CHM 110: FUNDAMENTALS OF CHEMISTRY

CHM 110 Fundamentals of Chemistry

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CHAPTER OVERVIEW

1: The Process of Science

1.1: Welcome!

Welcome! This course seeks to help you contextualize chemistry concepts through the lens of sustainability and climate change. The text will help to support your understanding of chemistry concepts along with topics of climate change.

The chemistry concepts are divided into four sections

Module 1: Process of Science

Module 2: Radiation - Pros and Cons

Module 3: Energy Production

Module 4: Water

The final module will help you bring together the chemical concepts and how they are related to climate change (currently under construction)

1.2: The Scope of Chemistry

Learning Objectives

- To recognize the breadth, depth, and scope of chemistry.
- Define chemistry in relation to other sciences.
- Identify the main disciplines of chemistry.

Chemistry is the study of matter—what it consists of, what its properties are, and how it changes. Matter is anything that has mass and takes up space—that is, anything that is physically real. Some things are easily identified as matter—the screen on which you are reading this book, for example. Others are not so obvious. Because we move so easily through air, we sometimes forget that it, too, is matter. Because of this, chemistry is a science that has its fingers in just about everything. Being able to describe the ingredients in a cake and how they change when the cake is baked, for example, is chemistry!

Chemistry is one branch of science. Science is the process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations. Because the physical universe is so vast, there are many different branches of science (Figure 1.2.1). Thus, chemistry is the study of matter, biology is the study of living things, and geology is the study of rocks and the earth. Mathematics is the language of science, and we will use it to communicate some of the ideas of chemistry.

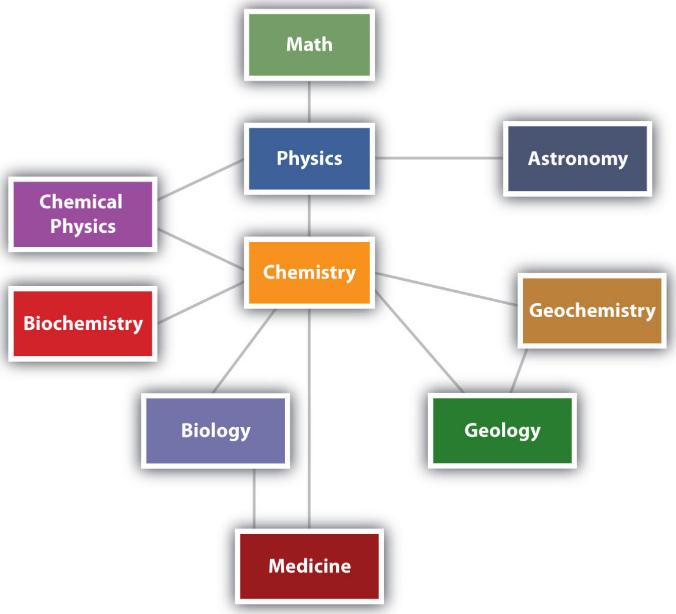

Figure 1.2.1: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

Although we divide science into different fields, there is much overlap among them. For example, some biologists and chemists work in both fields so much that their work is called biochemistry. Similarly, geology and chemistry overlap in the field called geochemistry. Figure 1.2.1 shows how many of the individual fields of science are related. At some level, all of these fields depend on matter because they all involve "stuff"; because of this, chemistry has been called the "central science", linking them all together.

> *There are many other fields of science, in addition to the ones (biology, medicine, etc.) listed here.*

Example 1.2.1: Science Fields

Which fields of study are branches of science? Explain.

a. sculpture
b. astronomy

Solution

a. Sculpture is not considered a science because it is not a study of some aspect of the natural universe.
b. Astronomy is the study of stars and planets, which are part of the natural universe. Astronomy is therefore a field of science.

? Exercise 1.2.1

Which fields of study are branches of science?

a. physiology (the study of the function of an animal's or a plant's body)
b. geophysics
c. agriculture
d. politics

Answer a:

yes

Answer b:

yes

Answer c:

yes

Answer d:

no

Areas of Chemistry

The study of modern chemistry has many branches, but can generally be broken down into five main disciplines, or areas of study:

- **Physical chemistry:** Physical chemistry is the study of macroscopic properties, atomic properties, and phenomena in chemical systems. A physical chemist may study such things as the rates of chemical reactions, the energy transfers that occur in reactions, or the physical structure of materials at the molecular level.
- **Organic chemistry:** Organic chemistry is the study of chemicals containing carbon. Carbon is one of the most abundant elements on Earth and is capable of forming a tremendously vast number of chemicals (over twenty million so far). Most of the chemicals found in all living organisms are based on carbon.
- **Inorganic chemistry:** Inorganic chemistry is the study of chemicals that, in general, are not primarily based on carbon. Inorganic chemicals are commonly found in rocks and minerals. One current important area of inorganic chemistry deals with the design and properties of materials involved in energy and information technology.
- **Analytical chemistry:** Analytical chemistry is the study of the composition of matter. It focuses on separating, identifying, and quantifying chemicals in samples of matter. An analytical chemist may use complex instruments to analyze an unknown material in order to determine its various components.
- **Biochemistry:** Biochemistry is the study of chemical processes that occur in living things. Research may cover anything from basic cellular processes up to understanding disease states so that better treatments can be developed.

