules can be made.

PROBLEM 5.3.1.14

The phosphorus pentoxide used to produce phosphoric acid for cola soft drinks is prepared by burning phosphorus in oxygen.

a. What is the limiting reactant when 0.200 mol of P_4 and 0.200 mol of O_2 react according to

$$P_4 + 5\,O_2 \rightarrow P_4O_{10} \tag{5.3.1.9}$$

b. Calculate the percent yield if 10.0 g of P_4O_{10} is isolated from the reaction.

Answer a

O_2 is the limiting reactant

Answer b

$\text{percent yield} = 88\%$

PROBLEM 5.3.1.15

Would you agree to buy 1 trillion (1,000,000,000,000) gold atoms for $5? Explain why or why not. Find the current price of gold at http://money.cnn.com/data/commodities/ $(1\ \text{troy ounce} = 31.1\ \text{g})$

Answer

This amount cannot be weighted by ordinary balances and is worthless.

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CHAPTER OVERVIEW

6: Aqueous Reactions

6.1: Solutions and Solution Concentration

6.1.1: Practice Problems- Solution Concentration

6.2: Solutions Chemistry

6.2.1: Practice Problems- Solutions Chemistry

6.3: Precipitation Reactions

6.3.1: Practice Problems- Precipitation

6.4: Classifying Chemical Reactions (Acids and Bases)

6.4.1: Classifying Chemical Reactions (Acids and Bases) (Problems)

6.5: Classifying Chemical Reactions (Redox)

6.5.1: Classifying Chemical Reactions (Redox) (Problems)

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6.1: Solutions and Solution Concentration

Skills to Develop

- Describe the fundamental properties of solutions
- Calculate solution concentrations using molarity and mass-mass-percent
- Perform dilution calculations using the dilution equation

In preceding sections, we focused on the composition of substances: samples of matter that contain only one type of element or compound. However, mixtures—samples of matter containing two or more substances physically combined—are more commonly encountered in nature than are pure substances. Similar to a pure substance, the relative composition of a mixture plays an important role in determining its properties. The relative amount of oxygen in a planet's atmosphere determines its ability to sustain aerobic life. The relative amounts of iron, carbon, nickel, and other elements in steel (a mixture known as an "alloy") determine its physical strength and resistance to corrosion. The relative amount of the active ingredient in a medicine determines its effectiveness in achieving the desired pharmacological effect. The relative amount of sugar in a beverage determines its sweetness (Figure **6.1.1**). In this section, we will describe one of the most common ways in which the relative compositions of mixtures may be quantified.



Figure* 6.1.1*: *Sugar is one of many components in the complex mixture known as coffee. The amount of sugar in a given amount of coffee is an important determinant of the beverage's sweetness. (credit: Jane Whitney)*

Solutions



Video* 6.1.1*: *An introduction to solutions from Crash Course Chemistry.*

We have previously defined solutions as homogeneous mixtures, meaning that the composition of the mixture (and therefore its properties) is uniform throughout its entire volume. Solutions occur frequently in nature and have also been implemented in many forms of manmade technology. We will explore a more thorough treatment of solution properties in the chapter on solutions and colloids, but here we will introduce some of the basic properties of solutions.

The relative amount of a given solution component is known as its concentration. Often, though not always, a solution contains one component with a concentration that is significantly greater than that of all other components. This component is called the solvent and may be viewed as the medium in which the other components are dispersed, or dissolved. Solutions in which water is the solvent are, of course, very common on our planet. A solution in which water is the solvent is called an aqueous solution.

A solute is a component of a solution that is typically present at a much lower concentration than the solvent. Solute concentrations are often described with qualitative terms such as dilute (of relatively low concentration) and concentrated (of relatively high concentration).



***Video* 6.1.2:** *An introduction to molarity from Crash Course Chemistry.*

Concentrations of Solutions

Concentrations may be quantitatively assessed using a wide variety of measurement units, each convenient for particular applications. Some of the common concetrations measurements seen in the science labs include, parts-per-million (ppm), parts-per-billion (ppb), mass-mass-percent (m/m %), molarity (M), molality (m) and mole percent. In this unit we will use two of the more common units of measure, Molarity, and mass-mass percent.

Molarity

Molarity (*M*) is a useful concentration unit for many applications in chemistry. Molarity is defined as the number of moles of solute in exactly 1 liter (1 L) of the solution:

$$M = \frac{\text{mol solute}}{\text{L solution}} \tag{6.1.1}$$

The units of molarity are therefore moles per liter of solution (mol/L), abbreviated as M. An aqueous solution that contains 1 mol (342 g) of sucrose in enough water to give a final volume of 1.00 L has a sucrose concentration of 1.00 mol/L or 1.00 M. In chemical notation, square brackets around the name or formula of the solute represent the molar concentration of a solute. Therefore,

$$[\text{sucrose}] = 1.00\ \text{M} \tag{6.1.2}$$

is read as “the concentration of sucrose is 1.00 molar.”

Example 6.1.1: Calculating Molar Concentrations

A 355-mL soft drink sample contains 0.133 mol of sucrose (table sugar). What is the molar concentration of sucrose in the beverage?

Solution

Since the molar amount of solute and the volume of solution are both given, the molarity can be calculated using the definition of molarity. Per this definition, the solution volume must be converted from mL to L:

$$\begin{aligned} M &= \frac{mol\ solute}{L\ solution} \\ &= \frac{0.133\ mol}{355\ mL \times \frac{1\ L}{1000\ mL}} \\ &= 0.375\ M \end{aligned}$$

Exercise 6.1.1

A teaspoon of table sugar contains about 0.01 mol sucrose. What is the molarity of sucrose if a teaspoon of sugar has been dissolved in a cup of tea with a volume of 200 mL?

Answer

0.05 M

Example 6.1.2: Deriving Moles and Volumes from Molar Concentrations

How much sugar (mol) is contained in a modest sip (~10 mL) of the soft drink from Example **6.1.1**?

Solution

In this case, we can rearrange the definition of molarity to isolate the quantity sought, moles of sugar. We then substitute the value for molarity that we derived in Example 3.4.2, 0.375 *M*:

$$M = \frac{\text{mol solute}}{\text{L solution}}$$

Multiple \label

Exercise 6.1.2

What volume (mL) of the sweetened tea described in Example **6.1.1** contains the same amount of sugar (mol) as 10 mL of the soft drink in this example?

Answer

80 mL

Example 6.1.3: Calculating Molar Concentrations from the Mass of Solute

Distilled white vinegar (Figure **6.1.2**) is a solution of acetic acid, CH_3CO_2H, in water. A 0.500-L vinegar solution contains 25.2 g of acetic acid. What is the concentration of the acetic acid solution in units of molarity?



Figure* 6.1.3*: *Distilled white vinegar is a solution of acetic acid in water.*

Solution

As in previous examples, the definition of molarity is the primary equation used to calculate the quantity sought. In this case, the mass of solute is provided instead of its molar amount, so we must use the solute's molar mass to obtain the amount of solute in moles:

$$M = \frac{\text{mol solute}}{\text{L solution}} = \frac{25.2\ \text{g}\ \ce{CH3CO2H} \times \frac{1\ \text{mol}\ \ce{CH3CO2H}}{60.052\ \text{g}\ \ce{CH3CO2H}}}{0.500\ \text{L solution}} = 0.839\ M \tag{6.1.3}$$

$$M = \frac{\text{mol solute}}{\text{L solution}} = 0.839\ M \tag{6.1.4}$$

$$M = \frac{0.839\ \text{mol solute}}{1.00\ \text{L solution}} \tag{6.1.5}$$

Exercise 6.1.3

Calculate the molarity of 6.52 g of $\ce{CoCl2}$ (128.9 g/mol) dissolved in an aqueous solution with a total volume of 75.0 mL.

Answer

0.674 M

Example 6.1.4: Determining the Mass of Solute in a Given Volume of Solution

How many grams of NaCl are contained in 0.250 L of a 5.30-*M* solution?

Solution

The volume and molarity of the solution are specified, so the amount (mol) of solute is easily computed as demonstrated in Example **6.1.3**:

$$M = \frac{\text{mol solute}}{\text{L solution}} \tag{6.1.6}$$

$$\text{mol solute} = M \times \text{L solution} \tag{6.1.7}$$

$$\text{mol solute} = 5.30\ \frac{\text{mol NaCl}}{\text{L}} \times 0.250\ \text{L} = 1.325\ \text{mol NaCl} \tag{6.1.8}$$

Finally, this molar amount is used to derive the mass of NaCl:

$$1.325\ \text{mol NaCl} \times \frac{58.44\ \text{g NaCl}}{\text{mol NaCl}} = 77.4\ \text{g NaCl} \tag{6.1.9}$$

Exercise 6.1.4

How many grams of $\ce{CaCl2}$ (110.98 g/mol) are contained in 250.0 mL of a 0.200-M solution of calcium chloride?

Answer

5.55 g $\ce{CaCl2}$

When performing calculations stepwise, as in Example **6.1.3**, it is important to refrain from rounding any intermediate calculation results, which can lead to rounding errors in the final result. In Example **6.1.4**, the molar amount of NaCl computed in the first step, 1.325 mol, would be properly rounded to 1.32 mol if it were to be reported; however, although the last digit (5) is not significant, it must be retained as a guard digit in the intermediate calculation. If we had not retained this guard digit, the final calculation for the mass of NaCl would have been 77.1 g, a difference of 0.3 g.

In addition to retaining a guard digit for intermediate calculations, we can also avoid rounding errors by performing computations in a single step (Example **6.1.5**). This eliminates intermediate steps so that only the final result is rounded.

Example 6.1.5: Determining the Volume of Solution

In Example **6.1.3**, we found the typical concentration of vinegar to be 0.839 *M*. What volume of vinegar contains 75.6 g of acetic acid?

Solution

First, use the molar mass to calculate moles of acetic acid from the given mass:

$$\text{g solute} \times \frac{\text{mol solute}}{\text{g solute}} = \text{mol solute} \tag{6.1.10}$$

Then, use the molarity of the solution to calculate the volume of solution containing this molar amount of solute:

$$\text{mol solute} \times \frac{\text{L solution}}{\text{mol solute}} = \text{L solution} \tag{6.1.11}$$

Combining these two steps into one yields:

$$\text{g solute} \times \frac{\text{mol solute}}{\text{g solute}} \times \frac{\text{L solution}}{\text{mol solute}} = \text{L solution} \tag{6.1.12}$$

$$75.6\ \text{g}\ CH_3CO_2H \left(\frac{\text{mol}\ CH_3CO_2H}{60.05\ \text{g}}\right)\left(\frac{\text{L solution}}{0.839\ \text{mol}\ CH_3CO_2H}\right) = 1.50\ \text{L solution} \tag{6.1.13}$$

Exercise 6.1.5:

What volume of a 1.50-M KBr solution contains 66.0 g KBr?

Answer

0.370 L

Mass-mass percent

Percentages are also commonly used to express the composition of mixtures, including solutions. The **mass percentage** (m/m%)of a solution component is defined as the ratio of the component's mass to the solution's mass, expressed as a percentage:

$$\%\ \text{m/m} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \tag{6.1.14}$$

mass of solution = mass of solute + mass solvent

If you can measure the masses of the solute and the solution, determining the mass/mass percent is easy. Each mass must be expressed in the same units to determine the proper concentration.

Suppose that a solution was prepared by dissolving **25.0 g** of sugar into **100.0 g** of water.

The mass of the solution is

mass of solution = 25.0g sugar + 100.0g water = 125.0 g

The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25.0\ \text{g sugar}}{125.0\ \text{g solution}} \times 100\% = 20.0\%\ \text{sugar} \tag{6.1.15}$$

Mass percentage is also referred to by similar names such as *percent mass*, *percent weight*, *weight/weight percent*, and other variations on this theme. The most common symbol for mass percentage is simply the percent sign, %, although more detailed symbols are often used including %mass, %weight, and (w/w)%. Use of these more detailed symbols can prevent confusion of mass percentages with other types of percentages, such as volume percentages (to be discussed later in this section).

Mass percentages are popular concentration units for consumer products. The label of a typical liquid bleach bottle (Figure 3.17) cites the concentration of its active ingredient, sodium hypochlorite (NaOCl), as being 7.4%. A 100.0-g sample of bleach would

therefore contain 7.4 g of NaOCl.



Figure* 6.1.4*: *Distilled white vinegar is a solution of acetic acid in water.*

{6}\): Determining the Volume of Solution

A 5.0-g sample of spinal fluid contains 3.75 mg (0.00375 g) of glucose. What is the percent by mass of glucose in spinal fluid?

Solution

The spinal fluid sample contains roughly 4 mg of glucose in 5000 mg of fluid, so the mass fraction of glucose should be a bit less than one part in 1000, or about 0.1%. Substituting the given masses into the equation defining mass percentage yields:

$$\%\,\text{glucose} = \frac{3.75\ \text{mg glucose} \times \frac{1\ \text{g}}{1000\ \text{mg}}}{5.0\ \text{g spinal fluid}} = 0.075\% \tag{6.1.16}$$

The computed mass percentage agrees with our rough estimate (it’s a bit less than 0.1%).

Note that while any mass unit may be used to compute a mass percentage (mg, g, kg, oz, and so on), the same unit must be used for both the solute and the solution so that the mass units cancel, yielding a dimensionless ratio. In this case, we converted the units of solute in the numerator from mg to g to match the units in the denominator. We could just as easily have converted the denominator from g to mg instead. As long as identical mass units are used for both solute and solution, the computed mass percentage will be correct.

Exercise 6.1.6:

A bottle of a tile cleanser contains 135 g of HCl and 775 g of water. What is the percent by mass of HCl in this cleanser?

Answer

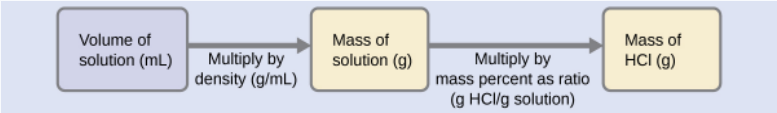
14.8%

Example 6.1.7: Calculations using Mass Percentage

“Concentrated” hydrochloric acid is an aqueous solution of 37.2% HCl that is commonly used as a laboratory reagent. The density of this solution is 1.19 g/mL. What mass of HCl is contained in 0.500 L of this solution?

Solution

The HCl concentration is near 40%, so a 100-g portion of this solution would contain about 40 g of HCl. Since the solution density isn’t greatly different from that of water (1 g/mL), a reasonable estimate of the HCl mass in 500 g (0.5 L) of the solution is about five times greater than that in a 100 g portion, or $5 \times 40 = 200\ \text{g}$. To derive the mass of solute in a solution from its mass percentage, we need to know the corresponding mass of the solution. Using the solution density given, we can convert the solution’s volume to mass, and then use the given mass percentage to calculate the solute mass. This mathematical approach is outlined in this flowchart:



For proper unit cancellation, the 0.500-L volume is converted into 500 mL, and the mass percentage is expressed as a ratio, 37.2 g HCl/g solution:

$$500 \text{ mL solution} \left(\frac{1.19 \text{ g solution}}{\text{mL solution}} \right) \left(\frac{37.2 \text{ g HCl}}{100 \text{ g solution}} \right) = 221 \text{ g HCl} \tag{6.1.17}$$

This mass of HCl is consistent with our rough estimate of approximately 200 g.

Exercise 6.1.7

What volume of concentrated HCl solution contains 125 g of HCl?

Answer

282 mL

Dilution of Solutions



Video* 6.1.3*: *An introduction to dilutions from Crash Course Chemistry.*

Dilution is the process whereby the concentration of a solution is lessened by the addition of solvent. For example, we might say that a glass of iced tea becomes increasingly diluted as the ice melts. The water from the melting ice increases the volume of the solvent (water) and the overall volume of the solution (iced tea), thereby reducing the relative concentrations of the solutes that give the beverage its taste (Figure 6.1.2).

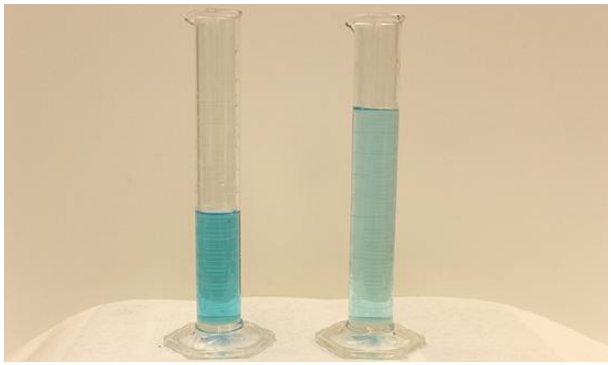


Figure* 6.1.5*: *Both solutions contain the same mass of copper nitrate. The solution on the right is more dilute because the copper nitrate is dissolved in more solvent. (credit: Mark Ott).*

Dilution is also a common means of preparing solutions of a desired concentration. By adding solvent to a measured portion of a more concentrated *stock solution*, we can achieve a particular concentration. For example, commercial pesticides are typically sold as solutions in which the active ingredients are far more concentrated than is appropriate for their application. Before they can be used on crops, the pesticides must be diluted. This is also a very common practice for the preparation of a number of common laboratory reagents (Figure 6.1.3).



Figure* 6.1.6: A solution of* $KMnO_4$ *is prepared by mixing water with 4.74 g of KMnO4 in a flask. (credit: modification of work by Mark Ott)*

A simple mathematical relationship can be used to relate the volumes and concentrations of a solution before and after the dilution process. According to the definition of molarity, the molar amount of solute in a solution is equal to the product of the solution's molarity and its volume in liters:

$$n = ML \tag{6.1.18}$$

Expressions like these may be written for a solution before and after it is diluted:

$$n_1 = M_1L_1 \tag{6.1.19}$$

$$n_2 = M_2L_2 \tag{6.1.20}$$

where the subscripts "1" and "2" refer to the solution before and after the dilution, respectively. Since the dilution process *does not change the amount of solute in the solution*, $n_1 = n_2$. Thus, these two equations may be set equal to one another:

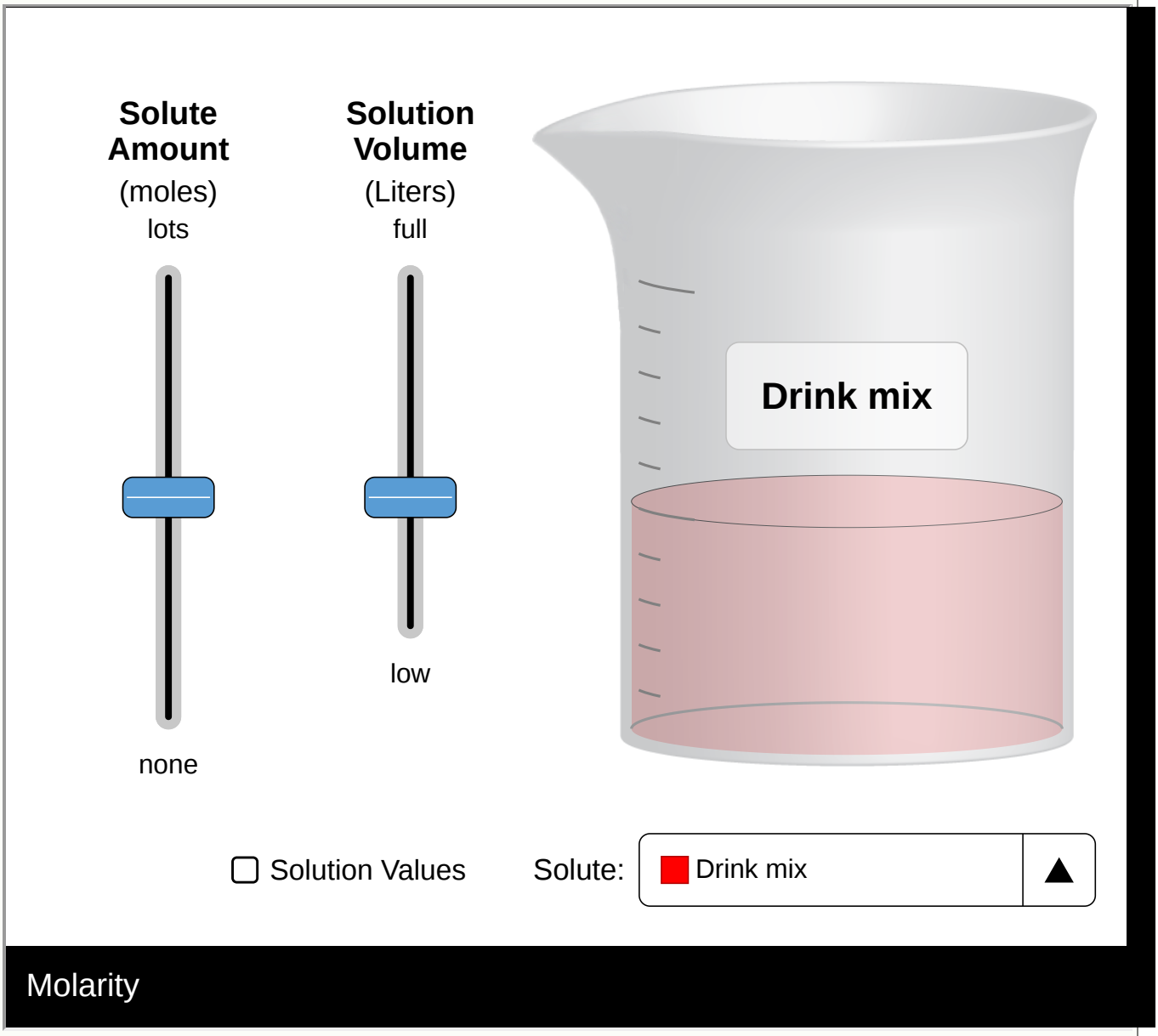
$$M_1L_1 = M_2L_2 \tag{6.1.21}$$

This relation is commonly referred to as the dilution equation. Although we derived this equation using molarity as the unit of concentration and liters as the unit of volume, other units of concentration and volume may be used, so long as the units properly cancel per the factor-label method. Reflecting this versatility, the dilution equation is often written in the more general form:

$$C_1V_1 = C_2V_2 \tag{6.1.22}$$

where C and V are concentration and volume, respectively.

Phet Simulation



Example 6.1.8: Determining the Concentration of a Diluted Solution

If 0.850 L of a 5.00-*M* solution of copper nitrate, $Cu(NO_3)_2$, is diluted to a volume of 1.80 L by the addition of water, what is the molarity of the diluted solution?

Solution

We are given the volume and concentration of a stock solution, V_1 and C_1, and the volume of the resultant diluted solution, V_2. We need to find the concentration of the diluted solution, C_2. We thus rearrange the dilution equation in order to isolate C_2:

$$C_1V_1 = C_2V_2 \tag{6.1.23}$$

$$C_2 = \frac{C_1V_1}{V_2} \tag{6.1.24}$$

Since the stock solution is being diluted by more than two-fold (volume is increased from 0.85 L to 1.80 L), we would expect the diluted solution's concentration to be less than one-half 5 *M*. We will compare this ballpark estimate to the calculated result

to check for any gross errors in computation (for example, such as an improper substitution of the given quantities). Substituting the given values for the terms on the right side of this equation yields:

$$C_2 = \frac{0.850\ \text{L} \times 5.00\ \frac{\text{mol}}{\text{L}}}{1.80\ \text{L}} = 2.36\ M \tag{6.1.25}$$

This result compares well to our ballpark estimate (it's a bit less than one-half the stock concentration, 5 *M*).

Exercise 6.1.8

What is the concentration of the solution that results from diluting 25.0 mL of a 2.04-M solution of CH3OH to 500.0 mL?

Answer

0.102 M CH_3OH

Example 6.1.9: Volume of a Diluted Solution

What volume of 0.12 *M* HBr can be prepared from 11 mL (0.011 L) of 0.45 *M* HBr?

Solution

We are given the volume and concentration of a stock solution, V_1 and C_1, and the concentration of the resultant diluted solution, C_2. We need to find the volume of the diluted solution, V_2. We thus rearrange the dilution equation in order to isolate V_2:

$$C_1V_1 = C_2V_2 \tag{6.1.26}$$

$$V_2 = \frac{C_1V_1}{C_2} \tag{6.1.27}$$

Since the diluted concentration (0.12 *M*) is slightly more than one-fourth the original concentration (0.45 *M*), we would expect the volume of the diluted solution to be roughly four times the original volume, or around 44 mL. Substituting the given values and solving for the unknown volume yields:

$$V_2 = \frac{(0.45\ M)(0.011\ \text{L})}{(0.12\ M)} \tag{6.1.28}$$

$$V_2 = 0.041\ \text{L} \tag{6.1.29}$$

The volume of the 0.12-*M* solution is 0.041 L (41 mL). The result is reasonable and compares well with our rough estimate.

Exercise 6.1.9

A laboratory experiment calls for 0.125 M HNO_3. What volume of 0.125 M HNO_3 can be prepared from 0.250 L of 1.88 M HNO_3?

Answer

3.76 L

Example 6.1.10: Volume of a Concentrated Solution Needed for Dilution

What volume of 1.59 *M* KOH is required to prepare 5.00 L of 0.100 *M* KOH?

Solution

We are given the concentration of a stock solution, C_1, and the volume and concentration of the resultant diluted solution, V_2 and C_2. We need to find the volume of the stock solution, V_1. We thus rearrange the dilution equation in order to isolate V_1:

$$C_1V_1 = C_2V_2 \tag{6.1.30}$$

$$V_1 = \frac{C_2 V_2}{C_1} \tag{6.1.31}$$

Since the concentration of the diluted solution 0.100 *M* is roughly one-sixteenth that of the stock solution (1.59 *M*), we would expect the volume of the stock solution to be about one-sixteenth that of the diluted solution, or around 0.3 liters. Substituting the given values and solving for the unknown volume yields:

$$V_1 = \frac{(0.100\ M)(5.00\ \text{L})}{1.59\ M} \tag{6.1.32}$$

$$V_1 = 0.314\ \text{L} \tag{6.1.33}$$

Thus, we would need 0.314 L of the 1.59-*M* solution to prepare the desired solution. This result is consistent with our rough estimate.

Exercise 6.1.10

What volume of a 0.575-M solution of glucose, $C_6H_{12}O_6$, can be prepared from 50.00 mL of a 3.00-M glucose solution?

Answer

0.261 L

Summary

Solutions are homogeneous mixtures. Many solutions contain one component, called the solvent, in which other components, called solutes, are dissolved. An aqueous solution is one for which the solvent is water. The concentration of a solution is a measure of the relative amount of solute in a given amount of solution. Concentrations may be measured using various units, with one very useful unit being molarity, defined as the number of moles of solute per liter of solution. The solute concentration of a solution may be decreased by adding solvent, a process referred to as dilution. The dilution equation is a simple relation between concentrations and volumes of a solution before and after dilution.



Video* 6.1.4*: *An review of calculating molarity.*



Video* 6.1.5*: *An review of calculating dilutions. Also, orange squash is a drink in England.*

Key Equations

- $M = \frac{\text{mol solute}}{\text{L solution}}$
- $C_1V_1 = C_2V_2$

Glossary

aqueous solution
solution for which water is the solvent

concentrated
qualitative term for a solution containing solute at a relatively high concentration

concentration
quantitative measure of the relative amounts of solute and solvent present in a solution

dilute
qualitative term for a solution containing solute at a relatively low concentration

dilution
process of adding solvent to a solution in order to lower the concentration of solutes

dissolved
describes the process by which solute components are dispersed in a solvent

molarity (*M*)
unit of concentration, defined as the number of moles of solute dissolved in 1 liter of solution

solute
solution component present in a concentration less than that of the solvent

solvent
solution component present in a concentration that is higher relative to other components

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6.1.1: Practice Problems- Solution Concentration

PROBLEM 6.1.1.1

Explain what changes and what stays the same when 1.00 L of a solution of NaCl is diluted to 1.80 L.

Answer

The number of moles always stays the same in a dilution.

The concentration and the volumes change in a dilution.

PROBLEM 6.1.1.2

What does it mean when we say that a 200-mL sample and a 400-mL sample of a solution of salt have the same molarity? In what ways are the two samples identical? In what ways are these two samples different?

Answer

The two samples contain the same proportion of moles of salt to liters of solution, but have different numbers of actual moles.

PROBLEM 6.1.1.3

Determine the molarity for each of the following solutions:

a. 0.444 mol of $CoCl_2$ in 0.654 L of solution
b. 98.0 g of phosphoric acid, H_3PO_4, in 1.00 L of solution
c. 0.2074 g of calcium hydroxide, $Ca(OH)_2$, in 40.00 mL of solution
d. 10.5 kg of $Na_2SO_4 \cdot 10H_2O$ in 18.60 L of solution
e. 7.0×10^{-3} mol of I_2 in 100.0 mL of solution
f. 1.8×10^4 mg of HCl in 0.075 L of solution

Answer a

0.679 *M*

Answer b

1.00 *M*

Answer c

0.06998 *M*

Answer d

1.75 *M*

Answer e

0.070 *M*

Answer f

6.6 *M*

Click here to see a video of the solution





PROBLEM 6.1.1.4

Determine the molarity of ea

a. 1.457 mol KCl in 1.500 L
b. 0.515 g of H_2SO_4 in 1.00
c. 20.54 g of $Al(NO_3)_3$ in 1
d. 2.76 kg of $CuSO_4 \cdot 5H_2O$
e. 0.005653 mol of Br_2 in 1
f. 0.000889 g of glycine, C

Answer a

0.9713 M

Answer b

5.25×10^{-3} M

Answer c

6.122×10^{-2} M

Answer d

7.62 M

Answer e

0.5653 M

Answer f

1.13×10^{-2} M

PROBLEM 6.1.1.5

Calculate the number of moles and the mass of the solute in each of the following solutions:

(a) 2.00 L of 18.5 *M* H_2SO_4, concentrated sulfuric acid
(b) 100.0 mL of 3.8×10^{-5} *M* NaCN, the minimum lethal concentration of sodium cyanide in blood serum
(c) 5.50 L of 13.3 *M* H_2CO, the formaldehyde used to “fix” tissue samples
(d) 325 mL of 1.8×10^{-6} *M* $FeSO_4$, the minimum concentration of iron sulfate detectable by taste in drinking water

Answer a

37.0 mol H_2SO_4

3.63×10^3 g H_2SO_4

Answer b

3.8×10^{-6} mol NaCN

1.9×10^{-4} g NaCN

Answer c

73.2 mol H_2CO

2.20 kg H_2CO

Answer d

5.9×10^{-7} mol $FeSO_4$

8.9×10^{-5} g $FeSO_4$

Click here to see a video of the solution



PROBLEM 6.1.1.6

Calculate the molarity of each of the following solutions:

(a) 0.195 g of cholesterol, $C_{27}H_{46}O$, in 0.100 L of serum, the average concentration of cholesterol in human serum
(b) 4.25 g of NH_3 in 0.500 L of solution, the concentration of NH_3 in household ammonia
(c) 1.49 kg of isopropyl alcohol, C_3H_7OH, in 2.50 L of solution, the concentration of isopropyl alcohol in rubbing alcohol
(d) 0.029 g of I_2 in 0.100 L of solution, the solubility of I_2 in water at 20 °C

Answer a

5.04×10^{-3} *M*

Answer b

0.499 *M*

Answer c

9.92 *M*

Answer d

1.1×10^{-3} *M*

PROBLEM 6.1.1.7

There is about 1.0 g of calcium, as Ca^{2+}, in 1.0 L of milk. What is the molarity of Ca^{2+} in milk?

Answer

0.025 *M*

Click here to see a video of the solution



PROBLEM 6.1.1.8

What volume of a 1.00-*M* $Fe(NO_3)_3$ solution can be diluted to prepare 1.00 L of a solution with a concentration of 0.250 *M*?

Answer

0.250 L

PROBLEM 6.1.1.9

If 0.1718 L of a 0.3556-*M* C_3H_7OH solution is diluted to a concentration of 0.1222 *M*, what is the volume of the resulting solution?

Answer

0.5000 L

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PROBLEM 6.1.1.10

What volume of a 0.33-*M* $C_{12}H_{22}O_{11}$ solution can be diluted to prepare 25 mL of a solution with a concentration of 0.025 *M*?

Answer

1.9 mL

PROBLEM 6.1.1.11

What is the concentration of the NaCl solution that results when 0.150 L of a 0.556-*M* solution is allowed to evaporate until the volume is reduced to 0.105 L?

Answer

0.794 M

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PROBLEM 6.1.1.12

What is the molarity of the diluted solution when each of the following solutions is diluted to the given final volume?

a. 1.00 L of a 0.250-*M* solution of $Fe(NO_3)_3$ is diluted to a final volume of 2.00 L
b. 0.5000 L of a 0.1222-*M* solution of C_3H_7OH is diluted to a final volume of 1.250 L
c. 2.35 L of a 0.350-*M* solution of H_3PO_4 is diluted to a final volume of 4.00 L
d. 22.50 mL of a 0.025-*M* solution of $C_{12}H_{22}O_{11}$ is diluted to 100.0 mL

Answer a

0.125 *M*

Answer b

0.04888 *M*

Answer c

0.206 *M*

Answer d

0.0056 *M*

PROBLEM 6.1.1.13

What is the final concentration of the solution produced when 225.5 mL of a 0.09988-*M* solution of Na_2CO_3 is allowed to evaporate until the solution volume is reduced to 45.00 mL?

Answer

0.5005 M

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PROBLEM 6.1.1.14

A 2.00-L bottle of a solution of concentrated HCl was purchased for the general chemistry laboratory. The solution contained 868.8 g of HCl. What is the molarity of the solution?

Answer

11.9 *M*

PROBLEM 6.1.1.15

An experiment in a general chemistry laboratory calls for a 2.00-*M* solution of HCl. How many mL of 11.9 *M* HCl would be required to make 250 mL of 2.00 *M* HCl?

Answer

42.0 mL

Click here to see a video of the solution



PROBLEM 6.1.1.16

What volume of a 0.20-*M* K_2SO_4 solution contains 57 g of K_2SO_4?

Answer

1.6 L

PROBLEM 6.1.1.17

The US Environmental Protection Agency (EPA) places limits on the quantities of toxic substances that may be discharged into the sewer system. Limits have been established for a variety of substances, including hexavalent chromium, which is limited to 0.50 mg/L. If an industry is discharging hexavalent chromium as potassium dichromate ($K_2Cr_2O_7$), what is the maximum permissible molarity of that substance?

Answer

$4.8 \times 10^{-6}\ M$

Click here to see a video of the solution



Contributors

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6.2: Solutions Chemistry

Skills to Develop

- Describe the basic properties of solutions and how they form
- Define and give examples of electrolytes
- Predict the solubility of common inorganic compounds by using solubility rules
- Write and balance chemical equations in molecular, total ionic, and net ionic formats.

The previous unit introduced *solutions*, defined as homogeneous mixtures of two or more substances. Often, one component of a solution is present at a significantly greater concentration, in which case it is called the solvent. The other components of the solution present in relatively lesser concentrations are called solutes. Sugar is a covalent solid composed of sucrose molecules, $C_{12}H_{22}O_{11}$. When this compound dissolves in water, its molecules become uniformly distributed among the molecules of water:

$$C_{12}H_{22}O_{11(s)} \longrightarrow C_{12}H_{22}O_{11(aq)} \tag{6.2.1}$$

The subscript "aq" in the equation signifies that the sucrose molecules are solutes and are therefore *individually dispersed* throughout the *aqueous solution* (water is the solvent). Although sucrose molecules are heavier than water molecules, they remain dispersed throughout the solution; gravity does not cause them to "settle out" over time.

Potassium dichromate, $K_2Cr_2O_7$, is an ionic compound composed of colorless potassium ions, K^+, and orange dichromate ions, $Cr_2O_7^{2-}$. When a small amount of solid potassium dichromate is added to water, the compound dissolves and dissociates to yield potassium ions and dichromate ions uniformly distributed throughout the mixture (Figure **6.2.1**), as indicated in this equation:

$$K_2Cr_2O_{7(s)} \longrightarrow 2K^+{}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} \tag{6.2.2}$$

As with the mixture of sugar and water, this mixture is also an aqueous solution. Its solutes, potassium and dichromate ions, remain individually dispersed among the solvent (water) molecules.



Figure* 6.2.1*: *When potassium dichromate ($K_2Cr_2O_7$) is mixed with water, it forms a homogeneous orange solution. (credit: modification of work by Mark Ott).*

Learn More about The dissolution of covalent and ionic compounds

Click the link to use a simulator.



Click to Run

Water is used so often as a solvent that the word solution has come to imply an aqueous solution to many people. However, almost any gas, liquid, or solid can act as a solvent. Many alloys are solid solutions of one metal dissolved in another; for example, US

five-cent coins contain nickel dissolved in copper. Air is a gaseous solution, a homogeneous mixture of nitrogen, oxygen, and several other gases. Oxygen (a gas), alcohol (a liquid), and sugar (a solid) all dissolve in water (a liquid) to form liquid solutions. Table **6.2.1** gives examples of several different solutions and the phases of the solutes and solvents.

***Table* 6.2.1:** *Different Types of Solutions*

Solution	Solute	Solvent
air	$O_2(g)$	$N_2(g)$
soft drinks	$CO_2(g)$	$H_2O(l)$
hydrogen in palladium	$H_2(g)$	$Pd(s)$
rubbing alcohol	$H_2O(l)$	$C_3H_8O(l)$ (2-propanol)
saltwater	$NaCl(s)$	$H_2O(l)$
brass	$Zn(s)$	$Cu(s)$

Solutions exhibit these defining traits:

- They are homogeneous; that is, after a solution is mixed, it has the same composition at all points throughout (its composition is uniform).
- The physical state of a solution—solid, liquid, or gas—is typically the same as that of the solvent, as demonstrated by the examples in Table **6.2.1**.
- The components of a solution are dispersed on a molecular scale; that is, they consist of a mixture of separated molecules, atoms, and/or ions.
- The dissolved solute in a solution will not settle out or separate from the solvent.
- The composition of a solution, or the concentrations of its components, can be varied continuously, within limits.

Electrolytes & Nonelectrolytes

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called electrolytes. Substances that do not yield ions when dissolved are called nonelectrolytes. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte. If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a weak electrolyte.

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure **6.2.6**).

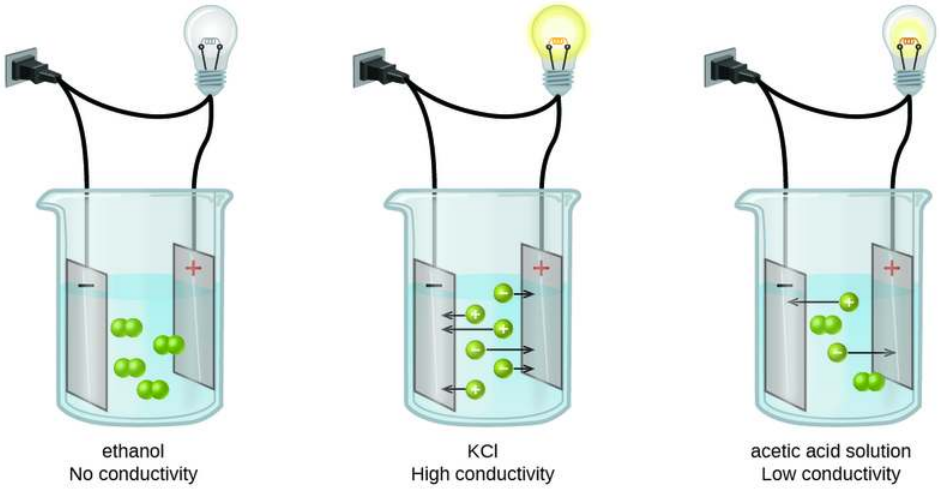


***Figure* 6.2.2**: *Solutions of nonelectrolytes such as ethanol do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.*

Ionic Electrolytes

Water and other polar molecules are attracted to ions, as shown in Figure **6.2.2**.

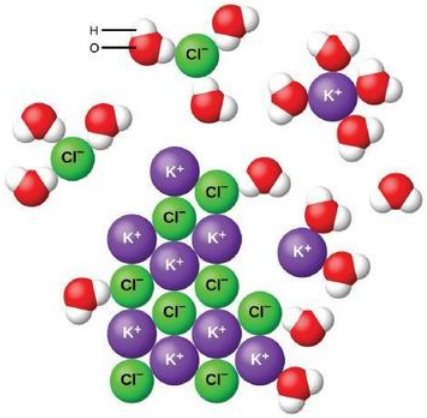


***Figure* 6.2.3:** *As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown.*

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

Let us consider what happens at the microscopic level when we add solid KCl to water. Forces attract the positive (hydrogen) end of the polar water molecules to the negative chloride ions at the surface of the solid, and they attract the negative (oxygen) ends to the positive potassium ions. The water molecules penetrate between individual K^+ and Cl^- ions and surround them, reducing the strong interionic forces that bind the ions together and letting them move off into solution as solvated ions, as Figure shows. The reduction of the electrostatic attraction permits the independent motion of each hydrated ion in a dilute solution, resulting in an increase in the disorder of the system as the ions change from their fixed and ordered positions in the crystal to mobile and much more disordered states in solution. This increased disorder is responsible for the dissolution of many ionic compounds, including KCl, which dissolve with absorption of heat.

In other cases, the electrostatic attractions between the ions in a crystal are so large that the increase in disorder cannot compensate for the energy required to separate the ions, and the crystal is insoluble. Such is the case for compounds such as calcium carbonate (limestone), calcium phosphate (the inorganic component of bone), and iron oxide (rust).

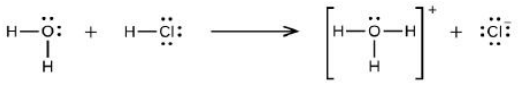
Covalent Electrolytes

Pure water is an extremely poor conductor of electricity because it is only very slightly ionized—only about two out of every 1 billion molecules ionize at 25 °C. Water ionizes when one molecule of water gives up a proton to another molecule of water, yielding hydronium and hydroxide ions.

$$\mathrm{H_2O\ (l)^+\ H_2O\,(l)\ \rightleftharpoons H_3O^+\,(aq) + OH^-\,(aq)} \tag{6.2.3}$$

In some cases, we find that solutions prepared from covalent compounds conduct electricity because the solute molecules react chemically with the solvent to produce ions. For example, pure hydrogen chloride is a gas consisting of covalent HCl molecules. This gas contains no ions. However, when we dissolve hydrogen chloride in water, we find that the solution is a very good conductor. The water molecules play an essential part in forming ions: Solutions of hydrogen chloride in many other solvents, such as benzene, do not conduct electricity and do not contain ions.

Hydrogen chloride is an *acid*, and so its molecules react with water, transferring H^+ ions to form hydronium ions (H_3O^+) and chloride ions (Cl^-):



This reaction is essentially 100% complete for HCl (i.e., it is a *strong acid* and, consequently, a strong electrolyte). Likewise, weak acids and bases that only react partially generate relatively low concentrations of ions when dissolved in water and are classified as weak electrolytes. The reader may wish to review the discussion of strong and weak acids provided in the earlier chapter of this text on reaction classes and stoichiometry.

Solubility

Solubility is the tendency of one substance to dissolve in another. In chemistry, the term solubility is often used to denote a substance's aqueous solubility, its ability to dissolve in water.In general if a substance does not dissolve in water it is labeled **insoluble** and if it does dissolve in water it is called as **soluble**. In fact, solubility is much more nuanced than the general terms soluble and insoluble describe. These nuances will be explored later in this course, for now, we will describe substances as soluble or insoluble.

A substance's solubility is determined by several factors, including the types and relative strengths of intermolecular attractive forces that may exist between the substances' atoms, ions, or molecules. For example, oil and water do not mix because the intermolecular forces that attract oil particles to one another and the intermolecular forces that attract water molecules together are much stronger than the intermolecular forces that attract oil particles to water molecules. Intermolecular forces is a subject that will be expanded on in greater detail later in this course.



Video* 6.2.1*: *An overview of solubility.*

Aqueous Solubility of Ionic Compounds

When studying chemical reactions, it is important to determine if a substance is soluble or insoluble in the reaction solvent. The next sections in this unit will cover common chemical reactions in aqueous solutions, therefore we must learn how to assess the soluble of substances in water. Table 6.2.1 outlines a set of rules that have been written based upon empirical evidence.

Table* 6.2.1*: *Solubilities of Common Ionic Compounds in Water*

Soluble compounds contain	**Exceptions** to these solubility rules include
• group 1 metal cations (Li^+, Na^+, K^+, Rb^+, and Cs^+) and ammonium ion (NH_4^+) • the halide ions (Cl^-, Br^-, and I^-) • the acetate ($C_2H_3O_2^-$), bicarbonate (HCO_3^-), nitrate (NO_3^-), and chlorate (ClO_3^-) ions • the sulfate (SO_4^-) ion	• halides of Ag^+, Hg_2^{2+}, and Pb^{2+} • sulfates of Ag^+, Ba^{2+}, Ca^{2+}, Hg_2^{2+}, Pb^{2+}, and Sr^{2+}
Insoluble compounds contain	**Exceptions** to these insolubility rules include
• carbonate (CO_3^{2-}), chromate (CrO_4^{2-}), phosphate (PO_4^{3-}), and sulfide (S^{2-}) ions • hydroxide ion (OH^-)	• compounds of these anions with group 1 metal cations and ammonium ion • hydroxides of group 1 metal cations and Ba^{2+}

Representing Aqueous Reactions

Given the abundance of water on earth, it stands to reason that a great many chemical reactions take place in aqueous media. When ions are involved in these reactions, the chemical equations may be written with various levels of detail appropriate to their intended use. To illustrate this, consider a reaction between ionic compounds taking place in an aqueous solution. When aqueous solutions of $CaCl_2$ and $AgNO_3$ are mixed, a reaction takes place producing aqueous $Ca(NO_3)_2$ and solid $AgCl$:

$$CaCl_2(aq) + 2\,AgNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2\,AgCl(s) \tag{6.2.4}$$

This balanced equation, derived in the usual fashion, is called a molecular equation because it doesn't explicitly represent the ionic species that are present in solution. When ionic compounds dissolve in water, they may *dissociate* into their constituent ions, which are subsequently dispersed homogenously throughout the resulting solution (a thorough discussion of this important process is provided in the chapter on solutions). Ionic compounds dissolved in water are, therefore, more realistically represented as dissociated ions, in this case:

$$CaCl_2(aq) \rightarrow Ca^{2+}(aq) + 2\,Cl^-(aq) \tag{6.2.5}$$

$$2\,AgNO_3(aq) \rightarrow 2\,Ag^+(aq) + 2\,NO_3{}^-(aq) \tag{6.2.6}$$

$$Ca(NO_3)_2(aq) \rightarrow Ca^{2+}(aq) + 2\,NO_3{}^-(aq) \tag{6.2.7}$$

Unlike these three ionic compounds, AgCl does not dissolve in water to a significant extent, as signified by its physical state notation, *(s)*.

Explicitly representing all dissolved ions results in a complete ionic equation. In this particular case, the formulas for the dissolved ionic compounds are replaced by formulas for their dissociated ions:

$$\begin{aligned} &Ca^{2+}(aq) + 2\,Cl^-(aq) + 2\,Ag^+(aq) + 2\,NO_3{}^-(aq) \\ &\rightarrow Ca^{2+}(aq) + 2\,NO_3{}^-(aq) + 2\,AgCl(s) \end{aligned} \tag{6.2.8}$$

Examining this equation shows that two chemical species are present in identical form on both sides of the arrow, $Ca^{2+}(aq)$ and $NO_3{}^-(aq)$. These spectator ions—ions whose presence is required to maintain charge neutrality—are neither chemically nor physically changed by the process, and so they may be eliminated from the equation to yield a more succinct representation called a net ionic equation:

$$\begin{aligned} &\cancel{Ca^{2+}(aq)} + 2\,Cl^-(aq) + 2\,Ag^+(aq) + \cancel{2\,NO_3{}^-(aq)} \\ &\rightarrow \cancel{Ca^{2+}(aq)} + \cancel{2\,NO_3{}^-(aq)} + 2\,AgCl(s) \end{aligned} \tag{6.2.9}$$

$$2\,Cl^-(aq) + 2\,Ag^+(aq) \rightarrow 2\,AgCl(s) \tag{6.2.10}$$

Following the convention of using the smallest possible integers as coefficients, this equation is then written:

$$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s) \tag{6.2.11}$$

This net ionic equation indicates that solid silver chloride may be produced from dissolved chloride and silver(I) ions, regardless of the source of these ions. These molecular and complete ionic equations provide additional information, namely, the ionic compounds used as sources of Cl^- and Ag^+.

Example 6.2.1: Molecular and Ionic Equations

When carbon dioxide is dissolved in an aqueous solution of sodium hydroxide, the mixture reacts to yield aqueous sodium carbonate and liquid water. Write balanced molecular, complete ionic, and net ionic equations for this process.