Figure 1.2.2: (left) Measurement of trace metals using atomic spectroscopy. (right) Measurement of hormone concentrations.

In practice, chemical research is often not limited to just one of the five major disciplines. A particular chemist may use biochemistry to isolate a particular chemical found in the human body such as hemoglobin, the oxygen carrying component of red blood cells. He or she may then proceed to analyze the hemoglobin using methods that would pertain to the areas of physical or analytical chemistry. Many chemists specialize in areas that are combinations of the main disciplines, such as bioinorganic chemistry or physical organic chemistry.

History of Chemistry

The **history** of chemistry is an interesting and challenging one. Very early chemists were often motivated mainly by the achievement of a specific goal or product. Making perfume or soaps did not need a lot of **theory**, just a good recipe and careful attention to detail. There was no standard way of naming materials (and no periodic table that we could all agree on). It is often difficult to figure out exactly what a particular person was using. However, the science developed over the centuries by trial and error.

Major progress was made toward putting chemistry on a solid foundation when Robert Boyle (1637-1691) began his **research** in chemistry (Figure $1.2.3$). He developed the basic ideas about the behavior of gases. He could then describe gases mathematically. Boyle also helped form the idea that small particles could combine to form molecules. Many years later, John Dalton used these ideas to develop the atomic theory.

Figure $1.2.3$: Robert Boyle.

The field of chemistry began to develop rapidly in the 1700's. Joseph Priestley (1733-1804) isolated and characterized several gases: oxygen, carbon monoxide, and nitrous oxide. It was later discovered that nitrous oxide ("laughing gas") worked as an anesthetic. This gas was used for that purpose for the first time in 1844 during a tooth extraction. Other gases discovered during that time were chlorine, by C.W. Scheele (1742-1786) and nitrogen, by Antoine Lavoisier (1743-1794). Lavoisier has been considered by many scholars to be the "father of chemistry". Among other accomplishments, he discovered the role of oxygen in combustion and definitively formulated the law of conservation of matter.

Chemists continued to discover new compounds in the 1800's. The science also began to develop a more theoretical foundation. John Dalton (1766-1844) put forth his atomic theory in 1807. This idea allowed scientists to think about chemistry in a much more systematic way. Amadeo Avogadro (1776-1856) laid the groundwork for a more quantitative approach to chemistry by calculating the number of particles in a given amount of a gas. A lot of effort was put forth in studying chemical reactions. These efforts led to new materials being produced. Following the invention of the battery by Alessandro Volta (1745-1827), the field of electrochemistry (both theoretical and applications) developed through major contributions by Humphry Davy (1778-1829) and Michael Faraday (1791-1867). Other areas of the discipline also progressed rapidly.

It would take a large book to cover developments in chemistry during the twentieth century and up to today. One major area of expansion was in the area of the chemistry of living processes. Research in photosynthesis in plants, the discovery and characterization of enzymes as biochemical catalysts, elucidation of the structures of biomolecules such as insulin and DNA—these efforts gave rise to an explosion of information in the field of biochemistry.

The practical aspects of chemistry were not ignored. The work of Volta, Davy, and Faraday eventually led to the development of batteries that provided a source of electricity to power a number of devices (Figure $1.2.4$).

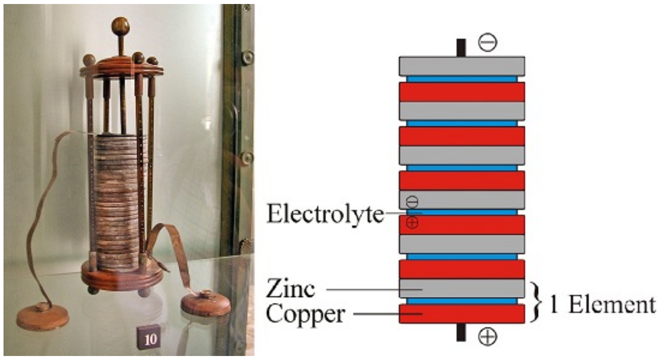

Figure 1.2.4: Battery developed by Volta. (CC BY-SA 3.0; (left) GuidoB and (right) Kkkdc).

Charles Goodyear (1800-1860) discovered the process of vulcanization, allowing a stable rubber product to be produced for the tires of all the vehicles we have today. Louis Pasteur (1822-1895) pioneered the use of heat sterilization to eliminate unwanted microorganisms in wine and milk. Alfred Nobel (1833-1896) invented dynamite (Figure 1.2.5). After his death, the fortune he made from this product was used to fund the Nobel Prizes in science and the humanities. J.W. Hyatt (1837-1920) developed the first plastic. Leo Baekeland (1863-1944) developed the first synthetic resin, widely used for inexpensive and sturdy dinnerware.

Figure 1.2.5: Dynamite explosion in Panama, Central America (1908).

Today, chemistry continues to be essential to the development of new materials and technologies, from semiconductors for electronics to powerful new medicines, and beyond.

Summary

- Chemistry is the study of matter and the changes it undergoes and considers both *macroscopic* and *microscopic* information.
- Matter is anything that has mass and occupies space.
- The five main disciplines of chemistry are physical chemistry, organic chemistry, inorganic chemistry, analytical chemistry and biochemistry.
- Many civilizations contributed to the growth of chemistry. A lot of early chemical research focused on practical uses. Basic chemistry theories were developed during the nineteenth century. New materials and batteries are a few of the products of modern chemistry.

1.3: Chemicals Compose Ordinary Things

Chemistry is the branch of science dealing with the structure, composition, properties, and the reactive characteristics of matter. Matter is anything that has mass and occupies space. Thus, chemistry is the study of literally everything around us—the liquids that we drink, the gases we breathe, the composition of everything from the plastic case on your phone to the earth beneath your feet. Moreover, chemistry is the study of the transformation of matter. Crude oil is transformed into more useful petroleum products, such as gasoline and kerosene, by the process of refining. Some of these products are further transformed into plastics. Crude metal ores are transformed into metals, that can then be fashioned into everything from foil to automobiles. Potential drugs are identified from natural sources, isolated and then prepared in the laboratory. Their structures are systematically modified to produce the pharmaceuticals that have led to vast advances in modern medicine. Chemistry is at the center of all of these processes; chemists are the people that study the nature of matter and learn to design, predict, and control these chemical transformations. Within the branches of chemistry you will find several apparent subdivisions. Inorganic chemistry, historically, focused on minerals and metals found in the earth, while organic chemistry dealt with carbon-containing compounds that were first identified in living things. Biochemistry is an outgrowth of the application of organic chemistry to biology and relates to the chemical basis for living things. In the later chapters of this text we will explore organic and biochemistry in a bit more detail and you will notice examples of organic compounds scattered throughout the text. Today, the lines between the various fields have blurred significantly and a contemporary chemist is expected to have a broad background in all of these areas.

In this chapter, we will discuss some of the properties of matter and how chemists measure those properties. We will introduce some of the vocabulary that is used throughout chemistry and the other physical sciences.

Let’s begin with matter. **Matter** is defined as any substance that has mass. It is important to distinguish here between weight and mass. Weight is the result of the pull of gravity on an object. On the Moon, an object will weigh less than the same object on Earth because the pull of gravity is less on the Moon. The mass of an object, however, is an inherent property of that object and does not change, regardless of location, gravitational pull, or anything else. It is a property that is solely dependent on the quantity of matter within the object.

Contemporary theories suggests that matter is composed of **atoms**. Atoms themselves are constructed from neutrons, protons and electrons, along with an ever-increasing array of other subatomic particles. We will focus on the neutron, a particle having no charge; the proton, which carries a positive charge; and the electron, which has a negative charge. Atoms are incredibly small. To give you an idea of the size of an atom, a single copper penny contains approximately 28,000,000,000,000,000,000,000 atoms (that’s 28 sextillion). Because atoms and subatomic particles are so small, their mass is not readily measured using pounds, ounces, grams or any other scale that we would use on larger objects. Instead, the mass of atoms and subatomic particles is measured using atomic mass units (abbreviated amu). The atomic mass unit is based on a scale that relates the mass of different types of atoms to each other (using the most common form of the element carbon as a standard). The amu scale gives us a convenient means to describe the masses of individual atoms and to do quantitative measurements concerning atoms and their reactions. Within an atom, the neutron and proton both have a mass of one amu; the electron has a much smaller mass (about 0.0005 amu).

Figure 1.3.1: Atoms are incredible small. To give you an idea of the size of an atom, a single copper penny contains approximately 28,000,000,000,000,000,000,000 atoms (that’s 28 sextillion).

Atomic theory places the neutron and the proton in the center of the atom in the nucleus. In an atom, the nucleus is very small, very dense, carries a positive charge (from the protons) and contains virtually all of the mass of the atom. Electrons are placed in a diffuse cloud surrounding the nucleus. The electron cloud carries a net negative charge (from the charge on the electrons) and in a neutral atom there are always as many electrons in this cloud as there are protons in the nucleus (the positive charges in the nucleus are balanced by the negative charges of the electrons, making the atom neutral).

An atom is characterized by the number of neutrons, protons and electrons that it possesses. Today, we recognize at least 116 different types of atoms, each type having a different number of protons in its nucleus. These different types of atoms are called elements. The neutral element hydrogen (the lightest element) will always have one proton in its nucleus and one electron in the cloud surrounding the nucleus. The element helium will always have two protons in its nucleus. It is the number of protons in the nucleus of an atom that defines the identity of an element. Elements can, however, have differing numbers of neutrons in their nucleus. For example, stable helium nuclei exist that contain one, or two neutrons (but they all have two protons). These different types of helium atoms have different masses (3 or 4 amu) and they are called isotopes. For any given isotope, the sum of the numbers of protons and neutrons in the nucleus is called the mass number. All elements exist as a collection of isotopes, and the mass of an element that we use in chemistry, the atomic mass, is the average of the masses of these isotopes. For helium, there is approximately one isotope of Helium-3 for every one million isotopes of Helium-4, hence the average atomic mass is very close to 4 (4.002602).