Solution

Begin by identifying formulas for the reactants and products and arranging them properly in chemical equation form:

$$CO_2\,(aq) + NaOH\,(aq) \rightarrow Na_2CO_3\,(aq) + H_2O\,(l) \quad \text{(unbalanced)}$$

Balance is achieved easily in this case by changing the coefficient for NaOH to 2, resulting in the molecular equation for this reaction:

$$CO_2(aq) + 2\,NaOH(aq) \rightarrow Na_2CO_3\,(aq) + H_2O(l)$$

The two dissolved ionic compounds, NaOH and Na_2CO_3, can be represented as dissociated ions to yield the complete ionic equation:

$$CO_2\,(aq) + 2\,Na^+\,(aq) + 2\,OH^-\,(aq) \rightarrow 2\,Na^+\,(aq) + CO_3{}^{2-}\,(aq) + H_2O\,(l)$$

Finally, identify the spectator ion(s), in this case $Na^+(aq)$, and remove it from each side of the equation to generate the net ionic equation:

$$\begin{aligned} &CO_2(aq) + \cancel{2\,Na^+(aq)} + 2\,OH^-(aq) \rightarrow \cancel{2\,Na^+(aq)} + CO_3{}^{2-}(aq) + H_2O(l) \\ &CO_2(aq) + 2\,OH^-(aq) \rightarrow CO_3{}^{2-}(aq) + H_2O(l) \end{aligned}$$

Exercise 6.2.1

Diatomic chlorine and sodium hydroxide (lye) are commodity chemicals produced in large quantities, along with diatomic hydrogen, via the electrolysis of brine, according to the following unbalanced equation:

$$\mathrm{NaCl\,(aq) + H_2O\,(l) \xrightarrow{electricity} NaOH\,(aq) + H_2\,(g) + Cl_2\,(g)}$$

Write balanced molecular, complete ionic, and net ionic equations for this process.

Answer

Balanced molecular equation:

$$\mathrm{2\,NaCl\,(aq) + 2\,H_2O\,(l) \rightarrow 2\,NaOH\,(aq) + H_2\,(g) + Cl_2\,(g)}$$

Balanced ionic equation:

$$\mathrm{2\,Na^+\,(aq) + 2\,Cl^-\,(aq) + 2\,H_2O\,(l) \rightarrow 2\,Na^+\,(aq) + 2\,OH^-\,(aq) + H_2\,(g) + Cl_2\,(g)}$$

Balanced net ionic equation:

$$\mathrm{2\,Cl^-\,(aq) + 2\,H_2O\,(l) \rightarrow 2\,OH^-\,(aq) + H_2\,(g) + Cl_2\,(g)}$$

Summary



Video* 6.2.5*: *An overview of solutions and electrolytes.*

A solution forms when two or more substances combine physically to yield a mixture that is homogeneous at the molecular level. The solvent is the most concentrated component and determines the physical state of the solution. The solutes are the other components typically present at concentrations less than that of the solvent. Solutions may form endothermically or exothermically, depending upon the relative magnitudes attractive forces we will discuss in the second term of this course. Ideal solutions form with no appreciable change in energy.

Substances that dissolve in water to yield ions are called electrolytes. Electrolytes may be covalent compounds that chemically react with water to produce ions (for example, acids and bases), or they may be ionic compounds that dissociate to yield their constituent cations and anions, when dissolved. Soluble ionic substances and strong acids ionize completely and are strong electrolytes, while weak acids and bases ionize to only a small extent and are weak electrolytes. Nonelectrolytes are substances that do not produce ions when dissolved in water.

Glossary

alloy

solid mixture of a metallic element and one or more additional elements

dissociation

physical process accompanying the dissolution of an ionic compound in which the compound's constituent ions are solvated and dispersed throughout the solution

complete ionic equation

chemical equation in which all dissolved ionic reactants and products, including spectator ions, are explicitly represented by formulas for their dissociated ions

electrolyte

substance that produces ions when dissolved in water

insoluble

of relatively low solubility; dissolving only to a slight extent

molecular equation

chemical equation in which all reactants and products are represented as neutral substances

net ionic equation

chemical equation in which only those dissolved ionic reactants and products that undergo a chemical or physical change are represented (excludes spectator ions)

nonelectrolyte

substance that does not produce ions when dissolved in water

soluble

of relatively high solubility; dissolving to a relatively large extent

solubility

the extent to which a substance may be dissolved in water, or any solvent

strong electrolyte

substance that dissociates or ionizes completely when dissolved in water

supersaturated

of concentration that exceeds solubility; a nonequilibrium state

weak electrolyte

substance that ionizes only partially when dissolved in water

Contributors

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6.2.1: Practice Problems- Solutions Chemistry

PROBLEM 6.2.1.1

How do solutions differ from compounds? From other mixtures?

Answer

A solution can vary in composition, while a compound cannot vary in composition. Solutions are homogeneous at the molecular level, while other mixtures are heterogeneous.

PROBLEM 6.2.1.2

Give an example of each of the following types of solutions:

a. a gas in a liquid
b. a gas in a gas
c. a solid in a solid

Answer a

CO_2 in water

Answer b

O_2 in N_2 (air)

Answer c

bronze (solution of tin or other metals in copper)

PROBLEM 6.2.1.3

Solutions of hydrogen in palladium may be formed by exposing Pd metal to H_2 gas. The concentration of hydrogen in the palladium depends on the pressure of H_2 gas applied, but in a more complex fashion than can be described by Henry's law. Under certain conditions, 0.94 g of hydrogen gas is dissolved in 215 g of palladium metal. Determine the molarity of this solution (solution density = 1.8 g/cm^3).

Answer

3.91 M

Click here to see a video of the solution



PROBLEM 6.2.1.4

Explain why solutions of HBr in benzene (a nonpolar solvent) are nonconductive, while solutions in water (a polar solvent) are conductive.

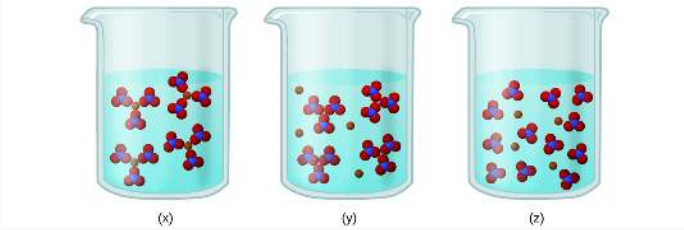
Answer

HBr is an acid and so its molecules react with water molecules to form H_3O^+ and Br^- ions that provide conductivity to the solution. Though HBr is soluble in benzene, it does not react chemically but remains dissolved as neutral HBr molecules. With no ions present in the benzene solution, it is electrically nonconductive.

PROBLEM 6.2.1.5

Consider the solutions presented:

(a) Which of the following sketches best represents the ions in a solution of $Fe(NO_3)_3(aq)$?



(b) Write a balanced chemical equation showing the products of the dissolution of $Fe(NO_3)_3$.

Answer a

$Fe(NO_3)_3$ is a strong electrolyte, thus it should completely dissociate into Fe^{3+} and (NO_3^-) ions. Therefore, (z) best represents the solution.

Answer b

$$Fe(NO_3)_3(s) \longrightarrow Fe^{3+}(aq) + 3\,NO_3^-(aq)$$

PROBLEM 6.2.1.6

What is the expected electrical conductivity of the following solutions?

a. NaOH(aq)
b. HCl(aq)
c. $C_6H_{12}O_6(aq)$ (glucose)
d. $NH_3(l)$

Answer a

high conductivity (solute is an ionic compound that will dissociate when dissolved)

Answer b

high conductivity (solute is a strong acid and will ionize completely when dissolved)

Answer c

nonconductive (solute is a covalent compound, neither acid nor base, unreactive towards water)

Answer d

low conductivity (solute is a weak base and will partially ionize when dissolved)

PROBLEM 6.2.1.7

Why are most *solid* ionic compounds electrically nonconductive, whereas aqueous solutions of ionic compounds are good conductors? Would you expect a *liquid* (molten) ionic compound to be electrically conductive or nonconductive? Explain.

Answer

A medium must contain freely mobile, charged entities to be electrically conductive. The ions present in a typical ionic solid are immobilized in a crystalline lattice and so the solid is not able to support an electrical current. When the ions are mobilized, either by melting the solid or dissolving it in water to dissociate the ions, current may flow and these forms of the ionic compound are conductive.

PROBLEM 6.2.1.8

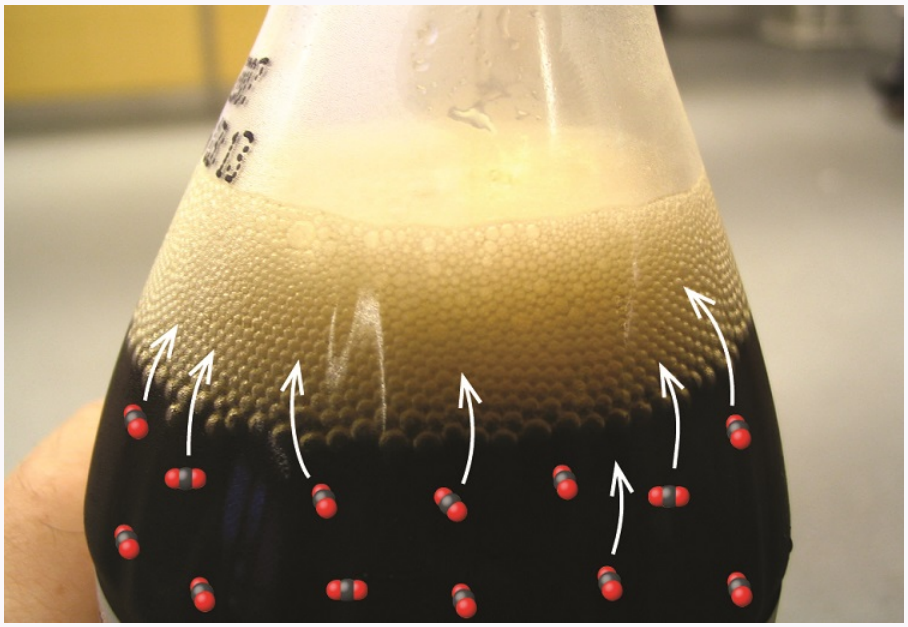
Suppose you are presented with a clear solution of sodium thiosulfate, $Na_2S_2O_3$. How could you determine whether the solution is unsaturated, saturated, or supersaturated?

Answer

Add a small crystal of $\boldsymbol{Na_2S_2O_3}$. It will dissolve in an unsaturated solution, remain apparently unchanged in a saturated solution, or initiate precipitation in a supersaturated solution.

PROBLEM 6.2.1.9

Refer to following figure for the following three questions:



a. How did the concentration of dissolved CO_2 in the beverage change when the bottle was opened?
b. What caused this change?
c. Is the beverage unsaturated, saturated, or supersaturated with CO_2?

Answer a

It decreased as some of the CO_2 gas left the solution (evidenced by effervescence).

Answer b

Opening the bottle released the high-pressure CO_2 gas above the beverage. The reduced CO_2 gas pressure, per Henry's law, lowers the solubility for CO_2.

Answer c

The dissolved CO_2 concentration will continue to slowly decrease until equilibrium is reestablished between the beverage and the very low CO_2 gas pressure in the opened bottle. Immediately after opening, the beverage, therefore, contains dissolved CO_2 at a concentration greater than its solubility, a nonequilibrium condition, and is said to be supersaturated.

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6.3: Precipitation Reactions

Skills to Develop

- Define precipitation reactions
- Classify chemical reactions as one of these three types given appropriate descriptions or chemical equations
- Write and balance precipitation equations in molecular, total ionic, and net ionic formats.

Faced with a wide range of varied interactions between chemical substances, scientists have found it convenient (or even necessary) to classify chemical interactions by identifying common patterns of reactivity. The following three modules will provide an introduction to three of the most prevalent types of chemical reactions: precipitation, acid-base, and oxidation-reduction.

Precipitation Reactions and Solubility Rules

A precipitation reaction is one in which dissolved substances react to form one (or more) solid products. Many reactions of this type involve the exchange of ions between ionic compounds in aqueous solution and are sometimes referred to as *double displacement*, *double replacement*, or *metathesis* reactions. These reactions are common in nature and are responsible for the formation of coral reefs in ocean waters and kidney stones in animals. They are used widely in industry for production of a number of commodity and specialty chemicals. Precipitation reactions also play a central role in many chemical analysis techniques, including spot tests used to identify metal ions and *gravimetric methods* for determining the composition of matter (see the last module of this chapter).

The extent to which a substance may be dissolved in water, or any solvent, is quantitatively expressed as its solubility, defined as the maximum concentration of a substance that can be achieved under specified conditions. Substances with relatively large solubilities are said to be soluble. A substance will precipitate when solution conditions are such that its concentration exceeds its solubility. Substances with relatively low solubilities are said to be insoluble, and these are the substances that readily precipitate from solution. More information on these important concepts is provided in the text chapter on solutions. For purposes of predicting the identities of solids formed by precipitation reactions, one may simply refer to patterns of solubility that have been observed for many ionic compounds (Table **6.3.1**).

***Table* 6.3.1:** *Solubilities of Common Ionic Compounds in Water*

Soluble compounds contain	**Exceptions** to these solubility rules include
• group 1 metal cations (Li^+, Na^+, K^+, Rb^+, and Cs^+) and ammonium ion (NH_4^+) • the halide ions (Cl^-, Br^-, and I^-) • the acetate ($C_2H_3O_2^-$), bicarbonate (HCO_3^-), nitrate (NO_3^-), and chlorate (ClO_3^-) ions • the sulfate (SO_4^-) ion	• halides of Ag^+, Hg_2^{2+}, and Pb^{2+} • sulfates of Ag^+, Ba^{2+}, Ca^{2+}, Hg_2^{2+}, Pb^{2+}, and Sr^{2+}
Insoluble compounds contain	**Exceptions** to these insolubility rules include
• carbonate (CO_3^{2-}), chromate (CrO_4^{2-}), phosphate (PO_4^{3-}), and sulfide (S^{2-}) ions • hydroxide ion (OH^-)	• compounds of these anions with group 1 metal cations and ammonium ion • hydroxides of group 1 metal cations and Ba^{2+}

A vivid example of precipitation is observed when solutions of potassium iodide and lead nitrate are mixed, resulting in the formation of solid lead iodide:

$$2\,KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\,KNO_3(aq) \tag{6.3.1}$$

This observation is consistent with the solubility guidelines: The only insoluble compound among all those involved is lead iodide, one of the exceptions to the general solubility of iodide salts.

The net ionic equation representing this reaction is:

$$Pb^{2+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s) \tag{6.3.2}$$

Lead iodide is a bright yellow solid that was formerly used as an artist's pigment known as iodine yellow (Figure **6.3.1**). The properties of pure PbI_2 crystals make them useful for fabrication of X-ray and gamma ray detectors.



***Video* 6.3.1**: *A precipitate of PbI_2 forms when solutions containing Pb^{2+} and I^- are mixed*. *Credit: MrLundScience*

The solubility guidelines in Table **6.3.1** may be used to predict whether a precipitation reaction will occur when solutions of soluble ionic compounds are mixed together. One merely needs to identify all the ions present in the solution and then consider if possible cation/anion pairing could result in an insoluble compound. For example, mixing solutions of silver nitrate and sodium fluoride will yield a solution containing Ag^+, NO_3^-, Na^+, and F^- ions. Aside from the two ionic compounds originally present in the solutions, $AgNO_3$ and NaF, two additional ionic compounds may be derived from this collection of ions: $NaNO_3$ and AgF. The solubility guidelines indicate all nitrate salts are soluble but that AgF is one of the exceptions to the general solubility of fluoride salts. A precipitation reaction, therefore, is predicted to occur, as described by the following equations:

$$NaF(aq) + AgNO_3(aq) \rightarrow AgF(s) + NaNO_3(aq) \quad \text{(molecular)} \tag{6.3.3}$$

$$Ag^+(aq) + F^-(aq) \rightarrow AgF(s) \quad \text{(net ionic)} \tag{6.3.4}$$

Example 6.3.1: Predicting Precipitation Reactions

Predict the result of mixing reasonably concentrated solutions of the following ionic compounds. If precipitation is expected, write a balanced net ionic equation for the reaction.

a. potassium sulfate and barium nitrate
b. lithium chloride and silver acetate
c. lead nitrate and ammonium carbonate

Solution

(a) The two possible products for this combination are KNO_3 and $BaSO_4$. The solubility guidelines indicate $BaSO_4$ is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

$$Ba^{2+}(aq) + {SO_4}^{2-}(aq) \rightarrow BaSO_4(s)$$

(b) The two possible products for this combination are $LiC_2H_3O_2$ and AgCl. The solubility guidelines indicate AgCl is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

(c) The two possible products for this combination are $PbCO_3$ and NH_4NO_3. The solubility guidelines indicate $PbCO_3$ is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is

$$Pb^{2+}(aq) + CO_3^{2-}(aq) \rightarrow PbCO_3(s)$$

Exercise 6.3.1

Which solution could be used to precipitate the barium ion, Ba^{2+}, in a water sample: sodium chloride, sodium hydroxide, or sodium sulfate? What is the formula for the expected precipitate?

Answer

sodium sulfate, $BaSO_4$

Double Displacement Reactions

Precipitation reactions are an example of a double-displacement reaction.

In a **double replacement reaction**, which can also be referred to as a **double displacement reaction**, either the cationic or the anionic portions of two compounds exchange their relative positions. The unique characteristic of a double replacement reaction is that the components that are repositioned must both exist within *compounds*.

A double replacement reaction can be represented symbolically, as shown below.

AD+QZ→QD+AZAD+QZ→QD+AZ

The following pattern *also* accurately reflects the description that is provided above.

AD+QZ→AZ+QDAD+QZ→AZ+QD

The inclusion of two letters within each of the symbolic representations that are shown above indicates that the reactants and products for both reactions are all compounds. Furthermore, these equations share a common pair of reactants *and* products. The only discernable difference between these patterns is the order in which the products, "QD" and "AZ," are written. However, the *order* in which chemicals are written has no impact on the classification of a reaction, as long as the position of those substances is consistent, relative to the reaction arrow.

A characteristic of a double-replacement equation is that there are two compounds as reactants and two different compounds as products. An example is

$$CuCl_2(aq) + 2\,AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2\,AgCl(s) \tag{6.3.5}$$

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.

Summary



Video* 6.3.2: A review of precipitation reactions.*

Chemical reactions are classified according to similar patterns of behavior. A large number of important reactions are included in three categories: precipitation, acid-base, and oxidation-reduction (redox). Precipitation reactions involve the formation of one or more insoluble products.

Glossary

insoluble
of relatively low solubility; dissolving only to a slight extent

precipitate
insoluble product that forms from reaction of soluble reactants

precipitation reaction
reaction that produces one or more insoluble products; when reactants are ionic compounds, sometimes called double-displacement or metathesis

double-displacement reaction
a type of reaction that occurs when the cations and anions switch between two reactants to form new products

soluble
of relatively high solubility; dissolving to a relatively large extent

solubility
the extent to which a substance may be dissolved in water, or any solvent

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- Adelaide Clark, Oregon Institute of Technology
- Crash Course Chemistry: Crash Course is a division of Complexly and videos are free to stream for educational purposes.
- Types of Chemical Reactions Rachael Harper Delupio, Lost Angles Trade Technical College

- Classifying Chemical Reactions,Nicola Burrmann, Heartland Community College

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6.3.1: Practice Problems- Precipitation

PROBLEM 6.3.1.1

Aqueous hydrogen fluoride (hydrofluoric acid) is used to etch glass and to analyze minerals for their silicon content. Hydrogen fluoride will also react with sand (silicon dioxide).

a. Write an equation for the reaction of solid silicon dioxide with hydrofluoric acid to yield gaseous silicon tetrafluoride and liquid water.
b. The mineral fluorite (calcium fluoride) occurs extensively in Illinois. Solid calcium fluoride can also be prepared by the reaction of aqueous solutions of calcium chloride and sodium fluoride, yielding aqueous sodium chloride as the other product. Write complete and net ionic equations for this reaction.

Answer a

$$4\,HF(aq) + SiO_2(s) \rightarrow SiF_4(g) + 2\,H_2O(l)$$

Answer b

complete ionic equation: $2\,Na^+(aq) + 2\,F^-(aq) + Ca^{2+}(aq) + 2\,Cl^-(aq) \rightarrow CaF_2(s) + 2\,Na^+(aq) + 2\,Cl^-(aq)$

net ionic equation: $2\,F^-(aq) + Ca^{2+}(aq) \rightarrow CaF_2(s)$

PROBLEM 6.3.1.2

From the balanced molecular equations, write the complete ionic and net ionic equations for the following:

a. $K_2C_2O_4(aq) + Ba(OH)_2(aq) \rightarrow 2\,KOH(aq) + BaC_2O_2(s)$
b. $Pb(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow PbSO_4(s) + 2\,HNO_3(aq)$
c. $CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$

Answer a

$2\,K^+(aq) + C_2O_4{}^{2-}(aq) + Ba^{2+}(aq) + 2\,OH^-(aq) \rightarrow 2\,K^+(aq) + 2\,OH^-(aq) + BaC_2O_4(s)$ (complete)

$Ba^{2+}(aq) + C_2O_4{}^{2-}(aq) \rightarrow BaC_2O_4(s)$ (net)

Answer b

$Pb^{2+}(aq) + 2\,NO_3{}^-(aq) + 2\,H^+(aq) + SO_4{}^{2-}(aq) \rightarrow PbSO_4(s) + 2\,H^+(aq) + 2\,NO_3{}^-(aq)$ (complete)

$Pb^{2+}(aq) + SO_4{}^{2-}(aq) \rightarrow PbSO_4(s)$ (net)

Answer c

$CaCO_3(s) + 2\,H^+(aq) + SO_4{}^{2-}(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$ (complete)

$CaCO_3(s) + 2\,H^+(aq) + SO_4{}^{2-}(aq) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l)$ (net)

PROBLEM 6.3.1.3

Use the following equations to answer the next five questions:

i. $H_2O(s) \rightarrow H_2O(l)$
ii. $Na^+(aq) + Cl^-(aq)Ag^+(aq) + NO_3{}^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3{}^-(aq)$
iii. $CH_3OH(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
iv. $2\,H_2O(l) \rightarrow 2\,H_2(g) + O_2(g)$
v. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

a. Which equation describes a physical change?
b. Which equation identifies the reactants and products of a combustion reaction?

c. Which equation is not balanced?
d. Which is a net ionic equation?

Answer a

i. $H_2O(solid) \rightarrow H_2O(liquid)$

Answer b

iii.

Answer c

iii. $2\,CH_3OH(g) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 4\,H_2O(g)$

Answer d

v.

PROBLEM 6.3.1.4

Write the molecular, total ionic, and net ionic equations for the following reactions:

a. $Ca(OH)_2(aq) + HC_2H_3O_2(aq) \rightarrow$
b. $H_3PO_4(aq) + CaCl_2(aq) \rightarrow$

Answer a

molecular: $Ca(OH)_2(aq) + 2\,HC_2H_3O_2(aq) \rightarrow Ca(C_2H_3O_2)_2(aq) + 2\,H_2O(l)$

complete ionic: $Ca^{2+}(aq) + 2\,OH^-(aq) + 2\,H^+(aq) + 2\,C_2H_3O_2{}^-(aq) \rightarrow Ca^{2+}(aq) + 2\,C_2H_3O_2{}^-(aq) + 2\,H_2O(l)$

net ionic: $2\,OH^-(aq) + 2\,H^+(aq) \rightarrow 2\,H_2O(l)$

Answer b

molecular: $2\,H_3PO_4(aq) + 3\,CaCl_2(aq) \rightarrow Ca_3(PO_4)_2(s) + 6\,HCl(aq)$

complete ionic: $6\,H^+(aq) + 2\,PO_4{}^{3-}(aq) + 3\,Ca^{2+}(aq) + 6\,Cl^-(aq) \rightarrow Ca_3(PO_4)_2(s) + 6\,H^+(aq) + 6\,Cl^-(aq)$

net ionic: $2\,PO_4{}^{3-}(aq) + 3\,Ca^{2+}(aq) \rightarrow Ca_3(PO_4)_2(s)$

Click here to see a video of the solution



PROBLEM 6.3.1.5

Great Lakes Chemical Company produces bromine, Br_2, from bromide salts such as NaBr, in Arkansas brine by treating the brine with chlorine gas. Write a balanced equation for the reaction of NaBr with Cl_2.

Answer

$$2\,NaBr(aq) + Cl_2(g) \rightarrow 2\,NaCl(aq) + Br_2(l)$$

Contributors

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6.4: Classifying Chemical Reactions (Acids and Bases)

Skills to Develop

- Define three common types of chemical reactions (precipitation, acid-base, and oxidation-reduction)
- Classify chemical reactions as one of these three types given appropriate descriptions or chemical equations
- Identify common acids and bases

Acid-Base Reactions



Video* 6.4.1*: *An introduction to acids and bases.*

An acid-base reaction is one in which a hydrogen ion, H^+, is transferred from one chemical species to another. Such reactions are of central importance to numerous natural and technological processes, ranging from the chemical transformations that take place within cells and the lakes and oceans, to the industrial-scale production of fertilizers, pharmaceuticals, and other substances essential to society. The subject of acid-base chemistry, therefore, is worthy of thorough discussion, which is unfortunately beyond the scope of this course.

For purposes of this brief introduction, we will consider only the more common types of acid-base reactions that take place in aqueous solutions. In this context, an acid is a substance that will dissolve in water to yield hydronium ions, H_3O^+. As an example, consider the equation shown here:

$$\mathbf{HCl}(aq) + \mathbf{H_2O}(aq) \rightarrow \mathbf{Cl^-}(aq) + \mathbf{H_3O^+}(aq) \tag{6.4.1}$$

The process represented by this equation confirms that hydrogen chloride is an acid. When dissolved in water, H_3O^+ ions are produced by a chemical reaction in which H^+ ions are transferred from HCl molecules to H_2O molecules (Figure **6.4.1**).

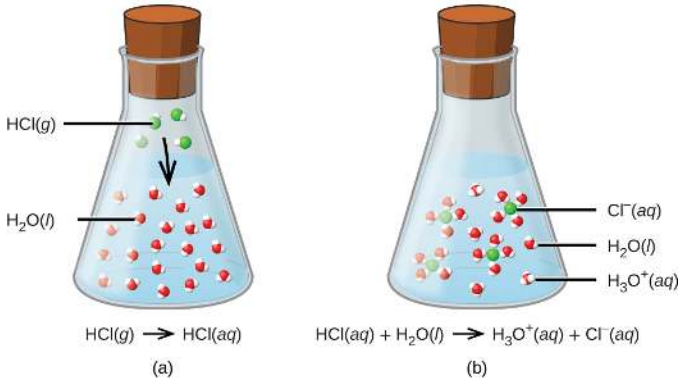


Figure* 6.4.1*: *When hydrogen chloride gas dissolves in water, (a) it reacts as an acid, transferring protons to water molecules to yield (b) hydronium ions (and solvated chloride ions)*

The nature of HCl is such that its reaction with water as just described is essentially 100% efficient: Virtually every HCl molecule that dissolves in water will undergo this reaction. Acids that completely react in this fashion are called strong acids, and HCl is one among just a handful of common acid compounds that are classified as strong (Table **6.4.1**). A far greater number of compounds behave as weak acids and only partially react with water, leaving a large majority of dissolved molecules in their original form and generating a relatively small amount of hydronium ions. Weak acids are commonly encountered in nature, being the substances partly responsible for the tangy taste of citrus fruits, the stinging sensation of insect bites, and the unpleasant smells associated with body odor. A familiar example of a weak acid is acetic acid, the main ingredient in food vinegars:

$$\mathrm{CH_3CO_2H}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{CH_3CO_2^-}(aq) + \mathrm{H_3O^+}(aq) \tag{6.4.2}$$

When dissolved in water under typical conditions, only about 1% of acetic acid molecules are present in the ionized form, $\mathrm{CH_3CO_2^-}$ (Figure **6.4.2**). (The use of a double-arrow in the equation above denotes the partial reaction aspect of this process, a concept addressed fully in the chapters on chemical equilibrium.)



(a) (b)

Figure* 6.4.2*: (a) Fruits such as oranges, lemons, and grapefruit contain the weak acid citric acid. (b) Vinegars contain the weak acid acetic acid. (credit a: modification of work by Scott Bauer; credit b: modification of work by Brücke-Osteuropa/Wikimedia Commons)

Table* 6.4.1*: Common Strong Acids

Compound Formula	Name in Aqueous Solution
HBr	hydrobromic acid
HCl	hydrochloric acid
HI	hydroiodic acid
HNO_3	nitric acid
$HClO_4$	perchloric acid
H_2SO_4	sulfuric acid

A base is a substance that will dissolve in water to yield hydroxide ions, OH^-. The most common bases are ionic compounds composed of alkali or alkaline earth metal cations (groups 1 and 2) combined with the hydroxide ion—for example, NaOH and $Ca(OH)_2$. When these compounds dissolve in water, hydroxide ions are released directly into the solution. For example, KOH and $Ba(OH)_2$ dissolve in water and dissociate completely to produce cations (K^+ and Ba^{2+}, respectively) and hydroxide ions, OH^-. These bases, along with other hydroxides that completely dissociate in water, are considered strong bases.

Consider as an example the dissolution of lye (sodium hydroxide) in water:

$$\mathrm{NaOH}(s) \rightarrow \mathrm{Na^+}(aq) + \mathrm{OH^-}(aq) \tag{6.4.3}$$

This equation confirms that sodium hydroxide is a base. When dissolved in water, NaOH dissociates to yield Na^+ and OH^- ions. This is also true for any other ionic compound containing hydroxide ions. Since the dissociation process is essentially complete when ionic compounds dissolve in water under typical conditions, NaOH and other ionic hydroxides are all classified as strong bases.

Unlike ionic hydroxides, some compounds produce hydroxide ions when dissolved by chemically reacting with water molecules. In all cases, these compounds react only partially and so are classified as weak bases. These types of compounds are also abundant in nature and important commodities in various technologies. For example, global production of the weak base ammonia is typically well over 100 metric tons annually, being widely used as an agricultural fertilizer, a raw material for chemical synthesis of other compounds, and an active ingredient in household cleaners (Figure **6.4.3**). When dissolved in water, ammonia reacts partially to yield hydroxide ions, as shown here:

$$\mathrm{NH_3}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{NH_4^+}(aq) + \mathrm{OH^-}(aq) \tag{6.4.4}$$

This is, by definition, an acid-base reaction, in this case involving the transfer of H^+ ions from water molecules to ammonia molecules. Under typical conditions, only about 1% of the dissolved ammonia is present as $\mathrm{NH_4^+}$ ions.



Figure* 6.4.3*: *Ammonia is a weak base used in a variety of applications. (a) Pure ammonia is commonly applied as an agricultural fertilizer. (b) Dilute solutions of ammonia are effective household cleansers. (credit a: modification of work by National Resources Conservation Service; credit b: modification of work by pat00139)*

The chemical reactions described in which acids and bases dissolved in water produce hydronium and hydroxide ions, respectively, are, by definition, acid-base reactions. In these reactions, water serves as both a solvent *and* a reactant. A neutralization reaction is a specific type of acid-base reaction in which the reactants are an acid and a base, the products are often a salt and water, and neither reactant is the water itself:

$$\mathrm{acid} + \mathrm{base} \rightarrow \mathrm{salt} + \mathrm{water} \tag{6.4.5}$$

To illustrate a neutralization reaction, consider what happens when a typical antacid such as milk of magnesia (an aqueous suspension of solid $Mg(OH)_2$) is ingested to ease symptoms associated with excess stomach acid (HCl):

$$\mathrm{Mg(OH)_2}(s) + 2\,\mathrm{HCl}(aq) \rightarrow \mathrm{MgCl_2}(aq) + 2\,\mathrm{H_2O}(l). \tag{6.4.6}$$

Note that in addition to water, this reaction produces a salt, magnesium chloride.

Example 6.4.1: Writing Equations for Acid-Base Reactions

Write balanced chemical equations for the acid-base reactions described here:

a. the weak acid hydrogen hypochlorite reacts with water
b. a solution of barium hydroxide is neutralized with a solution of nitric acid

Solution

(a) The two reactants are provided, HOCl and H_2O. Since the substance is reported to be an acid, its reaction with water will involve the transfer of H^+ from HOCl to H_2O to generate hydronium ions, H_3O^+ and hypochlorite ions, OCl^-.

$$\mathrm{HOCl}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{OCl^-}(aq) + \mathrm{H_3O^+}(aq)$$

A double-arrow is appropriate in this equation because it indicates the HOCl is a weak acid that has not reacted completely.

(b) The two reactants are provided, $Ba(OH)_2$ and HNO_3. Since this is a neutralization reaction, the two products will be water and a salt composed of the cation of the ionic hydroxide (Ba^{2+}) and the anion generated when the acid transfers its hydrogen ion (NO_3^-).

$$Ba(OH)_2(aq) + 2\,HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + 2\,H_2O(l)$$

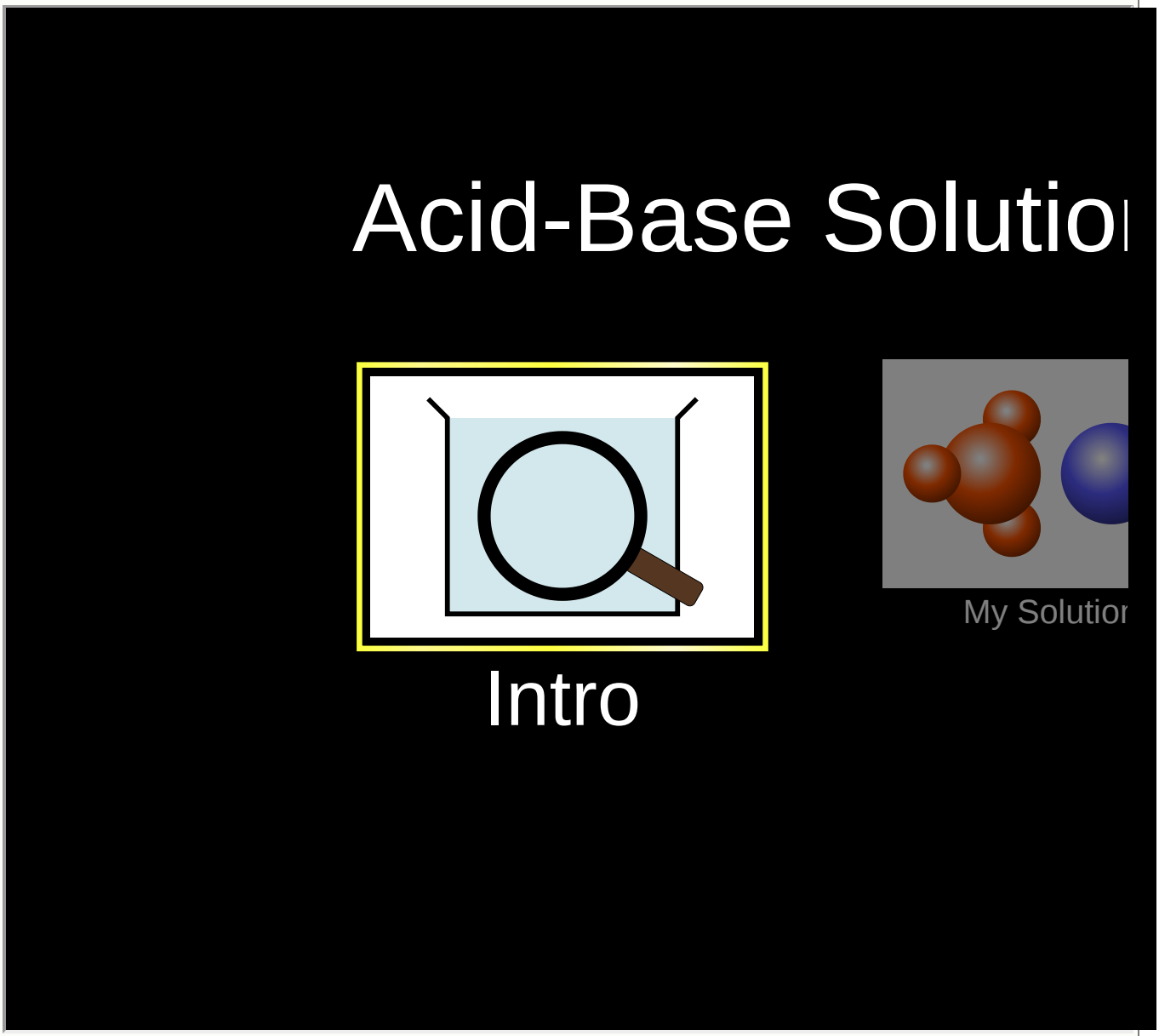
Exercise 6.4.1

Write the net ionic equation representing the neutralization of any strong acid with an ionic hydroxide. (Hint: Consider the ions produced when a strong acid is dissolved in water.)

Answer

$$H_3O^+(aq) + OH^-(aq) \rightarrow 2\,H_2O(l)$$

Phet Simulation



Chemistry in Everyday Life

Stomach Antacids

Our stomachs contain a solution of roughly 0.03 *M* HCl, which helps us digest the food we eat. The burning sensation associated with heartburn is a result of the acid of the stomach leaking through the muscular valve at the top of the stomach into the lower reaches of the esophagus. The lining of the esophagus is not protected from the corrosive effects of stomach acid the way the lining of the stomach is, and the results can be very painful. When we have heartburn, it feels better if we reduce the excess acid in the esophagus by taking an antacid. As you may have guessed, antacids are bases. One of the most common antacids is calcium carbonate, $CaCO_3$. The reaction,

$$\ce{CaCO3}(s)+\ce{2HCl}(aq)\rightleftharpoons\ce{CaCl2}(aq)+\ce{H2O}(l)+\ce{CO2}(g)$$

not only neutralizes stomach acid, it also produces $CO_2(g)$, which may result in a satisfying belch.

Milk of Magnesia is a suspension of the sparingly soluble base magnesium hydroxide, $Mg(OH)_2$. It works according to the reaction:

$$\ce{Mg(OH)2}(s)\rightleftharpoons Mg^{2+}(aq)+2OH^{-}(aq)$$

The hydroxide ions generated in this equilibrium then go on to react with the hydronium ions from the stomach acid, so that:

$$H_3O^{+}+OH^{-}\rightleftharpoons 2H_2O(l)$$

This reaction does not produce carbon dioxide, but magnesium-containing antacids can have a laxative effect. Several antacids have aluminum hydroxide, $Al(OH)_3$, as an active ingredient. The aluminum hydroxide tends to cause constipation, and some antacids use aluminum hydroxide in concert with magnesium hydroxide to balance the side effects of the two substances.

Culinary Aspects of Chemistry

Examples of acid-base chemistry are abundant in the culinary world. One example is the use of baking soda, or sodium bicarbonate in baking. $NaHCO_3$ is a base. When it reacts with an acid such as lemon juice, buttermilk, or sour cream in a batter, bubbles of carbon dioxide gas are formed from decomposition of the resulting carbonic acid, and the batter “rises.” Baking powder is a combination of sodium bicarbonate, and one or more acid salts that react when the two chemicals come in contact with water in the batter.

Many people like to put lemon juice or vinegar, both of which are acids, on cooked fish (Figure **6.4.4**). It turns out that fish have volatile amines (bases) in their systems, which are neutralized by the acids to yield involatile ammonium salts. This reduces the odor of the fish, and also adds a “sour” taste that we seem to enjoy.



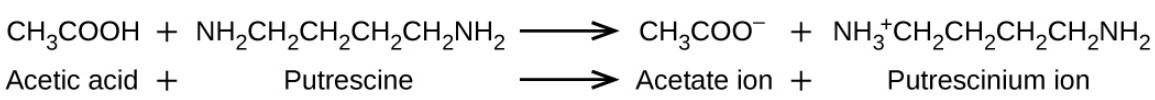


***Figure* 6.4.4:** *A neutralization reaction takes place between citric acid in lemons or acetic acid in vinegar, and the bases in the flesh of fish.*

Pickling is a method used to preserve vegetables using a naturally produced acidic environment. The vegetable, such as a cucumber, is placed in a sealed jar submerged in a brine solution. The brine solution favors the growth of beneficial bacteria and suppresses the growth of harmful bacteria. The beneficial bacteria feed on starches in the cucumber and produce lactic acid as a waste product in a process called fermentation. The lactic acid eventually increases the acidity of the brine to a level that kills any harmful bacteria, which require a basic environment. Without the harmful bacteria consuming the cucumbers they are able to last much longer than if they were unprotected. A byproduct of the pickling process changes the flavor of the vegetables with the acid making them taste sour.

Summary



***Video* 6.4.2:** *A summary of acid/base reactions.*

Chemical reactions are classified according to similar patterns of behavior. A large number of important reactions are included in three categories: precipitation, acid-base, and oxidation-reduction (redox). Precipitation reactions involve the formation of one or more insoluble products. Acid-base reactions involve the transfer of hydrogen ions between reactants. Redox reactions involve a change in oxidation number for one or more reactant elements. Writing balanced equations for some redox reactions that occur in aqueous solutions is simplified by using a systematic approach called the half-reaction method.

Glossary

acid
substance that produces H_3O^+ when dissolved in water

acid-base reaction
reaction involving the transfer of a hydrogen ion between reactant species

base
substance that produces OH^- when dissolved in water

salt
ionic compound that can be formed by the reaction of an acid with a base that contains a cation and an anion other than hydroxide or oxide

strong acid
acid that reacts completely when dissolved in water to yield hydronium ions

strong base
base that reacts completely when dissolved in water to yield hydroxide ions

weak acid
acid that reacts only to a slight extent when dissolved in water to yield hydronium ions

weak base
base that reacts only to a slight extent when dissolved in water to yield hydroxide ions

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6.4.1: Classifying Chemical Reactions (Acids and Bases) (Problems)

PROBLEM 6.4.1.1

Complete and balance the following acid-base equations:

a. HCl gas reacts with solid $Ca(OH)_2(s)$.
b. A solution of $Sr(OH)_2$ is added to a solution of HNO_3.

Answer a

$$2\,HCl(g) + Ca(OH)_2(s) \rightarrow CaCl_2(s) + 2\,H_2O(l)$$

Answer b

$$Sr(OH)_2(aq) + 2\,HNO_3(aq) \rightarrow Sr(NO_3)_2(aq) + 2\,H_2O(l)$$

PROBLEM 6.4.1.2

Complete and balance the following acid-base equations:

a. A solution of $HClO_4$ is added to a solution of LiOH.
b. Aqueous H_2SO_4 reacts with NaOH.
c. $Ba(OH)_2$ reacts with HF gas.

Answer a

$$HClO_4(aq) + LiOH(aq) \rightarrow LiClO_4(aq) + H_2O(l)$$

Answer b

$$H_2SO_4(aq) + 2\,NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\,H_2O(l)$$

Answer c

$$2\,HF(aq) + Ba(OH)_2(aq) \rightarrow BaF_2(aq) + 2\,H_2O(l)$$

Click here to see a video of the solution



(This video is improperly titled, but is the correct problem solution)

PROBLEM 6.4.1.3

Complete and balance the equations for the following acid-base neutralization reactions. If water is used as a solvent, write the reactants and products as aqueous ions. In some cases, there may be more than one correct answer, depending on the amounts of reactants used.

a. $Mg(OH)_2(s) + HClO_4(aq) \rightarrow$

b. $SO_3(g) + H_2O(l) \rightarrow$ (assume an excess of water and that the product dissolves)
c. $SrO(s) + H_2SO_4(l) \rightarrow$

Answer a

$$Mg(OH)_2(s) + 2\,HClO_4(aq) \rightarrow Mg^{2+}(aq) + 2\,ClO_4^-(aq) + 2\,H_2O(l)$$

Answer b

$SO_3(g) + 2\,H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$, (a solution of H_2SO_4)

Answer c

$$SrO(s) + H_2SO_4(l) \rightarrow SrSO_4(s) + H_2O$$

PROBLEM 6.4.1.4

Complete and balance the equations of the following reactions, each of which could be used to remove hydrogen sulfide from natural gas:

a. $Ca(OH)_2(s) + H_2S(g) \rightarrow$
b. $Na_2CO_3(aq) + H_2S(g) \rightarrow$

Answer

$$Ca(OH)_2(s) + H_2S(g) \rightarrow CaS(s) + 2\,H_2O(l)$$

Answer

$$Na_2CO_3(aq) + H_2S(g) \rightarrow Na_2S(aq) + CO_2(g) + H_2O(l)$$

PROBLEM 6.4.1.5

Calcium cyclamate $Ca(C_6H_{11}NHSO_3)_2$ is an artificial sweetener used in many countries around the world but is banned in the United States. It can be purified industrially by converting it to the barium salt through reaction of the acid $C_6H_{11}NHSO_3H$ with barium carbonate, treatment with sulfuric acid (barium sulfate is very insoluble), and then neutralization with calcium hydroxide. Write the balanced equations for these reactions.

Answer

$$2\,C_6H_{11}NHSO_3H + BaCO_3 \rightarrow Ba(C_6H_{11}NHSO_3)_2 + H_2CO_3$$

$$Ba(C_6H_{11}NHSO_3)_2 + H_2SO_4 \rightarrow BaSO_4 + 2\,C_6H_{11}NHSO_3H$$

$$2\,C_6H_{11}NHSO_3H + Ca(OH)_2 \rightarrow Ca(C_6H_{11}NHSO_3)_2 + 2\,H_2O$$

Contributors

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- Adelaide Clark, Oregon Institute of Technology

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6.5: Classifying Chemical Reactions (Redox)

Skills to Develop

- Define three common types of chemical reactions (precipitation, acid-base, and oxidation-reduction)
- Classify chemical reactions as one of these three types given appropriate descriptions or chemical equations
- Compute the oxidation states for elements in compounds

Oxidation-Reduction Reactions

Earth's atmosphere contains about 20% molecular oxygen, O_2, a chemically reactive gas that plays an essential role in the metabolism of aerobic organisms and in many environmental processes that shape the world. The term oxidation was originally used to describe chemical reactions involving O_2, but its meaning has evolved to refer to a broad and important reaction class known as *oxidation-reduction (redox) reactions*. A few examples of such reactions will be used to develop a clear picture of this classification.

Some redox reactions involve the transfer of electrons between reactant species to yield ionic products, such as the reaction between sodium and chlorine to yield sodium chloride:

$$2\,\mathrm{Na}(s) + \mathrm{Cl_2}(g) \rightarrow 2\,\mathrm{NaCl}(s) \tag{6.5.1}$$

It is helpful to view the process with regard to each individual reactant, that is, to represent the fate of each reactant in the form of an equation called a half-reaction:

$$2\,\mathrm{Na}(s) \rightarrow 2\,\mathrm{Na^+}(s) + 2\,\mathrm{e^-} \tag{6.5.2}$$

$$\mathrm{Cl_2}(g) + 2\,\mathrm{e^-} \rightarrow 2\,\mathrm{Cl^-}(s) \tag{6.5.3}$$

These equations show that Na atoms *lose electrons* while Cl atoms (in the Cl_2 molecule) *gain electrons*, the "*s*" subscripts for the resulting ions signifying they are present in the form of a solid ionic compound. For redox reactions of this sort, the loss and gain of electrons define the complementary processes that occur:

oxidation = loss of electrons (6.5.4)
reduction = gain of electrons (6.5.5)

Remembering Oxidation and Reduction

It is common to remember the difference between oxidation and reduction using one of two pneumonic devices:

1. "LeO goes GeR"
   In Greek Mythology, Leo was a lion (like the zodiac sign). And what do lion's do? Well, they roar but that doesn't really fit, so we go with "Ger".
   Loses **e**lectrons = **O**xidation
   Gains **e**lectrons = **R**eduction



Figure 6.5.1: *Lion image by Tambako, reused under the Creative Commons Attribution 2.0 Generic license.*

2. "Oil Rig"
   Oxidation **i**s **l**oss
   Reduction **i**s **g**ain



***Figure* 6.5.2:** *Image by Peter Grima, reused under the Creative Commons Attribution 2.0 Generic license.*

In this reaction, then, sodium is *oxidized* and chlorine undergoes reduction. Viewed from a more active perspective, sodium functions as a reducing agent (reductant), since it provides electrons to (or reduces) chlorine. Likewise, chlorine functions as an oxidizing agent (oxidant), as it effectively removes electrons from (oxidizes) sodium.

$$\textbf{reducing agent} = \textbf{species that is oxidized} \tag{6.5.6}$$
$$\textbf{oxidizing agent} = \textbf{species that is reduced} \tag{6.5.7}$$

Some redox processes, however, do not involve the transfer of electrons. Consider, for example, a reaction similar to the one yielding NaCl:

$$\mathrm{H_2}(g) + \mathrm{Cl_2}(g) \rightarrow 2\,\mathrm{HCl}(g) \tag{6.5.8}$$

The product of this reaction is a covalent compound, so transfer of electrons in the explicit sense is not involved. To clarify the similarity of this reaction to the previous one and permit an unambiguous definition of redox reactions, a property called *oxidation number* has been defined. The oxidation number (or oxidation state) of an element in a compound is the charge its atoms would possess *if the compound was ionic*. The following guidelines are used to assign oxidation numbers to each element in a molecule or ion. These rules are *hierarchical;* if two rules conflict, the rule that is higher up on the list takes precedence:

1. The oxidation number of an atom in an elemental substance (a free element) is zero.
2. The oxidation number of a monatomic ion is equal to the ion's charge.
3. The sum of oxidation numbers for all atoms in a molecule or polyatomic ion equals the charge on the molecule or ion.
4. Metals always have positive oxidation states when in a compound.
    - Group 1 (hydrogen) is ALWAYS +1
    - Group 2 (beryllium) is ALWAYS +2
5. Oxidation numbers for common nonmetals are usually assigned as follows:
    - Fluorine: always -1
    - Hydrogen: +1 when combined with nonmetals, −1 when combined with metals
    - Oxygen: −2 in most compounds, sometimes −1 (so-called peroxides, $\mathrm{O_2}^{2-}$), very rarely $-\frac{1}{2}$ (so-called superoxides, $\mathrm{O_2}^{-}$), positive values when combined with F (values vary)

- Halogens: −1 for other halogens except when combined with oxygen or other halogens (positive oxidation numbers in these cases, varying values)
- Group 6 (oxygen) is -2
- Group 5 (nitrogen) is -3

Note: The proper convention for reporting charge is to write the number first, followed by the sign (e.g., 2+), while oxidation number is written with the reversed sequence, sign followed by number (e.g., +2). This convention aims to emphasize the distinction between these two related properties.

*It is also important to note that it is possible to have **fractional** oxidation states.*

Example 6.5.2: Assigning Oxidation Numbers

Follow the guidelines in this section of the text to assign oxidation numbers to all the elements in the following species:

a. H_2S
b. SO_3^{2-}
c. Na_2SO_4

Solution

(a) According to guideline 1, the oxidation number for H is +1.

Using this oxidation number and the compound's formula, guideline 4 may then be used to calculate the oxidation number for sulfur:

$$\text{charge on } H_2S = 0 = (2 \times +1) + (1 \times x)$$

$$x = 0 - (2 \times +1) = -2$$

(b) Guideline 3 suggests the oxidation number for oxygen is −2.

Using this oxidation number and the ion's formula, guideline 4 may then be used to calculate the oxidation number for sulfur:

$$\text{charge on } SO_3^{2-} = -2 = (3 \times -2) + (1 \times x)$$
$$x = -2 - (3 \times -2) = +4$$

(c) For ionic compounds, it's convenient to assign oxidation numbers for the cation and anion separately.

According to guideline 2, the oxidation number for sodium is +1.

Assuming the usual oxidation number for oxygen (−2 per guideline 3), the oxidation number for sulfur is calculated as directed by guideline 4:

$$\text{charge on } SO_4^{2-} = -2 = (4 \times -2) + (1 \times x)$$

$$x = -2 - (4 \times -2) = +6$$

Exercise 6.5.2

Assign oxidation states to the elements whose atoms are underlined in each of the following compounds or ions:

a. $K\underline{N}O_3$
b. $\underline{Al}H_3$
c. $\underline{N}H_4^+$
d. msubsup

Answer a

N, +5

Answer b

Al, +3

Answer c

N, −3

Answer d

P, +5

Using the oxidation number concept, an all-inclusive definition of redox reaction has been established. Oxidation-reduction (redox) reactions are those in which one or more elements involved undergo a change in oxidation number. While the vast majority of redox reactions involve changes in oxidation number for two or more elements, a few interesting exceptions to this rule do exist as shown below\). Definitions for the complementary processes of this reaction class are correspondingly revised as shown here:

$$\textbf{oxidation} = \textbf{increase in oxidation number} \tag{6.5.9}$$

$$\textbf{reduction} = \textbf{decrease in oxidation number} \tag{6.5.10}$$

Returning to the reactions used to introduce this topic, they may now both be identified as redox processes. In the reaction between sodium and chlorine to yield sodium chloride, sodium is oxidized (its oxidation number increases from 0 in Na to +1 in NaCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl_2 to −1 in NaCl). In the reaction between molecular hydrogen and chlorine, hydrogen is oxidized (its oxidation number increases from 0 in H_2 to +1 in HCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl_2 to −1 in HCl).

Several subclasses of redox reactions are recognized, including combustion reactions in which the reductant (also called a *fuel*) and oxidant (often, but not necessarily, molecular oxygen) react vigorously and produce significant amounts of heat, and often light, in the form of a flame. Solid rocket-fuel reactions such as the one depicted below are combustion processes. A typical propellant reaction in which solid aluminum is oxidized by ammonium perchlorate is represented by this equation:

$$10\,Al(s) + 6\,NH_4ClO_4(s) \rightarrow 4\,Al_2O_3(s) + 2\,AlCl_3(s) + 12\,H_2O(g) + 3\,N_2(g) \tag{6.5.11}$$

Watch a brief video showing the test firing of a small-scale, prototype, hybrid rocket engine planned for use in the new Space Launch System being developed by NASA. The first engines firing at 3 s (green flame) use a liquid fuel/oxidant mixture, and the second, more powerful engines firing at 4 s (yellow flame) use a solid mixture.



Single-displacement (replacement) reactions are redox reactions in which an ion in solution is displaced (or replaced) via the oxidation of a metallic element. One common example of this type of reaction is the acid oxidation of certain metals:

$$Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g) \tag{6.5.12}$$

Metallic elements may also be oxidized by solutions of other metal salts; for example:

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\,Ag(s) \tag{6.5.13}$$

This reaction may be observed by placing copper wire in a solution containing a dissolved silver salt. Silver ions in solution are reduced to elemental silver at the surface of the copper wire, and the resulting Cu^{2+} ions dissolve in the solution to yield a characteristic blue color (Figure **6.5.6**).

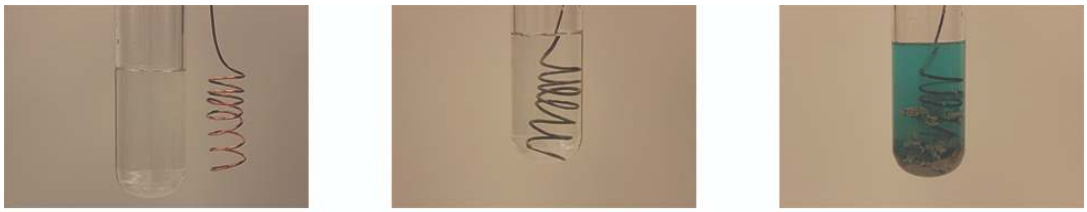


***Figure* 6.5.6:** *(a) A copper wire is shown next to a solution containing silver(I) ions. (b) Displacement of dissolved silver ions by copper ions results in (c) accumulation of gray-colored silver metal on the wire and development of a blue color in the solution, due to dissolved copper ions. (credit: modification of work by Mark Ott)*

Example 6.5.3: Describing Redox Reactions

Identify which equations represent redox reactions, providing a name for the reaction if appropriate. For those reactions identified as redox, name the oxidant and reductant.

a. $\mathrm{ZnCO_3}(s) \rightarrow \mathrm{ZnO}(s) + \mathrm{CO_2}(g)$
b. $2\,\mathrm{Ga}(l) + 3\,\mathrm{Br_2}(l) \rightarrow 2\,\mathrm{GaBr_3}(s)$
c. $2\,\mathrm{H_2O_2}(aq) \rightarrow 2\,\mathrm{H_2O}(l) + \mathrm{O_2}(g)$
d. $\mathrm{BaCl_2}(aq) + \mathrm{K_2SO_4}(aq) \rightarrow \mathrm{BaSO_4}(s) + 2\,\mathrm{KCl}(aq)$
e. $\mathrm{C_2H_4}(g) + 3\,\mathrm{O_2}(g) \rightarrow 2\,\mathrm{CO_2}(g) + 2\,\mathrm{H_2O}(l)$

Solution

Redox reactions are identified per definition if one or more elements undergo a change in oxidation number.

a. This is not a redox reaction, since oxidation numbers remain unchanged for all elements.
b. This is a redox reaction. Gallium is oxidized, its oxidation number increasing from 0 in Ga(l) to +3 in $GaBr_3$(s). The reducing agent is Ga(l). Bromine is reduced, its oxidation number decreasing from 0 in Br_2(l) to −1 in $GaBr_3$(s). The oxidizing agent is Br_2(l).
c. This is a redox reaction. It is a particularly interesting process, as it involves the same element, oxygen, undergoing both oxidation and reduction (a so-called *disproportionation reaction)*. Oxygen is oxidized, its oxidation number increasing from −1 in H_2O_2(aq) to 0 in O_2(g). Oxygen is also reduced, its oxidation number decreasing from −1 in H_2O_2(aq) to −2 in H_2O(l). For disproportionation reactions, the same substance functions as an oxidant and a reductant.
d. This is not a redox reaction, since oxidation numbers remain unchanged for all elements.
e. This is a redox reaction (combustion). Carbon is oxidized, its oxidation number increasing from −2 in C_2H_4(g) to +4 in CO_2(g). The reducing agent (fuel) is C_2H_4(g). Oxygen is reduced, its oxidation number decreasing from 0 in O_2(g) to −2 in H_2O(l). The oxidizing agent is O_2(g).

Exercise 6.5.3

This equation describes the production of tin(II) chloride:

$$\mathrm{Sn}(s) + 2\,\mathrm{HCl}(g) \rightarrow \mathrm{SnCl_2}(s) + \mathrm{H_2}(g) \tag{6.5.14}$$

Is this a redox reaction? If so, provide a more specific name for the reaction if appropriate, and identify the oxidant and reductant.

Answer

Yes, a single-replacement reaction. Sn(s)is the reductant, HCl(g) is the oxidant.

Summary



Video* 6.5.1*: *A summary of redox reactions.*

One of your professors was a script editor for this video. Hint: It was Dr. Clark

Chemical reactions are classified according to similar patterns of behavior. A large number of important reactions are included in three categories: precipitation, acid-base, and oxidation-reduction (redox). Precipitation reactions involve the formation of one or more insoluble products. Acid-base reactions involve the transfer of hydrogen ions between reactants. Redox reactions involve a change in oxidation number for one or more reactant elements. Writing balanced equations for some redox reactions that occur in aqueous solutions is simplified by using a systematic approach called the half-reaction method.

Glossary

combustion reaction

vigorous redox reaction producing significant amounts of energy in the form of heat and, sometimes, light

oxidation

process in which an element’s oxidation number is increased by loss of electrons

oxidation-reduction reaction

(also, redox reaction) reaction involving a change in oxidation number for one or more reactant elements

oxidation number

(also, oxidation state) the charge each atom of an element would have in a compound if the compound were ionic

oxidizing agent

(also, oxidant) substance that brings about the oxidation of another substance, and in the process becomes reduced

reduction

process in which an element’s oxidation number is decreased by gain of electrons

reducing agent

(also, reductant) substance that brings about the reduction of another substance, and in the process becomes oxidized

single-displacement reaction

(also, replacement) redox reaction involving the oxidation of an elemental substance by an ionic species

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- Adelaide Clark, Oregon Institute of Technology
- Crash Course Chemistry: Crash Course is a division of Complexly and videos are free to stream for educational purposes.

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6.5.1: Classifying Chemical Reactions (Redox) (Problems)

Some of these problems represent a culmination of all the reaction types we've covered in Unit 6!

PROBLEM 6.5.1.1

Indicate what type, or types, of reaction each of the following represents:

a. $\mathrm{Ca}(s) + \mathrm{Br_2}(l) \rightarrow \mathrm{CaBr_2}(s)$
b. $\mathrm{Ca(OH)_2}(aq) + 2\,\mathrm{HBr}(aq) \rightarrow \mathrm{CaBr_2}(aq) + 2\,\mathrm{H_2O}(l)$
c. $\mathrm{C_6H_{12}}(l) + 9\,\mathrm{O_2}(g) \rightarrow 6\,\mathrm{CO_2}(g) + 6\,\mathrm{H_2O}(g)$

Answer a

oxidation-reduction (addition)

Answer b

acid-base (neutralization)

Answer c

oxidation-reduction (combustion)

PROBLEM 6.5.1.2

Indicate what type, or types, of reaction each of the following represents:

a. $\mathrm{H_2O}(g) + \mathrm{C}(s) \rightarrow \mathrm{CO}(g) + \mathrm{H_2}(g)$
b. $2\,\mathrm{KClO_3}(s) \rightarrow 2\,\mathrm{KCl}(s) + 3\,\mathrm{O_2}(g)$
c. $\mathrm{Al(OH)_3}(aq) + 3\,\mathrm{HCl}(aq) \rightarrow \mathrm{AlBr_3}(aq) + 3\,\mathrm{H_2O}(l)$
d. $\mathrm{Pb(NO_3)_2}(aq) + \mathrm{H_2SO_4}(aq) \rightarrow \mathrm{PbSO_4}(s) + 2\,\mathrm{HNO_3}(aq)$

Answer a

oxidation-reduction (single displacement)

Answer b

oxidation-reduction (dissociation)

Answer c

acid-base (neutralization)

Answer d

precipitation (double replacement)

PROBLEM 6.5.1.3

Silver can be separated from gold because silver dissolves in nitric acid while gold does not. Is the dissolution of silver in nitric acid an acid-base reaction or an oxidation-reduction reaction? Explain your answer.

Answer

It is an oxidation-reduction reaction because the oxidation state of the silver changes during the reaction.

PROBLEM 6.5.1.4

Determine the oxidation states of the elements in the following compounds:

a. NaI
b. $\mathrm{GdCl_3}$

c. $LiNO_3$
d. H_2Se
e. Mg_2Si
f. RbO_2 (rubidium superoxide)
g. HF

Answer a

Na +1, I -1

Answer b

Gd +3, Cl -1

Answer c

Li +1, N +5, O -2

Answer d

H +1, Se -2

Answer e

Mg +2, Si -4

Answer f

Rb +1, O -1/2

Answer g

H +1, F -1

Click here to see a video of the solution



PROBLEM 6.5.1.5

Determine the oxidation states of the elements in the compounds listed. None of the oxygen-containing compounds are peroxides or superoxides.

a. H_3PO_4
b. $Al(OH)_3$
c. SeO_2
d. KNO_2
e. In_2S_3
f. P_4O_6

Answer a

H +1, P +5, O −2

Answer b

Al +3, H +1, O −2

Answer c

Se +4, O −2

Answer d

K +1, N +3, O −2

Answer e

In +3, S −2

Answer f

P +3, O −2

PROBLEM 6.5.1.6

Determine the oxidation states of the elements in the compounds listed. None of the oxygen-containing compounds are peroxides or superoxides.

a. H_2SO_4
b. $Ca(OH)_2$
c. BrOH
d. $ClNO_2$
e. $TiCl_4$
f. NaH

Answer a

H +1, O -2, S +6

Answer b

H +1, O -2, Ca +2

Answer c

H +1, O -2, Br +1,

Answer d

O -2, Cl -1, N +5

Answer e

Cl -1, Ti +4

Answer f

H +1, Na -1

PROBLEM 6.5.1.7

Classify the following as acid-base reactions or oxidation-reduction reactions:

a. $\mathbf{Na_2S}(aq) + \mathbf{2\,HCl}(aq) \rightarrow \mathbf{2\,NaCl}(aq) + \mathbf{H_2S}(g)$

b. $2\,Na(s) + 2\,HCl(aq) \rightarrow 2\,NaCl(aq) + H_2(g)$
c. $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$
d. $MgO(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$
e. $K_3P(s) + 2\,O_2(g) \rightarrow K_3PO_4(s)$
f. $3\,KOH(aq) + H_3PO_4(aq) \rightarrow K_3PO_4(aq) + 3\,H_2O(l)$

Answer a

acid-base

Answer b

oxidation-reduction: Na is oxidized, H^+ is reduced

Answer c

oxidation-reduction: Mg is oxidized, Cl_2 is reduced

Answer d

acid-base

Answer e

oxidation-reduction: P^{3-} is oxidized, O_2 is reduced

Answer f

acid-base

PROBLEM 6.5.1.8

Identify the atoms that are oxidized and reduced, the change in oxidation state for each, and the oxidizing and reducing agents in each of the following equations:

a. $Mg(s) + NiCl_2(aq) \rightarrow MgCl_2(aq) + Ni(s)$
b. $PCl_3(l) + Cl_2(g) \rightarrow PCl_5(s)$
c. $C_2H_4(g) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 2\,H_2O(g)$
d. $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
e. $2\,K_2S_2O_3(s) + I_2(s) \rightarrow K_2S_4O_6(s) + 2\,KI(s)$
f. $3\,Cu(s) + 8\,HNO_3(aq) \rightarrow 3\,Cu(NO_3)_2(aq) + 2\,NO(g) + 4\,H_2O(l)$

Answer a

Mg: 0 → +2; loses electrons; oxidized; reducing agent

Ni: +2 → 0; gains electrons; reduced; oxidizing agent

Answer b

P: +3 → +5; loses electrons; oxidized; reducing agent

Cl: 0 → -1; gains electrons; reduced; oxidizing agent

Answer c

C: -2 → +4; loses electrons; oxidized; reducing agent

O: 0 → -2; gains electrons; reduced; oxidizing agent

Answer d

Zn: 0 → +2; loses electrons; oxidized; reducing agent

H: +1 → 0; gains electrons; reduced; oxidizing agent

Answer e

S: +2 → +5/2; loses electrons; oxidized; reducing agent

I: 0 → -1; gains electrons; reduced; oxidizing agent

Answer f

Cu: 0 → +2; loses electrons; oxidized; reducing agent

N: +5 → +2; gains electrons; reduced; oxidizing agent

Click here to see a video of the solution



PROBLEM 6.5.1.10

When heated to 700–800 °C, diamonds, which are pure carbon, are oxidized by atmospheric oxygen. (They burn!) Write the balanced equation for this reaction.

Answer

$$C_{diamond}(s) + O_2(g) \rightarrow CO_2(g)$$

PROBLEM 6.5.1.11

The military has experimented with lasers that produce very intense light when fluorine combines explosively with hydrogen. What is the balanced equation for this reaction?

Answer

$$H_2(g) + F_2(g) \rightarrow 2\,HF(g)$$

PROBLEM 6.5.1.12

In a common experiment in the general chemistry laboratory (that you will do if you haven't already), magnesium metal is heated in air to produce MgO. MgO is a white solid, but in these experiments it often looks gray, due to small amounts of Mg_3N_2, a compound formed as some of the magnesium reacts with nitrogen. Write a balanced equation for each reaction.

Answer

$$2\,Mg(s) + O_2(g) \rightarrow 2\,MgO(s)$$

$$3\,Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$$

PROBLEM 6.5.1.13

Copper(II) sulfide is oxidized by molecular oxygen to produce gaseous sulfur trioxide and solid copper(II) oxide. The gaseous product then reacts with liquid water to produce liquid hydrogen sulfate as the only product. Write the two balanced equations which represent these reactions.

Answer

$$\mathrm{CuS}(s) + 2\,\mathrm{O_2}(g) \rightarrow \mathrm{SO_3}(g) + \mathrm{CuO}(s)$$

$$\mathrm{SO_3}(g) + \mathrm{H_2O}(l) \rightarrow \mathrm{H_2SO_4}(l)$$

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CHAPTER OVERVIEW

7: Ideal Gas Behavior

Unit 7 Objectives

By the end of this unit, you will be able to:

- Define the property of temperature and pressure
- Define and convert among the units of temperature and pressure
- Describe the operation of common tools for measuring gas pressure
- Identify the mathematical relationships between the various properties of gases
- Use the ideal gas law, and related gas laws, to compute the values of various gas properties under specified conditions
- Use the ideal gas law to compute gas densities and molar masses
- Perform stoichiometric calculations involving gaseous substances
- State Dalton’s law of partial pressures and use it in calculations involving gaseous mixtures
- State the postulates of the kinetic-molecular theory
- Use this theory’s postulates to explain the gas laws

7.1: Temperature and Pressure

7.1.1: Practice Problems- Temperature and Pressure

7.2: The Gas Laws

7.2.1: Practice Problems- The Gas Laws

7.3: Applications of the Ideal Gas Law and Partial Pressures

7.3.1: Practice Problems- Applications of the Ideal Gas Law

7.4: Kinetic Molecular Theory

7.4.1: Practice Problems- Kinetic Molecular Theory

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7.1: Temperature and Pressure

Skills to Develop

- Define the property of temperature and pressure
- Define and convert among the units of temperature and pressure
- Describe the operation of common tools for measuring gas pressure

Temperature



Video 7.1.1: *An introduction to temperature in terms of the Ideal Gas Law, which we will cover later in this Unit.*

Temperature is an intensive property. The SI unit of temperature is the kelvin (K). The IUPAC convention is to use kelvin (all lowercase) for the word, K (uppercase) for the unit symbol, and neither the word “degree” nor the degree symbol (°). The degree Celsius (°C) is also allowed in the SI system, with both the word “degree” and the degree symbol used for Celsius measurements. Celsius degrees are the same magnitude as those of kelvin, but the two scales place their zeros in different places. Water freezes at 273.15 K (0 °C) and boils at 373.15 K (100 °C) by definition, and normal human body temperature is approximately 310 K (37 °C).

Conversion of Temperature Units

We use the word temperature to refer to the hotness or coldness of a substance. One way we measure a change in temperature is to use the fact that most substances expand when their temperature increases and contract when their temperature decreases. The mercury or alcohol in a common glass thermometer changes its volume as the temperature changes. Because the volume of the liquid changes more than the volume of the glass, we can see the liquid expand when it gets warmer and contract when it gets cooler.

To mark a scale on a thermometer, we need a set of reference values: Two of the most commonly used are the freezing and boiling temperatures of water at a specified atmospheric pressure. On the Celsius scale, 0 °C is defined as the freezing temperature of water and 100 °C as the boiling temperature of water. The space between the two temperatures is divided into 100 equal intervals, which we call degrees. On the Fahrenheit scale, the freezing point of water is defined as 32 °F and the boiling temperature as 212 °F. The space between these two points on a Fahrenheit thermometer is divided into 180 equal parts (degrees).

Defining the Celsius and Fahrenheit temperature scales as described in the previous paragraph results in a slightly more complex relationship between temperature values on these two scales than for different units of measure for other properties. Most measurement units for a given property are directly proportional to one another (y = mx). Using familiar length units as one example:

$$\text{length in feet} = \left(\frac{1\ \text{ft}}{12\ \text{in.}}\right) \times \text{length in inches} \tag{7.1.1}$$

where

- y = length in feet,

- x = length in inches, and
- the proportionality constant, m, is the conversion factor.

The Celsius and Fahrenheit temperature scales, however, do not share a common zero point, and so the relationship between these two scales is a linear one rather than a proportional one ($y = mx + b$). Consequently, converting a temperature from one of these scales into the other requires more than simple multiplication by a conversion factor, m, it also must take into account differences in the scales' zero points (b).

The linear equation relating Celsius and Fahrenheit temperatures is easily derived from the two temperatures used to define each scale. Representing the Celsius temperature as x and the Fahrenheit temperature as y, the slope, m, is computed to be:

$$m = \frac{\Delta y}{\Delta x} \tag{7.1.2}$$

$$= \frac{212\ ^\circ\text{F} - 32\ ^\circ\text{F}}{100\ ^\circ\text{C} - 0\ ^\circ\text{C}} \tag{7.1.3}$$

$$= \frac{180\ ^\circ\text{F}}{100\ ^\circ\text{C}} \tag{7.1.4}$$

$$= \frac{9\ ^\circ\text{F}}{5\ ^\circ\text{C}} \tag{7.1.5}$$

The y-intercept of the equation, b, is then calculated using either of the equivalent temperature pairs, (100 °C, 212 °F) or (0 °C, 32 °F), as:

$$b = y - mx \tag{7.1.6}$$

$$= 32\ ^\circ\text{F} - \frac{9\ ^\circ\text{F}}{5\ ^\circ\text{C}} \times 0\ ^\circ\text{C} \tag{7.1.7}$$

$$= 32\ ^\circ\text{F} \tag{7.1.8}$$

The equation relating the temperature scales is then:

$$T_{^\circ\text{F}} = \left(\frac{9\ ^\circ\text{F}}{5\ ^\circ\text{C}} \times T_{^\circ\text{C}}\right) + 32\ ^\circ\text{C} \tag{7.1.9}$$

An abbreviated form of this equation that omits the measurement units is:

$$T_{^\circ\text{F}} = \frac{9}{5} \times T_{^\circ\text{C}} + 32 \tag{7.1.10}$$

Rearrangement of this equation yields the form useful for converting from Fahrenheit to Celsius:

$$T_{^\circ\text{C}} = \frac{5}{9}(T_{^\circ\text{F}} + 32) \tag{7.1.11}$$

As mentioned earlier in this chapter, the SI unit of temperature is the kelvin (K). Unlike the Celsius and Fahrenheit scales, the kelvin scale is an absolute temperature scale in which 0 (zero) K corresponds to the lowest temperature that can theoretically be achieved. The early 19th-century discovery of the relationship between a gas's volume and temperature suggested that the volume of a gas would be zero at −273.15 °C. In 1848, British physicist William Thompson, who later adopted the title of Lord Kelvin, proposed an absolute temperature scale based on this concept (further treatment of this topic is provided in this text's chapter on gases).

The freezing temperature of water on this scale is 273.15 K and its boiling temperature 373.15 K. Notice the numerical difference in these two reference temperatures is 100, the same as for the Celsius scale, and so the linear relation between these two temperature scales will exhibit a slope of $1\ \frac{\text{K}}{^\circ\text{C}}$. Following the same approach, the equations for converting between the kelvin and Celsius temperature scales are derived to be:

$$T_\text{K} = T_{^\circ\text{C}} + 273.15 \tag{7.1.12}$$

$$T_{^\circ\text{C}} = T_\text{K} - 273.15 \tag{7.1.13}$$

The 273.15 in these equations has been determined experimentally, so it is not exact. Figure 7.1.1 shows the relationship among the three temperature scales. Recall that we do not use the degree sign with temperatures on the kelvin scale.

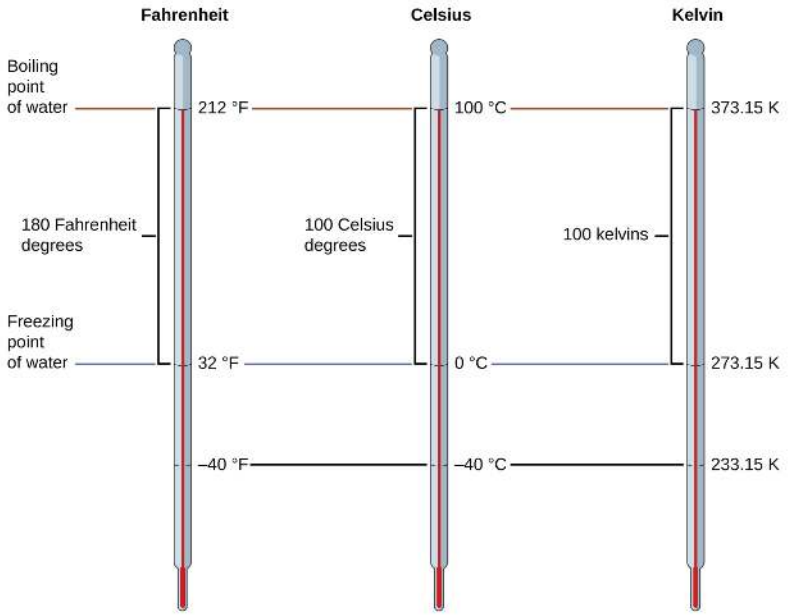


Figure* 7.1.1*: *The Fahrenheit, Celsius, and kelvin temperature scales are compared.*

Although the kelvin (absolute) temperature scale is the official SI temperature scale, Celsius is commonly used in many scientific contexts and is the scale of choice for nonscience contexts in almost all areas of the world. Very few countries (the U.S. and its territories, the Bahamas, Belize, Cayman Islands, and Palau) still use Fahrenheit for weather, medicine, and cooking.

Example 7.1.1: Conversion from Celsius

Normal body temperature has been commonly accepted as 37.0 °C (although it varies depending on time of day and method of measurement, as well as among individuals). What is this temperature on the kelvin scale and on the Fahrenheit scale?

Solution

$$\mathrm{K} = {}^\circ\mathrm{C} + 273.15 = 37.0 + 273.2 = 310.2\ \mathrm{K}$$

$${}^\circ\mathrm{F} = \frac{9}{5}\,{}^\circ\mathrm{C} + 32.0 = \left(\frac{9}{5} \times 37.0\right) + 32.0 = 66.6 + 32.0 = 98.6\ {}^\circ\mathrm{F}$$

Exercise 7.1.1

Convert 80.92 °C to K and °F.

Answer

354.07 K, 177.7 °F

Example 7.1.2: Conversion from Fahrenheit

Baking a ready-made pizza calls for an oven temperature of 450 °F. If you are in Europe, and your oven thermometer uses the Celsius scale, what is the setting? What is the kelvin temperature?

Solution

$${}^\circ\mathrm{C} = \frac{5}{9}({}^\circ\mathrm{F} - 32) = \frac{5}{9}(450 - 32) = \frac{5}{9} \times 418 = 232{}^\circ\mathrm{C} \rightarrow \text{set oven to } 230{}^\circ\mathrm{C}$$
$$\text{(two significant figures)}$$

$$\mathrm{K} = {}^\circ\mathrm{C} + 273.15 = 230 + 273 = 503\ \mathrm{K} \rightarrow 5.0 \times 10^2\ \mathrm{K} \quad \text{(two significant figures)}$$

Exercise 7.1.2

Convert 50 °F to °C and K.

Answer

10 °C, 280 K

Pressure

The earth's atmosphere exerts a pressure, as does any other gas. Although we do not normally notice atmospheric pressure, we are sensitive to pressure changes—for example, when your ears "pop" during take-off and landing while flying, or when you dive underwater. Gas pressure is caused by the force exerted by gas molecules colliding with the surfaces of objects (Figure **7.1.2**). Although the force of each collision is very small, any surface of appreciable area experiences a large number of collisions in a short time, which can result in a high pressure. In fact, normal air pressure is strong enough to crush a metal container when not balanced by equal pressure from inside the container.

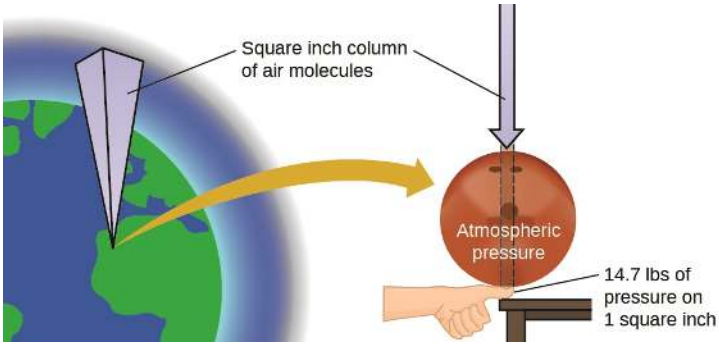
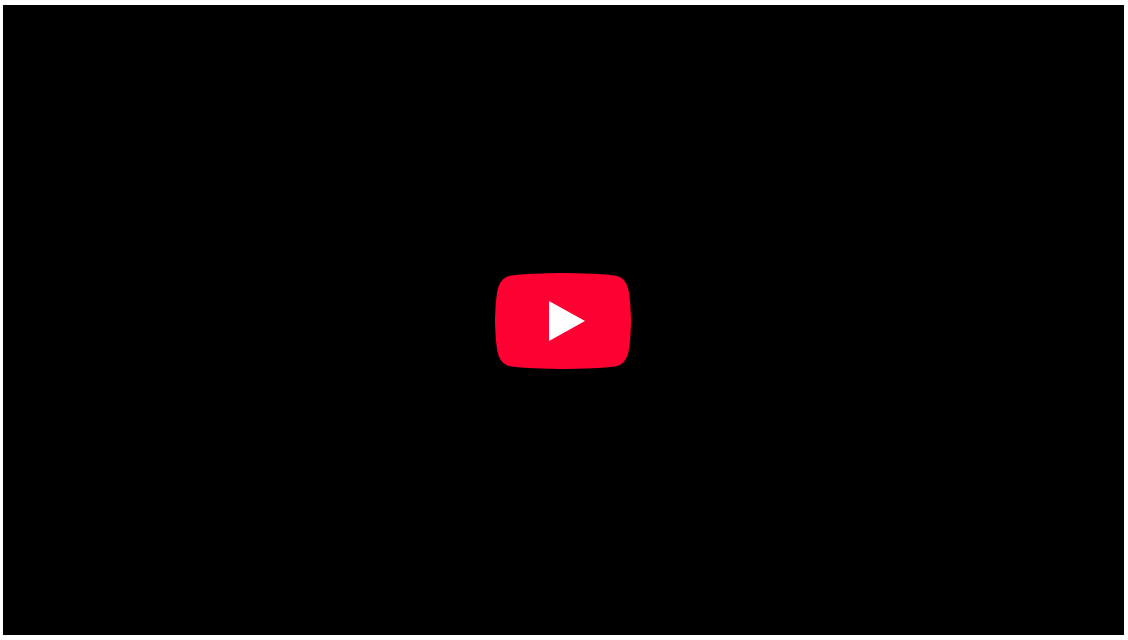


***Figure* 7.1.2**: *The atmosphere above us exerts a large pressure on objects at the surface of the earth, roughly equal to the weight of a bowling ball pressing on an area the size of a human thumbnail.*

Atmospheric pressure is caused by the weight of the column of air molecules in the atmosphere above an object, such as the tanker car. At sea level, this pressure is roughly the same as that exerted by a full-grown African elephant standing on a doormat, or a typical bowling ball resting on your thumbnail. These may seem like huge amounts, and they are, but life on earth has evolved under such atmospheric pressure. If you actually perch a bowling ball on your thumbnail, the pressure experienced is twice the usual pressure, and the sensation is unpleasant.





Video* 7.1.2*: *A dramatic illustration of atmospheric pressure is provided in this brief video, which shows a railway tanker car imploding when its internal pressure is decreased.*





Video* 7.1.3*: *A smaller scale demonstration of this phenomenon is briefly explained.*

Pressure is defined as the force exerted on a given area:

$$P = \frac{F}{A} \tag{7.1.14}$$

Since pressure is directly proportional to force and inversely proportional to area (Equation 7.1.14), pressure can be increased either by either **increasing** the amount of force or by **decreasing** the area over which it is applied. Correspondingly, pressure can be decreased by either **decreasing** the force or **increasing** the area.

Let's apply the definition of pressure (Equation 7.1.14) to determine which would be more likely to fall through thin ice in Figure 7.1.3.—the elephant or the figure skater?



***Figure 7.1.3**: Although (a) an elephant's weight is large, creating a very large force on the ground, (b) the figure skater exerts a much higher pressure on the ice due to the small surface area of her skates. (credit a: modification of work by Guido da Rozze; credit b: modification of work by Ryosuke Yagi).*

A large African elephant can weigh 7 tons, supported on four feet, each with a diameter of about 1.5 ft (footprint area of 250 in^2), so the pressure exerted by each foot is about 14 lb/in^2:

$$\text{pressure per elephant foot} = 14,000\frac{\text{lb}}{\text{elephant}} \times \frac{1\text{ elephant}}{4\text{ feet}} \times \frac{1\text{ foot}}{250\text{ in}^2} = 14\text{ lb/in}^2 \tag{7.1.15}$$

The figure skater weighs about 120 lbs, supported on two skate blades, each with an area of about 2 in^2, so the pressure exerted by each blade is about 30 lb/in^2:

$$\text{pressure per skate blade} = 120\frac{\text{lb}}{\text{skater}} \times \frac{1\text{ skater}}{2\text{ blades}} \times \frac{1\text{ blade}}{2\text{ in}^2} = 30\text{ lb/in}^2 \tag{7.1.16}$$

Even though the elephant is more than one hundred-times heavier than the skater, it exerts less than one-half of the pressure and would therefore be less likely to fall though thin ice. On the other hand, if the skater removes her skates and stands with bare feet (or regular footwear) on the ice, the larger area over which her weight is applied greatly reduces the pressure exerted:

$$\text{pressure per human foot} = 120\frac{\text{lb}}{\text{skater}} \times \frac{1\text{ skater}}{2\text{ feet}} \times \frac{1\text{ foot}}{30\text{ in}^2} = 2\text{ lb/in}^2 \tag{7.1.17}$$

The SI unit of pressure is the pascal (Pa), with 1 Pa = 1 N/m^2, where N is the newton, a unit of force defined as 1 kg m/s^2. One pascal is a small pressure; in many cases, it is more convenient to use units of kilopascal (1 kPa = 1000 Pa) or bar (1 bar = 100,000 Pa). In the United States, pressure is often measured in pounds of force on an area of one square inch—pounds per square inch (psi)—for example, in car tires. Pressure can also be measured using the unit atmosphere (atm), which originally represented the average sea level air pressure at the approximate latitude of Paris (45°). Table 7.1.1 provides some information on these and a few other common units for pressure measurements

***Table 7.1.1:** Pressure Units*

Unit Name and Abbreviation	Definition or Relation to Other Unit	Comment
pascal (Pa)	1 Pa = 1 N/m^2	recommended IUPAC unit
kilopascal (kPa)	1 kPa = 1000 Pa	
pounds per square inch (psi)	air pressure at sea level is ~14.7 psi	
atmosphere (atm)	1 atm = 101,325 Pa	air pressure at sea level is ~1 atm
bar (bar, or b)	1 bar = 100,000 Pa (exactly)	commonly used in meteorology
millibar (mbar, or mb)	1000 mbar = 1 bar	
inches of mercury (in. Hg)	1 in. Hg = 3386 Pa	used by aviation industry, also some weather reports
torr	$1\ \text{torr} = \frac{1}{760}\ \text{atm}$	named after Evangelista Torricelli, inventor of the barometer
millimeters of mercury (mm Hg)	1 mm Hg ~1 torr	

Example 7.1.3: Conversion of Pressure Units

The United States National Weather Service reports pressure in both inches of Hg and millibars. Convert a pressure of 29.2 in. Hg into:

a. torr
b. atm
c. kPa
d. mbar

Solution

This is a unit conversion problem. The relationships between the various pressure units are given in Table **7.1.1**.

a. $29.2\ \text{in Hg} \times \frac{25.4\ \text{mm}}{1\ \text{in}} \times \frac{1\ \text{torr}}{1\ \text{mm Hg}} = 742\ \text{torr}$

b. $742\ \text{torr} \times \frac{1\ \text{atm}}{760\ \text{torr}} = 0.976\ \text{atm}$

c. $742\ \text{torr} \times \frac{101.325\ \text{kPa}}{760\ \text{torr}} = 98.9\ \text{kPa}$

d. $98.9\ \text{kPa} \times \frac{1000\ \text{Pa}}{1\ \text{kPa}} \times \frac{1\ \text{bar}}{100,000\ \text{Pa}} \times \frac{1000\ \text{mbar}}{1\ \text{bar}} = 989\ \text{mbar}$

Exercise 7.1.3

A typical barometric pressure in Kansas City is 740 torr. What is this pressure in atmospheres, in millimeters of mercury, in kilopascals, and in bar?

Answer

0.974 atm; 740 mm Hg; 98.7 kPa; 0.987 bar

We can measure atmospheric pressure, the force exerted by the atmosphere on the earth's surface, with a barometer (Figure **7.1.4**). A barometer is a glass tube that is closed at one end, filled with a nonvolatile liquid such as mercury, and then inverted and immersed in a container of that liquid. The atmosphere exerts pressure on the liquid outside the tube, the column of liquid exerts

pressure inside the tube, and the pressure at the liquid surface is the same inside and outside the tube. The height of the liquid in the tube is therefore proportional to the pressure exerted by the atmosphere.

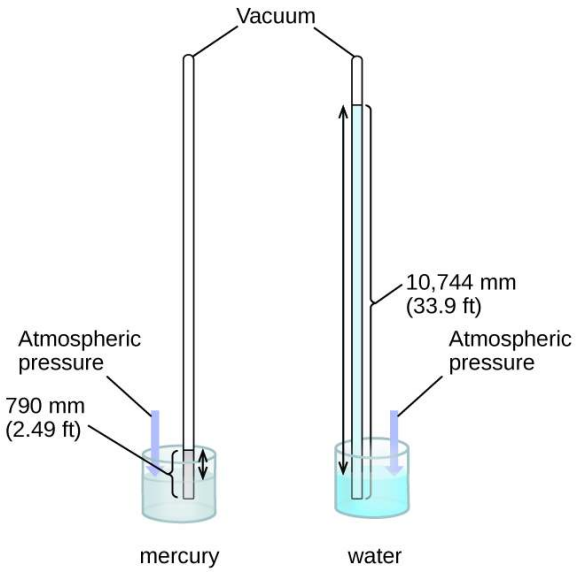


***Figure* 7.1.4**: *In a barometer, the height, h, of the column of liquid is used as a measurement of the air pressure. Using very dense liquid mercury (left) permits the construction of reasonably sized barometers, whereas using water (right) would require a barometer more than 30 feet tall.*



***Video* 7.1.2**: *The invention of the barometer.*

If the liquid is water, normal atmospheric pressure will support a column of water over 10 meters high, which is rather inconvenient for making (and reading) a barometer. Because mercury (Hg) is about 13.6-times denser than water, a mercury barometer only needs to be $\frac{1}{13.6}$ as tall as a water barometer—a more suitable size. Standard atmospheric pressure of 1 atm at sea level (101,325 Pa) corresponds to a column of mercury that is about 760 mm (29.92 in.) high. The torr was originally intended to be a unit equal to one millimeter of mercury, but it no longer corresponds exactly. The pressure exerted by a fluid due to gravity is known as hydrostatic pressure, p:

$$p = h\rho g \tag{7.1.18}$$

where

- h is the height of the fluid,
- ρ is the density of the fluid, and
- g is acceleration due to gravity.

A manometer is a device similar to a barometer that can be used to measure the pressure of a gas trapped in a container. A closed-end manometer is a U-shaped tube with one closed arm, one arm that connects to the gas to be measured, and a nonvolatile liquid (usually mercury) in between. As with a barometer, the distance between the liquid levels in the two arms of the tube (*h* in the diagram) is proportional to the pressure of the gas in the container. An open-end manometer (Figure **7.1.5**) is the same as a closed-end manometer, but one of its arms is open to the atmosphere. In this case, the distance between the liquid levels corresponds to the difference in pressure between the gas in the container and the atmosphere.

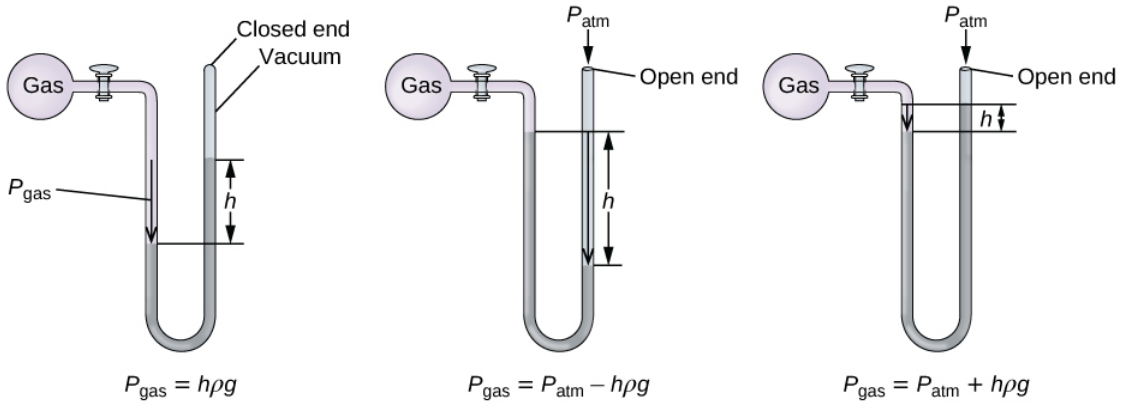


Figure* 7.1.5*: *A manometer can be used to measure the pressure of a gas. The (difference in) height between the liquid levels (h) is a measure of the pressure. Mercury is usually used because of its large density.*

Application: Measuring Blood Pressure

Blood pressure is measured using a device called a sphygmomanometer (Greek *sphygmos* = "pulse"). It consists of an inflatable cuff to restrict blood flow, a manometer to measure the pressure, and a method of determining when blood flow begins and when it becomes impeded (Figure **7.1.6**). Since its invention in 1881, it has been an essential medical device. There are many types of sphygmomanometers: manual ones that require a stethoscope and are used by medical professionals; mercury ones, used when the most accuracy is required; less accurate mechanical ones; and digital ones that can be used with little training but that have limitations. When using a sphygmomanometer, the cuff is placed around the upper arm and inflated until blood flow is completely blocked, then slowly released. As the heart beats, blood forced through the arteries causes a rise in pressure. This rise in pressure at which blood flow begins is the *systolic pressure*—the peak pressure in the cardiac cycle. When the cuff's pressure equals the arterial systolic pressure, blood flows past the cuff, creating audible sounds that can be heard using a stethoscope. This is followed by a decrease in pressure as the heart's ventricles prepare for another beat. As cuff pressure continues to decrease, eventually sound is no longer heard; this is the *diastolic pressure*—the lowest pressure (resting phase) in the cardiac cycle. Blood pressure units from a sphygmomanometer are in terms of millimeters of mercury (mm Hg).



Figure* 7.1.6*: *(a) A medical technician prepares to measure a patient's blood pressure with a sphygmomanometer. (b) A typical sphygmomanometer uses a valved rubber bulb to inflate the cuff and a diaphragm gauge to measure pressure. (credit a: modification of work by Master Sgt. Jeffrey Allen)*

How the Sciences interconnect: Meteorology, Climatology, and Atmospheric Science

Throughout the ages, people have observed clouds, winds, and precipitation, trying to discern patterns and make predictions: when it is best to plant and harvest; whether it is safe to set out on a sea voyage; and much more. We now face complex weather and atmosphere-related challenges that will have a major impact on our civilization and the ecosystem. Several different scientific disciplines use chemical principles to help us better understand weather, the atmosphere, and climate. These are meteorology, climatology, and atmospheric science. Meteorology is the study of the atmosphere, atmospheric phenomena, and atmospheric effects on earth's weather. Meteorologists seek to understand and predict the weather in the short term, which can save lives and benefit the economy. Weather forecasts (Figure **7.1.7**) are the result of thousands of measurements of air pressure, temperature, and the like, which are compiled, modeled, and analyzed in weather centers worldwide.

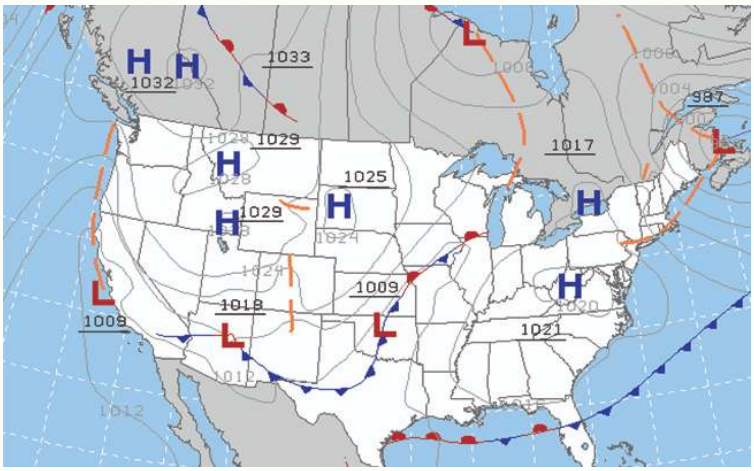


Figure* 7.1.7*: *Meteorologists use weather maps to describe and predict weather. Regions of high (H) and low (L) pressure have large effects on weather conditions. The gray lines represent locations of constant pressure known as isobars. (credit: modification of work by National Oceanic and Atmospheric Administration)*

In terms of weather, low-pressure systems occur when the earth's surface atmospheric pressure is lower than the surrounding environment: Moist air rises and condenses, producing clouds. Movement of moisture and air within various weather fronts instigates most weather events.