As different elements were discovered and named, abbreviations of their names were developed to allow for a convenient chemical shorthand. The abbreviation for an element is called its chemical symbol. A chemical symbol consists of one or two letters, and the relationship between the symbol and the name of the element is generally apparent. Thus helium has the chemical symbol He, nitrogen is N, and lithium is Li. Sometimes the symbol is less apparent but is decipherable; magnesium is Mg, strontium is Sr, and manganese is Mn. Symbols for elements that have been known since ancient times, however, are often based on Latin or Greek names and appear somewhat obscure from their modern English names. For example, copper is Cu (from cuprum), silver is Ag (from argentum), gold is Au (from aurum), and iron is Fe (from ferrum). Throughout your study of chemistry, you will routinely use chemical symbols and it is important that you begin the process of learning the names and chemical symbols for the common elements. By the time you complete General Chemistry, you will find that you are adept at naming and identifying virtually all of the 116 known elements. Table $1.3.1$ contains a starter list of common elements that you should begin learning now!

Table $1.3.1$: Names and Chemical Symbols for Common Elements

Element	**Chemical Symbol**	**Element**	**Chemical Symbol**
Hydrogen	H	Phosphorus	P
Helium	He	Sulfur	S
Lithium	Li	Chlorine	Cl
Beryllium	Be	Argon	Ar
Boron	B	Potassium	K
Carbon	C	Calcium	Ca
Nitrogen	N	Iron	Fe
Oxygen	O	Copper	Cu
Fluorine	F	Zinc	Zn
Neon	Ne	Bromine	Br
Sodium	Na	Silver	Ag
Magnesium	Mg	Iodine	I
Aluminum	Al	Gold	Au
Silicon	Si	Lead	Pb

The chemical symbol for an element is often combined with information regarding the number of protons and neutrons in a particular isotope of that atom to give the atomic symbol. To write an atomic symbol, begin with the chemical symbol, then write the atomic number for the element (the number of protons in the nucleus) as a subscript, preceding the chemical symbol. Directly above this, as a superscript, write the mass number for the isotope, that is, the total number of protons and neutrons in the nucleus. Thus, for helium, the atomic number is 2 and there are two neutrons in the nucleus for the most common isotope, making the atomic symbol $^{4}_{2}\text{He}$. In the definition of the atomic mass unit, the “most common isotope of carbon”, $^{12}_{6}\text{C}$, is defined as having a mass of exactly 12 amu and the atomic masses of the remaining elements are based on their masses relative to this isotope. Chlorine (chemical symbol Cl) consists of two major isotopes, one with 18 neutrons (the most common, comprising 75.77% of natural chlorine atoms) and one with 20 neutrons (the remaining 24.23%). The atomic number of chlorine is 17 (it has 17 protons in its nucleus), therefore the chemical symbols for the two isotopes are $^{35}_{17}\text{Cl}$ and $^{37}_{17}\text{Cl}$.

When data is available regarding the natural abundance of various isotopes of an element, it is simple to calculate the average atomic mass. In the example above, $^{35}_{17}\mathrm{Cl}$ was the most common isotope with an abundance of 75.77% and $^{37}_{17}\mathrm{Cl}$ had an abundance of the remaining 24.23%. To calculate the average mass, first convert the percentages into fractions; that is, simply divide them by 100. Now, chlorine-35 represents a fraction of natural chlorine of 0.7577 and has a mass of 35 (the mass number). Multiplying these, we get (0.7577 × 35) = 26.51. To this, we need to add the fraction representing chlorine-37, or (0.2423 × 37) = 8.965; adding, (26.51 + 8.965) = 35.48, which is the weighted average atomic mass for chlorine. Whenever we do mass calculations involving elements or compounds (combinations of elements), we always need to use average atomic masses.

Contributions & Attributions

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1.4: What is an Argument?

How to write an argument using Claim, Evidence, Reasoning (CER)

Definitions

Claim: A statement or conclusion that answers the original question or problem.

Evidence: Scientific data that supports the claim. The data need to be appropriate, sufficient, and free of inference to support the claim. Evidence does not include generalizations and beliefs about the data.

Reasoning: A justification that connects the evidence to the claim. It shows why the data count as evidence by using appropriate and sufficient scientific principles.

***Rebuttal:** A description of alternate explanations which provides counter evidence and reasoning why the alternate explanation is not the most appropriate explanation for the question or problem.

After completing the laboratory investigation and analyzing the data, answer the question or problem posed at the beginning of the investigation (claim). Determine what data from the experiment support the claim. Do not simply state the data support the claim but write several sentences with specific, appropriate and sufficient data – the actual numbers or observations from the experiment. Then, using appropriate and sufficient scientific principles, explain why the evidence supports the claim.

If you have more than one question to answer (and thus claim to make), be sure to include evidence and reasoning for each claim.