The atmosphere is the gaseous layer that surrounds a planet. Earth's atmosphere, which is roughly 100–125 km thick, consists of roughly 78.1% nitrogen and 21.0% oxygen, and can be subdivided further into the regions shown in Figure **7.1.8**: the exosphere (furthest from earth, > 700 km above sea level), the thermosphere (80–700 km), the mesosphere (50–80 km), the stratosphere (second lowest level of our atmosphere, 12–50 km above sea level), and the troposphere (up to 12 km above sea level, roughly 80% of the earth's atmosphere by mass and the layer where most weather events originate). As you go higher in the troposphere, air density and temperature both decrease.

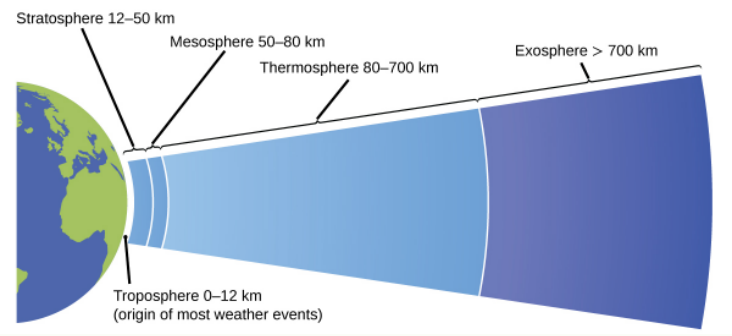


Figure* 7.1.8*: *Earth's atmosphere has five layers: the troposphere, the stratosphere, the mesosphere, the thermosphere, and the exosphere.*

Climatology is the study of the climate, averaged weather conditions over long time periods, using atmospheric data. However, climatologists study patterns and effects that occur over decades, centuries, and millennia, rather than shorter time frames of

hours, days, and weeks like meteorologists. Atmospheric science is an even broader field, combining meteorology, climatology, and other scientific disciplines that study the atmosphere.

Summary



Video* 7.1.3: A brief overview of pressure.*

Gases exert pressure, which is force per unit area. The pressure of a gas may be expressed in the SI unit of pascal or kilopascal, as well as in many other units including torr, atmosphere, and bar. Atmospheric pressure is measured using a barometer; other gas pressures can be measured using one of several types of manometers.

Key Equations

- $T_{°C} = \frac{5}{9} \times T_{°F} - 32$
- $T_{°F} = \frac{9}{5} \times T_{°C} + 32$
- $T_K = °C + 273.15$
- $T_{°C} = K - 273.15$
- $P = \frac{F}{A}$
- $p = h\rho g$

Glossary

atmosphere (atm)
unit of pressure; 1 atm = 101,325 Pa

bar
(bar or b) unit of pressure; 1 bar = 100,000 Pa

barometer
device used to measure atmospheric pressure

Celsius (°C)
unit of temperature; water freezes at 0 °C and boils at 100 °C on this scale

Fahrenheit
unit of temperature; water freezes at 32 °F and boils at 212 °F on this scale

hydrostatic pressure
pressure exerted by a fluid due to gravity

kelvin (K)
SI unit of temperature; 273.15 K = 0 ºC

manometer

device used to measure the pressure of a gas trapped in a container

pascal (Pa)

SI unit of pressure; 1 Pa = 1 N/m^2

pounds per square inch (psi)

unit of pressure common in the US

pressure

force exerted per unit area

torr

unit of pressure; $1\ \text{torr} = \frac{1}{760}\ \text{atm}$

Contributors

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7.1.1: Practice Problems- Temperature and Pressure

PROBLEM 7.1.1.1

Convert the boiling temperature of gold, 2966 °C, into degrees Fahrenheit and kelvin.

Answer

5371 °F

3239 K

PROBLEM 7.1.1.2

Convert the temperature of scalding water, 54 °C, into degrees Fahrenheit and kelvin.

Answer

129.2 °F

327 K

Click here to see a video solution



PROBLEM 7.1.1.3

Convert the temperature of the coldest area in a freezer, −10 °F, to degrees Celsius and kelvin.

Answer

−23 °C

250 K

PROBLEM 7.1.1.4

Convert the temperature of dry ice, −77 °C, into degrees Fahrenheit and kelvin.

Answer

-106.6 °F

196 K

Click here to see a video solution



PROBLEM 7.1.1.5

Convert the boiling temperature of liquid ammonia, −28.1 °F, into degrees Celsius and kelvin.

Answer

−33.4 °C

239.8 K

PROBLEM 7.1.1.6

The label on a pressurized can of spray disinfectant warns against heating the can above 130 °F. What are the corresponding temperatures on the Celsius and kelvin temperature scales?

Answer

90 °C

363 K

Click here to see a video solution



PROBLEM 7.1.1.7

The weather in Europe was unusually warm during the summer of 1995. The TV news reported temperatures as high as 45 °C. What was the temperature on the Fahrenheit scale?

Answer

113 °F

PROBLEM 7.1.1.8

Why do some small bridges have weight limits that depend on how many wheels or axles the crossing vehicle has?

Answer

Pressure is defined as force per unit area. Similarly to the example in the text of the elephant and the figure skater, the more wheels or axles on the vehicle, the more area the weight is spread over, causing less pressure to be exerted on any one point on the bridge.

PROBLEM 7.1.1.9

A typical barometric pressure in Redding, California, is about 750 mm Hg. Calculate this pressure in atm and kPa

Answer

0.987 atm

99.97 kPa

PROBLEM 7.1.1.10

A typical barometric pressure in Denver, Colorado, is 615 mm Hg. What is this pressure in atmospheres and kilopascals?

Answer

0.809 atm

82.0 kPa

Click here to see a video solution



PROBLEM 7.1.1.11

Canadian tire pressure gauges are marked in units of kilopascals. What reading on such a gauge corresponds to 32 psi?

Answer

2.2×10^2 kPa

PROBLEM 7.1.1.12

During the Viking mission landings on Mars, the atmospheric pressure was determined to be on the average about 6.50 millibars (1 bar = 0.987 atm). What is that pressure in torr and kPa?

Answer

4.88 torr

0.650 kPa

Click here to see a video solution



PROBLEM 7.1.1.13

The pressure of the atmosphere on the surface of the planet Venus is about 88.8 atm. Compare that pressure in psi to the normal pressure on earth at sea level in psi.

Answer

Earth: 14.7 psi

Venus: 13.1×10^3 psi

PROBLEM 7.1.1.14

Consider this scenario and answer the following questions: On a mid-August day in the northeastern United States, the following information appeared in the local newspaper: atmospheric pressure at sea level 29.97 in. Hg, 1013.9 mbar.

a. What was the pressure in kPa?
b. The pressure near the seacoast in the northeastern United States is usually reported near 30.0 in. Hg. During a hurricane, the pressure may fall to near 28.0 in. Hg. Calculate the drop in pressure in torr.

Answer a

101.5 kPa

Answer b

51 torr drop

PROBLEM 7.1.1.15

The pressure of a sample of gas is measured at sea level is 264 mmHg. Determine the pressure of the gas in:

a. torr
b. Pa

c. bar

Answer a

264 torr

Answer b

35197 Pa

Answer c

0.352 bar

Contributors

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- Adelaide Clark, Oregon Institute of Technology

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7.2: The Gas Laws

Skills to Develop

- Identify the mathematical relationships between the various properties of gases
- Use the ideal gas law, and related gas laws, to compute the values of various gas properties under specified conditions

During the seventeenth and especially eighteenth centuries, driven both by a desire to understand nature and a quest to make balloons in which they could fly (Figure **7.2.1**), a number of scientists established the relationships between the macroscopic physical properties of gases, that is, pressure, volume, temperature, and amount of gas. Although their measurements were not precise by today's standards, they were able to determine the mathematical relationships between pairs of these variables (e.g., pressure and temperature, pressure and volume) that hold for an *ideal* gas—a hypothetical construct that real gases approximate under certain conditions. Eventually, these individual laws were combined into a single equation—the *ideal gas law*—that relates gas quantities for gases and is quite accurate for low pressures and moderate temperatures. We will consider the key developments in individual relationships (for pedagogical reasons not quite in historical order), then put them together in the ideal gas law.

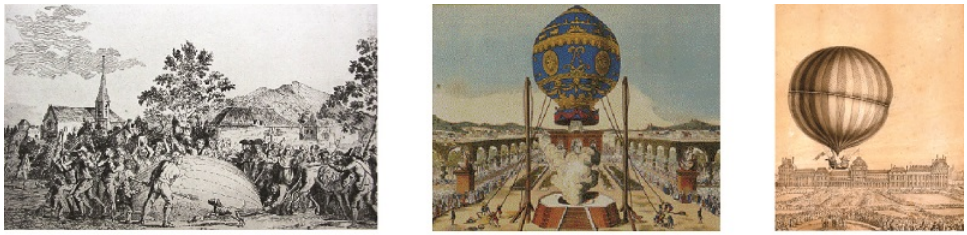


***Figure* 7.2.1**: *In 1783, the first (a) hydrogen-filled balloon flight, (b) manned hot air balloon flight, and (c) manned hydrogen-filled balloon flight occurred. When the hydrogen-filled balloon depicted in (a) landed, the frightened villagers of Gonesse reportedly destroyed it with pitchforks and knives. The launch of the latter was reportedly viewed by 400,000 people in Paris.*

Pressure and Temperature: Amontons's Law

Imagine filling a rigid container attached to a pressure gauge with gas and then sealing the container so that no gas may escape. If the container is cooled, the gas inside likewise gets colder and its pressure is observed to decrease. Since the container is rigid and tightly sealed, both the volume and number of moles of gas remain constant. If we heat the sphere, the gas inside gets hotter (Figure **7.2.2**) and the pressure increases.

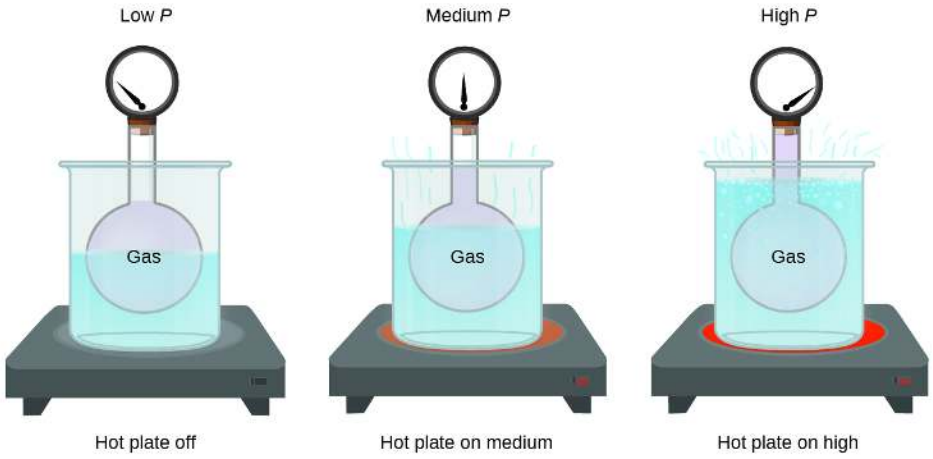


***Figure* 7.2.2**: *The effect of temperature on gas pressure: When the hot plate is off, the pressure of the gas in the sphere is relatively low. As the gas is heated, the pressure of the gas in the sphere increases.*

This relationship between temperature and pressure is observed for any sample of gas confined to a constant volume. An example of experimental pressure-temperature data is shown for a sample of air under these conditions in Figure **7.2.3**. We find that

temperature and pressure are linearly related, and if the temperature is on the kelvin scale, then P and T are directly proportional (again, when *volume and moles of gas are held constant*); if the temperature on the kelvin scale increases by a certain factor, the gas pressure increases by the same factor.

Temperature (°C)	Temperature (K)	Pressure (kPa)
−100	173	36.0
−50	223	46.4
0	273	56.7
50	323	67.1
100	373	77.5
150	423	88.0

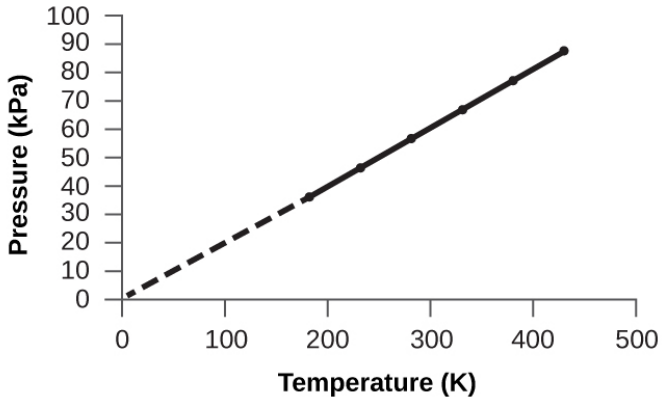


***Figure* 7.2.3**: *For a constant volume and amount of air, the pressure and temperature are directly proportional, provided the temperature is in kelvin. (Measurements cannot be made at lower temperatures because of the condensation of the gas.) When this line is extrapolated to lower pressures, it reaches a pressure of 0 at –273 °C, which is 0 on the kelvin scale and the lowest possible temperature, called absolute zero.*

Guillaume Amontons was the first to empirically establish the relationship between the pressure and the temperature of a gas (~1700), and Joseph Louis Gay-Lussac determined the relationship more precisely (~1800). Because of this, the *P*-*T* relationship for gases is known as either Amontons's law or Gay-Lussac's law. Under either name, it states that *the pressure of a given amount of gas is directly proportional to its temperature on the kelvin scale when the volume is held constant*. Mathematically, this can be written:

$$P \propto T \text{ or } P = \text{constant} \times T \text{ or } P = k \times T \tag{7.2.1}$$

where ∝ means "is proportional to," and k is a proportionality constant that depends on the identity, amount, and volume of the gas. For a confined, constant volume of gas, the ratio $\frac{P}{T}$ is therefore constant (i.e., $\frac{P}{T} = k$). If the gas is initially in "Condition 1" (with $P = P_1$ and $T = T_1$), and then changes to "Condition 2" (with $P = P_2$ and $T = T_2$), we have that $\frac{P_1}{T_1} = k$ and $\frac{P_2}{T_2} = k$, which reduces to $\frac{P_1}{T_1} = \frac{P_2}{T_2}$. This equation is useful for pressure-temperature calculations for a confined gas at constant volume. Note that temperatures must be on the kelvin scale for any gas law calculations (0 on the kelvin scale and the lowest possible temperature is called absolute zero). (Also note that there are at least three ways we can describe how the pressure of a gas changes as its temperature changes: We can use a table of values, a graph, or a mathematical equation.)

Example 7.2.1: Predicting Change in Pressure with Temperature

A can of hair spray is used until it is empty except for the propellant, isobutane gas.

a. On the can is the warning "Store only at temperatures below 120 °F (48.8 °C). Do not incinerate." Why?
b. The gas in the can is initially at 24 °C and 360 kPa, and the can has a volume of 350 mL. If the can is left in a car that reaches 50 °C on a hot day, what is the new pressure in the can?

Solution

a. The can contains an amount of isobutane gas at a constant volume, so if the temperature is increased by heating, the pressure will increase proportionately. High temperature could lead to high pressure, causing the can to burst. (Also, isobutane is combustible, so incineration could cause the can to explode.)
b. We are looking for a pressure change due to a temperature change at constant volume, so we will use Amontons's/Gay-Lussac's law. Taking P_1 and T_1 as the initial values, T_2 as the temperature where the pressure is unknown and P_2 as the unknown pressure, and converting °C to K, we have:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ which means that } \frac{360\ k\text{Pa}}{297\ \text{K}} = \frac{P_2}{323\ \text{K}}$$

Rearranging and solving gives:

$$P_2 = \frac{360\ \text{kPa} \times 323\ \cancel{\text{K}}}{297\ \cancel{\text{K}}} = 390\ \text{kPa}$$

Exercise 7.2.1

A sample of nitrogen, N_2, occupies 45.0 mL at 27 °C and 600 torr. What pressure will it have if cooled to –73 °C while the volume remains constant?

Answer

400 torr

Volume and Temperature: Charles's Law

If we fill a balloon with air and seal it, the balloon contains a specific amount of air at atmospheric pressure, let's say 1 atm. If we put the balloon in a refrigerator, the gas inside gets cold and the balloon shrinks (although both the amount of gas and its pressure remain constant). If we make the balloon very cold, it will shrink a great deal, and it expands again when it warms up.



Video* 7.2.1*: *This video shows how cooling and heating a gas causes its volume to decrease or increase, respectively.*

These examples of the effect of temperature on the volume of a given amount of a confined gas at constant pressure are true in general: The volume increases as the temperature increases, and decreases as the temperature decreases. Volume-temperature data for a 1-mole sample of methane gas at 1 atm are listed and graphed in Figure 7.2.2.

Temperature (°C)	Temperature (K)	Volume (L)
–3	270	22
–23	250	21
–53	220	18
–162	111	9

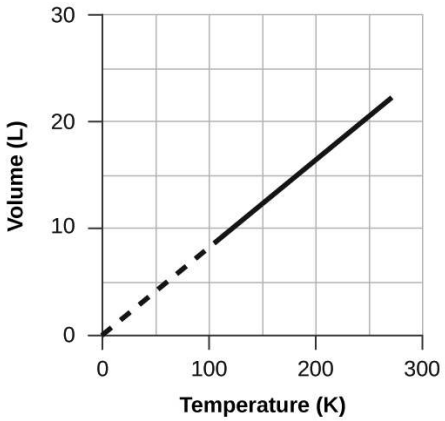


Figure* 7.2.4*: *The volume and temperature are linearly related for 1 mole of methane gas at a constant pressure of 1 atm. If the temperature is in kelvin, volume and temperature are directly proportional. The line stops at 111 K because methane liquefies at*

this temperature; when extrapolated, it intersects the graph's origin, representing a temperature of absolute zero.

The relationship between the volume and temperature of a given amount of gas at constant pressure is known as Charles's law in recognition of the French scientist and balloon flight pioneer Jacques Alexandre César Charles. Charles's law states that *the volume of a given amount of gas is directly proportional to its temperature on the kelvin scale when the pressure is held constant.*

Mathematically, this can be written as:

$$V \alpha T \text{ or } V = \text{constant} \cdot T \text{ or } V = k \cdot T \text{ or } V_1/T_1 = V_2/T_2 \tag{7.2.2}$$

with *k* being a proportionality constant that depends on the amount and pressure of the gas.

For a confined, constant pressure gas sample, $\frac{V}{T}$ is constant (i.e., the ratio = *k*), and as seen with the *P*-*T* relationship, this leads to another form of Charles's law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$.

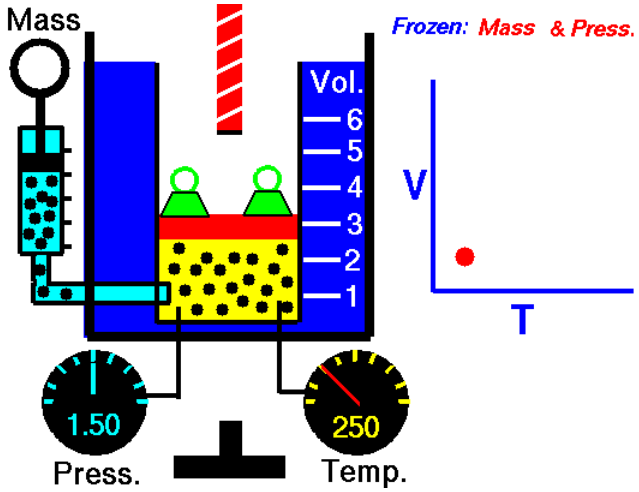


***Figure* 7.2.5:** *An illustration of Charles's Law, showing that as volume decreases, so does temperature, and vice versa.* This file is in the **public domain** because it was created by NASA . NASA's copyright policy states that "NASA material is not copyrighted **unless otherwise specified** ". (*NASA copyright policy* and *JPL Image Use Policy*).

Example 7.2.2: Predicting Change in Volume with Temperature

A sample of carbon dioxide, CO_2, occupies 0.300 L at 10 °C and 750 torr. What volume will the gas have at 30 °C and 750 torr?

Solution

Because we are looking for the volume change caused by a temperature change at constant pressure, this is a job for Charles's law. Taking V_1 and T_1 as the initial values, T_2 as the temperature at which the volume is unknown and V_2 as the unknown volume, and converting °C into K we have:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ which means that } \frac{0.300\ \text{L}}{283\ \text{K}} = \frac{V_2}{303\ \text{K}}$$

Rearranging and solving gives: $V_2 = \frac{0.300\ \text{L} \times 303\ \cancel{\text{K}}}{283\ \cancel{\text{K}}} = 0.321\ \text{L}$

This answer supports our expectation from Charles's law, namely, that raising the gas temperature (from 283 K to 303 K) at a constant pressure will yield an increase in its volume (from 0.300 L to 0.321 L).

Exercise 7.2.2

A sample of oxygen, O_2, occupies 32.2 mL at 30 °C and 452 torr. What volume will it occupy at –70 °C and the same pressure?

Answer

21.6 mL

Example 7.2.3: Measuring Temperature with a Volume

Change Temperature is sometimes measured with a gas thermometer by observing the change in the volume of the gas as the temperature changes at constant pressure. The hydrogen in a particular hydrogen gas thermometer has a volume of 150.0 cm^3 when immersed in a mixture of ice and water (0.00 °C). When immersed in boiling liquid ammonia, the volume of the hydrogen, at the same pressure, is 131.7 cm^3. Find the temperature of boiling ammonia on the kelvin and Celsius scales.

Solution

A volume change caused by a temperature change at constant pressure means we should use Charles's law. Taking V_1 and T_1 as the initial values, T_2 as the temperature at which the volume is unknown and V_2 as the unknown volume, and converting °C into K we have:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ which means that } \frac{150.0\ \text{cm}^3}{273.15\ \text{K}} = \frac{131.7\ \text{cm}^3}{T_2} \tag{7.2.3}$$

Rearrangement gives $T_2 = \dfrac{131.7\ \cancel{\text{cm}^3} \times 273.15\ \text{K}}{150.0\ \text{cm}^3} = 239.8\ \text{K}$

Subtracting 273.15 from 239.8 K, we find that the temperature of the boiling ammonia on the Celsius scale is –33.4 °C.

Exercise 7.2.3

What is the volume of a sample of ethane at 467 K and 1.1 atm if it occupies 405 mL at 298 K and 1.1 atm?

Answer

635 mL

Volume and Pressure: Boyle's Law

If we partially fill an airtight syringe with air, the syringe contains a specific amount of air at constant temperature, say 25 °C. If we slowly push in the plunger while keeping temperature constant, the gas in the syringe is compressed into a smaller volume and its pressure increases; if we pull out the plunger, the volume increases and the pressure decreases. This example of the effect of volume on the pressure of a given amount of a confined gas is true in general. Decreasing the volume of a contained gas will increase its pressure, and increasing its volume will decrease its pressure. In fact, if the volume increases by a certain factor, the pressure decreases by the same factor, and vice versa. Volume-pressure data for an air sample at room temperature are graphed in Figure 7.2.6.



This figure contains a diagram and two graphs. The diagram shows a syringe labeled with a scale in m l or c c with multiples of 5 labeled beginning at 5 and ending at 30. The markings halfway between these measurements are also provided. Attached at the top of the syringe is a pressure gauge with a scale marked by fives from 40 on the left to 5 on the right. The gauge needle rests between 10 and 15, slightly closer to 15. The syringe plunger position indicates a volume measurement about halfway between 10 and 15 m l or c c. The first graph is labeled "V (m L)" on the horizontal axis and "P (p s i)" on the vertical axis. Points are labeled at 5, 10, 15, 20, and 25 m L with corresponding values of 39.0, 19.5, 13.0, 9.8, and 6.5 p s i. The points are connected with a smooth curve that is declining at a decreasing rate of change. The second graph is labeled "V (m L)" on the horizontal axis and "1 divided by P (p s i)" on the vertical axis. The horizontal axis is labeled at multiples of 5, beginning at zero and extending up to 35 m L. The vertical axis is labeled by multiples of 0.02, beginning at 0 and extending up to 0.18. Six points indicated by black dots on this graph are connected with a black line segment showing a positive linear trend.

Figure 7.2.6: *When a gas occupies a smaller volume, it exerts a higher pressure; when it occupies a larger volume, it exerts a lower pressure (assuming the amount of gas and the temperature do not change). Since P and V are inversely proportional, a graph of* $\frac{1}{P}$ *vs. V is linear.*

Unlike the *P*-*T* and *V*-*T* relationships, pressure and volume are not directly proportional to each other. Instead, *P* and *V* exhibit inverse proportionality: Increasing the pressure results in a decrease of the volume of the gas. Mathematically this can be written:

$$P \alpha 1/V \text{ or } P = k \cdot 1/V \text{ or } P \cdot V = k \text{ or } P_1V_1 = P_2V_2 \tag{7.2.4}$$

with *k* being a constant. Graphically, this relationship is shown by the straight line that results when plotting the inverse of the pressure $\left(\frac{1}{P}\right)$ versus the volume (*V*), or the inverse of volume $\left(\frac{1}{V}\right)$ versus the pressure (*P*). Graphs with curved lines are difficult to read accurately at low or high values of the variables, and they are more difficult to use in fitting theoretical equations and parameters to experimental data. For those reasons, scientists often try to find a way to "linearize" their data. If we plot *P* versus *V*, we obtain a hyperbola (Figure 7.2.7).

This diagram shows two graphs. In a, a graph is shown with volume on the horizontal axis and pressure on the vertical axis. A curved line is shown on the graph showing a decreasing trend with a decreasing rate of change. In b, a graph is shown with volume on the horizontal axis and one divided by pressure on the vertical axis. A line segment, beginning at the origin of the graph, shows a positive, linear trend.

***Figure* 7.2.7:** *The relationship between pressure and volume is inversely proportional. (a) The graph of P vs. V is a hyperbola, whereas (b) the graph of* $\left(\frac{1}{P}\right)$ *vs. V is linear.*

The relationship between the volume and pressure of a given amount of gas at constant temperature was first published by the English natural philosopher Robert Boyle over 300 years ago. It is summarized in the statement now known as Boyle's law: *The volume of a given amount of gas held at constant temperature is inversely proportional to the pressure under which it is measured.*

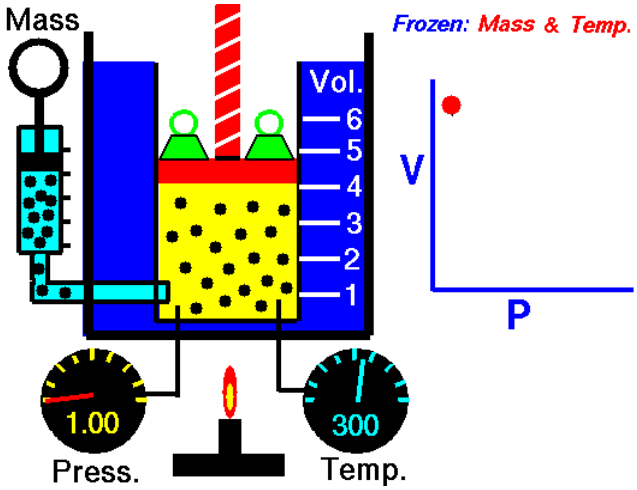


***Figure* 7.2.7:** *An illustration of Boyle's Law, showing that as volume decreases, pressure increases, and vice versa.* This file is in the **public domain** because it was created by NASA . NASA's copyright policy states that "NASA material is not copyrighted **unless otherwise specified** ". (*NASA copyright policy* and *JPL Image Use Policy*).

Example 7.2.4: Volume of a Gas Sample

The sample of gas has a volume of 15.0 mL at a pressure of 13.0 psi. Determine the pressure of the gas at a volume of 7.5 mL, using:

a. the *P*-*V* graph in Figure **7.2.6*a***
b. the $\frac{1}{P}$ vs. *V* graph in Figure **7.2.6*b***
c. the Boyle's law equation

Comment on the likely accuracy of each method.

Solution

a. Estimating from the *P*-*V* graph gives a value for *P* somewhere around 27 psi.
b. Estimating from the $\frac{1}{P}$ versus *V* graph give a value of about 26 psi.
c. From Boyle's law, we know that the product of pressure and volume (*PV*) for a given sample of gas at a constant temperature is always equal to the same value. Therefore we have $P_1V_1 = k$ and $P_2V_2 = k$ which means that $P_1V_1 = P_2V_2$.

Using P_1 and V_1 as the known values 13.0 psi and 15.0 mL, P_2 as the pressure at which the volume is unknown, and V_2 as the unknown volume, we have:

$$P_1V_1 = P_2V_2 \text{ or } 13.0 \text{ psi} \times 15.0 \text{ mL} = P_2 \times 7.5 \, m\text{L} \tag{7.2.5}$$

Solving:

$$P_2 = \frac{13.0 \text{ psi} \times 15.0 \cancel{\text{mL}}}{7.5 \cancel{\text{mL}}} = 26 \text{ psi} \tag{7.2.6}$$

It was more difficult to estimate well from the *P*-*V* graph, so (a) is likely more inaccurate than (b) or (c). The calculation will be as accurate as the equation and measurements allow.

Exercise 7.2.4

The sample of gas has a volume of 30.0 mL at a pressure of 6.5 psi. Determine the volume of the gas at a pressure of 11.0 psi, using:

a. the *P*-*V* graph in Figure **7.2.6*a***
b. the $\frac{1}{P}$ vs. *V* graph in Figure **7.2.6*b***
c. the Boyle's law equation

Comment on the likely accuracy of each method.

Answer a

about 17–18 mL

Answer b

~18 mL

Answer c

17.7 mL; it was more difficult to estimate well from the *P*-*V* graph, so (a) is likely more inaccurate than (b); the calculation will be as accurate as the equation and measurements allow

> *The ideal gas law is easy to remember and apply in solving problems, as long as you get the **proper values** and units for the gas constant, R.*

Chemistry in Everyday Life: Breathing and Boyle's Law

What do you do about 20 times per minute for your whole life, without break, and often without even being aware of it? The answer, of course, is respiration, or breathing. How does it work? It turns out that the gas laws apply here. Your lungs take in gas that your body needs (oxygen) and get rid of waste gas (carbon dioxide). Lungs are made of spongy, stretchy tissue that expands and contracts while you breathe. When you inhale, your diaphragm and intercostal muscles (the muscles between your ribs) contract, expanding your chest cavity and making your lung volume larger. The increase in volume leads to a decrease in pressure (Boyle's law). This causes air to flow into the lungs (from high pressure to low pressure). When you exhale, the process reverses: Your diaphragm and rib muscles relax, your chest cavity contracts, and your lung volume decreases, causing the pressure to increase (Boyle's law again), and air flows out of the lungs (from high pressure to low pressure). You then breathe in and out again, and again, repeating this Boyle's law cycle for the rest of your life (Figure **7.2.8**).

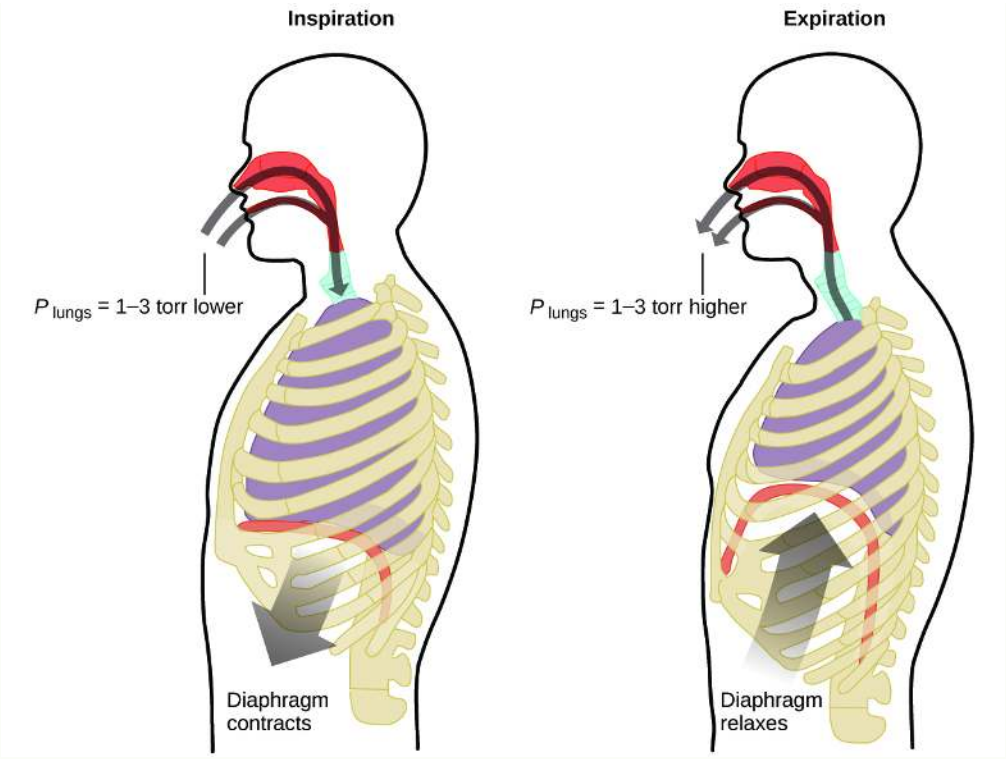


Figure 7.2.8: *Breathing occurs because expanding and contracting lung volume creates small pressure differences between your lungs and your surroundings, causing air to be drawn into and forced out of your lungs.*

Moles of Gas and Volume: Avogadro's Law

The Italian scientist Amedeo Avogadro advanced a hypothesis in 1811 to account for the behavior of gases, stating that equal volumes of all gases, measured under the same conditions of temperature and pressure, contain the same number of molecules. Over time, this relationship was supported by many experimental observations as expressed by Avogadro's law: *For a confined gas, the volume (V) and number of moles (n) are directly proportional if the pressure and temperature both remain constant*.

In equation form, this is written as:

$$V \propto n \text{ or } V = k \times n \text{ or } \frac{V_1}{n_1} = \frac{V_2}{n_2} \tag{7.2.7}$$

Mathematical relationships can also be determined for the other variable pairs, such as P versus n, and n versus T.

The Ideal Gas Law

To this point, four separate laws have been discussed that relate pressure, volume, temperature, and the number of moles of the gas:

- Boyle's law: PV = constant at constant T and n
- Amontons's law: $\frac{P}{T}$ = constant at constant V and n
- Charles's law: $\frac{V}{T}$ = constant at constant P and n
- Avogadro's law: $\frac{V}{n}$ = constant at constant P and T

Combining these four laws yields the ideal gas law, a relation between the pressure, volume, temperature, and number of moles of a gas:

$$PV = nRT \tag{7.2.8}$$

where P is the pressure of a gas, V is its volume, n is the number of moles of the gas, T is its temperature on the kelvin scale, and R is a constant called the ideal gas constant or the universal gas constant. The units used to express pressure, volume, and temperature will determine the proper form of the gas constant as required by dimensional analysis, the most commonly encountered values being 0.08206 L atm mol^{-1} K^{-1} and 8.314 kPa L mol^{-1} K^{-1}.

Gases whose properties of P, V, and T are accurately described by the ideal gas law (or the other gas laws) are said to exhibit *ideal behavior* or to approximate the traits of an ideal gas. An ideal gas is a hypothetical construct that may be used along with *kinetic molecular theory* to effectively explain the gas laws as will be described in a later module of this chapter. Although all the calculations presented in this module assume ideal behavior, this assumption is only reasonable for gases under conditions of relatively low pressure and high temperature. In the final module of this chapter, a modified gas law will be introduced that accounts for the *non-ideal* behavior observed for many gases at relatively high pressures and low temperatures.

The ideal gas equation contains five terms, the gas constant R and the variable properties P, V, n, and T. Specifying any four of these terms will permit use of the ideal gas law to calculate the fifth term as demonstrated in the following example exercises.

Example 7.2.5: Using the Ideal Gas Law

Methane, CH_4, is being considered for use as an alternative automotive fuel to replace gasoline. One gallon of gasoline could be replaced by 655 g of CH_4. What is the volume of this much methane at 25 °C and 745 torr?

Solution

We must rearrange $PV = nRT$ to solve for V: $V = \frac{nRT}{P}$

If we choose to use R = 0.08206 L atm mol^{-1} K^{-1}, then the amount must be in moles, temperature must be in kelvin, and pressure must be in atm.

Converting into the "right" units:

$$n = 655 \text{ g } \cancel{CH_4} \times \frac{1 \text{ mol}}{16.043 \text{ g } \cancel{CH_4}} = 40.8 \text{ mol} \tag{7.2.9}$$

$$T = 25°\text{C} + 273 = 298 \text{ K} \tag{7.2.10}$$

$$P = 745 \cancel{\text{ torr}} \times \frac{1\text{atm}}{760 \cancel{\text{ torr}}} = 0.980 \text{ atm} \tag{7.2.11}$$

$$V = \frac{nRT}{P} = \frac{(40.8 \cancel{\text{ mol}})(0.08206 \text{ L } \cancel{\text{atm}}\cancel{\text{mol}^{-1}}\cancel{\text{K}^{-1}})(298 \text{ K})}{0.980 \cancel{\text{ atm}}} = 1.02 \times 10^3 \text{ L} \tag{7.2.12}$$

It would require 1020 L (269 gal) of gaseous methane at about 1 atm of pressure to replace 1 gal of gasoline. It requires a large container to hold enough methane at 1 atm to replace several gallons of gasoline.

Exercise 7.2.5

Calculate the pressure in bar of 2520 moles of hydrogen gas stored at 27 °C in the 180-L storage tank of a modern hydrogen-powered car.

Answer

350 bar

If the number of moles of an ideal gas are kept constant under two different sets of conditions, a useful mathematical relationship called the combined gas law is obtained: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ using units of atm, L, and K. Both sets of conditions are equal to the product of $n \times R$ (where n = the number of moles of the gas and R is the ideal gas law constant).

Example 7.2.6: Using the Combined Gas Law

When filled with air, a typical scuba tank with a volume of 13.2 L has a pressure of 153 atm (Figure **7.2.9**). If the water temperature is 27 °C, how many liters of air will such a tank provide to a diver's lungs at a depth of approximately 70 feet in the ocean where the pressure is 3.13 atm?



Figure 7.2.9: *Scuba divers use compressed air to breathe while underwater. (credit: modification of work by Mark Goodchild)*

Letting *1* represent the air in the scuba tank and *2* represent the air in the lungs, and noting that body temperature (the temperature the air will be in the lungs) is 37 °C, we have:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \longrightarrow \frac{(153\ \text{atm})(13.2\ \text{L})}{(300\ \text{K})} = \frac{(3.13\ \text{atm})(V_2)}{(310\ \text{K})} \tag{7.2.13}$$

Solving for V_2:

$$V_2 = \frac{(153\ \cancel{\text{atm}})(13.2\ \text{L})(310\ \cancel{\text{K}})}{(300\ \cancel{\text{K}})(3.13\ \cancel{\text{atm}})} = 667\ \text{L} \tag{7.2.14}$$

(Note: Be advised that this particular example is one in which the assumption of ideal gas behavior is not very reasonable, since it involves gases at relatively high pressures and low temperatures. Despite this limitation, the calculated volume can be viewed as a good "ballpark" estimate.)

Exercise 7.2.6

A sample of ammonia is found to occupy 0.250 L under laboratory conditions of 27 °C and 0.850 atm. Find the volume of this sample at 0 °C and 1.00 atm.

Answer

0.193 L

chemistry in everyday life: The Interdependence between Ocean Depth and Pressure in Scuba Diving

Whether scuba diving at the Great Barrier Reef in Australia (shown in Figure **7.2.10**) or in the Caribbean, divers must understand how pressure affects a number of issues related to their comfort and safety.



Figure* 7.2.10*: *Scuba divers, whether at the Great Barrier Reef or in the Caribbean, must be aware of buoyancy, pressure equalization, and the amount of time they spend underwater, to avoid the risks associated with pressurized gases in the body. (credit: Kyle Taylor).*

Pressure increases with ocean depth, and the pressure changes most rapidly as divers reach the surface. The pressure a diver experiences is the sum of all pressures above the diver (from the water and the air). Most pressure measurements are given in units of atmospheres, expressed as “atmospheres absolute” or ATA in the diving community: Every 33 feet of salt water represents 1 ATA of pressure in addition to 1 ATA of pressure from the atmosphere at sea level. As a diver descends, the increase in pressure causes the body’s air pockets in the ears and lungs to compress; on the ascent, the decrease in pressure causes these air pockets to expand, potentially rupturing eardrums or bursting the lungs. Divers must therefore undergo equalization by adding air to body airspaces on the descent by breathing normally and adding air to the mask by breathing out of the nose or adding air to the ears and sinuses by equalization techniques; the corollary is also true on ascent, divers must release air from the body to maintain equalization. Buoyancy, or the ability to control whether a diver sinks or floats, is controlled by the buoyancy compensator (BCD). If a diver is ascending, the air in his BCD expands because of lower pressure according to Boyle’s law (decreasing the pressure of gases increases the volume). The expanding air increases the buoyancy of the diver, and she or he begins to ascend. The diver must vent air from the BCD or risk an uncontrolled ascent that could rupture the lungs. In descending, the increased pressure causes the air in the BCD to compress and the diver sinks much more quickly; the diver must add air to the BCD or risk an uncontrolled descent, facing much higher pressures near the ocean floor. The pressure also impacts how long a diver can stay underwater before ascending. The deeper a diver dives, the more compressed the air that is breathed because of increased pressure: If a diver dives 33 feet, the pressure is 2 ATA and the air would be compressed to one-half of its original volume. The diver uses up available air twice as fast as at the surface.

Standard Conditions of Temperature and Pressure

We have seen that the volume of a given quantity of gas and the number of molecules (moles) in a given volume of gas vary with changes in pressure and temperature. Chemists sometimes make comparisons against a standard temperature and pressure (STP) for reporting properties of gases: 273.15 K and 1 atm (101.325 kPa). At STP, an ideal gas has a volume of about 22.4 L—this is referred to as the standard molar volume (Figure 7.2.11).

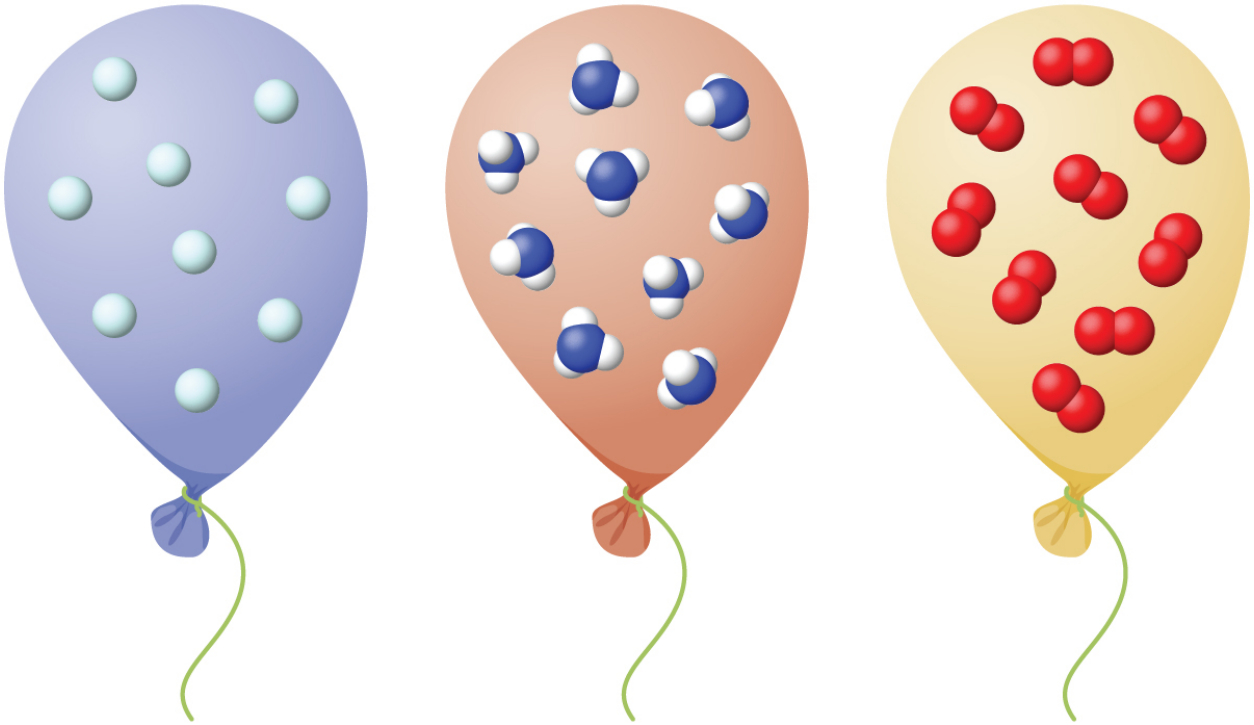


***Figure* 7.2.11**: *Since the number of moles in a given volume of gas varies with pressure and temperature changes, chemists use standard temperature and pressure (273.15 K and 1 atm or 101.325 kPa) to report properties of gases.*

Summary



***Video* 7.2.2**: *An overview of the ideal gas law and STP and a demonstration of the ideal gas law.*

The behavior of gases can be described by several laws based on experimental observations of their properties. The pressure of a given amount of gas is directly proportional to its absolute temperature, provided that the volume does not change (Amontons's law). The volume of a given gas sample is directly proportional to its absolute temperature at constant pressure (Charles's law). The volume of a given amount of gas is inversely proportional to its pressure when temperature is held constant (Boyle's law). Under the same conditions of temperature and pressure, equal volumes of all gases contain the same number of molecules (Avogadro's law).

The equations describing these laws are special cases of the ideal gas law, $PV = nRT$, where P is the pressure of the gas, V is its volume, n is the number of moles of the gas, T is its kelvin temperature, and R is the ideal (universal) gas constant.

Key Equations

- $PV = nRT$

Glossary

absolute zero
temperature at which the volume of a gas would be zero according to Charles's law.

Amontons's law
(also, Gay-Lussac's law) pressure of a given number of moles of gas is directly proportional to its kelvin temperature when the volume is held constant

Avogadro's law
volume of a gas at constant temperature and pressure is proportional to the number of gas molecules

Boyle's law
volume of a given number of moles of gas held at constant temperature is inversely proportional to the pressure under which it is measured

Charles's law
volume of a given number of moles of gas is directly proportional to its kelvin temperature when the pressure is held constant

ideal gas
hypothetical gas whose physical properties are perfectly described by the gas laws

ideal gas constant (R)
constant derived from the ideal gas equation $R = 0.08226$ L atm mol^{-1} K^{-1} or 8.314 L kPa mol^{-1} K^{-1}

ideal gas law
relation between the pressure, volume, amount, and temperature of a gas under conditions derived by combination of the simple gas laws

standard conditions of temperature and pressure (STP)
273.15 K (0 °C) and 1 atm (101.325 kPa)

standard molar volume
volume of 1 mole of gas at STP, approximately 22.4 L for gases behaving ideally

Contributors

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7.2.1: Practice Problems- The Gas Laws

PROBLEM 7.2.1.1

Sometimes leaving a bicycle in the sun on a hot day will cause a blowout. Why?

Answer

As temperature of a gas increases, pressure will also increase based on the ideal gas law. The volume of the tire can only expand so much before the rubber gives and releases the build up of pressure.

PROBLEM 7.2.1.2

Explain how the volume of the bubbles exhausted by a scuba diver change as they rise to the surface, assuming that they remain intact.

Answer

As the bubbles rise, the pressure decreases, so their volume increases as suggested by Boyle's law.

PROBLEM 7.2.1.3

One way to state Boyle's law is "All other things being equal, the pressure of a gas is inversely proportional to its volume."

(a) What is the meaning of the term "inversely proportional?"

(b) What are the "other things" that must be equal?

Answer a

The pressure of the gas *increases* as the volume *decreases*

Answer b

amount of gas, temperature

PROBLEM 7.2.1.4

An alternate way to state Avogadro's law is "All other things being equal, the number of molecules in a gas is directly proportional to the volume of the gas."

a. What is the meaning of the term "directly proportional?"
b. What are the "other things" that must be equal?

Answer a

The number of particles in the gas *increases* as the volume *increases*

Answer b

temperature, pressure

PROBLEM 7.2.1.5

A spray can is used until it is empty except for the propellant gas, which has a pressure of 1344 torr at 23 °C. If the can is thrown into a fire (T = 475 °C), what will be the pressure in the hot can?

Answer

3.40×10^3 torr

Click here to see a video solution





PROBLEM 7.2.1.6

What is the temperature of a[illegible]t 55 °C and 744 torr?

Answer

276°K

3°C

PROBLEM 7.2.1.7

A 2.50-L volume of hydrogen measured at –196 °C is warmed to 100 °C. Calculate the volume of the gas at the higher temperature, assuming no change in pressure.

Answer

12.1 L

Click here to see a video solution



PROBLEM 7.2.1.8

A balloon inflated with three breaths of air has a volume of 1.7 L. At the same temperature and pressure, what is the volume of the balloon if five more same-sized breaths are added to the balloon?

Answer

4.5 L

PROBLEM 7.2.1.9

A weather balloon contains 8.80 moles of helium at a pressure of 0.992 atm and a temperature of 25 °C at ground level. What is the volume of the balloon under these conditions?

Answer

217 L

Click here to see a video solution





PROBLEM 7.2.1.10

How many grams of gas are

a. 0.100 L of CO_2 at
b. 8.75 L of C_2H_4, at
c. 221 mL of Ar at 0.

Answer a

7.24×10^{-2} g

Answer b

23.1 g

Answer c

1.5×10^{-4} g

PROBLEM 7.2.1.11

A high altitude balloon is filled with 1.41×10^4 L of hydrogen at a temperature of 21 °C and a pressure of 745 torr. What is the volume of the balloon at a height of 20 km, where the temperature is –48 °C and the pressure is 63.1 torr?

Answer

1.2741×10^5 L or more correctly to 3 significant figures 1.27×10^5 L

Click here to see a video solution



PROBLEM 7.2.1.12

While resting, the average 70-kg human male consumes 14 L of pure O_2 per hour at 25 °C and 100 kPa. How many moles of O_2 are consumed by a 70 kg man while resting for 1.0 h?

Answer

0.565 mol

PROBLEM 7.2.1.13

A balloon that is 100.21 L at 21 °C and 0.981 atm is released and just barely clears the top of Mount Crumpet in British Columbia. If the final volume of the balloon is 144.53 L at a temperature of 5.24 °C, what is the pressure experienced by the

balloon as it clears Mount Crumpet?

Answer

0.644 atm

Click here to see a video solution



PROBLEM 7.2.1.14

If the temperature of a fixed amount of a gas is doubled at constant volume, what happens to the pressure?

Answer

Temperature and Pressure are directly proportional. Pressure will also have to **increase (doubling).**

PROBLEM 7.2.1.15

If the volume of a fixed amount of a gas is tripled at constant temperature, what happens to the pressure?

Answer

Volume and pressure are inversely proportional. The pressure **decreases** by a factor of 3.

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7.3: Applications of the Ideal Gas Law and Partial Pressures

Skills to Develop

- Use the ideal gas law to compute gas densities and molar masses
- Perform stoichiometric calculations involving gaseous substances
- State Dalton's law of partial pressures and use it in calculations involving gaseous mixtures

The study of the chemical behavior of gases was part of the basis of perhaps the most fundamental chemical revolution in history. French nobleman Antoine Lavoisier, widely regarded as the "father of modern chemistry," changed chemistry from a qualitative to a quantitative science through his work with gases. He discovered the law of conservation of matter, discovered the role of oxygen in combustion reactions, determined the composition of air, explained respiration in terms of chemical reactions, and more. He was a casualty of the French Revolution, guillotined in 1794. Of his death, mathematician and astronomer Joseph-Louis Lagrange said, "It took the mob only a moment to remove his head; a century will not suffice to reproduce it."

As described in an earlier chapter of this text, we can turn to chemical stoichiometry for answers to many of the questions that ask "How much?" We can answer the question with masses of substances or volumes of solutions. However, we can also answer this question another way: with volumes of gases. We can use the ideal gas equation to relate the pressure, volume, temperature, and number of moles of a gas. Here we will combine the ideal gas equation with other equations to find gas density and molar mass. We will deal with mixtures of different gases, and calculate amounts of substances in reactions involving gases. This section will not introduce any new material or ideas, but will provide examples of applications and ways to integrate concepts we have already discussed.

Density of a Gas

Recall that the density of a gas is its mass to volume ratio, $\rho = \frac{m}{V}$. Therefore, if we can determine the mass of some volume of a gas, we will get its density. The density of an unknown gas can used to determine its molar mass and thereby assist in its identification. The ideal gas law, $PV = nRT$, provides us with a means of deriving such a mathematical formula to relate the density of a gas to its volume in the proof shown in Example 7.3.1.

Example 7.3.1: Derivation of a Density Formula from the Ideal Gas Law

Use $PV = nRT$ to derive a formula for the density of gas in g/L

Solution

$$PV = nRT \tag{7.3.1}$$

Rearrange to get (mol/L):

$$\frac{n}{v} = \frac{P}{RT} \tag{7.3.2}$$

Multiply each side of the equation by the molar mass, $\mathcal{M}$. When moles are multiplied by $\mathcal{M}$ in g/mol, g are obtained:

$$(\mathcal{M})\left(\frac{n}{V}\right) = \left(\frac{P}{RT}\right)(\mathcal{M}) \tag{7.3.3}$$

$$\mathcal{M}/V = \rho = \frac{P\mathcal{M}}{RT} \tag{7.3.4}$$

Exercise 7.3.1

A gas was found to have a density of 0.0847 g/L at 17.0 °C and a pressure of 760 torr. What is its molar mass? What is the gas?

Answer

$$\rho = \frac{P\mathcal{M}}{RT}$$

$$0.0847\ \text{g/L} = 760\ \cancel{\text{torr}} \times \frac{1\ \cancel{\text{atm}}}{760\ \cancel{\text{torr}}} \times \frac{\mathcal{M}}{0.0821\ \text{L}\ \cancel{\text{atm}}/\text{mol K}} \times 290\ \text{K} \tag{7.3.5}$$

$\mathcal{M}$ = 2.02 g/mol; therefore, the gas must be hydrogen (H_2, 2.02 g/mol)

We must specify both the temperature and the pressure of a gas when calculating its density because the number of moles of a gas (and thus the mass of the gas) in a liter changes with temperature or pressure. Gas densities are often reported at STP.

Example 7.3.2: Empirical/Molecular Formula Problems Using the Ideal Gas Law and Density of a Gas

Cyclopropane, a gas once used with oxygen as a general anesthetic, is composed of 85.7% carbon and 14.3% hydrogen by mass. Find the empirical formula. If 1.56 g of cyclopropane occupies a volume of 1.00 L at 0.984 atm and 50 °C, what is the molecular formula for cyclopropane?

Solution

Strategy:

First solve the empirical formula problem using methods discussed earlier. Assume 100 g and convert the percentage of each element into grams. Determine the number of moles of carbon and hydrogen in the 100-g sample of cyclopropane. Divide by the smallest number of moles to relate the number of moles of carbon to the number of moles of hydrogen. In the last step, realize that the smallest whole number ratio is the empirical formula:

$$85.7\ \text{g C} \times \frac{1\ \text{mol C}}{12.01\ \text{g C}} = 7.136\ \text{mol C} \quad \frac{7.136}{7.136} = 1.00\ \text{mol C} \tag{7.3.6}$$

$$14.3\ \text{g H} \times \frac{1\ \text{mol H}}{1.01\ \text{g H}} = 14.158\ \text{mol H} \quad \frac{14.158}{7.136} = 1.98\ \text{mol H} \tag{7.3.7}$$

Empirical formula is CH_2 [empirical mass (EM) of 14.03 g/empirical unit].

Next, use the density equation related to the ideal gas law to determine the molar mass:

$$d = \frac{P\mathcal{M}}{RT} \quad \frac{1.56\ \text{g}}{1.00\ \text{L}} = 0.984\ \text{atm} \times \frac{\mathcal{M}}{0.0821\ \text{L atm/mol K}} \times 323\ \text{K} \tag{7.3.8}$$

$\mathcal{M}$ = 42.0 g/mol, $\frac{\mathcal{M}}{EM} = \frac{42.0}{14.03} = 2.99$, so (3)($CH_2$) = C_3H_6 (molecular formula)

Exercise 7.3.2

Acetylene, a fuel used welding torches, is comprised of 92.3% C and 7.7% H by mass. Find the empirical formula. If 1.10 g of acetylene occupies of volume of 1.00 L at 1.15 atm and 59.5 °C, what is the molecular formula for acetylene?

Answer

Empirical formula, CH; Molecular formula, C_2H_2

Molar Mass of a Gas

Another useful application of the ideal gas law involves the determination of molar mass. By definition, the molar mass of a substance is the ratio of its mass in grams, *m*, to its amount in moles, *n*:

$$\mathcal{M} = \frac{\text{grams of substance}}{\text{moles of substance}} = \frac{m}{n} \tag{7.3.9}$$

The ideal gas equation can be rearranged to isolate n:

$$n = \frac{PV}{RT} \tag{7.3.10}$$

and then combined with the molar mass equation to yield:

$$\mathcal{M} = \frac{mRT}{PV} \tag{7.3.11}$$

This equation can be used to derive the molar mass of a gas from measurements of its pressure, volume, temperature, and mass.

Example 7.3.3: Determining the Molar Mass of a Volatile Liquid

The approximate molar mass of a volatile liquid can be determined by:

1. Heating a sample of the liquid in a flask with a tiny hole at the top, which converts the liquid into gas that may escape through the hole
2. Removing the flask from heat at the instant when the last bit of liquid becomes gas, at which time the flask will be filled with only gaseous sample at ambient pressure
3. Sealing the flask and permitting the gaseous sample to condense to liquid, and then weighing the flask to determine the sample's mass (Figure **7.3.1**)

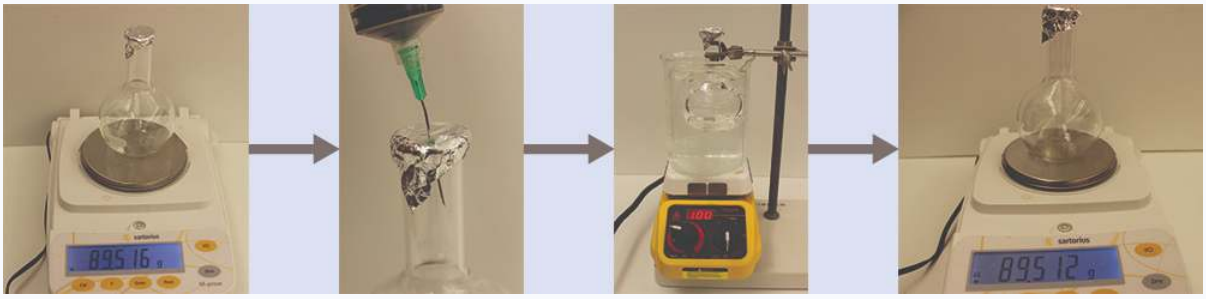


Figure* 7.3.1*: *When the volatile liquid in the flask is heated past its boiling point, it becomes gas and drives air out of the flask. At $t_{l \longrightarrow g}$, the flask is filled with volatile liquid gas at the same pressure as the atmosphere. If the flask is then cooled to room temperature, the gas condenses and the mass of the gas that filled the flask, and is now liquid, can be measured. (credit: modification of work by Mark Ott)*

Using this procedure, a sample of chloroform gas weighing 0.494 g is collected in a flask with a volume of 129 cm^3 at 99.6 °C when the atmospheric pressure is 742.1 mm Hg. What is the approximate molar mass of chloroform?

Solution

Since

$$\mathcal{M} = \frac{m}{n}$$

and

$$n = \frac{PV}{RT}$$

substituting and rearranging gives

$$\mathcal{M} = \frac{mRT}{PV} \tag{7.3.12}$$

then

$$\mathcal{M} = \frac{mRT}{PV} = \frac{(0.494\ \text{g}) \times 0.08206\ \text{L} \cdot \text{atm/mol K} \times 372.8\ \text{K}}{0.976\ \text{atm} \times 0.129\ \text{L}} = 120\ \text{g/mol} \tag{7.3.13}$$

Exercise 7.3.3

A sample of phosphorus that weighs 3.243×10^{-2} g exerts a pressure of 31.89 kPa in a 56.0-mL bulb at 550 °C. What are the molar mass and molecular formula of phosphorus vapor?

Answer

124 g/mol P_4

The Pressure of a Mixture of Gases: Dalton's Law

Unless they chemically react with each other, the individual gases in a mixture of gases do not affect each other's pressure. Each individual gas in a mixture exerts the same pressure that it would exert if it were present alone in the container (Figure **7.3.2**). The pressure exerted by each individual gas in a mixture is called its partial pressure. This observation is summarized by Dalton's law of partial pressures: *The total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of the component gases*:

$$P_{Total} = P_A + P_B + P_C + ... = \sum_i P_i \tag{7.3.14}$$

In the equation P_{Total} is the total pressure of a mixture of gases, P_A is the partial pressure of gas A; P_B is the partial pressure of gas B; P_C is the partial pressure of gas C; and so on.

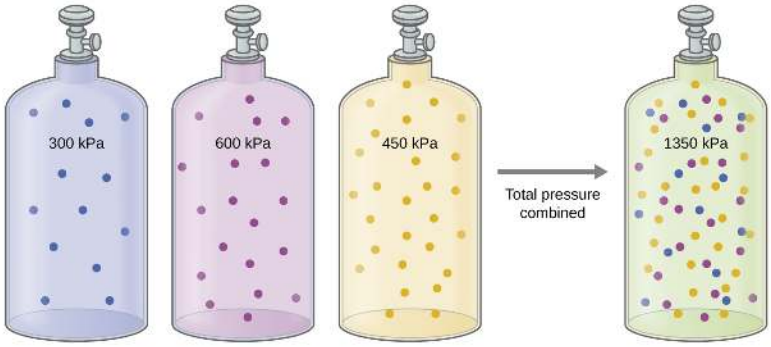


Figure 7.3.2: *If equal-volume cylinders containing gas A at a pressure of 300 kPa, gas B at a pressure of 600 kPa, and gas C at a pressure of 450 kPa are all combined in the same-size cylinder, the total pressure of the mixture is 1350 kPa.*

The partial pressure of gas A is related to the total pressure of the gas mixture via its mole fraction (X), a unit of concentration defined as the number of moles of a component of a solution divided by the total number of moles of all components:

$$P_A = X_A \times P_{Total} \quad \text{where} \quad X_A = \frac{n_A}{n_{Total}} \tag{7.3.15}$$

where P_A, X_A, and n_A are the partial pressure, mole fraction, and number of moles of gas A, respectively, and n_{Total} is the number of moles of all components in the mixture.

Example 7.3.4: The Pressure of a Mixture of Gases

A 10.0-L vessel contains 2.50×10^{-3} mol of H_2, 1.00×10^{-3} mol of He, and 3.00×10^{-4} mol of Ne at 35 °C.

a. What are the partial pressures of each of the gases?
b. What is the total pressure in atmospheres?

Solution

The gases behave independently, so the partial pressure of each gas can be determined from the ideal gas equation, using $P = \frac{nRT}{V}$:

$$P_{H_2} = \frac{(2.50 \times 10^{-3}\ \text{mol})(0.08206\ \text{L atm mol}^{-1}\text{K}^{-1})(308\ \text{K})}{10.0\ \text{L}} = 6.32 \times 10^{-3}\ \text{atm} \tag{7.3.16}$$

$$P_{He} = \frac{(1.00 \times 10^{-3}\ \text{mol})(0.08206\ \text{L atm mol}^{-1}\text{K}^{-1})(308\ \text{K})}{10.0\ \text{L}} = 2.53 \times 10^{-3}\ \text{atm} \tag{7.3.17}$$

$$P_{Ne} = \frac{(3.00 \times 10^{-4} \cancel{mol})(0.08206 \cancel{L}\ \cancel{atm\ mol^{-1}\ K^{-1}})(308 \cancel{K})}{10.0 \cancel{L}} = 7.58 \times 10^{-4}\ atm \tag{7.3.18}$$

The total pressure is given by the sum of the partial pressures:

$$P_T = P_{H_2} + P_{He} + P_{Ne} = (0.00632 + 0.00253 + 0.00076)\ atm = 9.61 \times 10^{-3}\ atm \tag{7.3.19}$$

Exercise 7.3.4

A 5.73-L flask at 25 °C contains 0.0388 mol of N_2, 0.147 mol of CO, and 0.0803 mol of H_2. What is the total pressure in the flask in atmospheres?

Answer

1.137 atm

Here is another example of this concept, but dealing with mole fraction calculations.

Example 7.3.5: The Pressure of a Mixture of Gases

A gas mixture used for anesthesia contains 2.83 mol oxygen, O_2, and 8.41 mol nitrous oxide, N_2O. The total pressure of the mixture is 192 kPa.

a. What are the mole fractions of O_2 and N_2O?
b. What are the partial pressures of O_2 and N_2O?

Solution

The mole fraction is given by

$$X_A = \frac{n_A}{n_{Total}}$$

and the partial pressure is

$$P_A = X_A \times P_{Total}$$

For O_2,

$$X_{O_2} = \frac{n_{O_2}}{n_{Total}} = \frac{2.83 mol}{(2.83 + 8.41)\ mol} = 0.252$$

and

$$P_{O_2} = X_{O_2} \times P_{Total} = 0.252 \times 192\ kPa = 48.4\ kPa$$

For N_2O,

$$X_{N_2O} = \frac{n_{N_2O}}{n_{Total}} = \frac{8.41\ mol}{(2.83 + 8.41)\ mol} = 0.748$$

and

$$P_{N_2O} = X_{N_2O} \times P_{Total} = (0.748) \times 192\ kPa = 143.6\ kPa$$

Exercise 7.3.5

What is the pressure of a mixture of 0.200 g of H_2, 1.00 g of N_2, and 0.820 g of Ar in a container with a volume of 2.00 L at 20 °C?

Answer

1.87 atm

Collection of Gases over Water

A simple way to collect gases that do not react with water is to capture them in a bottle that has been filled with water and inverted into a dish filled with water. The pressure of the gas inside the bottle can be made equal to the air pressure outside by raising or lowering the bottle. When the water level is the same both inside and outside the bottle (Figure **7.3.3**), the pressure of the gas is equal to the atmospheric pressure, which can be measured with a barometer.

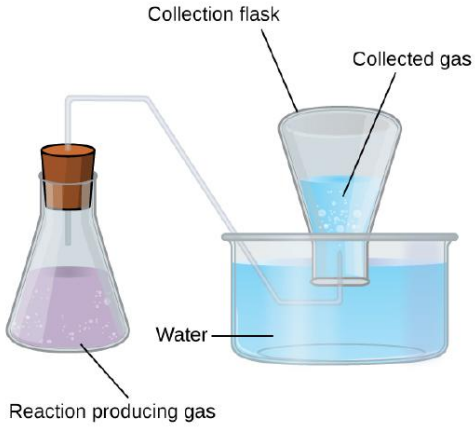


Figure* 7.3.3*: *When a reaction produces a gas that is collected above water, the trapped gas is a mixture of the gas produced by the reaction and water vapor. If the collection flask is appropriately positioned to equalize the water levels both within and outside the flask, the pressure of the trapped gas mixture will equal the atmospheric pressure outside the flask (see the earlier discussion of manometers).*

However, there is another factor we must consider when we measure the pressure of the gas by this method. Water evaporates and there is always gaseous water (water vapor) above a sample of liquid water. As a gas is collected over water, it becomes saturated with water vapor and the total pressure of the mixture equals the partial pressure of the gas plus the partial pressure of the water vapor. The pressure of the pure gas is therefore equal to the total pressure minus the pressure of the water vapor—this is referred to as the "dry" gas pressure, that is, the pressure of the gas only, without water vapor.

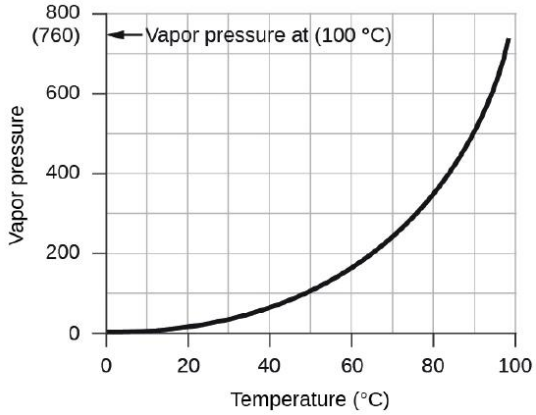


Figure* 7.3.4*: *This graph shows the vapor pressure of water at sea level as a function of temperature.*

The vapor pressure of water, which is the pressure exerted by water vapor in equilibrium with liquid water in a closed container, depends on the temperature (Figure **7.3.4**); more detailed information on the temperature dependence of water vapor can be found in Table **7.3.1**, and vapor pressure will be discussed in more detail in the next chapter on liquids.

Table **7.3.1**: *Vapor Pressure of Ice and Water in Various Temperatures at Sea Level*

Temperature (°C)	Pressure (torr)	Temperature (°C)	Pressure (torr)	Temperature (°C)	Pressure (torr)
–10	1.95	18	15.5	30	31.8
–5	3.0	19	16.5	35	42.2
–2	3.9	20	17.5	40	55.3
0	4.6	21	18.7	50	92.5
2	5.3	22	19.8	60	149.4
4	6.1	23	21.1	70	233.7
6	7.0	24	22.4	80	355.1
8	8.0	25	23.8	90	525.8
10	9.2	26	25.2	95	633.9
12	10.5	27	26.7	99	733.2
14	12.0	28	28.3	100.0	760.0
16	13.6	29	30.0	101.0	787.6

Example 7.3.6: Pressure of a Gas Collected Over Water

If 0.200 L of argon is collected over water at a temperature of 26 °C and a pressure of 750 torr in a system like that shown in Figure **7.3.3**, what is the partial pressure of argon?

Solution

According to Dalton's law, the total pressure in the bottle (750 torr) is the sum of the partial pressure of argon and the partial pressure of gaseous water:

$$P_T = P_{Ar} + P_{H_2O} \tag{7.3.20}$$

Rearranging this equation to solve for the pressure of argon gives:

$$P_{Ar} = P_T - P_{H_2O} \tag{7.3.21}$$

The pressure of water vapor above a sample of liquid water at 26 °C is 25.2 torr (Appendix E), so:

$$P_{Ar} = 750 \text{ torr} - 25.2 \text{ torr} = 725 \text{ torr} \tag{7.3.22}$$

Exercise 7.3.6

A sample of oxygen collected over water at a temperature of 29.0 °C and a pressure of 764 torr has a volume of 0.560 L. What volume would the dry oxygen have under the same conditions of temperature and pressure?

Answer

0.583 L

Chemical Stoichiometry and Gases

Chemical stoichiometry describes the quantitative relationships between reactants and products in chemical reactions. We have previously measured quantities of reactants and products using masses for solids and volumes in conjunction with the molarity for solutions; now we can also use gas volumes to indicate quantities. If we know the volume, pressure, and temperature of a gas, we can use the ideal gas equation to calculate how many moles of the gas are present. If we know how many moles of a gas are involved, we can calculate the volume of a gas at any temperature and pressure.

Avogadro's Law Revisited

Sometimes we can take advantage of a simplifying feature of the stoichiometry of gases that solids and solutions do not exhibit: All gases that show ideal behavior contain the same number of molecules in the same volume (at the same temperature and pressure). Thus, the ratios of volumes of gases involved in a chemical reaction are given by the coefficients in the equation for the reaction, provided that the gas volumes are measured at the same temperature and pressure.

We can extend Avogadro's law (that the volume of a gas is directly proportional to the number of moles of the gas) to chemical reactions with gases: Gases combine, or react, in definite and simple proportions by volume, provided that all gas volumes are measured at the same temperature and pressure. For example, since nitrogen and hydrogen gases react to produce ammonia gas according to

$$\mathrm{N_2}(g) + 3\,\mathrm{H_2}(g) \longrightarrow 2\,\mathrm{NH_3}(g) \tag{7.3.23}$$

a given volume of nitrogen gas reacts with three times that volume of hydrogen gas to produce two times that volume of ammonia gas, if pressure and temperature remain constant.

The explanation for this is illustrated in Figure **7.3.4**. According to Avogadro's law, equal volumes of gaseous N_2, H_2, and NH_3, at the same temperature and pressure, contain the same number of molecules. Because one molecule of N_2 reacts with three molecules of H_2 to produce two molecules of NH_3, the volume of H_2 required is three times the volume of N_2, and the volume of NH_3 produced is two times the volume of N_2.

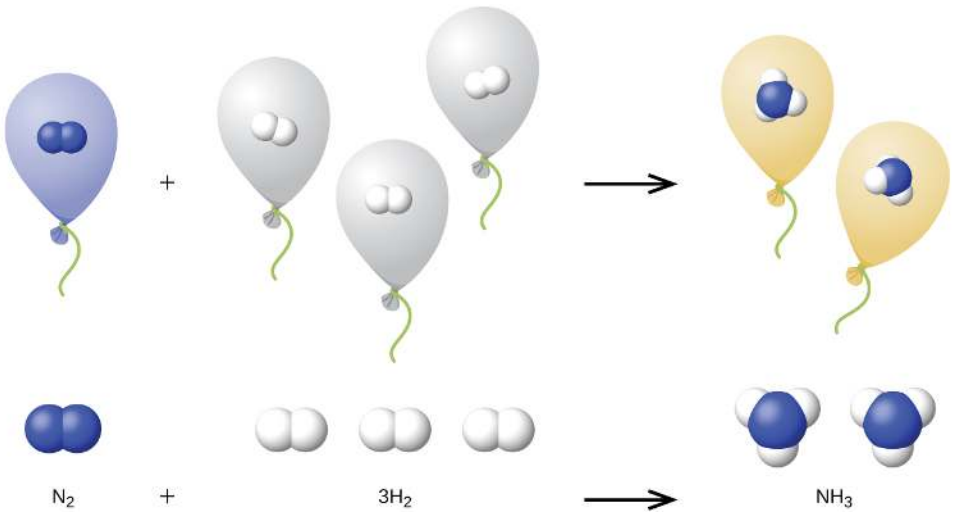


***Figure* 7.3.5**: *One volume of N_2 combines with three volumes of H_2 to form two volumes of NH_3.*

Example 7.3.7: Reaction of Gases

Propane, $C_3H_8(g)$, is used in gas grills to provide the heat for cooking. What volume of $O_2(g)$ measured at 25 °C and 760 torr is required to react with 2.7 L of propane measured under the same conditions of temperature and pressure? Assume that the propane undergoes complete combustion.

Solution

The ratio of the volumes of C_3H_8 and O_2 will be equal to the ratio of their coefficients in the balanced equation for the reaction:

$$\begin{array}{ll} \mathrm{C_3H_8}(g) + 5\,\mathrm{O_2}(g) \longrightarrow & 3\,\mathrm{CO_2}(g) + 4\,\mathrm{H_2O}(l) \quad (7.3.24) \\ 1\ \text{volume} + 5\ \text{volumes} & 3\ \text{volumes} + 4\ \text{volumes} \quad (7.3.25) \end{array}$$

From the equation, we see that one volume of C_3H_8 will react with five volumes of O_2:

$$2.7\ \mathrm{L}\ \mathrm{C_3H_8} \times \frac{5\ \mathrm{L}\ \mathrm{O_2}}{1\ \mathrm{L}\ \mathrm{C_3H_8}} = 13.5\ \mathrm{L}\ \mathrm{O_2} \quad (7.3.26)$$

A volume of 13.5 L of O_2 will be required to react with 2.7 L of C_3H_8.

Exercise 7.3.7

An acetylene tank for an oxyacetylene welding torch provides 9340 L of acetylene gas, C_2H_2, at 0 °C and 1 atm. How many tanks of oxygen, each providing 7.00×10^3 L of O_2 at 0 °C and 1 atm, will be required to burn the acetylene?

$$2\,\mathrm{C_2H_2} + 5\,\mathrm{O_2} \longrightarrow 4\,\mathrm{CO_2} + 2\,\mathrm{H_2O}$$

Answer

3.34 tanks (2.34×10^4 L)

Example 7.3.8: Volumes of Reacting Gases

Ammonia is an important fertilizer and industrial chemical. Suppose that a volume of 683 billion cubic feet of gaseous ammonia, measured at 25 °C and 1 atm, was manufactured. What volume of $H_2(g)$, measured under the same conditions, was required to prepare this amount of ammonia by reaction with N_2?

$$\mathrm{N_2}(g) + 3\,\mathrm{H_2}(g) \longrightarrow 2\,\mathrm{NH_3}(g)$$

Solution

Because equal volumes of H_2 and NH_3 contain equal numbers of molecules and each three molecules of H_2 that react produce two molecules of NH_3, the ratio of the volumes of H_2 and NH_3 will be equal to 3:2. Two volumes of NH_3, in this case in units of billion ft^3, will be formed from three volumes of H_2:

$$683\ \text{billion ft}^3\ \mathrm{NH_3} \times \frac{3\ \text{billion ft}^3\ \mathrm{H_2}}{2\ \text{billion ft}^3\ \mathrm{NH_3}} = 1.02 \times 10^3\ \text{billion ft}^3\ \mathrm{H_2} \quad (7.3.27)$$

The manufacture of 683 billion ft^3 of NH_3 required 1020 billion ft^3 of H_2. (At 25 °C and 1 atm, this is the volume of a cube with an edge length of approximately 1.9 miles.)

Exercise 7.3.8

What volume of $O_2(g)$ measured at 25 °C and 760 torr is required to react with 17.0 L of ethylene, $C_2H_4(g)$, measured under the same conditions of temperature and pressure? The products are CO_2 and water vapor.

Answer

51.0 L

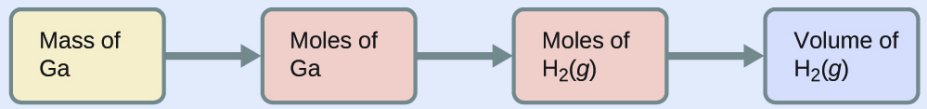
Example 7.3.9: Volume of Gaseous Product

What volume of hydrogen at 27 °C and 723 torr may be prepared by the reaction of 8.88 g of gallium with an excess of hydrochloric acid?

$$2\,\mathrm{Ga}(s) + 6\,\mathrm{HCl}(aq) \longrightarrow 2\,\mathrm{GaCl_3}(aq) + 3\,\mathrm{H_2}(g) \quad (7.3.28)$$

Solution

To convert from the mass of gallium to the volume of $H_2(g)$, we need to do something like this:



The first two conversions are:

$$8.88\ \cancel{\text{g Ga}} \times \frac{1\ \cancel{\text{mol Ga}}}{69.723\ \cancel{\text{g Ga}}} \times \frac{3\ \text{mol H}_2}{2\ \cancel{\text{mol Ga}}} = 0.191\ \text{mol H}_2 \tag{7.3.29}$$

Finally, we can use the ideal gas law:

$$V_{\text{H}_2} = \left(\frac{nRT}{P}\right)_{\text{H}_2} = \frac{0.191\ \cancel{\text{mol}} \times 0.08206\ \text{L}\ \cancel{\text{atm mol}^{-1}\ \text{K}^{-1}} \times 300\ \text{K}}{0.951\ \text{atm}} = 4.94\ \text{L} \tag{7.3.30}$$

Exercise 7.3.9

Sulfur dioxide is an intermediate in the preparation of sulfuric acid. What volume of SO_2 at 343 °C and 1.21 atm is produced by burning l.00 kg of sulfur in oxygen?

Answer

1.30×10^3 L

Greenhouse Gases and Climate Change

The thin skin of our atmosphere keeps the earth from being an ice planet and makes it habitable. In fact, this is due to less than 0.5% of the air molecules. Of the energy from the sun that reaches the earth, almost $\frac{1}{3}$ is reflected back into space, with the rest absorbed by the atmosphere and the surface of the earth. Some of the energy that the earth absorbs is re-emitted as infrared (IR) radiation, a portion of which passes back out through the atmosphere into space. However, most of this IR radiation is absorbed by certain substances in the atmosphere, known as greenhouse gases, which re-emit this energy in all directions, trapping some of the heat. This maintains favorable living conditions—without atmosphere, the average global average temperature of 14 °C (57 °F) would be about –19 °C (–2 °F). The major greenhouse gases (GHGs) are water vapor, carbon dioxide, methane, and ozone. Since the Industrial Revolution, human activity has been increasing the concentrations of GHGs, which have changed the energy balance and are significantly altering the earth's climate (Figure **7.3.6**).

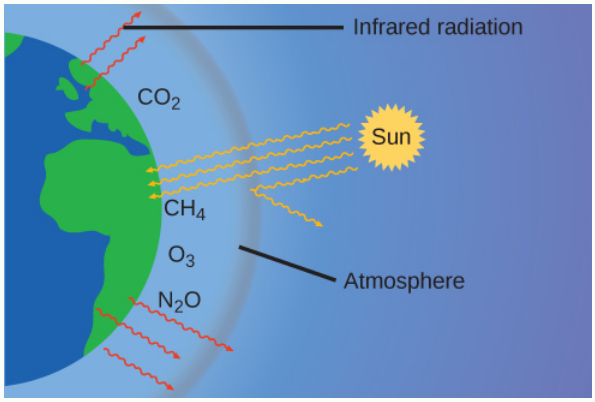


***Figure* 7.3.6**: Greenhouse gases trap enough of the sun's energy to make the planet habitable—this is known as the greenhouse effect. Human activities are increasing greenhouse gas levels, warming the planet and causing more extreme weather events.

There is strong evidence from multiple sources that higher atmospheric levels of CO_2 are caused by human activity, with fossil fuel burning accounting for about $\frac{3}{4}$ of the recent increase in CO_2. Reliable data from ice cores reveals that CO_2 concentration in the atmosphere is at the highest level in the past 800,000 years; other evidence indicates that it may be at its highest level in 20 million years. In recent years, the CO_2 concentration has increased from historical levels of below 300 ppm to almost 400 ppm today (Figure **7.3.7**).

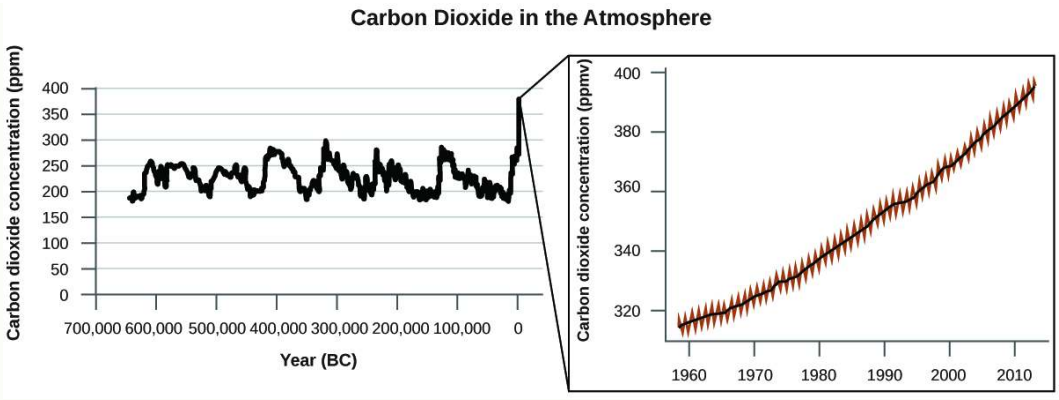


***Figure* 7.3.7**: CO_2 levels over the past 700,000 years were typically from 200–300 ppm, with a steep, unprecedented increase over the past 50 years.

Portrait of a Chemist: Susan Solomon

Atmospheric and climate scientist Susan Solomon (Figure **7.3.8**) is the author of one of *The New York Times*books of the year (*The Coldest March*, 2001), one of Time magazine's 100 most influential people in the world (2008), and a working group leader of the Intergovernmental Panel on Climate Change (IPCC), which was the recipient of the 2007 Nobel Peace Prize. She helped determine and explain the cause of the formation of the ozone hole over Antarctica, and has authored many important papers on climate change. She has been awarded the top scientific honors in the US and France (the National Medal of Science and the Grande Medaille, respectively), and is a member of the National Academy of Sciences, the Royal Society, the French Academy of Sciences, and the European Academy of Sciences. Formerly a professor at the University of Colorado, she is now at MIT, and continues to work at NOAA.

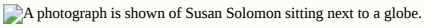


***Figure* 7.3.8**: Susan Solomon's research focuses on climate change and has been instrumental in determining the cause of the ozone hole over Antarctica. (credit: National Oceanic and Atmospheric Administration)

Summary



***Video* 7.3.1**: An overview of the topics of partial pressure, vapor pressure, and collecting gases over water.

The ideal gas law can be used to derive a number of convenient equations relating directly measured quantities to properties of interest for gaseous substances and mixtures. Appropriate rearrangement of the ideal gas equation may be made to permit the calculation of gas densities and molar masses. Dalton's law of partial pressures may be used to relate measured gas pressures for gaseous mixtures to their compositions. Avogadro's law may be used in stoichiometric computations for chemical reactions involving gaseous reactants or products.

Key Equations

- $P_{Total} = P_A + P_B + P_C + \ldots = \Sigma_i P_i$
- $P_A = X_A P_{Total}$
- $X_A = \frac{n_A}{n_{Total}}$

Footnotes

1. “Quotations by Joseph-Louis Lagrange,” last modified February 2006, accessed February 10, 2015, http://www-history.mcs.st-andrews.ac.../Lagrange.html

Glossary

Dalton’s law of partial pressures

total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of the component gases.

mole fraction (*X*)

concentration unit defined as the ratio of the molar amount of a mixture component to the total number of moles of all mixture components

partial pressure

pressure exerted by an individual gas in a mixture

vapor pressure of water

pressure exerted by water vapor in equilibrium with liquid water in a closed container at a specific temperature

Contributors

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7.3.1: Practice Problems- Applications of the Ideal Gas Law

PROBLEM 7.3.1.1

What is the density of laughing gas, dinitrogen monoxide, N_2O, at a temperature of 325 K and a pressure of 113.0 kPa?

Answer

1.84 g/L

PROBLEM 7.3.1.2

Calculate the density of Freon 12, CF_2Cl_2, at 30.0 °C and 0.954 atm.

Answer

4.64 g/L

Click here to see a video of the solution



PROBLEM 7.3.1.3

Which is denser at the same temperature and pressure, dry air or air saturated with water vapor? Explain.

Answer

air saturated with water vapor - it has a higher molar mass

PROBLEM 7.3.1.4

A cylinder of $O_2(g)$ used in breathing by emphysema patients has a volume of 3.00 L at a pressure of 10.0 atm. If the temperature of the cylinder is 28.0 °C, what mass of oxygen is in the cylinder?

Answer

38.8 g

Click here to see a video of the solution





PROBLEM 7.3.1.5

What is the molar mass of a 26 °C and a pressure of 307 torr?

Answer

30.0 g/mol

PROBLEM 7.3.1.6

What is the molar mass of a gas if 0.281 g of the gas occupies a volume of 125 mL at a temperature 126 °C and a pressure of 777 torr?

Answer

72.0 g/mol

Click here to see a video of the solution



PROBLEM 7.3.1.7

The density of a certain gaseous fluoride of phosphorus is 3.93 g/L at STP. Calculate the molar mass of this fluoride and determine its molecular formula.

Answer

88.1 g mol^{-1}; PF_3

PROBLEM 7.3.1.8

What is the molecular formula of a compound that contains 39% C, 45% N, and 16% H if 0.157 g of the compound occupies 125 mL with a pressure of 99.5 kPa at 22 °C?

Answer

H_5CN

Click here to see a video of the solution



PROBLEM 7.3.1.9

A sample of gas isolated from unrefined petroleum contains 90.0% CH_4, 8.9% C_2H_6, and 1.1% C_3H_8 at a total pressure of 307.2 kPa. What is the partial pressure of each component of this gas? (The percentages given indicate the percent of the total pressure that is due to each component.)

Answer

CH_4: 276 kPa; C_2H_6: 27 kPa; C_3H_8: 3.4 kPa

PROBLEM 7.3.1.10

Automobile air bags are inflated with nitrogen gas, which is formed by the decomposition of solid sodium azide (NaN_3). The other product is sodium metal. Calculate the volume of nitrogen gas at 27 °C and 756 torr formed by the decomposition of 125 g of sodium azide.

Answer

71.26 L

Click here to see a video of the solution



PROBLEM 7.3.1.11

A sample of a compound of xenon and fluorine was confined in a bulb with a pressure of 18 torr. Hydrogen was added to the bulb until the pressure was 72 torr. Passage of an electric spark through the mixture produced Xe and HF. After the HF was removed by reaction with solid KOH, the final pressure of xenon and unreacted hydrogen in the bulb was 36 torr. What is the empirical formula of the xenon fluoride in the original sample? (Note: Xenon fluorides contain only one xenon atom per molecule.)

Answer

XeF_2

Contributors

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- Adelaide Clark, Oregon Institute of Technology

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7.4: Kinetic Molecular Theory

Skills to Develop

- State the postulates of the kinetic-molecular theory
- Use this theory's postulates to explain the gas laws

The gas laws that we have seen to this point, as well as the ideal gas equation, are empirical, that is, they have been derived from experimental observations. The mathematical forms of these laws closely describe the macroscopic behavior of most gases at pressures less than about 1 or 2 atm. Although the gas laws describe relationships that have been verified by many experiments, they do not tell us why gases follow these relationships.

The kinetic molecular theory (KMT) is a simple microscopic model that effectively explains the gas laws described in previous modules of this chapter. This theory is based on the following five postulates described here. (Note: The term "molecule" will be used to refer to the individual chemical species that compose the gas, although some gases are composed of atomic species, for example, the noble gases.)

1. Gases are composed of molecules that are in continuous motion, travelling in straight lines and changing direction only when they collide with other molecules or with the walls of a container.
2. The molecules composing the gas are negligibly small compared to the distances between them.
3. The pressure exerted by a gas in a container results from collisions between the gas molecules and the container walls.
4. Gas molecules exert no attractive or repulsive forces on each other or the container walls; therefore, their collisions are *elastic* (do not involve a loss of energy).
5. The average kinetic energy of the gas molecules is proportional to the kelvin temperature of the gas.

The test of the KMT and its postulates is its ability to explain and describe the behavior of a gas. The various gas laws can be derived from the assumptions of the KMT, which have led chemists to believe that the assumptions of the theory accurately represent the properties of gas molecules. We will first look at the individual gas laws (Boyle's, Charles's, Amontons's, Avogadro's, and Dalton's laws) conceptually to see how the KMT explains them. Then, we will more carefully consider the relationships between molecular masses, speeds, and kinetic energies with temperature, and explain Graham's law.

The Kinetic-Molecular Theory Explains the Behavior of Gases, Part I

Recalling that gas pressure is exerted by rapidly moving gas molecules and depends directly on the number of molecules hitting a unit area of the wall per unit of time, we see that the KMT conceptually explains the behavior of a gas as follows:

- *Amontons's law.* If the temperature is increased, the average speed and kinetic energy of the gas molecules increase. If the volume is held constant, the increased speed of the gas molecules results in more frequent and more forceful collisions with the walls of the container, therefore increasing the pressure (Figure **7.4.1*a***).
- *Charles's law.* If the temperature of a gas is increased, a constant pressure may be maintained only if the volume occupied by the gas increases. This will result in greater average distances traveled by the molecules to reach the container walls, as well as increased wall surface area. These conditions will decrease both the frequency of molecule-wall collisions and the number of collisions per unit area, the combined effects of which balance the effect of increased collision forces due to the greater kinetic energy at the higher temperature.
- *Boyle's law.* If the gas volume is decreased, the container wall area decreases and the molecule-wall collision frequency increases, both of which increase the pressure exerted by the gas (Figure **7.4.1*b***).
- *Avogadro's law.* At constant pressure and temperature, the frequency and force of molecule-wall collisions are constant. Under such conditions, increasing the number of gaseous molecules will require a proportional increase in the container volume in order to yield a decrease in the number of collisions per unit area to compensate for the increased frequency of collisions (Figure **7.4.1*c***).
- *Dalton's Law.* Because of the large distances between them, the molecules of one gas in a mixture bombard the container walls with the same frequency whether other gases are present or not, and the total pressure of a gas mixture equals the sum of the (partial) pressures of the individual gases.

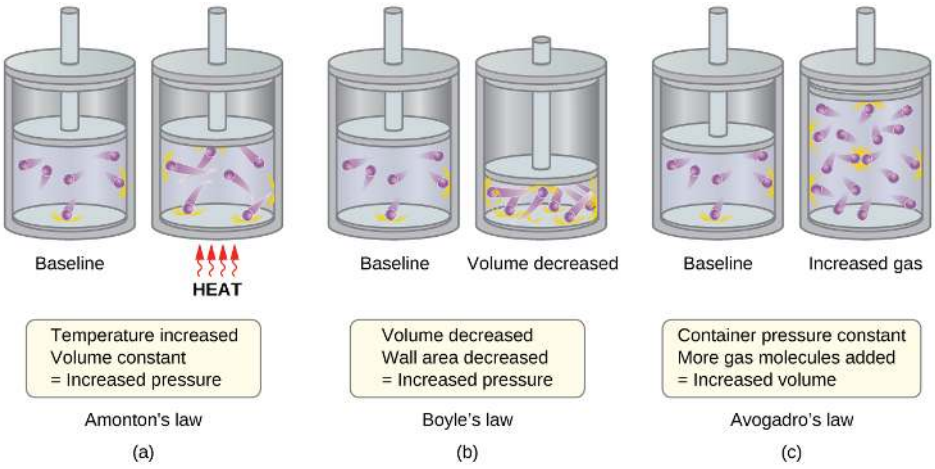


***Figure* 7.4.1:** *(a) When gas temperature increases, gas pressure increases due to increased force and frequency of molecular collisions. (b) When volume decreases, gas pressure increases due to increased frequency of molecular collisions. (c) When the amount of gas increases at a constant pressure, volume increases to yield a constant number of collisions per unit wall area per unit time.*

Molecular Velocities and Kinetic Energy

The previous discussion showed that the KMT qualitatively explains the behaviors described by the various gas laws. The postulates of this theory may be applied in a more quantitative fashion to derive these individual laws. To do this, we must first look at velocities and kinetic energies of gas molecules, and the temperature of a gas sample.

In a gas sample, individual molecules have widely varying speeds; however, because of the *vast* number of molecules and collisions involved, the molecular speed distribution and average speed are constant. This molecular speed distribution is known as a Maxwell-Boltzmann distribution, and it depicts the relative numbers of molecules in a bulk sample of gas that possesses a given speed (Figure **7.4.2**).

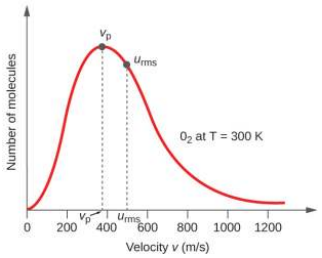


***Figure* 7.4.2:** *The molecular speed distribution for oxygen gas at 300 K is shown here. Very few molecules move at either very low or very high speeds. The number of molecules with intermediate speeds increases rapidly up to a maximum, which is the most probable speed, then drops off rapidly. Note that the most probable speed, v_p, is a little less than 400 m/s, while the root mean square speed, u_{rms}, is closer to 500 m/s.*

The kinetic energy (KE) of a particle of mass (m) and speed (u) is given by:

$$\mathrm{KE} = \frac{1}{2}mu^2 \tag{7.4.1}$$

Expressing mass in kilograms and speed in meters per second will yield energy values in units of joules (J = kg m^2 s^{-2}). To deal with a large number of gas molecules, we use averages for both speed and kinetic energy. In the KMT, the root mean square velocity of a particle, u_{rms}, is defined as the square root of the average of the squares of the velocities with n = the number of particles:

$$u_{rms} = \sqrt{\overline{u^2}} = \sqrt{\frac{u_1^2 + u_2^2 + u_3^2 + u_4^2 + \ldots}{n}} \tag{7.4.2}$$

The average kinetic energy, KE_{avg}, is then equal to:

$$\mathrm{KE_{avg}} = \frac{1}{2}mu_{rms}^2 \tag{7.4.3}$$

The KE_{avg} of a collection of gas molecules is also directly proportional to the temperature of the gas and may be described by the equation:

$$\mathrm{KE_{avg}} = \frac{3}{2}RT \tag{7.4.4}$$

where R is the gas constant and T is the kelvin temperature. When used in this equation, the appropriate form of the gas constant is 8.314 J $mol^{-1}K^{-1}$ (8.314 kg $m^2s^{-2}mol^{-1}K^{-1}$). These two separate equations for KE_{avg} may be combined and rearranged to yield a relation between molecular speed and

temperature:

$$\frac{1}{2}mu_{rms}^2 = \frac{3}{2}RT \tag{7.4.5}$$

$$u_{rms} = \sqrt{\frac{3RT}{m}} \tag{7.4.6}$$

Example 7.4.1: Calculation of u_{rms}

Calculate the root-mean-square velocity for a nitrogen molecule at 30 °C.