* A rebuttal is another valuable component to an argument, but will not be required in lab reports.

1.5: The Process of Science

In most textbooks, the scientific method is defined as the several steps generally starting with observations, followed by making a hypothesis, running an experiment and drawing a conclusion. This is the general form of a scientific method; however, there are so many pieces to the puzzle of running an experiment or understanding how something works. What if you don’t observe closely enough? What if we don’t examine all the variables? What if there are multiple steps that require you to make observations each time before you can come to a conclusion? What if you are collaborating among many different groups and collecting data? This linear scientific method doesn't really show what is going on. It might cause us to draw the wrong conclusion or an incomplete conclusion. Instead it is better to think of science as a process. The steps are iterative and cyclical. We are constantly questioning, gathering data, collaborating and receiving feedback. Experiments don’t often work “right” the first time. Many times in lab you will run an experiment you’ve never run before. The equipment, chemicals and/or the data analysis might be new to you. Running this experiment only once may cause you to draw an erroneous conclusion. However, if you can run the experiment multiple times, change variables, pool data, collaborate with other groups, you will likely draw a more complete conclusion and a more evidence based answer. Running an experiment requires that you have previous knowledge, that you understand all the steps, and that when you get a result, does that result make sense. Can you draw a conclusion from your results? How can you confirm that your result is correct? Scientists run experiments to answer a question, and then they run that same experiment again and again to check their result. They also run other experiments to confirm that other variables are not affecting the conclusion. In addition to confirming your result, consider what evidence do you have that your experiment did not work? Consider the errors that could have occurred during the experimental process.

Science is very methodical, but it does not always follow the same linear method.

You can visit this awesome site to learn more about how science works!

1.6: Hypothesis, Theories, and Laws

Learning Objectives

- Describe the difference between hypothesis and theory as scientific terms.
- Describe the difference between a theory and scientific law.

Although many have taken science classes throughout the course of their studies, people often have incorrect or misleading ideas about some of the most important and basic principles in science. Most students have heard of hypotheses, theories, and laws, but what do these terms really mean? Prior to reading this section, consider what you have learned about these terms before. What do these terms mean to you? What do you read that contradicts or supports what you thought?

What is a Fact?

A fact is a basic statement established by experiment or observation. All facts are true under the specific conditions of the observation.

What is a Hypothesis?

One of the most common terms used in science classes is a "hypothesis". The word can have many different definitions, depending on the context in which it is being used:

- An educated guess: a scientific hypothesis provides a suggested solution based on evidence.
- Prediction: if you have ever carried out a science experiment, you probably made this type of hypothesis when you predicted the outcome of your experiment.
- Tentative or proposed explanation: hypotheses can be suggestions about why something is observed. In order for it to be scientific, however, a scientist must be able to test the explanation to see if it works and if it is able to correctly predict what will happen in a situation. For example, "if my hypothesis is correct, we should see ___ result when we perform ___ test."

> *A hypothesis is very tentative; it can be easily changed.*

What is a Theory?

The *United States National Academy of Sciences* describes what a theory is as follows:

> "Some scientific explanations are so well established that no new evidence is likely to alter them. The explanation becomes a scientific theory. In everyday language a theory means a hunch or speculation. Not so in science. In science, the word **theory** refers to a comprehensive explanation of an important feature of nature supported by facts gathered over time. Theories also allow scientists to make predictions about as yet unobserved phenomena."
>
> "A scientific theory is a well-substantiated explanation of some aspect of the natural world, based on a body of facts that have been repeatedly confirmed through observation and experimentation. Such fact-supported theories are not "guesses" but reliable accounts of the real world. The theory of biological evolution is more than "just a theory." It is as factual an explanation of the universe as the atomic theory of matter (stating that everything is made of atoms) or the germ theory of disease (which states that many diseases are caused by germs). Our understanding of gravity is still a work in progress. But the phenomenon of gravity, like evolution, is an accepted fact.

Note some key features of theories that are important to understand from this description:

- Theories are explanations of natural phenomena. They aren't predictions (although we may use theories to make predictions). They are explanations as to why we observe something.
- Theories aren't likely to change. They have a large amount of support and are able to satisfactorily explain numerous observations. Theories can, indeed, be facts. Theories can change, but it is a long and difficult process. In order for a theory to change, there must be many observations or pieces of evidence that the theory cannot explain.
- Theories are not guesses. The phrase "just a theory" has no room in science. To be a scientific theory carries a lot of weight; it is not just one person's idea about something

> *Theories aren't likely to change.*

What is a Law?

Scientific laws are similar to scientific theories in that they are principles that can be used to predict the behavior of the natural world. Both scientific laws and scientific theories are typically well-supported by observations and/or experimental evidence. Usually scientific laws refer to rules for how nature will behave under certain conditions, frequently written as an equation. Scientific theories are more overarching explanations of how nature works and why it exhibits certain characteristics. As a comparison, theories explain why we observe what we do and laws describe what happens.