Solution

Convert the temperature into Kelvin:

$$30°C + 273 = 303\ K \tag{7.4.7}$$

Determine the mass of a nitrogen molecule in kilograms:

$$\frac{28.0\ \cancel{g}}{1\ \text{mol}} \times \frac{1\ \text{kg}}{1000\ \cancel{g}} = 0.028\ \text{kg/mol} \tag{7.4.8}$$

Replace the variables and constants in the root-mean-square velocity formula (Equation 7.4.6), replacing Joules with the equivalent kg m^2s^{-2}:

$$u_{rms} = \sqrt{\frac{3RT}{m}} \tag{7.4.9}$$

$$u_{rms} = \sqrt{\frac{3(8.314\ J/mol\ K)(303\ K)}{(0.028\ kg/mol)}} \tag{7.4.10}$$

$$= \sqrt{2.70 \times 10^5\ m^2s^{-2}} \tag{7.4.11}$$

$$= 519\ m/s \tag{7.4.12}$$

Exercise 7.4.1

Calculate the root-mean-square velocity for an oxygen molecule at –23 °C.

Answer

441 m/s

If the temperature of a gas increases, its KE_{avg} increases, more molecules have higher speeds and fewer molecules have lower speeds, and the distribution shifts toward higher speeds overall, that is, to the right. If temperature decreases, KE_{avg} decreases, more molecules have lower speeds and fewer molecules have higher speeds, and the distribution shifts toward lower speeds overall, that is, to the left. This behavior is illustrated for nitrogen gas in Figure 7.4.3.

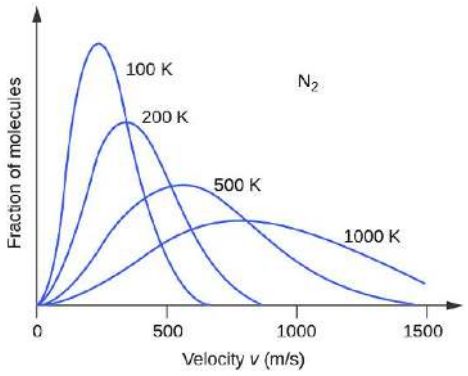


***Figure* 7.4.3**: *The molecular speed distribution for nitrogen gas (N_2) shifts to the right and flattens as the temperature increases; it shifts to the left and heightens as the temperature decreases.*

At a given temperature, all gases have the same KE_{avg} for their molecules. Gases composed of lighter molecules have more high-speed particles and a higher u_{rms}, with a speed distribution that peaks at relatively higher velocities. Gases consisting of heavier molecules have more low-speed particles, a lower u_{rms}, and a speed distribution that peaks at relatively lower velocities. This trend is demonstrated by the data for a series of noble gases shown in Figure 7.4.4.

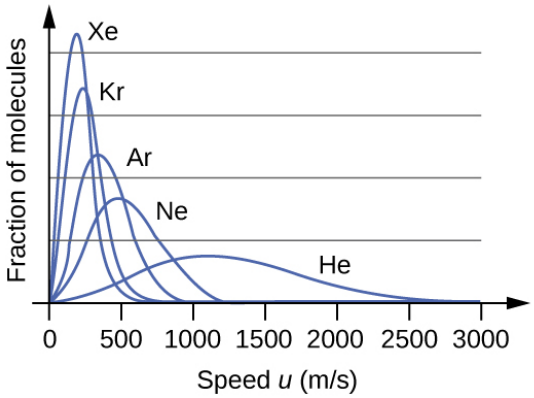
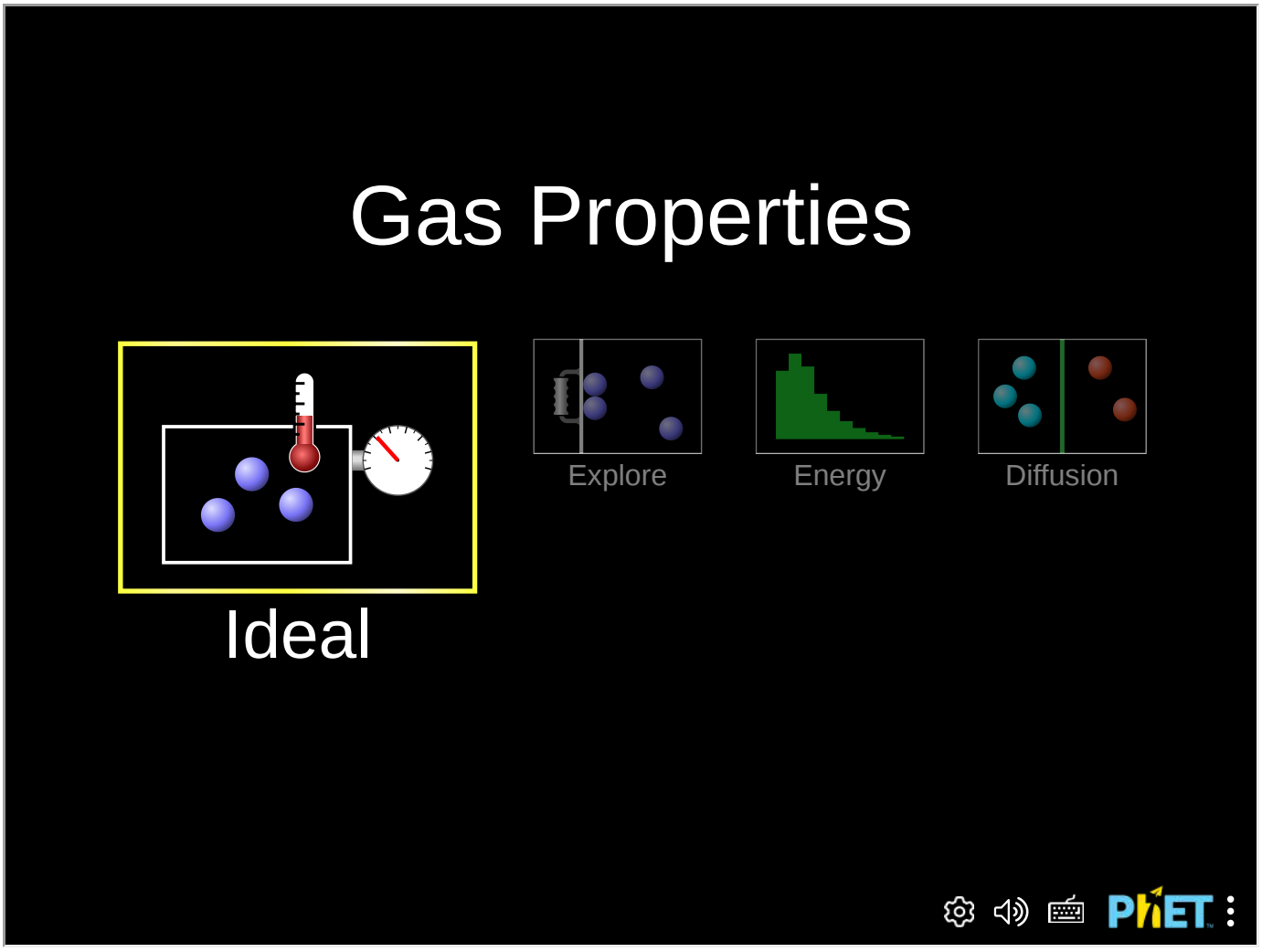


***Figure* 7.4.4:** *Molecular velocity is directly related to molecular mass. At a given temperature, lighter molecules move faster on average than heavier molecules.*



The Kinetic-Molecular Theory Explains the Behavior of Gases, Part II

According to Graham's law, the molecules of a gas are in rapid motion and the molecules themselves are small. The average distance between the molecules of a gas is large compared to the size of the molecules. As a consequence, gas molecules can move past each other easily and diffuse at relatively fast rates.

The rate of effusion of a gas depends directly on the (average) speed of its molecules:

$$\text{effusion rate} \propto u_{\text{rms}} \tag{7.4.13}$$

Using this relation, and the equation relating molecular speed to mass, Graham's law may be easily derived as shown here:

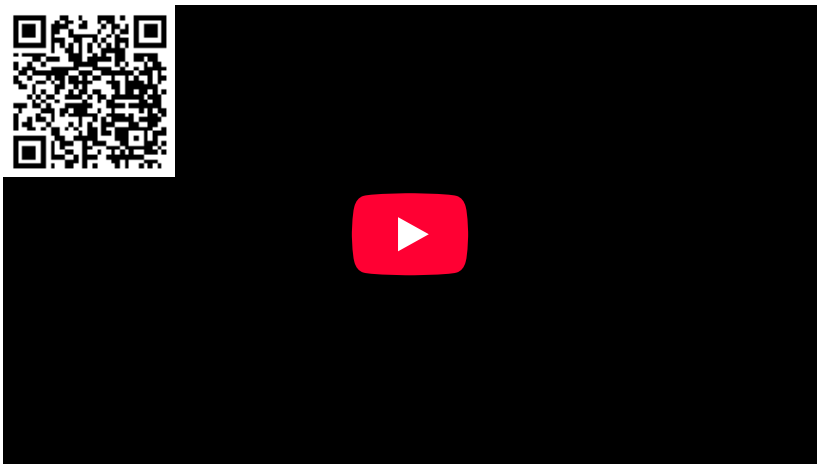
$$u_{\text{rms}} = \sqrt{\frac{3RT}{m}} \tag{7.4.14}$$

$$m = \frac{3RT}{u_{\text{rms}}^2} = \frac{3RT}{\overline{u}^2} \tag{7.4.15}$$

$$\frac{\text{effusion rate A}}{\text{effusion rate B}} = \frac{u_{\text{rms A}}}{u_{\text{rms B}}} = \frac{\sqrt{\frac{3RT}{m_{\text{A}}}}}{\sqrt{\frac{3RT}{m_{\text{B}}}}} = \sqrt{\frac{m_{\text{B}}}{m_{\text{A}}}} \tag{7.4.16}$$

The ratio of the rates of effusion is thus derived to be inversely proportional to the ratio of the square roots of their masses. This is the same relation observed experimentally and expressed as Graham's law.

Summary



***Video* 7.4.1:** *A summary of Kinetic Molecular Theory under the Ideal Gas Law.*

The kinetic molecular theory is a simple but very effective model that effectively explains ideal gas behavior. The theory assumes that gases consist of widely separated molecules of negligible volume that are in constant motion, colliding elastically with one another and the walls of their container with average velocities determined by their absolute temperatures. The individual molecules of a gas exhibit a range of velocities, the distribution of these velocities being dependent on the temperature of the gas and the mass of its molecules.

Key Equations

- $u_{\text{rms}} = \sqrt{\overline{u^2}} = \sqrt{\frac{u_1^2 + u_2^2 + u_3^2 + u_4^2 + \ldots}{n}}$
- $\text{KE}_{\text{avg}} = \frac{3}{2}RT$
- $u_{\text{rms}} = \sqrt{\frac{3RT}{m}}$

Glossary

kinetic molecular theory
: theory based on simple principles and assumptions that effectively explains ideal gas behavior

root mean square velocity (u_{rms})
: measure of average velocity for a group of particles calculated as the square root of the average squared velocity

Contributors

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7.4.1: Practice Problems- Kinetic Molecular Theory

PROBLEM 7.4.1.1

Using the postulates of the kinetic molecular theory, explain why a gas uniformly fills a container of any shape.

Answer

The first three parts of Kinetic Molecular Theory explain this:

Gases are composed of molecules that are in continuous motion, travelling in straight lines and changing direction only when they collide with other molecules or with the walls of a container.

The molecules composing the gas are negligibly small compared to the distances between them.

The pressure exerted by a gas in a container results from collisions between the gas molecules and the container walls.

PROBLEM 7.4.1.2

Can the speed of a given molecule in a gas double at constant temperature? Explain your answer.

Answer

Yes. At any given instant, there are a range of values of molecular speeds in a sample of gas. Any single molecule can speed up or slow down as it collides with other molecules. The average velocity of all the molecules is constant at constant temperature.

PROBLEM 7.4.1.3

Describe what happens to the average kinetic energy of ideal gas molecules when the conditions are changed as follows:

a. The pressure of the gas is increased by reducing the volume at constant temperature.
b. The pressure of the gas is increased by increasing the temperature at constant volume.
c. The average velocity of the molecules is increased by a factor of 2.

Answer a

Kinetic energy will not change (temperature is unchanged)

Answer b

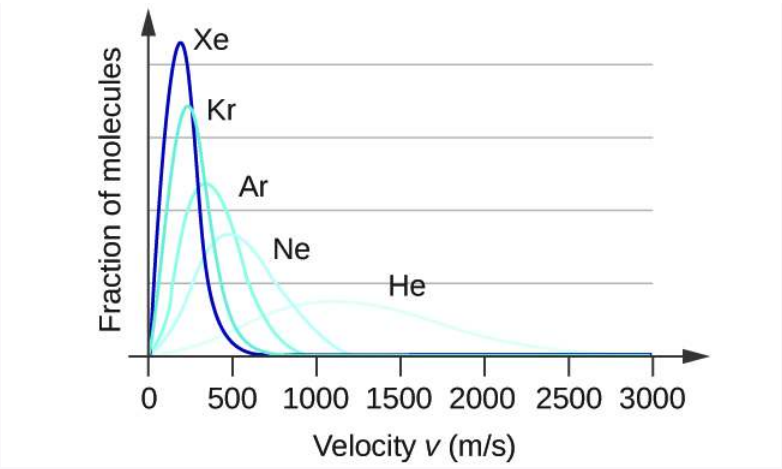
Kinetic energy will increase (temperature is increased)

Answer c

The kinetic energy will increase by a factor of four (velocity is squared in the equation)

PROBLEM 7.4.1.4

The distribution of molecular velocities in a sample of helium is shown in the photo below (Figure 7.4.4). If the sample is cooled, will the distribution of velocities look more like that of H_2 or of H_2O? Explain your answer.



Answer

H_2O. Cooling slows the velocities of the He atoms, causing them to behave as though they were heavier.

Contributors

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- Adelaide Clark, Oregon Institute of Technology

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CHAPTER OVERVIEW

8: Thermochemistry

Unit 8 Objectives

- Define energy, distinguish types of energy, and describe the nature of energy changes that accompany chemical and physical changes
- Distinguish the related properties of heat, thermal energy, and temperature
- Define and distinguish specific heat and heat capacity, and describe the physical implications of both
- Perform calculations involving heat, specific heat, and temperature change
- State the first law of thermodynamics
- Define enthalpy and explain its classification as a state function
- Write and balance thermochemical equations
- Calculate enthalpy changes for various chemical reactions
- Explain Hess's law and use it to compute reaction enthalpies
- Explain the technique of calorimetry
- Calculate and interpret heat and related properties using typical calorimetry data

8.1: The Basics of Energy

8.1.1: Practice Problems- The Basics of Energy

8.2: Measuring Energy and Heat Capacity

8.3: Enthalpy

8.4: Standard Enthalpy and Hess' Law

8.4.1: Practice Problems- Enthalpy and Hess' Law

8.5: Calorimetry

8.5.1: Practice Problems- Calorimetry

Standard Enthalpies of Formation Table (useful for Hess' Law HW)

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8.1: The Basics of Energy

Skills to Develop

- Define energy, distinguish types of energy, and describe the nature of energy changes that accompany chemical and physical changes
- Distinguish the related properties of heat, thermal energy, and temperature

Chemical changes and their accompanying changes in energy are important parts of our everyday world (Figure 8.1.1). The macronutrients in food (proteins, fats, and carbohydrates) undergo metabolic reactions that provide the energy to keep our bodies functioning. We burn a variety of fuels (gasoline, natural gas, coal) to produce energy for transportation, heating, and the generation of electricity. Industrial chemical reactions use enormous amounts of energy to produce raw materials (such as iron and aluminum). Energy is then used to manufacture those raw materials into useful products, such as cars, skyscrapers, and bridges.



Figure* 8.1.1*: *The energy involved in chemical changes is important to our daily lives: (a) A cheeseburger for lunch provides the energy you need to get through the rest of the day; (b) the combustion of gasoline provides the energy that moves your car (and you) between home, work, and school; and (c) coke, a processed form of coal, provides the energy needed to convert iron ore into iron, which is essential for making many of the products we use daily. (credit a: modification of work by "Pink Sherbet Photography"/Flickr; credit b: modification of work by Jeffery Turner).*

Over 90% of the energy we use comes originally from the sun. Every day, the sun provides the earth with almost 10,000 times the amount of energy necessary to meet all of the world's energy needs for that day. Our challenge is to find ways to convert and store incoming solar energy so that it can be used in reactions or chemical processes that are both convenient and nonpolluting. Plants and many bacteria capture solar energy through photosynthesis. We release the energy stored in plants when we burn wood or plant products such as ethanol. We also use this energy to fuel our bodies by eating food that comes directly from plants or from animals that got their energy by eating plants. Burning coal and petroleum also releases stored solar energy: These fuels are fossilized plant and animal matter.

This chapter will introduce the basic ideas of an important area of science concerned with the amount of heat absorbed or released during chemical and physical changes—an area called thermochemistry. The concepts introduced in this chapter are widely used in almost all scientific and technical fields. Food scientists use them to determine the energy content of foods. Biologists study the energetics of living organisms, such as the metabolic combustion of sugar into carbon dioxide and water. The oil, gas, and transportation industries, renewable energy providers, and many others endeavor to find better methods to produce energy for our commercial and personal needs. Engineers strive to improve energy efficiency, find better ways to heat and cool our homes, refrigerate our food and drinks, and meet the energy and cooling needs of computers and electronics, among other applications. Understanding thermochemical principles is essential for chemists, physicists, biologists, geologists, every type of engineer, and just about anyone who studies or does any kind of science.

Energy

Energy can be defined as the capacity to supply heat or do work. One type of work (w) is the process of causing matter to move against an opposing force. For example, we do work when we inflate a bicycle tire—we move matter (the air in the pump) against the opposing force of the air surrounding the tire.

Like matter, energy comes in different types. One scheme classifies energy into two types: potential energy, the energy an object has because of its relative position, composition, or condition, and kinetic energy, the energy that an object possesses because of its

motion. Water at the top of a waterfall or dam has potential energy because of its position; when it flows downward through generators, it has kinetic energy that can be used to do work and produce electricity in a hydroelectric plant (Figure 8.1.2). A battery has potential energy because the chemicals within it can produce electricity that can do work.



***Figure* 8.1.2**: *(a) Water that is higher in elevation, for example, at the top of Victoria Falls, has a higher potential energy than water at a lower elevation. As the water falls, some of its potential energy is converted into kinetic energy. (b) If the water flows through generators at the bottom of a dam, such as the Hoover Dam shown here, its kinetic energy is converted into electrical energy. (credit a: modification of work by Steve Jurvetson; credit b: modification of work by "curimedia"/Wikimedia commons).*

Energy can be converted from one form into another, but all of the energy present before a change occurs always exists in some form after the change is completed. This observation is expressed in the law of conservation of energy: during a chemical or physical change, energy can be neither created nor destroyed, although it can be changed in form. (This is also one version of the first law of thermodynamics, as you will learn later.)

When one substance is converted into another, there is always an associated conversion of one form of energy into another. Heat is usually released or absorbed, but sometimes the conversion involves light, electrical energy, or some other form of energy. For example, chemical energy (a type of potential energy) is stored in the molecules that compose gasoline. When gasoline is combusted within the cylinders of a car's engine, the rapidly expanding gaseous products of this chemical reaction generate mechanical energy (a type of kinetic energy) when they move the cylinders' pistons.

According to the law of conservation of matter (seen in an earlier chapter), there is no detectable change in the total amount of matter during a chemical change. When chemical reactions occur, the energy changes are relatively modest and the mass changes are too small to measure, so the laws of conservation of matter and energy hold well. However, in nuclear reactions, the energy changes are much larger (by factors of a million or so), the mass changes are measurable, and matter-energy conversions are significant. This will be examined in more detail in a later chapter on nuclear chemistry. To encompass both chemical and nuclear changes, we combine these laws into one statement: The total quantity of matter and energy in the universe is fixed.

Thermal Energy, Temperature, and Heat

Thermal energy is kinetic energy associated with the random motion of atoms and molecules. Temperature is a quantitative measure of "hot" or "cold." When the atoms and molecules in an object are moving or vibrating quickly, they have a higher average kinetic energy (KE), and we say that the object is "hot." When the atoms and molecules are moving slowly, they have lower KE, and we say that the object is "cold" (Figure 8.1.3). Assuming that no chemical reaction or phase change (such as melting or vaporizing) occurs, increasing the amount of thermal energy in a sample of matter will cause its temperature to increase. And, assuming that no chemical reaction or phase change (such as condensation or freezing) occurs, decreasing the amount of thermal energy in a sample of matter will cause its temperature to decrease.

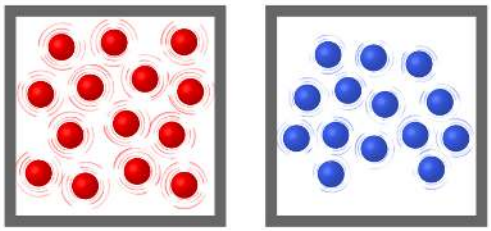


***Figure* 8.1.3**: *(a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water.*



Most substances expand as their temperature increases and contract as their temperature decreases. This property can be used to measure temperature changes, as shown in Figure **8.1.4**. The operation of many thermometers depends on the expansion and contraction of substances in response to temperature changes.

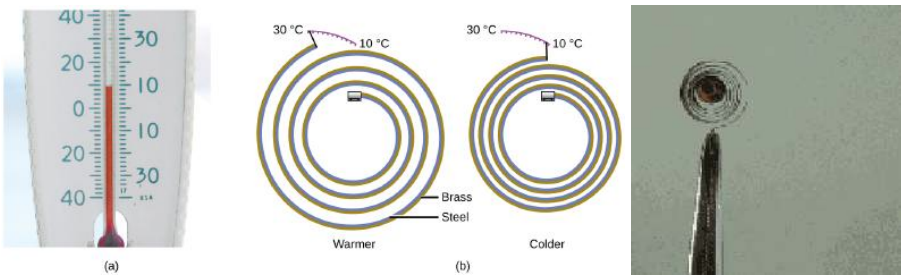


Figure 8.1.4***:*** *(a) In an alcohol or mercury thermometer, the liquid (dyed red for visibility) expands when heated and contracts when cooled, much more so than the glass tube that contains the liquid. (b) In a bimetallic thermometer, two different metals (such as brass and steel) form a two-layered strip. When heated or cooled, one of the metals (brass) expands or contracts more than the other metal (steel), causing the strip to coil or uncoil. Both types of thermometers have a calibrated scale that indicates the temperature. (credit a: modification of work by "dwstucke"/Flickr). (c) The demonstration allows one to view the effects of heating and cooling a coiled bimetallic strip.A bimetallic coil from a thermometer reacts to the heat from a lighter, by uncoiling and then coiling back up when the lighter is removed. Animation used with permission from Hustvedt (via Wikipedia)*

Heat (q) is the transfer of thermal energy between two bodies at different temperatures. Heat flow (a redundant term, but one commonly used) increases the thermal energy of one body and decreases the thermal energy of the other. Suppose we initially have a high temperature (and high thermal energy) substance (H) and a low temperature (and low thermal energy) substance (L). The atoms and molecules in H have a higher average KE than those in L. If we place substance H in contact with substance L, the thermal energy will flow spontaneously from substance H to substance L. The temperature of substance H will decrease, as will the average KE of its molecules; the temperature of substance L will increase, along with the average KE of its molecules. Heat flow will continue until the two substances are at the same temperature (Figure 8.1.5).

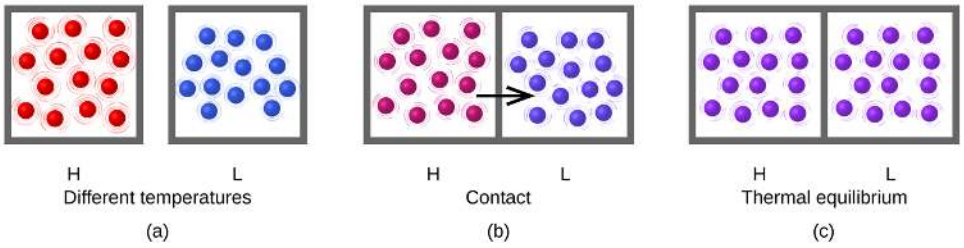


Figure 8.1.5*: (a) Substances H and L are initially at different temperatures, and their atoms have different average kinetic energies. (b) When they are put into contact with each other, collisions between the molecules result in the transfer of kinetic (thermal) energy from the hotter to the cooler matter. (c) The two objects reach "thermal equilibrium" when both substances are at the same temperature, and their molecules have the same average kinetic energy.*

Matter undergoing chemical reactions and physical changes can release or absorb heat. A change that releases heat is called an exothermic process. For example, the combustion reaction that occurs when using an oxyacetylene torch is an exothermic process—this process also releases energy in the form of light as evidenced by the torch's flame (Figure 8.1.6*a*). A reaction or change that absorbs heat is an endothermic process. A cold pack used to treat muscle strains provides an example of an endothermic process. When the substances in the cold pack (water and a salt like ammonium nitrate) are brought together, the resulting process absorbs heat, leading to the sensation of cold.



Figure 8.1.6*: (a) An oxyacetylene torch produces heat by the combustion of acetylene in oxygen. The energy released by this exothermic reaction heats and then melts the metal being cut. The sparks are tiny bits of the molten metal flying away. (b) A cold pack uses an endothermic process to create the sensation of cold. (credit a: modification of work by "Skatebiker"/Wikimedia commons).*

Units of Energy

Energy is measured in terms of its ability to perform work or to transfer heat. Mechanical work is done when a force f displaces an object by a distance *d*:

$$w = f \times d \tag{8.1.1}$$

The basic unit of energy is the *joule*. One joule is the amount of work done when a force of 1 newton acts over a distance of 1 m; thus 1 J = 1 N-m. The newton is the amount of force required to accelerate a 1-kg mass by 1 m/sec^2, so the basic dimensions of the joule are kg m^2 s^{-2}. Historically, energy was measured in units of calories (cal). A calorie is the amount of energy required to raise one gram of water by 1 degree C (1 kelvin). However, this quantity depends on the atmospheric pressure and the starting temperature of the water. The ease of measurement of energy changes in calories has meant that the calorie is still frequently used. The Calorie (with a capital C), or large calorie, commonly used in quantifying food energy content, is a kilocalorie. Another common unit of measurement of energy is the *BTU* (British thermal unit) which is defined in terms of the heating effect on water. Because of the many forms that energy can take, there are a correspondingly large number of units in which it can be expressed, a few of which are summarized below.

1 calorie will raise the temperature of 1 g of water by 1 C°. The "dietary" calorie is actually 1 kcal. An average young adult expends about 1800 kcal per day just to stay alive. (you should know this definition)	1 cal = 4.184 J
1 BTU (British Thermal Unit) will raise the temperature of 1 lb of water by 1F°.	1 BTU = 1055 J
The erg is the c.g.s. unit of energy and a very small one; the work done when a 1-dyne force acts over a distance of 1 cm.	1 J = 10^7 ergs 1 erg = 1 d-cm = 1 g cm^2 s^{-2}
The electron-volt is even tinier: 1 eV is the work required to move a unit electric charge (1 C) through a potential difference of 1 volt.	1 J = 6.24 × 1018 eV
The watt is a unit of power, which measures the rate of energy flow in J sec^{-1}. Thus the watt-hour is a unit of energy. An average human consumes energy at a rate of about 100 watts; the brain alone runs at about 5 watts.	1 J = 2.78 × 10^{-4}watt-hr 1 w-h = 3.6 kJ
The liter-atmosphere is a variant of force-displacement work associated with volume changes in gases.	1 L-atm = 101.325 J
The huge quantities of energy consumed by cities and countries are expressed in quads; the therm is a similar but smaller unit.	1 quad = 1015 Btu = 1.05 × 10^{18} J
If the object is to obliterate cities or countries with nuclear weapons, the energy unit of choice is the ton of TNT equivalent.	1 ton of TNT = 4.184 GJ (by definition)
In terms of fossil fuels, we have barrel-of-oil equivalent, cubic-meter-of-natural gas equivalent, and ton-of-coal equivalent.	1 bboe = 6.1 GJ 1 cmge = 37-39 mJ 1 toce = 29 GJ

Summary



Video* 8.1.1*: *A video summary of Energy and Chemistry.*

Energy is the capacity to do work (applying a force to move matter). Kinetic energy (KE) is the energy of motion; potential energy is energy due to relative position, composition, or condition. When energy is converted from one form into another, energy is neither created nor destroyed (law of conservation of energy or first law of thermodynamics). Matter has thermal energy due to the KE of its molecules and temperature that corresponds to the average KE of its molecules. Heat is energy that is transferred between

objects at different temperatures; it flows from a high to a low temperature. Chemical and physical processes can absorb heat (endothermic) or release heat (exothermic). The SI unit of energy, heat, and work is the joule (J). Specific heat and heat capacity are measures of the energy needed to change the temperature of a substance or object. The amount of heat absorbed or released by a substance depends directly on the type of substance, its mass, and the temperature change it undergoes.

Key Equations

- $q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$

Glossary

calorie (cal)

unit of heat or other energy; the amount of energy required to raise 1 gram of water by 1 degree Celsius; 1 cal is defined as 4.184 J

endothermic process

chemical reaction or physical change that absorbs heat

energy

capacity to supply heat or do work

exothermic process

chemical reaction or physical change that releases heat

joule (J)

SI unit of energy; 1 joule is the kinetic energy of an object with a mass of 2 kilograms moving with a velocity of 1 meter per second, 1 J = 1 kg m^2/s and 4.184 J = 1 cal

kinetic energy

energy of a moving body, in joules, equal to $\frac{1}{2}mv^2$ (where m = mass and v = velocity)

potential energy

energy of a particle or system of particles derived from relative position, composition, or condition

temperature

intensive property of matter that is a quantitative measure of “hotness” and “coldness”

thermal energy

kinetic energy associated with the random motion of atoms and molecules

thermochemistry

study of measuring the amount of heat absorbed or released during a chemical reaction or a physical change

work (w)

energy transfer due to changes in external, macroscopic variables such as pressure and volume; or causing matter to move against an opposing force

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- Adelaide E. Clark, Oregon Institute of Technology

- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook
- Crash Course Chemistry: Crash Course is a division of Complexly and videos are free to stream for educational purposes.

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8.1.1: Practice Problems- The Basics of Energy

PROBLEM 8.1.1.1

A burning match and a bonfire may have the same temperature, yet you would not sit around a burning match on a fall evening to stay warm. Why not?

Answer

The temperature of 1 gram of burning wood is approximately the same for both a match and a bonfire. This is an intensive property and depends on the material (wood). However, the overall amount of produced heat depends on the amount of material; this is an extensive property. The amount of wood in a bonfire is much greater than that in a match; the total amount of produced heat is also much greater, which is why we can sit around a bonfire to stay warm, but a match would not provide enough heat to keep us from getting cold.

PROBLEM 8.1.1.2

Explain the difference between heat capacity and specific heat of a substance.

Answer

Heat capacity refers to the heat required to raise the temperature of the mass of the substance 1 degree; specific heat refers to the heat required to raise the temperature of 1 gram of the substance 1 degree. Thus, heat capacity is an extensive property, and specific heat is an intensive one.

PROBLEM 8.1.1.3

How much heat, in joules and in calories, must be added to a 75.0–g iron block with a specific heat of 0.449 J/g °C to increase its temperature from 25 °C to its melting temperature of 1535 °C?

Answer

50,800 J

12,200 cal

Click here to see a video of the solution



PROBLEM 8.1.1.4

How much heat, in joules and in calories, is required to heat a 28.4-g (1-oz) ice cube from −23.0 °C to −1.0 °C?

Answer

1310 J

313 cal

PROBLEM 8.1.1.5

How much would the temperature of 275 g of water increase if 36.5 kJ of heat were added?

Answer

31.7° C

Click here to see a video of the solution



PROBLEM 8.1.1.6

If 14.5 kJ of heat were added to 485 g of liquid water, how much would its temperature increase?

Answer

7.15 °C

PROBLEM 8.1.1.7

A piece of unknown substance weighs 44.7 g and requires 2110 J to increase its temperature from 23.2 °C to 89.6 °C.

a. What is the specific heat of the substance?
b. If it is one of the substances found in Table 8.1.1, what is its likely identity?

Answer a

$$C = \frac{0.711\ J}{g\ °C}$$

Answer b

Silicon

Click here to see a video of the solution



PROBLEM 8.1.1.8

A piece of unknown solid substance weighs 437.2 g, and requires 8460 J to increase its temperature from 19.3 °C to 68.9 °C.

a. What is the specific heat of the substance?
b. If it is one of the substances found in Table 8.1.1, what is its likely identity?

Answer a

$$C = \frac{0.390\ J}{g\ °C}$$

Answer b

Copper

PROBLEM 8.1.1.9

An aluminum kettle weighs 1.05 kg.

a. What is the heat capacity of the kettle (Table 8.1.1)?
b. How much heat is required to increase the temperature of this kettle from 23.0 °C to 99.0 °C?
c. How much heat (in kJ) is required to heat this kettle from 23.0 °C to 99.0 °C if it contains 1.25 L of water (density of 0.997 g/mL and a specific heat of 4.184 J/g °C)?

Answer a

$$C = \frac{0.897\ J}{g\ °C}$$

Answer b

71580 J

Answer c

467.86 kJ

Click here to see a video of the solution



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- Adelaide Clark, Oregon Institute of Technology

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8.2: Measuring Energy and Heat Capacity

Skills to Develop

- Define and distinguish specific heat and heat capacity, and describe the physical implications of both
- Perform calculations involving heat, specific heat, and temperature change

Measuring Energy and Heat Capacity

We now introduce two concepts useful in describing heat flow and temperature change. The heat capacity (C) of a body of matter is the quantity of heat (q) it absorbs or releases when it experiences a temperature change (ΔT) of 1 degree Celsius (or equivalently, 1 kelvin)

$$C = \frac{q}{\Delta T} \tag{8.2.1}$$

Heat capacity is determined by both the type and amount of substance that absorbs or releases heat. It is therefore an extensive property—its value is proportional to the amount of the substance.

For example, consider the heat capacities of two cast iron frying pans. The heat capacity of the large pan is five times greater than that of the small pan because, although both are made of the same material, the mass of the large pan is five times greater than the mass of the small pan. More mass means more atoms are present in the larger pan, so it takes more energy to make all of those atoms vibrate faster. The heat capacity of the small cast iron frying pan is found by observing that it takes 18,150 J of energy to raise the temperature of the pan by 50.0 °C

$$C_{\text{small pan}} = \frac{18,140\ \text{J}}{50.0\ \text{°C}} = 363\ \text{J/°C} \tag{8.2.2}$$

The larger cast iron frying pan, while made of the same substance, requires 90,700 J of energy to raise its temperature by 50.0 °C. The larger pan has a (proportionally) larger heat capacity because the larger amount of material requires a (proportionally) larger amount of energy to yield the same temperature change:

$$C_{\text{large pan}} = \frac{90,700\ \text{J}}{50.0\ \text{°C}} = 1814\ \text{J/°C} \tag{8.2.3}$$

The specific heat capacity (c) of a substance, commonly called its “specific heat,” is the quantity of heat required to raise the temperature of 1 gram of a substance by 1 degree Celsius (or 1 kelvin):

$$c = \frac{q}{\text{m}\Delta\text{T}} \tag{8.2.4}$$

Specific heat capacity depends only on the kind of substance absorbing or releasing heat. It is an intensive property—the type, but not the amount, of the substance is all that matters. For example, the small cast iron frying pan has a mass of 808 g. The specific heat of iron (the material used to make the pan) is therefore:

$$c_{\text{iron}} = \frac{18,140\ \text{J}}{(808\ \text{g})(50.0\ \text{°C})} = 0.449\ \text{J/g °C} \tag{8.2.5}$$

The large frying pan has a mass of 4040 g. Using the data for this pan, we can also calculate the specific heat of iron:

$$c_{\text{iron}} = \frac{90,700\ \text{J}}{(4,040\ \text{g})(50.0\ \text{°C})} = 0.449\ \text{J/g °C} \tag{8.2.6}$$

Although the large pan is more massive than the small pan, since both are made of the same material, they both yield the same value for specific heat (for the material of construction, iron). Note that specific heat is measured in units of energy per temperature per mass and is an intensive property, being derived from a ratio of two extensive properties (heat and mass). The molar heat capacity, also an intensive property, is the heat capacity per mole of a particular substance and has units of J/mol °C (Figure **8.2.7**).



Figure* 8.2.1: Due to its larger mass, a large frying pan has a larger heat capacity than a small frying pan. Because they are made of the same material, both frying pans have the same specific heat. (credit: Mark Blaser).*

Liquid water has a relatively high specific heat (about 4.2 J/g °C); most metals have much lower specific heats (usually less than 1 J/g °C). The specific heat of a substance varies somewhat with temperature. However, this variation is usually small enough that we will treat specific heat as constant over the range of temperatures that will be considered in this chapter. Specific heats of some common substances are listed in Table 8.2.1.

Table* 8.2.1: Specific Heats of Common Substances at 25 °C and 1 bar*

Substance	Symbol (state)	Specific Heat (J/g °C)
helium	$He(g)$	5.193
water	$H_2O(l)$	4.184
ethanol	$C_2H_6O(l)$	2.376
ice	$H_2O(s)$	2.093 (at −10 °C)
water vapor	$H_2O(g)$	1.864
nitrogen	$N_2(g)$	1.040
air		1.007
oxygen	$O_2(g)$	0.918
aluminum	$Al(s)$	0.897
carbon dioxide	$CO_2(g)$	0.853
argon	$Ar(g)$	0.522
iron	$Fe(s)$	0.449
copper	$Cu(s)$	0.385
lead	$Pb(s)$	0.130

Substance	Symbol (state)	Specific Heat (J/g °C)
gold	Au(*s*)	0.129
silicon	Si(*s*)	0.712

If we know the mass of a substance and its specific heat, we can determine the amount of heat, q, entering or leaving the substance by measuring the temperature change before and after the heat is gained or lost:

$$q = (\text{specific heat}) \times (\text{mass of substance}) \times (\text{temperature change})$$
$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$

In this equation, c is the specific heat of the substance, m is its mass, and ΔT (which is read "delta T") is the temperature change, $T_{\text{final}} - T_{\text{initial}}$. If a substance gains thermal energy, its temperature increases, its final temperature is higher than its initial temperature, $T_{\text{final}} - T_{\text{initial}}$ has a positive value, and the value of q is positive. If a substance loses thermal energy, its temperature decreases, the final temperature is lower than the initial temperature, $T_{\text{final}} - T_{\text{initial}}$ has a negative value, and the value of q is negative.

Example 8.2.1: Measuring Heat

A flask containing 8.0×10^2 g of water is heated, and the temperature of the water increases from 21 °C to 85 °C. How much heat did the water absorb?

Solution

To answer this question, consider these factors:

- the specific heat of the substance being heated (in this case, water)
- the amount of substance being heated (in this case, 800 g)
- the magnitude of the temperature change (in this case, from 21 °C to 85 °C).

The specific heat of water is 4.184 J/g °C, so to heat 1 g of water by 1 °C requires 4.184 J. We note that since 4.184 J is required to heat 1 g of water by 1 °C, we will need 800 times as much to heat 800 g of water by 1 °C. Finally, we observe that since 4.184 J are required to heat 1 g of water by 1 °C, we will need 64 times as much to heat it by 64 °C (that is, from 21 °C to 85 °C).

This can be summarized using the equation:

$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$

$$= (4.184\ \text{J/}\cancel{\text{g}}\text{°C}) \times (800\ \cancel{\text{g}}) \times (85 - 21)\text{°C}$$
$$= (4.184\ \text{J/}\cancel{\text{g}}\text{°}\cancel{\text{C}}) \times (800\ \cancel{\text{g}}) \times (64)\text{°}\cancel{\text{C}}$$
$$= 210,000\ \text{J} (= 210\ \text{kJ})$$

Because the temperature increased, the water absorbed heat and q is positive.

Exercise 8.2.1

How much heat, in joules, must be added to a 5.00×10^2 g iron skillet to increase its temperature from 25 °C to 250 °C? The specific heat of iron is 0.451 J/g °C.

Answer

5.05×10^4 J

Note that the relationship between heat, specific heat, mass, and temperature change can be used to determine any of these quantities (not just heat) if the other three are known or can be deduced.

Example 8.2.2: Determining Other Quantities

A piece of unknown metal weighs 348 g. When the metal piece absorbs 6.64 kJ of heat, its temperature increases from 22.4 °C to 43.6 °C. Determine the specific heat of this metal (which might provide a clue to its identity).

Solution

Since mass, heat, and temperature change are known for this metal, we can determine its specific heat using the relationship:

$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}}) \tag{8.2.7}$$

Substituting the known values:

$$6,640\ \text{J} = c \times (348\ \text{g}) \times (43.6 - 22.4)\ ^\circ\text{C} \tag{8.2.8}$$

Solving:

$$c = \frac{6,640\ \text{J}}{(348\ \text{g}) \times (21.2^\circ\text{C})} = 0.900\ \text{J/g}\ ^\circ\text{C} \tag{8.2.9}$$

Comparing this value with the values in Table 8.2.1, this value matches the specific heat of aluminum, which suggests that the unknown metal may be aluminum.

Exercise 8.2.2

A piece of unknown metal weighs 217 g. When the metal piece absorbs 1.43 kJ of heat, its temperature increases from 24.5 °C to 39.1 °C. Determine the specific heat of this metal, and predict its identity.

Answer

$c = 0.45\ \text{J/g}\ ^\circ\text{C}$; the metal is likely to be iron from checking Table 8.2.1.

Solar Thermal Energy Power Plants

The sunlight that reaches the earth contains thousands of times more energy than we presently capture. Solar thermal systems provide one possible solution to the problem of converting energy from the sun into energy we can use. Large-scale solar thermal plants have different design specifics, but all concentrate sunlight to heat some substance; the heat “stored” in that substance is then converted into electricity.

The Solana Generating Station in Arizona’s Sonora Desert produces 280 megawatts of electrical power. It uses parabolic mirrors that focus sunlight on pipes filled with a heat transfer fluid (HTF) (Figure 8.2.8). The HTF then does two things: It turns water into steam, which spins turbines, which in turn produces electricity, and it melts and heats a mixture of salts, which functions as a thermal energy storage system. After the sun goes down, the molten salt mixture can then release enough of its stored heat to produce steam to run the turbines for 6 hours. Molten salts are used because they possess a number of beneficial properties, including high heat capacities and thermal conductivities.

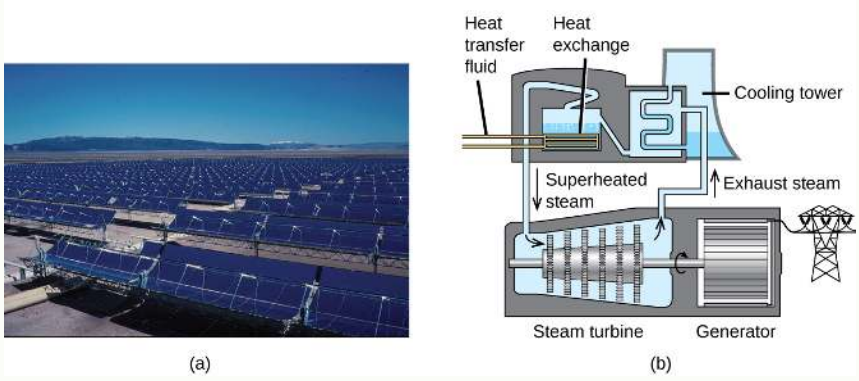


***Figure* 8.2.2:** *This solar thermal plant uses parabolic trough mirrors to concentrate sunlight. (credit a: modification of work by Bureau of Land Management)*

The 377-megawatt Ivanpah Solar Generating System, located in the Mojave Desert in California, is the largest solar thermal power plant in the world (Figure **8.2.9**). Its 170,000 mirrors focus huge amounts of sunlight on three water-filled towers, producing steam at over 538 °C that drives electricity-producing turbines. It produces enough energy to power 140,000 homes. Water is used as the working fluid because of its large heat capacity and heat of vaporization.



***Figure* 8.2.3:** *(a) The Ivanpah solar thermal plant uses 170,000 mirrors to concentrate sunlight on water-filled towers. (b) It covers 4000 acres of public land near the Mojave Desert and the California-Nevada border. (credit a: modification of work by Craig Dietrich; credit b: modification of work by "USFWS Pacific Southwest Region"/Flickr)*

Summary



***Video* 8.2.1:** *A video summary of Energy and Chemistry.*

Energy is the capacity to do work (applying a force to move matter). Kinetic energy (KE) is the energy of motion; potential energy is energy due to relative position, composition, or condition. When energy is converted from one form into another, energy is neither created nor destroyed (law of conservation of energy or first law of thermodynamics). Matter has thermal energy due to the KE of its molecules and temperature that corresponds to the average KE of its molecules. Heat is energy that is transferred between objects at different temperatures; it flows from a high to a low temperature. Chemical and physical processes can absorb heat (endothermic) or release heat (exothermic). The SI unit of energy, heat, and work is the joule (J). Specific heat and heat capacity are measures of the energy needed to change the temperature of a substance or object. The amount of heat absorbed or released by a substance depends directly on the type of substance, its mass, and the temperature change it undergoes.

Key Equations

- $q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$

Glossary

calorie (cal)
unit of heat or other energy; the amount of energy required to raise 1 gram of water by 1 degree Celsius; 1 cal is defined as 4.184 J

endothermic process
chemical reaction or physical change that absorbs heat

energy
capacity to supply heat or do work

exothermic process
chemical reaction or physical change that releases heat

heat (*q*)
transfer of thermal energy between two bodies

heat capacity (*C*)
extensive property of a body of matter that represents the quantity of heat required to increase its temperature by 1 degree Celsius (or 1 kelvin)

joule (J)
SI unit of energy; 1 joule is the kinetic energy of an object with a mass of 2 kilograms moving with a velocity of 1 meter per second, 1 J = 1 kg m^2/s and 4.184 J = 1 cal

kinetic energy
energy of a moving body, in joules, equal to $\frac{1}{2}mv^2$ (where m = mass and v = velocity)

potential energy
energy of a particle or system of particles derived from relative position, composition, or condition

specific heat capacity (*c*)
intensive property of a substance that represents the quantity of heat required to raise the temperature of 1 gram of the substance by 1 degree Celsius (or 1 kelvin)

temperature
intensive property of matter that is a quantitative measure of “hotness” and “coldness”

thermal energy
kinetic energy associated with the random motion of atoms and molecules

thermochemistry
study of measuring the amount of heat absorbed or released during a chemical reaction or a physical change

work (*w*)
energy transfer due to changes in external, macroscopic variables such as pressure and volume; or causing matter to move against an opposing force

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- Adelaide E. Clark, Oregon Institute of Technology
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8.3: Enthalpy

Skills to Develop

- State the first law of thermodynamics
- Define enthalpy and explain its classification as a state function

Thermochemistry is a branch of chemical thermodynamics, the science that deals with the relationships between heat, work, and other forms of energy in the context of chemical and physical processes. As we concentrate on thermochemistry in this chapter, we need to consider some widely used concepts of thermodynamics.

Substances act as reservoirs of energy, meaning that energy can be added to them or removed from them. Energy is stored in a substance when the kinetic energy of its atoms or molecules is raised. The greater kinetic energy may be in the form of increased translations (travel or straight-line motions), vibrations, or rotations of the atoms or molecules. When thermal energy is lost, the intensities of these motions decrease and the kinetic energy falls. The total of all possible kinds of energy present in a substance is called the internal energy (U), sometimes symbolized as E.

As a system undergoes a change, its internal energy can change, and energy can be transferred from the system to the surroundings, or from the surroundings to the system. Energy is transferred into a system when it absorbs heat (q) from the surroundings or when the surroundings do work (w) on the system. For example, energy is transferred into room-temperature metal wire if it is immersed in hot water (the wire absorbs heat from the water), or if you rapidly bend the wire back and forth (the wire becomes warmer because of the work done on it). Both processes increase the internal energy of the wire, which is reflected in an increase in the wire's temperature. Conversely, energy is transferred out of a system when heat is lost from the system, or when the system does work on the surroundings.

The relationship between internal energy, heat, and work can be represented by the equation:

$$\Delta U = q + w \tag{8.3.1}$$

as shown in Figure 8.3.1. This is one version of the first law of thermodynamics, and it shows that the internal energy of a system changes through heat flow into or out of the system (positive q is heat flow in; negative q is heat flow out) or work done on or by the system. The work, w, is positive if it is done on the system and negative if it is done by the system.