For example, around the year 1800, Jacques Charles and other scientists were working with gases to, among other reasons, improve the design of the hot air balloon. These scientists found, after many, many tests, that certain patterns existed in the observations on gas behavior. If the temperature of the gas is increased, the volume of the gas increased. This is known as a natural law. A law is a relationship that exists between variables in a group of data. Laws describe the patterns we see in large amounts of data, but do not describe why the patterns exist.

What is a Belief?

A **belief** is a statement that is not scientifically provable. Beliefs may or may not be incorrect; they just are outside the realm of science to explore.

Laws vs. Theories

A common misconception is that scientific theories are rudimentary ideas that will eventually graduate into scientific laws when enough data and evidence has accumulated. A theory does not change into a scientific law with the accumulation of new or better evidence. Remember, **theories are explanations** and **laws are patterns** we see in large amounts of data, frequently written as an equation. A theory will always remain a theory; a law will always remain a law.

Video $1.6.1$: What’s the difference between a scientific law and theory?

Summary

- A hypothesis is a tentative explanation that can be tested by further investigation.
- A theory is a well-supported explanation of observations.
- A scientific law is a statement that summarizes the relationship between variables.
- An experiment is a controlled method of testing a hypothesis.

Contributions & Attributions

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1.7: Classifying Matter According to Its Composition

Learning Objectives

- Explain the difference between a pure substance and a mixture.
- Explain the difference between an element and a compound.
- Explain the difference between a homogeneous mixture and a heterogeneous mixture.

One useful way of organizing our understanding of matter is to think of a hierarchy that extends down from the most general and complex to the simplest and most fundamental (Figure 1.7.1). Matter can be classified into two broad categories: pure substances and mixtures. A **pure substance** is a form of matter that has a constant composition (meaning that it is the same everywhere) and properties that are constant throughout the sample (meaning that there is only one set of properties such as melting point, color, boiling point, etc. throughout the matter). A material composed of two or more substances is a **mixture**. Elements and compounds are both examples of pure substances. A substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A substance that can be broken down into chemically simpler components (because it has more than one element) is a **compound**. For example, water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

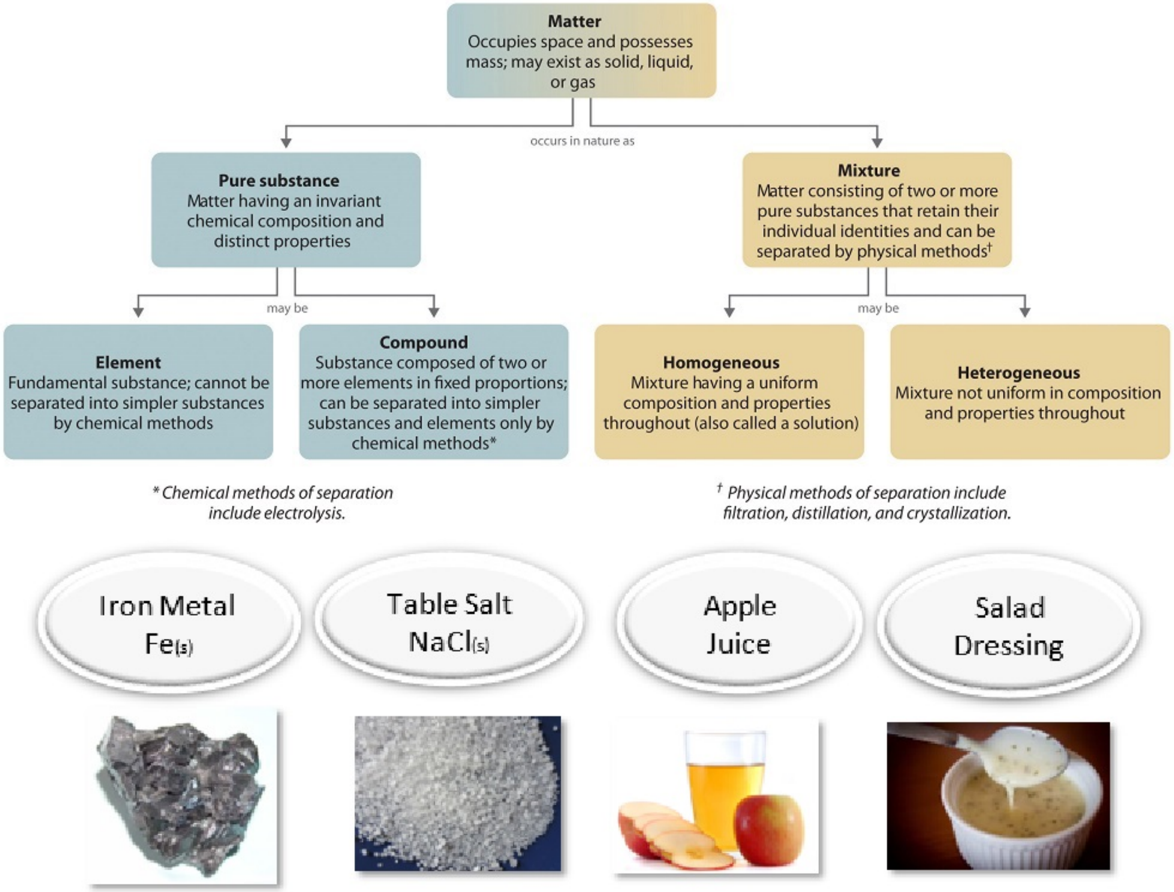

Figure 1.7.1: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

Ordinary table salt is called sodium chloride. It is considered a **substance** because it has a uniform and definite composition. All samples of sodium chloride are chemically identical. Water is also a pure substance. Salt easily dissolves in water, but salt water cannot be classified as a substance because its composition can vary. You may dissolve a small amount of salt or a large amount into a given amount of water. A mixture is a physical blend of two or more components, each of which retains its own identity and properties in the **mixture**. Only the form of the salt is changed when it is dissolved into water. It retains its composition and properties.

A **homogeneous** mixture is a mixture in which the composition is uniform throughout the mixture. The salt water described above is homogeneous because the dissolved salt is evenly distributed throughout the entire salt water sample. Often it is easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that the composition of the substance is always the same. The amount of salt in the salt water can vary from one sample to another. All solutions are considered homogeneous because the dissolved material is present in the same amount throughout the solution.

A **heterogeneous mixture** is a mixture in which the composition is not uniform throughout the mixture. Vegetable soup is a heterogeneous mixture. Any given spoonful of soup will contain varying amounts of the different vegetables and other components of the soup.

Phase

A phase is any part of a sample that has a uniform composition and properties. By definition, a pure substance or a homogeneous mixture consists of a single phase. A heterogeneous mixture consists of two or more phases. When oil and water are combined, they do not mix evenly, but instead form two separate layers. Each of the layers is called a phase.

✓ Example 1.7.1

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 1.7.1

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:

homogeneous mixture (solution)

Answer b:

element

Answer c:

heterogeneous mixture

Answer d:

compound

✓ Example 1.7.2

How would a chemist categorize each example of matter?

a. saltwater
b. soil
c. water
d. oxygen

Solution

a. Saltwater acts as if it were a single substance even though it contains two substances—salt and water. Saltwater is a homogeneous mixture, or a solution.
b. Soil is composed of small pieces of a variety of materials, so it is a heterogeneous mixture.
c. Water is a substance. More specifically, because water is composed of hydrogen and oxygen, it is a compound.
d. Oxygen, a substance, is an element.

? Exercise 1.7.2

How would a chemist categorize each example of matter?

a. coffee
b. hydrogen
c. an egg

Answer a:

a homogeneous mixture (solution), assuming it is filtered coffee

Answer b:

element

Answer c:

heterogeneous mixture

Summary

Matter can be classified into two broad categories: pure substances and mixtures. A pure substance is a form of matter that has a constant composition and properties that are constant throughout the sample. Mixtures are physical combinations of two or more elements and/or compounds. Mixtures can be classified as homogeneous or heterogeneous. Elements and compounds are both examples of pure substances. Compounds are substances that are made up of more than one type of atom. Elements are the simplest substances made up of only one type of atom.

Vocabulary

- Element: a substance that is made up of only one type of atom.
- Compound:a substance that is made up of more than one type of atom bonded together.
- Mixture: a combination of two or more elements or compounds which have not reacted to bond together; each part in the mixture retains its own properties.

Contributions & Attributions

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- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook

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1.8: Energy

Energy is a ubiquitous term we all use from a young age in a variety of contexts. We sometimes feel like we have no energy or we have a lot of excess energy. Many people now consume "energy drinks." We think about conserving energy by turning off lights or taking stairs instead of an elevator. Energy is something we encounter daily. We recharge our phones, burn gasoline in cars, feel the heat of the sun, watch as trees sway in the wind, glance up to see if a rock is poised to fall or an icicle is about to drop when we encounter the signs indicating we have entered a zone where that is possible. But what is energy and do the intuitive ideas we develop about energy from an early age align with what we mean by energy from a scientific perspective?

Energy is defined as the ability to do work or as the capacity to cause change. Work has a very specific definition in science, namely that it is a force applied through a distance (w = Fxd). Work is also one of the ways energy can be transferred to and from a system. It is not surprising then that energy and work have the same units – the Joule in the SI system. A joule is the amount of work done when a force of one Newton is applied over a distance of 1 meter. It is also the amount of energy expended during that process.

While the definition may seem circuitous, there are five major ideas associated with the concept of energy that allow us to explain many different phenomena including objects in motion and interactions between particles. One of these ideas is the Law of Conservation of energy which can be simply stated as energy is neither created or destroyed.

Another important idea is that all energy is fundamentally the same, but is manifested in different phenomena in different forms or types. Energy is not directly measured but rather is calculated, and often as a change in energy, based on properties that are measurable such as temperature or speed. Two major types of energy are kinetic energy and potential energy. Kinetic energy is referred to as energy of motion and can be calculated using the formula ½ mv2 where m = mass and v = velocity. Potential energy is referred to as stored energy due to position or composition.