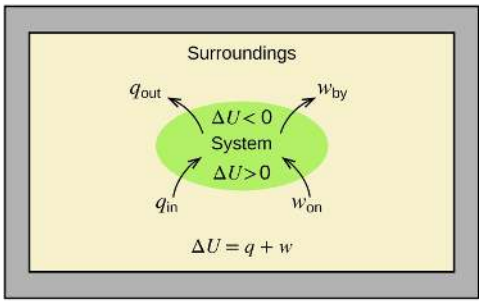


***Figure* 8.3.1:** *The internal energy, U, of a system can be changed by heat flow and work. If heat flows into the system, q_{in}, or work is done on the system, w_{on}, its internal energy increases, $\Delta U > 0$. If heat flows out of the system, q_{out}, or work is done by the system, w_{by}, its internal energy decreases, $\Delta U < 0$.*

A type of work called expansion work (or pressure-volume work) occurs when a system pushes back the surroundings against a restraining pressure, or when the surroundings compress the system. An example of this occurs during the operation of an internal combustion engine. The reaction of gasoline and oxygen is exothermic. Some of this energy is given off as heat, and some does work pushing the piston in the cylinder. The substances involved in the reaction are the system, and the engine and the rest of the universe are the surroundings. The system loses energy by both heating and doing work on the surroundings, and its internal energy decreases. (The engine is able to keep the car moving because this process is repeated many times per second while the engine is running.) We will consider how to determine the amount of work involved in a chemical or physical change in the chapter on thermodynamics.

As discussed, the relationship between internal energy, heat, and work can be represented as $\Delta U = q + w$. Internal energy is a type of quantity known as a state function (or state variable), whereas heat and work are not state functions. The value of a state function depends only on the state that a system is in, and not on how that state is reached. If a quantity is not a state function, then its value *does* depend on how the state is reached. An example of a state function is altitude or elevation. If you stand on the summit of Mt. Kilimanjaro, you are at an altitude of 5895 m, and it does not matter whether you hiked there or parachuted there. The distance you traveled to the top of Kilimanjaro, however, is not a state function. You could climb to the summit by a direct route or by a more roundabout, circuitous path (Figure **8.3.2**). The distances traveled would differ (distance is not a state function) but the elevation reached would be the same (altitude is a state function).

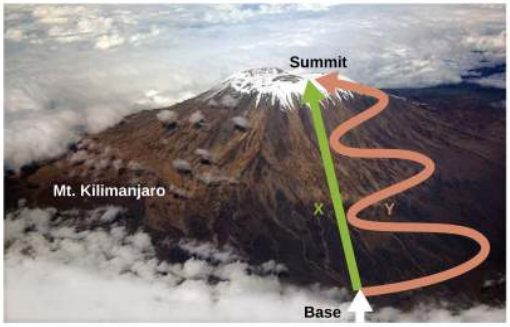


***Figure* 8.3.2:** *Paths X and Y represent two different routes to the summit of Mt. Kilimanjaro. Both have the same change in elevation (altitude or elevation on a mountain is a state function; it does not depend on path), but they have very different distances traveled (distance walked is not a state function; it depends on the path). (credit: modification of work by Paul Shaffner)*

Chemists ordinarily use a property known as enthalpy (H) to describe the thermodynamics of chemical and physical processes. Enthalpy is defined as the sum of a system's internal energy (U) and the mathematical product of its pressure (P) and volume (V):

$$H = U + PV \tag{8.3.2}$$

Since it is derived from three state functions (U, P, and V), enthalpy is also a state function. Enthalpy values for specific substances cannot be measured directly; only enthalpy changes for chemical or physical processes can be determined. For processes that take place at constant pressure (a common condition for many chemical and physical changes), the enthalpy change (ΔH) is:

$$\Delta H = \Delta U + P\Delta V \tag{8.3.3}$$

The mathematical product $P\Delta V$ represents work (w), namely, expansion or pressure-volume work as noted. By their definitions, the arithmetic signs of ΔV and w will always be opposite:

$$P\Delta V = -w \tag{8.3.4}$$

Substituting Equation 8.3.4 and the definition of internal energy (Equation 8.3.1) into Equation 8.3.3 yields:

$$\Delta H = \Delta U + P\Delta V \tag{8.3.5}$$

$$= q_p + \cancel{w} - \cancel{w} \tag{8.3.6}$$

$$= q_p \tag{8.3.7}$$

where q_p is the heat of reaction under conditions of constant pressure.

And so, if a chemical or physical process is carried out at constant pressure with the only work done caused by expansion or contraction, then the heat flow (q_p) and enthalpy change (ΔH) for the process are equal.

The heat given off when you operate a Bunsen burner is equal to the enthalpy change of the methane combustion reaction that takes place, since it occurs at the essentially constant pressure of the atmosphere. On the other hand, the heat produced by a reaction measured in a bomb calorimeter is not equal to ΔH because the closed, constant-volume metal container prevents expansion work from occurring. Chemists usually perform experiments under normal atmospheric conditions, at constant external pressure with $q = \Delta H$, which makes enthalpy the most convenient choice for determining heat.

The following conventions apply when we use ΔH:

1. Chemists use a thermochemical equation to represent the changes in both matter and energy. In a thermochemical equation, the enthalpy change of a reaction is shown as a ΔH value following the equation for the reaction. This ΔH value indicates the amount of heat associated with the reaction involving the number of moles of reactants and products as shown in the chemical equation. For example, consider this equation:

$$\mathrm{H_2\,(g) + \tfrac{1}{2}\,O_2\,(g) \longrightarrow H_2O\,(l)} \quad \Delta H = -286\ \mathrm{kJ} \tag{8.3.8}$$

   This equation indicates that when 1 mole of hydrogen gas and 12 mole of oxygen gas at some temperature and pressure change to 1 mole of liquid water at the same temperature and pressure, 286 kJ of heat are released to the surroundings. If the coefficients of the chemical equation are multiplied by some factor, the enthalpy change must be multiplied by that same factor (ΔH is an extensive property).

Multiple \label

2. The enthalpy change of a reaction depends on the physical state of the reactants and products of the reaction (whether we have gases, liquids, solids, or aqueous solutions), so these must be shown. For example, when 1 mole of hydrogen gas and 12 mole of oxygen gas change to 1 mole of liquid water at the same temperature and pressure, 286 kJ of heat are released. If gaseous water forms, only 242 kJ of heat are released.

$$\mathrm{H_2\,(g) + \tfrac{1}{2}\,O_2\,(g) \longrightarrow H_2O\,(g)} \quad \Delta H = -242\ k\,\mathrm{J} \tag{8.3.9}$$

3. A negative value of an enthalpy change, ΔH, indicates an exothermic reaction; a positive value of ΔH indicates an endothermic reaction. If the direction of a chemical equation is reversed, the arithmetic sign of its ΔH is changed (a process that is endothermic in one direction is exothermic in the opposite direction).

Example 8.3.1: Measurement of an Enthalpy Change

When 0.0500 mol of HCl(*aq*) reacts with 0.0500 mol of NaOH(*aq*) to form 0.0500 mol of NaCl(*aq*), 2.9 kJ of heat are produced. What is ΔH, the enthalpy change, per mole of acid reacting, for the acid-base reaction run under the conditions described ?

$$\mathrm{HCl\,(aq) + NaOH\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l)}$$

Solution

For the reaction of 0.0500 mol acid (HCl), q = −2.9 kJ. This ratio

$$\frac{-2.9\ \mathrm{kJ}}{0.0500\ \mathrm{mol\ HCl}}$$

can be used as a conversion factor to find the heat produced when 1 mole of HCl reacts:

$$\Delta H = 1\ \cancel{\mathrm{mol\ HCl}} \times \frac{-2.9\ \mathrm{kJ}}{0.0500\ \cancel{\mathrm{mol\ HCl}}} = -58\ \mathrm{kJ}$$

The enthalpy change when 1 mole of HCl reacts is −58 kJ. Since that is the number of moles in the chemical equation, we write the thermochemical equation as:

$$\mathrm{HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}} \quad \Delta H = -58\ \mathrm{kJ}$$

Exercise 8.3.1

When 1.34 g Zn(s) reacts with 60.0 mL of 0.750 M HCl(aq), 3.14 kJ of heat are produced. Determine the enthalpy change per mole of zinc reacting for the reaction:

$$\mathrm{Zn_{(s)} + 2\,HCl_{(aq)} \longrightarrow ZnCl_{(aq)} + H_{2(g)}}$$

Answer

ΔH = −153 kJ

Be sure to take both stoichiometry and limiting reactants into account when determining the ΔH for a chemical reaction.

Example 8.3.2: Another Example of the Measurement of an Enthalpy Change

A gummy bear contains 2.67 g sucrose, $C_{12}H_{22}O_{11}$. When it reacts with 7.19 g potassium chlorate, $KClO_3$, 43.7 kJ of heat are produced. Determine the enthalpy change for the reaction

$$C_{12}H_{22}O_{11}(aq) + 8\,KClO_3(aq) \longrightarrow 12\,CO_2(g) + 11\,H_2O(l) + 8\,KCl(aq)$$

Solution

We have $2.67\ \cancel{g} \times \frac{1\ \text{mol}}{342.3\ \cancel{g}} = 0.00780\ \text{mol}\ C_{12}H_{22}O_{11}$ available, and

$7.19\ \cancel{g} \times \frac{1\ \text{mol}}{122.5\ \cancel{g}} = 0.0587\ \text{mol}\ KClO_3$ available.

Since

$$0.0587\ \text{mol}\ KClO_3 \times \frac{1\ \text{mol}\ C_{12}H_{22}O_{11}}{8\ \text{mol}\ KClO_3} = 0.00734\ \text{mol}\ C_{12}H_{22}O_{11}$$

is needed, $C_{12}H_{22}O_{11}$ is the excess reactant and $KClO_3$ is the limiting reactant.

The reaction uses 8 mol $KClO_3$, and the conversion factor is $\frac{-43.7\ \text{kJ}}{0.0587\ \text{mol}\ KClO_3}$, so we have $\Delta H = 8\ \text{mol} \times \frac{-43.7\ \text{kJ}}{0.0587\ \text{mol}\ KClO_3} = -5960\ \text{kJ}$. The enthalpy change for this reaction is −5960 kJ, and the thermochemical equation is:

$$C_{12}H_{22}O_{11} + 8\,KClO_3 \longrightarrow 12\,CO_2 + 11\,H_2O + 8\,KCl \quad \Delta H = -5960\ kJ$$

Exercise 8.3.2

When 1.42 g of iron reacts with 1.80 g of chlorine, 3.22 g of $FeCl_{2(s)}$ and 8.60 kJ of heat is produced. What is the enthalpy change for the reaction when 1 mole of $FeCl_2\,(s)$ is produced?

Answer

ΔH = −338 kJ

Enthalpy changes are typically tabulated for reactions in which both the reactants and products are at the same conditions. A standard state is a commonly accepted set of conditions used as a reference point for the determination of properties under other different conditions. For chemists, the IUPAC standard state refers to materials under a pressure of 1 bar and solutions at 1 M, and **does not specify a temperature** (it used too). Many thermochemical tables list values with a standard state of 1 atm. Because the ΔH of a reaction changes very little with such small changes in pressure (1 bar = 0.987 atm), ΔH values (except for the most precisely measured values) are essentially the same under both sets of standard conditions. We will include a superscripted “o” in the enthalpy change symbol to designate standard state. Since the usual (but not technically standard) temperature is 298.15 K, we will use a subscripted “298” to designate this temperature. Thus, the symbol (ΔH°_{298}) is used to indicate an enthalpy change for a process occurring under these conditions. (The symbol ΔH is used to indicate an enthalpy change for a reaction occurring under nonstandard conditions.)

The enthalpy changes for many types of chemical and physical processes are available in the reference literature, including those for combustion reactions, phase transitions, and formation reactions. As we discuss these quantities, it is important to pay attention to the *extensive* nature of enthalpy and enthalpy changes. Since the enthalpy change for a given reaction is proportional to the amounts of substances involved, it may be reported on that basis (i.e., as the ΔH for specific amounts of reactants). However, we often find it more useful to divide one extensive property (ΔH) by another (amount of substance), and report a per-amount *intensive* value of ΔH, often “normalized” to a per-mole basis. (Note that this is similar to determining the intensive property specific heat from the extensive property heat capacity, as seen previously.)

Enthalpy of Combustion

Standard enthalpy of combustion (ΔH°_C) is the enthalpy change when 1 mole of a substance burns (combines vigorously with oxygen) under standard state conditions; it is sometimes called “heat of combustion.” For example, the enthalpy of combustion of

ethanol, −1366.8 kJ/mol, is the amount of heat produced when one mole of ethanol undergoes complete combustion at 25 °C and 1 atmosphere pressure, yielding products also at 25 °C and 1 atm.

$$\mathrm{C_2H_5OH}(l) + 3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{CO_2} + 3\,\mathrm{H_2O}(l) \quad \Delta H^\circ_{298} = -1366.8\ \mathrm{kJ} \tag{8.3.10}$$

Enthalpies of combustion for many substances have been measured; a few of these are listed in Table **8.3.1**. Many readily available substances with large enthalpies of combustion are used as fuels, including hydrogen, carbon (as coal or charcoal), and hydrocarbons (compounds containing only hydrogen and carbon), such as methane, propane, and the major components of gasoline.

Table* 8.3.1*: *Standard Molar Enthalpies of Combustion*

Substance	Combustion Reaction	Enthalpy of Combustion $\Delta H_c^\circ \left(\frac{\mathrm{kJ}}{\mathrm{mol}}\ \text{at } 25°\mathrm{C}\right)$
carbon	$\mathrm{C}(s) + \mathrm{O_2}(g) \longrightarrow \mathrm{CO_2}(g)$	−393.5
hydrogen	$\mathrm{H_2}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{H_2O}(l)$	−285.8
magnesium	$\mathrm{Mg}(s) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{MgO}(s)$	−601.6
sulfur	$\mathrm{S}(s) + \mathrm{O_2}(g) \longrightarrow \mathrm{SO_2}(g)$	−296.8
carbon monoxide	$\mathrm{CO}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{CO_2}(g)$	−283.0
methane	$\mathrm{CH_4}(g) + 2\,\mathrm{O_2}(g) \longrightarrow \mathrm{CO_2}(g) + 2\,\mathrm{H_2O}(l)$	−890.8
acetylene	$\mathrm{C_2H_2}(g) + \frac{5}{2}\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{CO_2}(g) + \mathrm{H_2O}(l)$	−1301.1
ethanol	$\mathrm{C_2H_5OH}(l) + 3\,\mathrm{O_2}(g) \longrightarrow \mathrm{CO_2}(g) + 3\,\mathrm{H_2O}(l)$	−1366.8
methanol	$\mathrm{CH_3OH}(l) + \frac{3}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{CO_2}(g) + 2\,\mathrm{H_2O}(l)$	−726.1
isooctane	$\mathrm{C_8H_{18}}(l) + \frac{25}{2}\mathrm{O_2}(g) \longrightarrow 8\,\mathrm{CO_2}(g) + 9\,\mathrm{H_2O}(l)$	−5461

Example 8.3.3: Using Enthalpy of Combustion

As Figure **8.3.3** suggests, the combustion of gasoline is a highly exothermic process. Let us determine the approximate amount of heat produced by burning 1.00 L of gasoline, assuming the enthalpy of combustion of gasoline is the same as that of isooctane, a common component of gasoline. The density of isooctane is 0.692 g/mL.



Figure* 8.3.3*: *The combustion of gasoline is very exothermic. (credit: modification of work by “AlexEagle”/Flickr)*

Solution

Starting with a known amount (1.00 L of isooctane), we can perform conversions between units until we arrive at the desired amount of heat or energy. The enthalpy of combustion of isooctane provides one of the necessary conversions. Table **8.3.1** gives this value as −5460 kJ per 1 mole of isooctane (C_8H_{18}).

Using these data,

$$1.00\ \cancel{\text{L C}_8\text{H}_{18}} \times \frac{1000\ \cancel{\text{mL C}_8\text{H}_{18}}}{1\ \cancel{\text{L C}_8\text{H}_{18}}} \times \frac{0.692\ \cancel{\text{g C}_8\text{H}_{18}}}{1\ \cancel{\text{mL C}_8\text{H}_{18}}} \times \frac{1\ \cancel{\text{mol C}_8\text{H}_{18}}}{114\ \cancel{\text{g C}_8\text{H}_{18}}} \times \frac{-5460\ \text{kJ}}{1\ \cancel{\text{mol C}_8\text{H}_{18}}} = -3.31 \times 10^4\ \text{kJ}$$

The combustion of 1.00 L of isooctane produces 33,100 kJ of heat. (This amount of energy is enough to melt 99.2 kg, or about 218 lbs, of ice.)

Note: If you do this calculation one step at a time, you would find:

$$1.00\ \text{L C}_8\text{H}_{18} \longrightarrow 1.00 \times 10^3\ \text{mL C}_8\text{H}_{18}$$
$$1.00 \times 10^3\ \text{mL C}_8\text{H}_{18} \longrightarrow 692\ \text{g C}_8\text{H}_{18}$$
$$692\ \text{g C}_8\text{H}_{18} \longrightarrow 6.07\ \text{mol C}_8\text{H}_{18}$$
$$692\ \text{g C}_8\text{H}_{18} \longrightarrow -3.31 \times 10^4\ \text{kJ}$$

Exercise 8.3.3

How much heat is produced by the combustion of 125 g of acetylene?

Answer

6.25×10^3 kJ

Emerging Algae-Based Energy Technologies (Biofuels)

As reserves of fossil fuels diminish and become more costly to extract, the search is ongoing for replacement fuel sources for the future. Among the most promising biofuels are those derived from algae (Figure **8.3.4**). The species of algae used are nontoxic, biodegradable, and among the world's fastest growing organisms. About 50% of algal weight is oil, which can be readily converted into fuel such as biodiesel. Algae can yield 26,000 gallons of biofuel per hectare—much more energy per acre than other crops. Some strains of algae can flourish in brackish water that is not usable for growing other crops. Algae can produce biodiesel, biogasoline, ethanol, butanol, methane, and even jet fuel.



Figure* 8.3.4*: *(a) Tiny algal organisms can be (b) grown in large quantities and eventually (c) turned into a useful fuel such as biodiesel. (credit a: modification of work by Micah Sittig; credit b: modification of work by Robert Kerton; credit c: modification of work by John F. Williams)*

According to the US Department of Energy, only 39,000 square kilometers (about 0.4% of the land mass of the US or less than $\frac{1}{7}$ of the area used to grow corn) can produce enough algal fuel to replace all the petroleum-based fuel used in the US. The cost of algal fuels is becoming more competitive—for instance, the US Air Force is producing jet fuel from algae at a total cost of under $5 per gallon. The process used to produce algal fuel is as follows: grow the algae (which use sunlight as their energy source and CO_2 as a raw material); harvest the algae; extract the fuel compounds (or precursor compounds); process as necessary (e.g., perform a transesterification reaction to make biodiesel); purify; and distribute (Figure **8.3.5**).

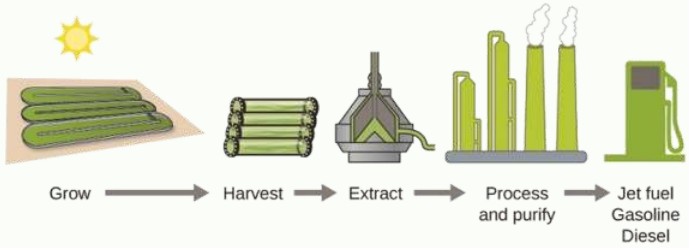


Figure* 8.3.5: Algae convert sunlight and carbon dioxide into oil that is harvested, extracted, purified, and transformed into a variety of renewable fuels.*

Summary



Video* 8.3.1*:* *A summary of Enthalpy and Hess' Law from Crash Course Chemistry.

If a chemical change is carried out at constant pressure and the only work done is caused by expansion or contraction, q for the change is called the enthalpy change with the symbol ΔH, or ΔH°_{298} for reactions occurring under standard state conditions. The value of ΔH for a reaction in one direction is equal in magnitude, but opposite in sign, to ΔH for the reaction in the opposite direction, and ΔH is directly proportional to the quantity of reactants and products. Examples of enthalpy changes include enthalpy of combustion, enthalpy of fusion, enthalpy of vaporization, and standard enthalpy of formation.

Key Equations

- $\Delta U = q + w$

Footnotes

1. For more on algal fuel, see http://www.theguardian.com/environme...n-fuel-problem.

Glossary

chemical thermodynamics

area of science that deals with the relationships between heat, work, and all forms of energy associated with chemical and physical processes

enthalpy (*H*)

sum of a system’s internal energy and the mathematical product of its pressure and volume

enthalpy change (ΔH)

heat released or absorbed by a system under constant pressure during a chemical or physical process

expansion work (pressure-volume work)

work done as a system expands or contracts against external pressure

first law of thermodynamics

internal energy of a system changes due to heat flow in or out of the system or work done on or by the system

hydrocarbon

compound composed only of hydrogen and carbon; the major component of fossil fuels

internal energy (U)

total of all possible kinds of energy present in a substance or substances

state function

property depending only on the state of a system, and not the path taken to reach that state

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- Adelaide E. Clark, Oregon Institute of Technology
- Crash Course Chemistry: Crash Course is a division of Complexly and videos are free to stream for educational purposes.

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8.4: Standard Enthalpy and Hess' Law

Skills to Develop

- Calculate enthalpy changes for various chemical reactions
- Explain Hess's law and use it to compute reaction enthalpies

Standard Enthalpy of Formation

A standard enthalpy of formation ΔH_f° is an enthalpy change for a reaction in which exactly 1 mole of a pure substance is formed from free elements in their most stable states under standard state conditions. These values are especially useful for computing or predicting enthalpy changes for chemical reactions that are impractical or dangerous to carry out, or for processes for which it is difficult to make measurements. If we have values for the appropriate standard enthalpies of formation, we can determine the enthalpy change for any reaction, which we will practice in the next section on Hess's law.

The standard enthalpy of formation of $CO_2(g)$ is −393.5 kJ/mol. This is the enthalpy change for the exothermic reaction:

$$\mathrm{C}(s) + \mathrm{O_2}(g) \longrightarrow \mathrm{CO_2}(g) \quad \Delta H_f^\circ = \Delta H_{298}^\circ = -393.5\ k\mathrm{J} \tag{8.4.1}$$

starting with the reactants at a pressure of 1 atm and 25 °C (with the carbon present as graphite, the most stable form of carbon under these conditions) and ending with one mole of CO_2, also at 1 atm and 25 °C. For nitrogen dioxide, $\mathrm{NO_2}(g)$, ΔH_f° is 33.2 kJ/mol. This is the enthalpy change for the reaction:

$$\frac{1}{2}\mathrm{N_2}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{NO_2}(g) \quad \Delta H_f^\circ = \Delta H_{298}^\circ = +33.2\ k\mathrm{J} \tag{8.4.2}$$

A reaction equation with $\frac{1}{2}$ mole of N_2 and 1 mole of O_2 is correct in this case because the standard enthalpy of formation always refers to 1 mole of product, $NO_2(g)$.

You will find a table of standard enthalpies of formation of many common substances in Tables T1 and T2. These values indicate that formation reactions range from highly exothermic (such as −2984 kJ/mol for the formation of P_4O_{10}) to strongly endothermic (such as +226.7 kJ/mol for the formation of acetylene, C_2H_2). By definition, the standard enthalpy of formation of an element in its most stable form is equal to zero under standard conditions, which is 1 atm for gases and 1 M for solutions.

Example 8.4.1: Evaluating an Enthalpy of Formation

Ozone, $O_3(g)$, forms from oxygen, $O_2(g)$, by an endothermic process. Ultraviolet radiation is the source of the energy that drives this reaction in the upper atmosphere. Assuming that both the reactants and products of the reaction are in their standard states, determine the standard enthalpy of formation, ΔH_f° of ozone from the following information:

$$3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{O_3}(g) \quad \Delta H_{298}^\circ = +286\ k\mathrm{J}$$

Solution

ΔH_f° is the enthalpy change for the formation of one mole of a substance in its standard state from the elements in their standard states. Thus, ΔH_f° for $O_3(g)$ is the enthalpy change for the reaction:

$$\frac{3}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{O_3}(g)$$

For the formation of 2 mol of $O_3(g)$, $\Delta H_{298}^\circ = +286\ k\mathrm{J}$. This ratio, $\left(\frac{286\ \mathrm{kJ}}{2\ \mathrm{mol\ O_3}}\right)$, can be used as a conversion factor to find the heat produced when 1 mole of $O_3(g)$ is formed, which is the enthalpy of formation for $O_3(g)$:

$$\Delta H^\circ \text{ for 1 mole of } \mathrm{O_3}(g) = 1\ \cancel{\mathrm{mol\ O_3}} \times \frac{286\ \mathrm{kJ}}{2\ \cancel{\mathrm{mol\ O_3}}} = 143\ \mathrm{kJ}$$

Therefore, $\Delta H_f^\circ[\mathrm{O_3}(g)] = +143\ \ k\mathrm{J/mol}$.

Exercise 8.4.1

Hydrogen gas, H_2, reacts explosively with gaseous chlorine, Cl_2, to form hydrogen chloride, HCl(g). What is the enthalpy change for the reaction of 1 mole of $H_2(g)$ with 1 mole of $Cl_2(g)$ if both the reactants and products are at standard state conditions? The standard enthalpy of formation of HCl(g) is −92.3 kJ/mol.

Answer

For the reaction

$$\mathrm{H_2}(g) + \mathrm{Cl_2}(g) \longrightarrow 2\,\mathrm{HCl}(g) \quad \Delta H^\circ_{298} = -184.6\ \mathrm{kJ}$$

Example 8.4.2: Writing Reaction Equations for ΔH_f°

Write the heat of formation reaction equations for:

a. $\mathrm{C_2H_5OH}_{(l)}$
b. $\mathrm{Ca_3(PO_4)_{2(s)}}$

Solution

Remembering that ΔH_f° reaction equations are for forming 1 mole of the compound from its constituent elements under standard conditions, we have:

a. $2\,\mathrm{C}(s,\ \mathrm{graphite}) + 3\,\mathrm{H_2}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{C_2H_5OH}(l)$
b. $3\,\mathrm{Ca}(s) + \frac{1}{2}\mathrm{P_4}(s) + 4\,\mathrm{O_2}(g) \longrightarrow \mathrm{Ca_3(PO_4)_2}(s)$

Note: The standard state of carbon is graphite, and phosphorus exists as P_4.

Exercise 8.4.2

Write the heat of formation reaction equations for:

a. $\mathrm{C_2H_5OC_2H_{5(l)}}$
b. $\mathrm{Na_2CO_{3(s)}}$

Answer a

$4\,\mathrm{C}(s,\ \mathrm{graphite}) + 5\,\mathrm{H_2}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{C_2H_5OC_2H_5}(l)$;

Answer b

$$2\,\mathrm{Na}(s) + \mathrm{C}(s,\ \mathrm{graphite}) + \frac{3}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{Na_2CO_3}(s)$$

Hess's Law

There are two ways to determine the amount of heat involved in a chemical change: measure it experimentally, or calculate it from other experimentally determined enthalpy changes. Some reactions are difficult, if not impossible, to investigate and make accurate measurements for experimentally. And even when a reaction is not hard to perform or measure, it is convenient to be able to determine the heat involved in a reaction without having to perform an experiment.

This type of calculation usually involves the use of Hess's law, which states: If a process can be written as the sum of several stepwise processes, the enthalpy change of the total process equals the sum of the enthalpy changes of the various steps. Hess's law is valid because enthalpy is a state function: Enthalpy changes depend only on where a chemical process starts and ends, but not on the path it takes from start to finish. For example, we can think of the reaction of carbon with oxygen to form carbon dioxide as occurring either directly or by a two-step process. The direct process is written:

$$\mathrm{C}_{(s)} + \mathrm{O}_{2(g)} \longrightarrow \mathrm{CO}_{2(g)} \quad \Delta H^\circ_{298} = -394\ \mathrm{kJ} \tag{8.4.3}$$

In the two-step process, first carbon monoxide is formed:

$$C_{(s)} + \frac{1}{2}O_{2(g)} \longrightarrow CO_{(g)} \quad \Delta H^\circ_{298} = -111\ \text{kJ} \tag{8.4.4}$$

Then, carbon monoxide reacts further to form carbon dioxide:

$$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g) \quad \Delta H^\circ_{298} = -283\ \text{kJ} \tag{8.4.5}$$

The equation describing the overall reaction is the sum of these two chemical changes:

$$\begin{aligned}
&\text{Step 1: } C(s) + \frac{1}{2}O_2(g) \longrightarrow CO(g)\\
&\text{Step 2: } CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)\\
&\text{Sum: } C(s) + \frac{1}{2}O_2(g) + CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO(g) + CO_2(g)
\end{aligned}$$

Because the CO produced in Step 1 is consumed in Step 2, the net change is:

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \tag{8.4.6}$$

According to Hess's law, the enthalpy change of the reaction will equal the sum of the enthalpy changes of the steps. We can apply the data from the experimental enthalpies of combustion in Table **8.4.1** to find the enthalpy change of the entire reaction from its two steps:

$$\begin{aligned}
&C(s) + \frac{1}{2}O_2(g) \longrightarrow CO(g) && \Delta H^\circ_{298} = -111\ \text{kJ}\\
&CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g) && \Delta H^\circ_{298} = -283\ \text{kJ}\\
&\overline{C(s) + O_2(g) \longrightarrow CO_2(g)} && \overline{\Delta H^\circ_{298} = -394\ \text{kJ}}
\end{aligned}$$

The result is shown in Figure **8.4.6**. We see that ΔH of the overall reaction is the same whether it occurs in one step or two. This finding (overall ΔH for the reaction = sum of Δ*H* values for reaction "steps" in the overall reaction) is true in general for chemical and physical processes.

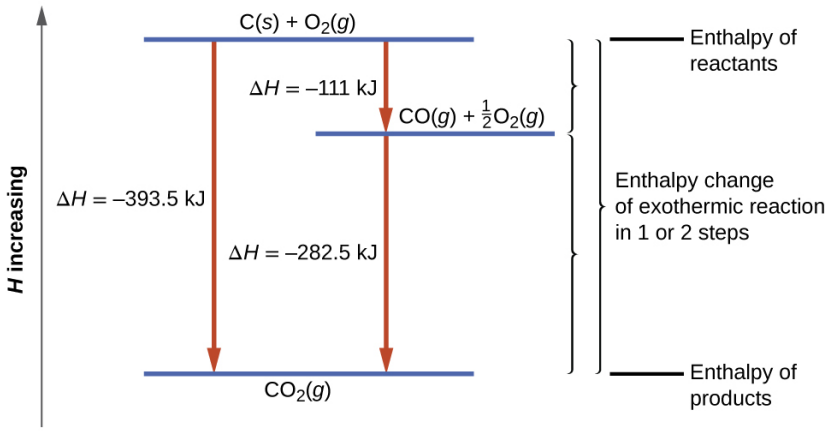


***Figure* 8.4.2:** *The formation of $CO_2(g)$ from its elements can be thought of as occurring in two steps, which sum to the overall reaction, as described by Hess's law. The horizontal blue lines represent enthalpies. For an exothermic process, the products are at lower enthalpy than are the reactants.*

Before we further practice using Hess's law, let us recall two important features of Δ*H*.

1. Δ*H* is directly proportional to the quantities of reactants or products. For example, the enthalpy change for the reaction forming 1 mole of $NO_2(g)$ is +33.2 kJ:

$$\frac{1}{2}N_2(g) + O_2(g) \longrightarrow NO_2(g) \quad \Delta H = +33.2\ \text{kJ} \tag{8.4.7}$$

When 2 moles of NO_2 (twice as much) are formed, the ΔH will be twice as large:

$$\mathrm{N_2}(g) + 2\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{NO_2}(g) \quad \Delta H = +66.4\ \mathrm{kJ} \tag{8.4.8}$$

In general, if we multiply or divide an equation by a number, then the enthalpy change should also be multiplied or divided by the same number.

2. ΔH for a reaction in one direction is equal in magnitude and opposite in sign to ΔH for the reaction in the reverse direction. For example, given that:

$$\mathrm{H_2}(g) + \mathrm{Cl_2}(g) \longrightarrow 2\,\mathrm{HCl}(g) \quad \Delta H = -184.6\ \mathrm{kJ} \tag{8.4.9}$$

Then, for the "reverse" reaction, the enthalpy change is also "reversed":

$$2\,\mathrm{HCl}(g) \longrightarrow \mathrm{H_2}(g) + \mathrm{Cl_2}(g) \quad \Delta H = +184.6\ \mathrm{kJ} \tag{8.4.10}$$

Example 8.4.3: Stepwise Calculation of ΔH_f°

Using Hess's Law Determine the enthalpy of formation, ΔH_f°, of $FeCl_3$(s) from the enthalpy changes of the following two-step process that occurs under standard state conditions:

$$\mathrm{Fe}(s) + \mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_2}(s) \quad \Delta H^\circ = -341.8\ \mathrm{kJ}$$

$$\mathrm{FeCl_2}(s) + \frac{1}{2}\mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_3}(s) \quad \Delta H^\circ = -57.7\ kJ$$

Solution

We are trying to find the standard enthalpy of formation of FeCl3(s), which is equal to ΔH° for the reaction:

$$\mathrm{Fe}(s) + \frac{3}{2}\mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_3}(s) \quad \Delta H_f^\circ = ?$$

Looking at the reactions, we see that the reaction for which we want to find ΔH° is the sum of the two reactions with known ΔH values, so we must sum their ΔHs:

$$\begin{array}{ll} \mathrm{Fe}(s) + \mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_2}(s) & \Delta H^\circ = -341.8\ \mathrm{kJ} \\ \mathrm{FeCl_2}(s) + \frac{1}{2}\mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_3}(s) & \Delta H^\circ = -57.7\ \mathrm{kJ} \\ \hline \mathrm{Fe}(s) + \frac{1}{2}\mathrm{Cl_2}(g) \longrightarrow \mathrm{FeCl_3}(s) & \Delta H^\circ = -399.5\ \mathrm{kJ} \end{array}$$

The enthalpy of formation, ΔH_f°, of $FeCl_3$(s) is −399.5 kJ/mol.

Exercise 8.4.3

Calculate ΔH for the process:

$$\mathrm{N_2}(g) + 2\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{NO_2}(g)$$

from the following information:

$$\mathrm{N_2}(g) + \mathrm{O_2}(g) \longrightarrow 2\,\mathrm{NO}(g) \quad \Delta H = 180.5\ \mathrm{kJ}$$

$$\mathrm{NO}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{NO_2}(g) \quad \Delta H = -57.06\ \mathrm{kJ}$$

Answer

66.4 kJ

Here is a less straightforward example that illustrates the thought process involved in solving many Hess's law problems. It shows how we can find many standard enthalpies of formation (and other values of ΔH) if they are difficult to determine experimentally.

Example 8.4.4: A More Challenging Problem

Using Hess's Law Chlorine monofluoride can react with fluorine to form chlorine trifluoride:

$$\text{(i) } \mathrm{ClF}(g) + \mathrm{F_2}(g) \longrightarrow \mathrm{ClF_3}(g) \quad \Delta H° = \,?$$

Use the reactions here to determine the ΔH° for reaction (i):

$$\text{(ii) } 2\,\mathrm{OF_2}(g) \longrightarrow \mathrm{O_2}(g) + 2\,\mathrm{F_2}(g) \quad \Delta H°_{(ii)} = -49.4\ \mathrm{kJ}$$

$$\text{(iii) } 2\,\mathrm{ClF}(g) + \mathrm{O_2}(g) \longrightarrow \mathrm{Cl_2O}(g) + \mathrm{OF_2}(g) \quad \Delta H°_{(iii)} = +205.6\ \mathrm{kJ}$$

$$\text{(iv) } \mathrm{ClF_3}(g) + \mathrm{O_2}(g) \longrightarrow \tfrac{1}{2}\mathrm{Cl_2O}(g) + \frac{3}{2}\mathrm{OF_2}(g) \quad \Delta H°_{(iv)} = +266.7\ \mathrm{kJ}$$

Solution

Our goal is to manipulate and combine reactions (ii), (iii), and (iv) such that they add up to reaction (i). Going from left to right in (i), we first see that $\mathrm{ClF}_{(g)}$ is needed as a reactant. This can be obtained by multiplying reaction (iii) by $\frac{1}{2}$, which means that the ΔH° change is also multiplied by $\frac{1}{2}$:

$$\mathrm{ClF}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \frac{1}{2}\mathrm{Cl_2O}(g) + \frac{1}{2}\mathrm{OF_2}(g) \quad \Delta H° = \frac{1}{2}(205.6) = +102.8\ k\,\mathrm{J}$$

Next, we see that $\mathrm{F_2}$ is also needed as a reactant. To get this, reverse and halve reaction (ii), which means that the ΔH° changes sign and is halved:

$$\frac{1}{2}\mathrm{O_2}(g) + \mathrm{F_2}(g) \longrightarrow \mathrm{OF_2}(g) \quad \Delta H° = +24.7\ k\,\mathrm{J}$$

To get ClF_3 as a product, reverse (iv), changing the sign of ΔH°:

$$\frac{1}{2}\mathrm{Cl_2O}(g) + \frac{3}{2}\mathrm{OF_2}(g) \longrightarrow \mathrm{ClF_3}(g) + \mathrm{O_2}(g) \quad \Delta H° = -266.7\ \mathrm{kJ}$$

Now check to make sure that these reactions add up to the reaction we want:

$$\begin{array}{ll}
\mathrm{ClF}(g) + \frac{1}{2}\mathrm{O_2}(g) \longrightarrow \frac{1}{2}\mathrm{Cl_2O}(g) + \frac{1}{2}\mathrm{OF_2}(g) & \Delta H° = +102.8\ \mathrm{kJ} \\
\frac{1}{2}\mathrm{O_2}(g) + \mathrm{F_2}(g) \longrightarrow \mathrm{OF_2}(g) & \Delta H° = +24.7\ \mathrm{kJ} \\
\frac{1}{2}\mathrm{Cl_2O}(g) + \frac{3}{2}\mathrm{OF_2}(g) \longrightarrow \mathrm{ClF_3}(g) + \mathrm{O_2}(g) & \Delta H° = -266.7\ \mathrm{kJ} \\
\hline
\mathrm{ClF}(g) + \mathrm{F_2} \longrightarrow \mathrm{ClF_3}(g) & \overline{\Delta H° = -139.2\ \mathrm{kJ}}
\end{array}$$

Reactants $\frac{1}{2}\mathrm{O_2}$ and $\frac{1}{2}\mathrm{O_2}$ cancel out product O_2; product $\frac{1}{2}\mathrm{Cl_2O}$ cancels reactant $\frac{1}{2}\mathrm{Cl_2O}$; and reactant $\frac{3}{2}\mathrm{OF_2}$ is cancelled by products $\frac{1}{2}\mathrm{OF_2}$ and OF_2. This leaves only reactants ClF(g) and F_2(g) and product ClF_3(g), which are what we want. Since summing these three modified reactions yields the reaction of interest, summing the three modified ΔH° values will give the desired ΔH°:

$$\Delta H° = (+102.8\ \mathrm{kJ}) + (24.7\ \mathrm{kJ}) + (-266.7\ \mathrm{kJ}) = -139.2\ \mathrm{kJ}$$

Exercise 8.4.4

Aluminum chloride can be formed from its elements:

$$\text{(i) } 2\,\mathrm{Al}(s) + 3\,\mathrm{Cl_2}(g) \longrightarrow 2\,\mathrm{AlCl_3}(s) \quad \Delta H° = \,?$$

Use the reactions here to determine the ΔH° for reaction (i):

$$\text{(ii) } \mathrm{HCl}(g) \longrightarrow \mathrm{HCl}(aq) \quad \Delta H°_{(ii)} = -74.8\ \mathrm{kJ}$$

$$\text{(iii) } \mathrm{H_2}(g) + \mathrm{Cl_2}(g) \longrightarrow 2\,\mathrm{HCl}(g) \quad \Delta H°_{(iii)} = -185\ \mathrm{kJ}$$

$$\text{(iv) } \mathrm{AlCl_3}(aq) \longrightarrow \mathrm{AlCl_3}(s) \quad \Delta H°_{(iv)} = +323\ \mathrm{kJ/mol}$$

$$\text{(v) } 2\,\mathrm{Al}(s) + 6\,\mathrm{HCl}(aq) \longrightarrow 2\,\mathrm{AlCl_3}(aq) + 3\,\mathrm{H_2}(g) \quad \Delta H°_{(v)} = -1049\ \mathrm{kJ}$$

Answer

−1407 kJ

We also can use Hess's law to determine the enthalpy change of any reaction if the corresponding enthalpies of formation of the reactants and products are available. The stepwise reactions we consider are: (i) decompositions of the reactants into their component elements (for which the enthalpy changes are proportional to the negative of the enthalpies of formation of the reactants), followed by (ii) re-combinations of the elements to give the products (with the enthalpy changes proportional to the enthalpies of formation of the products). The standard enthalpy change of the overall reaction is therefore equal to: (ii) the sum of the standard enthalpies of formation of all the products plus (i) the sum of the negatives of the standard enthalpies of formation of the reactants. This is usually rearranged slightly to be written as follows, with $\sum$ representing "the sum of" and n standing for the stoichiometric coefficients:

$$\Delta H^\circ_{\text{reaction}} = \sum n \times \Delta H^\circ_f(\text{products}) - \sum n \times \Delta H^\circ_f(\text{reactants}) \tag{8.4.11}$$

The following example shows in detail why this equation is valid, and how to use it to calculate the enthalpy change for a reaction of interest.

Example 8.4.5: Using Hess's Law

What is the standard enthalpy change for the reaction:

$$3\,NO_2(g) + H_2O(l) \longrightarrow 2\,HNO_3(aq) + NO(g) \quad \Delta H^\circ = ?$$

Solution 1: Using the Equation

Alternatively, we could use the special form of Hess's law given previously:

$$\Delta H^\circ_{\text{reaction}} = \sum n \times \Delta H^\circ_f(\text{products}) - \sum n \times \Delta H^\circ_f(\text{reactants})$$

$$= \left[2\ \cancel{\text{mol HNO}_3} \times \frac{-207.4\ \text{kJ}}{\cancel{\text{mol HNO}_3\ (aq)}} + 1\ \cancel{\text{mol NO}\ (g)} \times \frac{+90.2\ \text{kJ}}{\cancel{\text{mol NO}\ (g)}}\right]$$

$$- \left[3\ \cancel{\text{mol NO}_2(g)} \times \frac{+33.2\ \text{kJ}}{\cancel{\text{mol NO}_2\ (g)}} + 1\ \cancel{\text{mol H}_2\text{O}\ (l)} \times \frac{-285.8\ \text{kJ}}{\cancel{\text{mol H}_2\text{O}\ (l)}}\right]$$

$$= 2(-207.4\ \text{kJ}) + 1(+90.2\ \text{kJ}) - 3(+33.2\ \text{kJ}) - 1(-285.8\ \text{kJ})$$

$$= -138.4\ \text{kJ}$$

Solution 2: Supporting Why the General Equation Is Valid

We can write this reaction as the sum of the decompositions of $3NO_2(g)$ and $1H_2O(l)$ into their constituent elements, and the formation of 2 $HNO_3(aq)$ and 1 $NO(g)$ from their constituent elements. Writing out these reactions, and noting their relationships to the ΔH°_f values for these compounds (from Tables T1 and T2), we have:

$$3\,NO_2(g) \longrightarrow \frac{3}{2}N_2(g) + 3O_2(g) \quad \Delta H^\circ_1 = -99.6\ \text{kJ}$$

$$H_2O(l) \longrightarrow H_2(g) + \frac{1}{2}O_2(g) \quad \Delta H^\circ_2 = +285.8\ kJ\ [-1 \times \Delta H^\circ_f(H_2O)]$$

$$H_2(g) + N_2(g) + 3O_2(g) \longrightarrow 2\,HNO_3(aq) \quad \Delta H^\circ_3 = -414.8\ kJ\ [2 \times \Delta H^\circ_f(HNO_3)]$$

$$\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \longrightarrow NO(g) \quad \Delta H^\circ_4 = +90.2\ k\,\text{J}\ [1 \times (NO)]$$

Summing these reaction equations gives the reaction we are interested in:

$$3\,NO_2(g) + H_2O(l) \longrightarrow 2\,HNO_3(aq) + NO(g)$$

Summing their enthalpy changes gives the value we want to determine:

$$\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_1 + \Delta H^\circ_2 + \Delta H^\circ_3 + \Delta H^\circ_4 = (-99.6\ \text{kJ}) + (+285.8\ \text{kJ}) + (-414.8\ \text{kJ}) + (+90.2\ \text{kJ})$$
$$= -138.4\ \text{kJ}$$

So the standard enthalpy change for this reaction is $\Delta H^\circ = -138.4$ kJ.

Note that this result was obtained by:

1. multiplying the ΔH_f° of each product by its stoichiometric coefficient and summing those values,
2. multiplying the ΔH_f° of each reactant by its stoichiometric coefficient and summing those values, and then
3. subtracting the result found in step 2 from the result found in step 1.

This is also the procedure in using the general equation, as shown.

Exercise 8.4.5

Calculate the heat of combustion of 1 mole of ethanol, $C_2H_5OH(l)$, when $H_2O(l)$ and $CO_2(g)$ are formed. Use the following enthalpies of formation: $C_2H_5OH(l)$, −278 kJ/mol; $H_2O(l)$, −286 kJ/mol; and $CO_2(g)$, −394 kJ/mol.

Answer

−1368 kJ/mol

Summary



Video* 8.4.1*: *A summary of Enthalpy and Hess' Law from Crash Course Chemistry.*

If a chemical change is carried out at constant pressure and the only work done is caused by expansion or contraction, q for the change is called the enthalpy change with the symbol ΔH, or ΔH_{298}° for reactions occurring under standard state conditions. The value of ΔH for a reaction in one direction is equal in magnitude, but opposite in sign, to ΔH for the reaction in the opposite direction, and ΔH is directly proportional to the quantity of reactants and products. Examples of enthalpy changes include enthalpy of combustion, enthalpy of fusion, enthalpy of vaporization, and standard enthalpy of formation. The standard enthalpy of formation, ΔH_f°, is the enthalpy change accompanying the formation of 1 mole of a substance from the elements in their most stable states at 1 bar (standard state). Many of the processes are carried out at 298.15 K. If the enthalpies of formation are available for the reactants and products of a reaction, the enthalpy change can be calculated using Hess's law: If a process can be written as the sum of several stepwise processes, the enthalpy change of the total process equals the sum of the enthalpy changes of the various steps.

Key Equations

- $\Delta U = q + w$
- $\Delta H_{\text{reaction}}^\circ = \sum n \times \Delta H_f^\circ(\text{products}) - \sum n \times \Delta H_f^\circ(\text{reactants})$

Footnotes

1. For more on algal fuel, see http://www.theguardian.com/environme...n-fuel-problem.

Glossary

chemical thermodynamics

area of science that deals with the relationships between heat, work, and all forms of energy associated with chemical and physical processes

enthalpy (*H*)
sum of a system's internal energy and the mathematical product of its pressure and volume

enthalpy change (ΔH)
heat released or absorbed by a system under constant pressure during a chemical or physical process

expansion work (pressure-volume work)
work done as a system expands or contracts against external pressure

first law of thermodynamics
internal energy of a system changes due to heat flow in or out of the system or work done on or by the system

Hess's law
if a process can be represented as the sum of several steps, the enthalpy change of the process equals the sum of the enthalpy changes of the steps

hydrocarbon
compound composed only of hydrogen and carbon; the major component of fossil fuels

internal energy (*U*)
total of all possible kinds of energy present in a substance or substances

standard enthalpy of combustion (ΔH_c°)
heat released when one mole of a compound undergoes complete combustion under standard conditions

standard enthalpy of formation (ΔH_f°)
enthalpy change of a chemical reaction in which 1 mole of a pure substance is formed from its elements in their most stable states under standard state conditions

standard state
set of physical conditions as accepted as common reference conditions for reporting thermodynamic properties; 1 bar of pressure, and solutions at 1 molar concentrations, usually at a temperature of 298.15 K

state function
property depending only on the state of a system, and not the path taken to reach that state

Contributors

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- Adelaide E. Clark, Oregon Institute of Technology
- Crash Course Chemistry: Crash Course is a division of Complexly and videos are free to stream for educational purposes.

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8.4.1: Practice Problems- Enthalpy and Hess' Law

PROBLEM 8.4.1.1

How much heat is produced by burning 4.00 moles of acetylene under standard state conditions?

Answer

5204.4 kJ

PROBLEM 8.4.1.2

How much heat is produced by combustion of 125 g of methanol (CH_3OH) under standard state conditions?

Answer

2836.3 kJ

Click here to see a video of the solution



The section number changed after this video was made

PROBLEM 8.4.1.3

How many moles of isooctane must be burned to produce 100 kJ of heat under standard state conditions?

Answer

1.83×10^{-2} mol

PROBLEM 8.4.1.4

What mass of carbon monoxide must be burned to produce 175 kJ of heat under standard state conditions?

Answer

17.3 g

Click here to see a video of the solution



***The section number ch**

PROBLEM 8.4.1.5

When 2.50 g of methane bu ion per mole of methane under these conditions?

Answer

802 kJ mol^{-1}

PROBLEM 8.4.1.6

a. How much heat is produced when 100 mL of 0.250 M HCl (density, 1.00 g/mL) and 200 mL of 0.150 M NaOH (density, 1.00 g/mL) are mixed? (Refer to Example 8.2.1 for how the heat of this reaction was derived)

$$\mathrm{HCl}(aq) + \mathrm{NaOH}(aq) \longrightarrow \mathrm{NaCl}(aq) + \mathrm{H_2O}(l) \quad \Delta H^\circ_{298} = -58\ \mathrm{kJ/mol} \tag{8.4.1.1}$$

b. If both solutions are at the same temperature and the heat capacity of the products is 4.19 J/g °C, how much will the temperature increase? What assumption did you make in your calculation?

Answer a

-1.45 kJ

Answer b

-1.15 °C, assuming the products have the same density as the reactants (1.00 g/mL)

Click here to see a video of the solution

The section number changed after this video was made

PROBLEM 8.4.1.7

A sample of 0.562 g of carbon is burned in oxygen in a bomb calorimeter, producing carbon dioxide. Assume both the reactants and products are under standard state conditions, and that the heat released is directly proportional to the enthalpy of combustion of graphite. The temperature of the calorimeter increases from 26.74 °C to 27.93 °C. What is the heat capacity of the calorimeter and its contents?

Answer

15.5 kJ/ºC

PROBLEM 8.4.1.8

Homes may be heated by pumping hot water through radiators. What mass of water will provide the same amount of heat when cooled from 95.0 to 35.0 °C, as the heat provided when 100 g of steam is cooled from 110 °C to 100 °C.

Answer

7.43 g

PROBLEM 8.4.1.9

The following sequence of reactions occurs in the commercial production of aqueous nitric acid:

$$4\,NH_3(g) + 5\,O_2(g) \longrightarrow 4\,NO(g) + 6\,H_2O(l) \quad \Delta H = -907\,kJ$$

$$2\,NO(g) + O_2(g) \longrightarrow 2\,NO_2(g) \quad \Delta H = -113\,kJ$$

$$3\,NO_2 + H_2O(l) \longrightarrow 2\,HNO_2(aq) + NO(g) \quad \Delta H = -139\,kJ$$

Determine the total energy change for the production of one mole of aqueous nitric acid by this process.

Answer

495 kJ/mol

PROBLEM 8.4.1.10

Both graphite and diamond burn.

$$C(s,\ \text{diamond}) + O_2(g) \longrightarrow CO_2(g)$$

For the conversion of graphite to diamond:

$$C(s,\ \text{graphite}) \longrightarrow C(s,\ \text{diamond}) \quad \Delta H^\circ_{298} = 1.90\,kJ$$

Which produces more heat, the combustion of graphite or the combustion of diamond? (Hint: The heats of formation for all these compounds can be found in Table T1.

Answer

Diamond

PROBLEM 8.4.1.11

Calculate ΔH for the process

$$Hg_2Cl_2(s) \longrightarrow 2\,Hg(l) + Cl_2(g)$$

from the following information:

$$Hg(l) + Cl_2(g) \longrightarrow HgCl_2(s) \quad \Delta H = -224\,kJ$$

$$Hg(l) + HgCl_2(s) \longrightarrow Hg_2Cl_2(s) \quad \Delta H = -41.2\,kJ$$

Answer

265 kJ

PROBLEM 8.4.1.12

Using the data in Table T1, calculate the standard enthalpy change for each of the following reactions:

a. $N_2(g) + O_2(g) \longrightarrow 2\,NO(g)$
b. $Si(s) + 2\,Cl_2(g) \longrightarrow SiCl_4(g)$
c. $Fe_2O_3(s) + 3\,H_2(g) \longrightarrow 2\,Fe(s) + 3\,H_2O(l)$

d. $2\,LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(s) + H_2O(g)$ (Hint: For $LiOH_{(s)}$, ΔH_f = -487.5 kJ/mol; For $Li_2CO_{3\,(s)}$, ΔH_f = -1216.04 kJ/mol)

Answer

182.6 kJ

Answer

-657.0 kJ mol^{-1}

Answer

-33.2 kJ

Answer

-89.34 kJ

Click here to see a video of the solution



The section number changed after this video was made

PROBLEM 8.4.1.13

Using the data in Table T1, calculate the standard enthalpy change for each of the following reactions:

a. $Si(s) + 2\,F_2(g) \longrightarrow SiF_4(g)$ (Hint: For $SiF_{4\,(g)}$, ΔH_f = -1615.0 kJ/mol)
b. $2\,C(s) + 2\,H_2(g) + O_2(g) \longrightarrow CH_3CO_2H(l)$ (Hint: For $CH_3CO_2H_{(l)}$, ΔH_f = -484.3 kJ/mol)
c. $CH_4(g) + N_2(g) \longrightarrow HCN(g) + NH_3(g)$;
d. $CS_2(g) + 3\,Cl_2(g) \longrightarrow CCl_4(g) + S_2Cl_2(g)$ (Hint: For $S_2Cl_{2(g)}$, ΔH_f = -19.5 kJ/mol)

Answer a

$-1615.0\ kJ\ mol^{-1}$

Answer b

$-484.3\ kJ\ mol^{-1}$

Answer c

164.2 kJ

Answer d

−232.1 kJ

Contributors

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- Adelaide Clark, Oregon Institute of Technology

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8.5: Calorimetry

Skills to Develop

- Explain the technique of calorimetry
- Calculate and interpret heat and related properties using typical calorimetry data

One technique we can use to measure the amount of heat involved in a chemical or physical process is known as calorimetry. Calorimetry is used to measure amounts of heat transferred to or from a substance. To do so, the heat is exchanged with a calibrated object (calorimeter). The change in temperature of the measuring part of the calorimeter is converted into the amount of heat (since the previous calibration was used to establish its heat capacity). The measurement of heat transfer using this approach requires the definition of a system (the substance or substances undergoing the chemical or physical change) and its surroundings (the other components of the measurement apparatus that serve to either provide heat to the system or absorb heat from the system). Knowledge of the heat capacity of the surroundings, and careful measurements of the masses of the system and surroundings and their temperatures before and after the process allows one to calculate the heat transferred as described in this section.

A calorimeter is a device used to measure the amount of heat involved in a chemical or physical process. For example, when an exothermic reaction occurs in solution in a calorimeter, the heat produced by the reaction is absorbed by the solution, which increases its temperature. When an endothermic reaction occurs, the heat required is absorbed from the thermal energy of the solution, which decreases its temperature (Figure 8.5.1). The temperature change, along with the specific heat and mass of the solution, can then be used to calculate the amount of heat involved in either case.

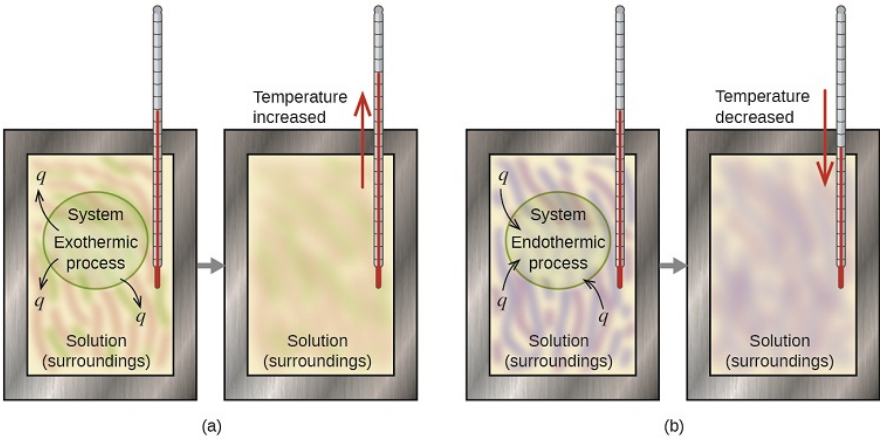


***Figure* 8.5.1:** *In a calorimetric determination, either (a) an exothermic process occurs and heat, q, is negative, indicating that thermal energy is transferred from the system to its surroundings, or (b) an endothermic process occurs and heat, q, is positive, indicating that thermal energy is transferred from the surroundings to the system.*

By convention, q is given a negative (-) sign when the system releases heat to the surroundings (exothermic); q is given a positive (+) sign when the system absorbs heat from the surroundings (endothermic).

Scientists use well-insulated calorimeters that all but prevent the transfer of heat between the calorimeter and its environment. This enables the accurate determination of the heat involved in chemical processes, the energy content of foods, and so on. General chemistry students often use simple calorimeters constructed from polystyrene cups (Figure 8.5.2). These easy-to-use “coffee cup” calorimeters allow more heat exchange with their surroundings, and therefore produce less accurate energy values.

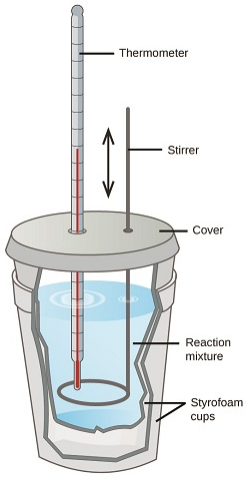


Figure* 8.5.2: A simple calorimeter can be constructed from two polystyrene cups. A thermometer and stirrer extend through the cover into the reaction mixture.*

Commercial solution calorimeters are also available. Relatively inexpensive calorimeters often consist of two thin-walled cups that are nested in a way that minimizes thermal contact during use, along with an insulated cover, handheld stirrer, and simple thermometer. More expensive calorimeters used for industry and research typically have a well-insulated, fully enclosed reaction vessel, motorized stirring mechanism, and a more accurate temperature sensor (Figure **8.5.3**).

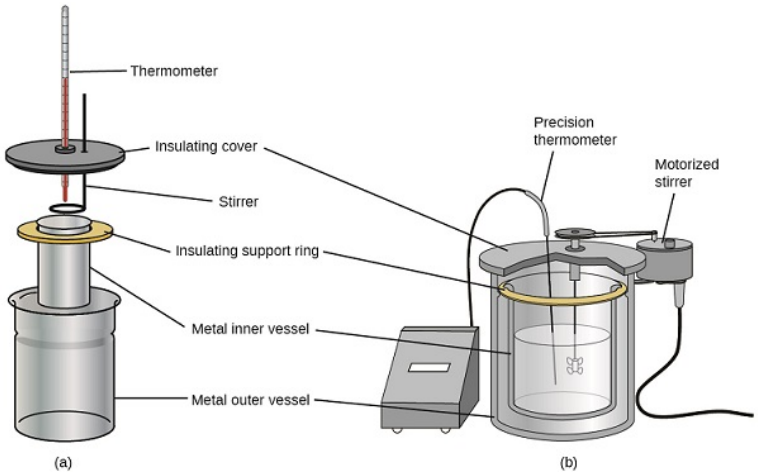


Figure* 8.5.3: Commercial solution calorimeters range from (a) simple, inexpensive models for student use to (b) expensive, more accurate models for industry and research.*

Before we practice calorimetry problems involving chemical reactions, consider a simple example that illustrates the core idea behind calorimetry. Suppose we initially have a high-temperature substance, such as a hot piece of metal (M), and a low-temperature substance, such as cool water (W). If we place the metal in the water, heat will flow from M to W. The temperature of M will decrease, and the temperature of W will increase, until the two substances have the same temperature—that is, when they reach thermal equilibrium (Figure **8.5.4**). If this occurs in a calorimeter, ideally all of this heat transfer occurs between the two substances, with no heat gained or lost by either the calorimeter or the calorimeter's surroundings. Under these ideal circumstances, the net heat change is zero:

$$q_{\text{substance M}} + q_{\text{substance W}} = 0 \tag{8.5.1}$$

This relationship can be rearranged to show that the heat gained by substance M is equal to the heat lost by substance W:

$$q_{\text{substance M}} = -q_{\text{substance W}} \tag{8.5.2}$$

The magnitude of the heat (change) is therefore the same for both substances, and the negative sign merely shows that $q_{\text{substance}}$ M and $q_{\text{substance}}$ W are opposite in direction of heat flow (gain or loss) but does not indicate the arithmetic sign of either q value (that is determined by whether the matter in question gains or loses heat, per definition). In the specific situation described, $q_{\text{substance}}$ M is a negative value and $q_{\text{substance}}$ W is positive, since heat is transferred from M to W.

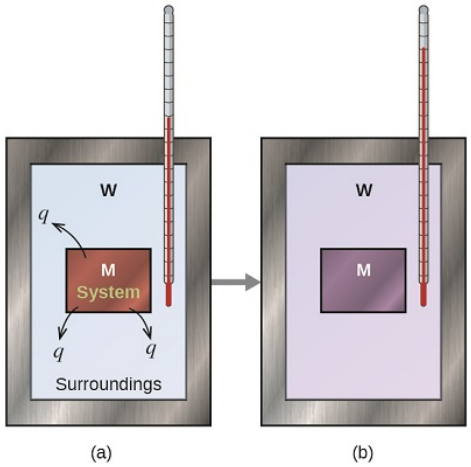


Figure* 8.5.4*: *In a simple calorimetry process, (a) heat, q, is transferred from the hot metal, M, to the cool water, W, until (b) both are at the same temperature.*

Example 8.5.1: Heat Transfer between Substances at Different Temperatures

A hot 360-g piece of rebar (a steel rod used for reinforcing concrete) is dropped into 425 mL of water at 24.0 °C. The final temperature of the water is measured as 42.7 °C. Calculate the initial temperature of the piece of rebar. Assume the specific heat of steel is approximately the same as that for iron (Table T4), and that all heat transfer occurs between the rebar and the water (there is no heat exchange with the surroundings).

Solution

The temperature of the water increases from 24.0 °C to 42.7 °C, so the water absorbs heat. That heat came from the piece of rebar, which initially was at a higher temperature. Assuming that all heat transfer was between the rebar and the water, with no heat "lost" to the surroundings, then heat given off by rebar = − heat taken in by water, or:

$$q_{\text{rebar}} = -q_{\text{water}} \tag{8.5.3}$$

Since we know how heat is related to other measurable quantities, we have:

$$(c \times m \times \Delta T)_{\text{rebar}} = -(c \times m \times \Delta T)_{\text{water}} \tag{8.5.4}$$

Letting f = final and i = initial, in expanded form, this becomes:

$$c_{\text{rebar}} \times m_{\text{rebar}} \times (T_{\text{f,rebar}} - T_{\text{i,rebar}}) = -c_{\text{water}} \times m_{\text{water}} \times (T_{\text{f,water}} - T_{\text{i,water}}) \tag{8.5.5}$$

The density of water is 1.0 g/mL, so 425 mL of water = 425 g. Noting that the final temperature of both the rebar and water is 42.7 °C, substituting known values yields:

$$(0.449\ \text{J/g °C})(360\text{g})(42.7\text{°C} - T_{\text{i,rebar}}) = -(4.184\ \text{J/g °C})(425\ \text{g})(42.7\text{°C} - 24.0\text{°C}) \tag{8.5.6}$$

$$T_{\text{i,rebar}} = \frac{(4.184\ \text{J/g °C})(425\ \text{g})(42.7\text{°C} - 24.0\text{°C})}{(0.449\ \text{J/g °C})(360\ \text{g})} + 42.7\text{°C} \tag{8.5.7}$$

Solving this gives Ti,rebar= 248 °C, so the initial temperature of the rebar was 248 °C.

Exercise 8.5.1*A*

A 248-g piece of copper is dropped into 390 mL of water at 22.6 °C. The final temperature of the water was measured as 39.9 °C. Calculate the initial temperature of the piece of copper. Assume that all heat transfer occurs between the copper and the water.

Answer

The initial temperature of the copper was 335.6 °C.

Exercise 8.5.1*B*

A 248-g piece of copper initially at 314 °C is dropped into 390 mL of water initially at 22.6 °C. Assuming that all heat transfer occurs between the copper and the water, calculate the final temperature.

Answer

The final temperature (reached by both copper and water) is 38.7 °C.

This method can also be used to determine other quantities, such as the specific heat of an unknown metal.

Example 8.5.2: Identifying a Metal by Measuring Specific Heat

A 59.7 g piece of metal that had been submerged in boiling water was quickly transferred into 60.0 mL of water initially at 22.0 °C. The final temperature is 28.5 °C. Use these data to determine the specific heat of the metal. Use this result to identify the metal.

Solution

Assuming perfect heat transfer, heat given off by metal = −heat taken in by water, or:

$$q_{\text{metal}} = -q_{\text{water}} \tag{8.5.8}$$

In expanded form, this is:

$$c_{\text{metal}} \times m_{\text{metal}} \times (T_{\text{f,metal}} - T_{\text{i,metal}}) = -c_{\text{water}} \times m_{\text{water}} \times (T_{\text{f,water}} - T_{\text{i,water}}) \tag{8.5.9}$$

Noting that since the metal was submerged in boiling water, its initial temperature was 100.0 °C; and that for water, 60.0 mL = 60.0 g; we have:

$$(c_{\text{metal}})(59.7\ \text{g})(28.5°\text{C} - 100.0°\text{C}) = -(4.18\ \text{J/g °C})(60.0\ \text{g})(28.5°\text{C} - 22.0°\text{C}) \tag{8.5.10}$$

Solving this:

$$c_{\text{metal}} = \frac{-(4.184\ \text{J/g °C})(60.0\ \text{g})(6.5°\text{C})}{(59.7\ \text{g})(-71.5°\text{C})} = 0.38\ \text{J/g °C} \tag{8.5.11}$$

Comparing this with values in Table T4, our experimental specific heat is closest to the value for copper (0.39 J/g °C), so we identify the metal as copper.

Exercise 8.5.2

A 92.9-g piece of a silver/gray metal is heated to 178.0 °C, and then quickly transferred into 75.0 mL of water initially at 24.0 °C. After 5 minutes, both the metal and the water have reached the same temperature: 29.7 °C. Determine the specific heat and the identity of the metal. (Note: You should find that the specific heat is close to that of two different metals. Explain how you can confidently determine the identity of the metal).

Answer

$c_{metal} = 0.13\ J/g\ °C$

This specific heat is close to that of either gold or lead. It would be difficult to determine which metal this was based solely on the numerical values. However, the observation that the metal is silver/gray in addition to the value for the specific heat indicates that the metal is lead.

When we use calorimetry to determine the heat involved in a chemical reaction, the same principles we have been discussing apply. The amount of heat absorbed by the calorimeter is often small enough that we can neglect it (though not for highly accurate measurements, as discussed later), and the calorimeter minimizes energy exchange with the surroundings. Because energy is neither created nor destroyed during a chemical reaction, there is no overall energy change during the reaction. The heat produced or consumed in the reaction (the "system"), q_{reaction}, plus the heat absorbed or lost by the solution (the "surroundings"), q_{solution}, must add up to zero:

$$q_{\text{reaction}} + q_{\text{solution}} = 0 \tag{8.5.12}$$

This means that the amount of heat produced or consumed in the reaction equals the amount of heat absorbed or lost by the solution:

$$q_{\text{reaction}} = -q_{\text{solution}} \tag{8.5.13}$$

This concept lies at the heart of all calorimetry problems and calculations.

Example 8.5.3: Heat Produced by an Exothermic Reaction

When 50.0 mL of 0.10 M HCl(aq) and 50.0 mL of 1.00 M NaOH(aq), both at 22.0 °C, are added to a coffee cup calorimeter, the temperature of the mixture reaches a maximum of 28.9 °C. What is the approximate amount of heat produced by this reaction?

$$\text{HCl}(aq) + \text{NaOH}(aq) \longrightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)$$

Solution

To visualize what is going on, imagine that you could combine the two solutions so quickly that no reaction took place while they mixed; then after mixing, the reaction took place. At the instant of mixing, you have 100.0 mL of a mixture of HCl and NaOH at 22.0 °C. The HCl and NaOH then react until the solution temperature reaches 28.9 °C.

The heat given off by the reaction is equal to that taken in by the solution. Therefore:

$$q_{\text{reaction}} = -q_{\text{solution}}$$

(It is important to remember that this relationship only holds if the calorimeter does not absorb any heat from the reaction, and there is no heat exchange between the calorimeter and its surroundings.)

Next, we know that the heat absorbed by the solution depends on its specific heat, mass, and temperature change:

$$q_{\text{solution}} = (c \times m \times \Delta T)_{\text{solution}}$$

To proceed with this calculation, we need to make a few more reasonable assumptions or approximations. Since the solution is aqueous, we can proceed as if it were water in terms of its specific heat and mass values. The density of water is approximately 1.0 g/mL, so 100.0 mL has a mass of about 1.0×10^2 g (two significant figures). The specific heat of water is approximately 4.18 J/g °C, so we use that for the specific heat of the solution. Substituting these values gives:

$$q_{\text{solution}} = (4.184\ \text{J/g °C})(1.0 \times 10^2\ \text{g})(28.9\text{°C} - 22.0\text{°C}) = 2.89 \times 10^3\ \text{J} \tag{8.5.14}$$

Finally, since we are trying to find the heat of the reaction, we have:

$$q_{\text{reaction}} = -q_{\text{solution}} = -2.89 \times 10^3\ J$$

The negative sign indicates that the reaction is exothermic. It produces 2.89 kJ of heat.

Exercise 8.5.3

When 100 mL of 0.200 M NaCl(aq) and 100 mL of 0.200 M $AgNO_3$(aq), both at 21.9 °C, are mixed in a coffee cup calorimeter, the temperature increases to 23.5 °C as solid AgCl forms. How much heat is produced by this precipitation reaction? What assumptions did you make to determine your value?

Answer

$1.34 \times 10^3\ J$; assume no heat is absorbed by the calorimeter, no heat is exchanged between the calorimeter and its surroundings, and that the specific heat and mass of the solution are the same as those for water

Thermochemistry of Hand Warmers

When working or playing outdoors on a cold day, you might use a hand warmer to warm your hands (Figure 8.5.5). A common reusable hand warmer contains a supersaturated solution of $NaC_2H_3O_2$ (sodium acetate) and a metal disc. Bending the disk creates nucleation sites around which the metastable $NaC_2H_3O_2$ quickly crystallizes (a later chapter on solutions will investigate saturation and supersaturation in more detail).

The process $\mathrm{NaC_2H_3O_2}(aq) \longrightarrow \mathrm{NaC_2H_3O_2}(s)$ is exothermic, and the heat produced by this process is absorbed by your hands, thereby warming them (at least for a while). If the hand warmer is reheated, the $NaC_2H_3O_2$ redissolves and can be reused.



***Figure* 8.5.5:** *Chemical hand warmers produce heat that warms your hand on a cold day. In this one, you can see the metal disc that initiates the exothermic precipitation reaction. (credit: modification of work by Science Buddies TV/YouTube)*

Another common hand warmer produces heat when it is ripped open, exposing iron and water in the hand warmer to oxygen in the air. One simplified version of this exothermic reaction is

$$2\,\mathrm{Fe}(s) + \frac{3}{2}\mathrm{O_2}(g) \longrightarrow \mathrm{Fe_2O_3}(s). \tag{8.5.15}$$

Salt in the hand warmer catalyzes the reaction, so it produces heat more rapidly; cellulose, vermiculite, and activated carbon help distribute the heat evenly. Other types of hand warmers use lighter fluid (a platinum catalyst helps lighter fluid oxidize exothermically), charcoal (charcoal oxidizes in a special case), or electrical units that produce heat by passing an electrical current from a battery through resistive wires.

Example 8.5.4: Heat Flow in an Instant Ice Pack

When solid ammonium nitrate dissolves in water, the solution becomes cold. This is the basis for an "instant ice pack" (Figure 8.5.5). When 3.21 g of solid NH_4NO_3 dissolves in 50.0 g of water at 24.9 °C in a calorimeter, the temperature decreases to 20.3 °C.

Calculate the value of *q* for this reaction and explain the meaning of its arithmetic sign. State any assumptions that you made.



***Figure* 8.5.5:** *An instant cold pack consists of a bag containing solid ammonium nitrate and a second bag of water. When the bag of water is broken, the pack becomes cold because the dissolution of ammonium nitrate is an endothermic process that removes thermal energy from the water. The cold pack then removes thermal energy from your body.*

Solution

We assume that the calorimeter prevents heat transfer between the solution and its external environment (including the calorimeter itself), in which case:

$$q_{rxn} = -q_{soln} \tag{8.5.16}$$

with "rxn" and "soln" used as shorthand for "reaction" and "solution," respectively.

Assuming also that the specific heat of the solution is the same as that for water, we have:

$$\begin{aligned} q_{rxn} &= -q_{soln} = -(c \times m \times \Delta T)_{soln} \\ &= -[(4.184 J/g\,°C) \times (53.2\,g) \times (20.3°C - 24.9°C)] \\ &= -[(4.184 J/g\,°C) \times (53.2\,g) \times (-4.6°C)] \\ &+ 1.0 \times 10^3\,J = +1.0\,kJ \end{aligned}$$

The positive sign for *q* indicates that the dissolution is an endothermic process.

Exercise 8.5.4

When a 3.00-g sample of KCl was added to 3.00 × 10^2 g of water in a coffee cup calorimeter, the temperature decreased by 1.05 °C. How much heat is involved in the dissolution of the KCl? What assumptions did you make?

Answer

1.33 kJ; assume that the calorimeter prevents heat transfer between the solution and its external environment (including the calorimeter itself) and that the specific heat of the solution is the same as that for water.

If the amount of heat absorbed by a calorimeter is too large to neglect or if we require more accurate results, then we must take into account the heat absorbed both by the solution and by the calorimeter.

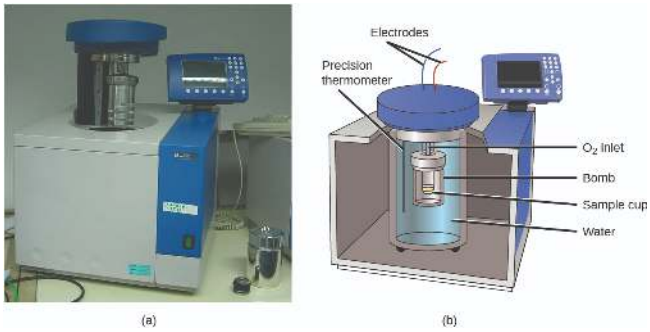
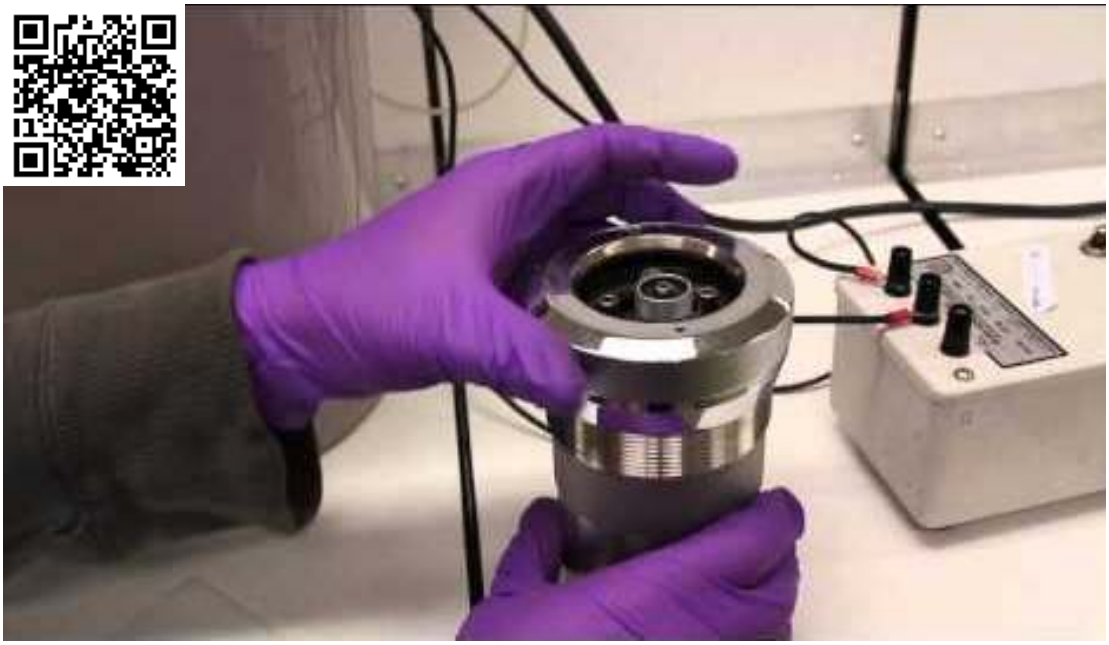


***Figure* 8.5.6:** *(a) A bomb calorimeter is used to measure heat produced by reactions involving gaseous reactants or products, such as combustion. (b) The reactants are contained in the gas-tight "bomb," which is submerged in water and surrounded by insulating materials. (credit a: modification of work by "Harbor1"/Wikimedia commons)*

The calorimeters described are designed to operate at constant (atmospheric) pressure and are convenient to measure heat flow accompanying processes that occur in solution. A different type of calorimeter that operates at constant volume, colloquially known as a bomb calorimeter, is used to measure the energy produced by reactions that yield large amounts of heat and gaseous products, such as combustion reactions. (The term "bomb" comes from the observation that these reactions can be vigorous enough to resemble explosions that would damage other calorimeters.) This type of calorimeter consists of a robust steel container (the "bomb") that contains the reactants and is itself submerged in water (Figure 8.5.6). The sample is placed in the bomb, which is then filled with oxygen at high pressure. A small electrical spark is used to ignite the sample. The energy produced by the reaction is trapped in the steel bomb and the surrounding water. The temperature increase is measured and, along with the known heat capacity of the calorimeter, is used to calculate the energy produced by the reaction. Bomb calorimeters require calibration to determine the heat capacity of the calorimeter and ensure accurate results. The calibration is accomplished using a reaction with a known q, such as a measured quantity of benzoic acid ignited by a spark from a nickel fuse wire that is weighed before and after the reaction. The temperature change produced by the known reaction is used to determine the heat capacity of the calorimeter. The calibration is generally performed each time before the calorimeter is used to gather research data.



***Video* 8.5.1:** *Video of view how a bomb calorimeter is prepared for action.*

Example 8.5.5: Bomb Calorimetry

When 3.12 g of glucose, $C_6H_{12}O_6$, is burned in a bomb calorimeter, the temperature of the calorimeter increases from 23.8 °C to 35.6 °C. The calorimeter contains 775 g of water, and the bomb itself has a heat capacity of 893 J/°C. How much heat was produced by the combustion of the glucose sample?

Solution

The combustion produces heat that is primarily absorbed by the water and the bomb. (The amounts of heat absorbed by the reaction products and the unreacted excess oxygen are relatively small and dealing with them is beyond the scope of this text. We will neglect them in our calculations.)

The heat produced by the reaction is absorbed by the water and the bomb:

$$\begin{aligned} q_{\text{rxn}} &= -(q_{\text{water}} + q_{\text{bomb}}) \\ &= -[(4.184\ \text{J/g °C}) \times (775\ \text{g}) \times (35.6\text{°C} - 23.8\text{°C}) + 893\ \text{J/°C} \times (35.6\text{°C} - 23.8\text{°C})] \\ &= -(38,300\ \text{J} + 10,500\ \text{J}) \\ &= -48,800\ \text{J} = -48.8\ \text{kJ} \end{aligned}$$

This reaction released 48.7 kJ of heat when 3.12 g of glucose was burned.

Exercise 8.5.5

When 0.963 g of benzene, C_6H_6, is burned in a bomb calorimeter, the temperature of the calorimeter increases by 8.39 °C. The bomb has a heat capacity of 784 J/°C and is submerged in 925 mL of water. How much heat was produced by the combustion of the glucose sample?

Answer

39.0 kJ

Since the first one was constructed in 1899, 35 calorimeters have been built to measure the heat produced by a living person.[1] These whole-body calorimeters of various designs are large enough to hold an individual human being. More recently, whole-room calorimeters allow for relatively normal activities to be performed, and these calorimeters generate data that more closely reflect the real world. These calorimeters are used to measure the metabolism of individuals under different environmental conditions, different dietary regimes, and with different health conditions, such as diabetes. In humans, metabolism is typically measured in Calories per day. A nutritional calorie (Calorie) is the energy unit used to quantify the amount of energy derived from the metabolism of foods; one Calorie is equal to 1000 calories (1 kcal), the amount of energy needed to heat 1 kg of water by 1 °C.

Measuring Nutritional Calories

In your day-to-day life, you may be more familiar with energy being given in Calories, or nutritional calories, which are used to quantify the amount of energy in foods. One calorie (cal) = exactly 4.184 joules, and one Calorie (note the capitalization) = 1000 cal, or 1 kcal. (This is approximately the amount of energy needed to heat 1 kg of water by 1 °C.)

The macronutrients in food are proteins, carbohydrates, and fats or oils. Proteins provide about 4 Calories per gram, carbohydrates also provide about 4 Calories per gram, and fats and oils provide about 9 Calories/g. Nutritional labels on food packages show the caloric content of one serving of the food, as well as the breakdown into Calories from each of the three macronutrients (Figure 8.5.7).

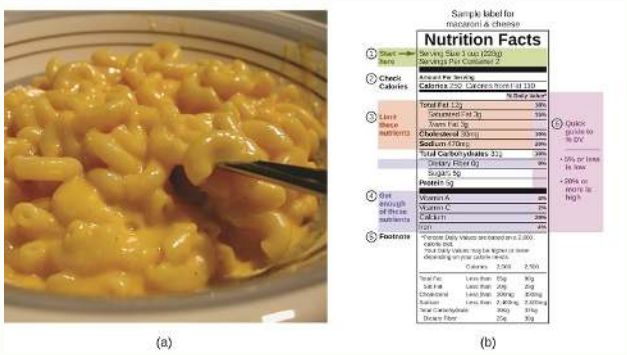


***Figure 8.5.7**: (a) Macaroni and cheese contain energy in the form of the macronutrients in the food. (b) The food's nutritional information is shown on the package label. In the US, the energy content is given in Calories (per serving); the rest of the world usually uses kilojoules. (credit a: modification of work by "Rex Roof"/Flickr)*

For the example shown in (b), the total energy per 228-g portion is calculated by:

$$(5\text{ g protein} \times 4\text{ Calories/g}) + (31\text{ g carb} \times 4\text{ Calories/g}) + (12\text{ g fat} \times 9\text{ Calories/g}) = 252\text{ Calories} \quad (8.5.17)$$

So, you can use food labels to count your Calories. But where do the values come from? And how accurate are they? The caloric content of foods can be determined by using bomb calorimetry; that is, by burning the food and measuring the energy it contains. A sample of food is weighed, mixed in a blender, freeze-dried, ground into powder, and formed into a pellet. The pellet is burned inside a bomb calorimeter, and the measured temperature change is converted into energy per gram of food.

Today, the caloric content on food labels is derived using a method called the Atwater system that uses the average caloric content of the different chemical constituents of food, protein, carbohydrate, and fats. The average amounts are those given in the equation and are derived from the various results given by bomb calorimetry of whole foods. The carbohydrate amount is discounted a certain amount for the fiber content, which is indigestible carbohydrate. To determine the energy content of a food, the quantities of carbohydrate, protein, and fat are each multiplied by the average Calories per gram for each and the products summed to obtain the total energy.

Summary



Video* 8.5.2: A Summary of Calorimetry.*

Calorimetry is used to measure the amount of thermal energy transferred in a chemical or physical process. This requires careful measurement of the temperature change that occurs during the process and the masses of the system and surroundings. These measured quantities are then used to compute the amount of heat produced or consumed in the process using known mathematical relations. Calorimeters are designed to minimize energy exchange between the system being studied and its surroundings. They range from simple coffee cup calorimeters used by introductory chemistry students to sophisticated bomb calorimeters used to determine the energy content of food.

Footnotes

1. 1 Francis D. Reardon et al. “The Snellen human calorimeter revisited, re-engineered and upgraded: Design and performance characteristics.” *Medical and Biological Engineering and Computing* 8 (2006)721–28, http://link.springer.com/article/10....517-006-0086-5.

Glossary

bomb calorimeter
: device designed to measure the energy change for processes occurring under conditions of constant volume; commonly used for reactions involving solid and gaseous reactants or products

calorimeter
: device used to measure the amount of heat absorbed or released in a chemical or physical process

calorimetry
: process of measuring the amount of heat involved in a chemical or physical process

nutritional calorie (Calorie)
: unit used for quantifying energy provided by digestion of foods, defined as 1000 cal or 1 kcal

surroundings
: all matter other than the system being studied

system
: portion of matter undergoing a chemical or physical change being studied

Contributors

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- Adelaide E. Clark, Oregon Institute of Technology
- Crash Course Chemistry: Crash Course is a division of Complexly and videos are free to stream for educational purposes.

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8.5.1: Practice Problems- Calorimetry

PROBLEM 8.5.1.1

A 500-mL bottle of water at room temperature and a 2-L bottle of water at the same temperature were placed in a refrigerator. After 30 minutes, the 500-mL bottle of water had cooled to the temperature of the refrigerator. An hour later, the 2-L of water had cooled to the same temperature. When asked which sample of water lost the most heat, Student A replied that both bottles lost the same amount of heat because they started at the same temperature and finished at the same temperature. Student B thought that the 2-L bottle of water lost more heat because there was more water. A third student believed that the 500-mL bottle of water lost more heat because it cooled more quickly. A fourth student thought that it was not possible to tell because we do not know the initial temperature and the final temperature of the water. Indicate which of these answers is correct and describe the error in each of the other answers.

Answer

Student A is incorrect because the mass of water in both containers is not the same.

Student C is incorrect because the bottle cooled quicker due to less mass of water.

Student D is incorrect because it doesn't matter what the change in temperature is as long as it is the same for both bottles.

Student B is correct: if the change in temperature is the same, the one with the more mass (the 2L bottle) had more heat loss. We could prove this using $q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$ from Section 8.1.

PROBLEM 8.5.1.2

How many milliliters of water at 23 °C with a density of 1.00 g/mL must be mixed with 180 mL (about 6 oz) of coffee at 95 °C so that the resulting combination will have a temperature of 60 °C? Assume that coffee and water have the same density and the same specific heat (4.184 J/g °C).

Answer

170 mL

Click here to see a video of the solution



The section number changed after this video was made

PROBLEM 8.5.1.3

How much will the temperature of a cup (180 g) of coffee at 95 °C be reduced when a 45 g silver spoon (specific heat 0.24 J/g °C) at 25 °C is placed in the coffee and the two are allowed to reach the same temperature? Assume that the coffee has the same density and specific heat as water.

Answer

The temperature of the coffee will drop 1 degree.

PROBLEM 8.5.1.4

A 45-g aluminum spoon (specific heat 0.88 J/g °C) at 24 °C is placed in 180 mL (180 g) of coffee at 85 °C and the temperature of the two become equal.

a. What is the final temperature when the two become equal? Assume that coffee has the same specific heat as water.
b. The first time a student solved this problem she got an answer of 88 °C. Explain why this is clearly an incorrect answer.

Answer a

81.95 °C

Answer b

This temperature is higher than the starting temperature of the coffee, which is impossible.

Click here to see a video of the solution



The section number changed after this video was made

PROBLEM 8.5.1.5

The temperature of the cooling water as it leaves the hot engine of an automobile is 240 °F. After it passes through the radiator it has a temperature of 175 °F. Calculate the amount of heat transferred from the engine to the surroundings by one gallon of water with a specific heat of 4.184 J/g °C.

Answer

$5.7\times 10^2\ kJ$

PROBLEM 8.5.1.6

When 50.0 g of 0.200 M NaCl(*aq*) at 24.1 °C is added to 100.0 g of 0.100 M $AgNO_3$(*aq*) at 24.1 °C in a calorimeter, the temperature increases to 25.2 °C as AgCl(*s*) forms. Assuming the specific heat of the solution and products is 4.20 J/g °C, calculate the approximate amount of heat in joules produced.

Answer

693 J

Click here to see a video of the solution





PROBLEM 8.5.1.7

The addition of 3.15 g of E a calorimeter caused the temperature to fall by 3.1 °C alculate the approximate amount of heat absorbed by t

E (8.5.1.1)

Answer

1.4 kJ

PROBLEM 8.5.1.8

When 1.0 g of fructose, $C_6H_{12}O_6(s)$, a sugar commonly found in fruits, is burned in oxygen in a bomb calorimeter, the temperature of the calorimeter increases by 1.58 °C. If the heat capacity of the calorimeter and its contents is 9.90 kJ/°C, what is *q* for this combustion?

Answer

15.64 kJ

Click here to see a video of the solution



PROBLEM 8.5.1.9

One method of generating electricity is by burning coal to heat water, which produces steam that drives an electric generator. To determine the rate at which coal is to be fed into the burner in this type of plant, the heat of combustion per ton of coal must be determined using a bomb calorimeter. When 1.00 g of coal is burned in a bomb calorimeter, the temperature increases by 1.48 °C. If the heat capacity of the calorimeter is 21.6 kJ/°C, determine the heat produced by combustion of a ton of coal (2000 pounds). Remember 1 kg = 2.2 pounds

Answer

2.91×10^7 kJ

PROBLEM 8.5.1.10

A teaspoon of the carbohydrate sucrose (common sugar) contains 16 Calories (16 kcal). What is the mass of one teaspoon of sucrose if the average number of Calories for carbohydrates is 4.1 Calories/g?

Answer

3.9 g

Click Here to see a video of the solution



This problem was renumbered after the video was made

PROBLEM 8.5.1.11

What is the maximum mass of carbohydrate in a 6-oz serving of diet soda that contains less than 1 Calorie per can if the average number of Calories for carbohydrates is 4.1 Calories/g?

Answer

0.24 g

PROBLEM 8.5.1.12

A pint of premium ice cream can contain 1100 Calories. What mass of fat, in grams and pounds, must be produced in the body to store an extra 1.1×10^3 Calories if the average number of Calories for fat is 9.1 Calories/g? Remember 1 kg = 2.2 pounds

Answer

120.87 g or 1.2 x 10^2 g with 2 significant figures

0.266 lbs or 0.27 lbs with 2 significant figures

Click here to see a video of the solution



PROBLEM 8.5.1.13

A serving of a breakfast cereal contains 3 g of protein, 18 g of carbohydrates, and 6 g of fat. What is the Calorie content of a serving of this cereal if the average number of Calories for fat is 9.1 Calories/g, for carbohydrates is 4.1 Calories/g, and for protein is 4.1 Calories/g?

Answer

1.4×10^2 Calories

Contributors

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- Adelaide Clark, Oregon Institute of Technology

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Index

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1.8: Significant Digits

Glossary

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By Page

- OT - PDX - Metro: General Chemistry I - *Undeclared*
  - Front Matter - *Undeclared*
    - TitlePage - *Undeclared*
    - InfoPage - *Undeclared*
    - Table of Contents - *Undeclared*
    - Licensing - *Undeclared*
  - 1: Matter and Measurement - *Undeclared*
    - 1.1: Atoms and Molecules - *Undeclared*
    - 1.2: The Scientific Approach to Knowledge - *Undeclared*
    - 1.3: The Classification of Matter - *Undeclared*
    - 1.4: Measurements - *Undeclared*
      - 1.4.1: Practice Problems on Measurements - *Undeclared*
    - 1.5: The Mole is a Measure of Amount - *Undeclared*
    - 1.6: Accuracy and Precision - *CC BY-NC 4.0*
    - 1.7: Dimensional Analysis - *Undeclared*
      - 1.7.1: Practice Problems on Dimensional Analysis - *Undeclared*
    - 1.8: Significant Digits - *Undeclared*
  - 2: Atoms and Elements - *Undeclared*
    - 2.1: A History of Atomic Theory - *Undeclared*
      - 2.1.1: Practice Problems- A History of Atomic Theory (Optional) - *Undeclared*
    - 2.2: The Structure of the Atom and How We Represent It - *Undeclared*
      - 2.2.1: Practice Problems- The Structure of the Atom and How We Represent It (Optional) - *Undeclared*
    - 2.3: Calculating Atomic Masses - *Undeclared*
      - 2.3.1: Practice Problems- Calculating Atomic Masses (Optional) - *Undeclared*
    - 2.4: Using the Mole in Calculations - *Undeclared*
    - 2.5: Highlighting Instrumentation- Mass Spectrometry - *Undeclared*
  - 3: Nuclei, Ions and the Periodic Table - *Undeclared*
    - 3.1: Nuclear Chemistry and Radioactive Decay - *Undeclared*
      - 3.1.1: Practice Problems- Nuclear Chemistry and Radioactive Decay (Optional) - *Undeclared*
    - 3.2: Organization of the Periodic Table - *Undeclared*
      - 3.2.1: Practice Problems - Organization of the Periodic Table (Optional) - *Undeclared*
    - 3.3: Predicting Charges of Ions - *Undeclared*
  - 4: Molecules, Compounds and Quantifying Chemicals - *Undeclared*
    - 4.1: Chemical Bonding - *Undeclared*
    - 4.2: Representing ionic and molecular compounds and molecules - *Undeclared*
      - 4.2.1: Practice Problems- Writing Chemical Formulas - *Undeclared*
    - 4.3: Chemical Nomenclature - *Undeclared*
      - 4.3.1: Practice Problems- Molecular and Ionic Compounds - *Undeclared*
    - 4.4: Formula Mass, Percent Composition, and the Mole - *Undeclared*
      - 4.4.1: Practice Problems- Formula Mass, Percent Composition, and the Mole - *Undeclared*
    - 4.5: Empirical and Molecular Formulas - *Undeclared*
      - 4.5.1: Practice Problems- Empirical and Molecular Formulas - *Undeclared*
  - 5: Transformations of Matter - *Undeclared*
    - 5.1: Writing and Balancing Chemical Equations - *Undeclared*
      - 5.1.1: Practice Problems- Writing and Balancing Chemical Equations - *Undeclared*

- 5.2: Reaction Stoichiometry - *Undeclared*
  - 5.2.1: Practice Problems- Reaction Stoichiometry - *Undeclared*
- 5.3: Calculating Reaction Yields - *Undeclared*
  - 5.3.1: Practice Problems Calculating Reaction Yields - *Undeclared*

6: Aqueous Reactions - *Undeclared*

- 6.1: Solutions and Solution Concentration - *Undeclared*
  - 6.1.1: Practice Problems- Solution Concentration - *Undeclared*
- 6.2: Solutions Chemistry - *Undeclared*
  - 6.2.1: Practice Problems- Solutions Chemistry - *Undeclared*
- 6.3: Precipitation Reactions - *Undeclared*
  - 6.3.1: Practice Problems- Precipitation - *Undeclared*
- 6.4: Classifying Chemical Reactions (Acids and Bases) - *Undeclared*
  - 6.4.1: Classifying Chemical Reactions (Acids and Bases) (Problems) - *Undeclared*
- 6.5: Classifying Chemical Reactions (Redox) - *Undeclared*
  - 6.5.1: Classifying Chemical Reactions (Redox) (Problems) - *Undeclared*

7: Ideal Gas Behavior - *Undeclared*

- 7.1: Temperature and Pressure - *Undeclared*
  - 7.1.1: Practice Problems- Temperature and Pressure - *Undeclared*
- 7.2: The Gas Laws - *Undeclared*
  - 7.2.1: Practice Problems- The Gas Laws - *Undeclared*
- 7.3: Applications of the Ideal Gas Law and Partial Pressures - *Undeclared*
  - 7.3.1: Practice Problems- Applications of the Ideal Gas Law - *Undeclared*
- 7.4: Kinetic Molecular Theory - *Undeclared*
  - 7.4.1: Practice Problems- Kinetic Molecular Theory - *Undeclared*

8: Thermochemistry - *Undeclared*

- 8.1: The Basics of Energy - *Undeclared*
  - 8.1.1: Practice Problems- The Basics of Energy - *Undeclared*
- 8.2: Measuring Energy and Heat Capacity - *Undeclared*
- 8.3: Enthalpy - *Undeclared*
- 8.4: Standard Enthalpy and Hess' Law - *Undeclared*
  - 8.4.1: Practice Problems- Enthalpy and Hess' Law - *Undeclared*
- 8.5: Calorimetry - *Undeclared*
  - 8.5.1: Practice Problems- Calorimetry - *Undeclared*

Back Matter - *Undeclared*

- Index - *Undeclared*
- Glossary - *Undeclared*
- Detailed Licensing - *Undeclared*