While energy is always conserved, it can be transformed from one type to another. For example, wind, which is really a form of solar energy, turns the blades of a turbine around a rotor connected to a generator, ultimately producing electrical energy that you can use to run your computer or light a room. Energy can be transferred between bodies or systems across boundaries. You may have experienced this holding a cup of a hot drink. Assuming the cup is warmer than your hands, the energy will transfer the cup to your hands - from the hotter region to the colder region. This transfer of energy from regions of different temperature is known as heat. Although related, heat, temperature and thermal energy are not the same. Thermal energy is the energy of the particles in a system due to the random motion of the particles and temperature is a measure of the average kinetic energy of a system.

One final important idea about energy is that you can't get something for nothing! There is no such thing as a perpetual motion machine. Energy cannot be completely converted to work – some energy is always dissipated as waste heat.

1.9: Looking for Patterns- The Periodic Table

Learning Objectives

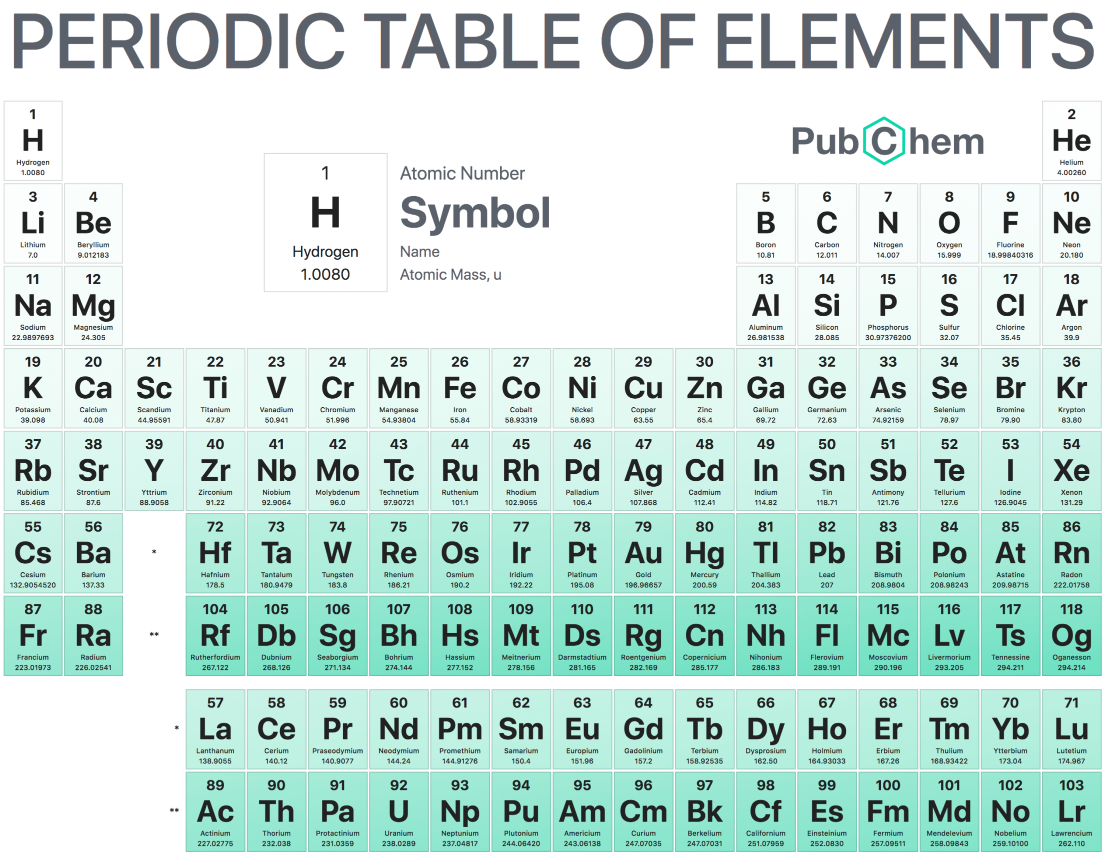
- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.

In the 19th century, many previously unknown elements were discovered, and scientists noted that certain sets of elements had similar chemical properties. For example, chlorine, bromine, and iodine react with other elements (such as sodium) to make similar compounds. Likewise, lithium, sodium, and potassium react with other elements (such as oxygen) to make similar compounds. Why is this so?

In 1864, Julius Lothar Meyer, a German chemist, organized the elements by atomic mass and grouped them according to their chemical properties. Later that decade, Dmitri Mendeleev, a Russian chemist, organized all the known elements according to similar properties. He left gaps in his table for what he thought were undiscovered elements, and he made some bold predictions regarding the properties of those undiscovered elements. When elements were later discovered whose properties closely matched Mendeleev's predictions, his version of the table gained favor in the scientific community. Because certain properties of the elements repeat on a regular basis throughout the table (that is, they are periodic), it became known as the periodic table.

> *Mendeleev had to list some elements out of the order of their atomic masses to group them with other elements that had similar properties.*

The periodic table is one of the cornerstones of chemistry because it organizes all the known elements on the basis of their chemical properties. A modern version is shown in Figure 1.9.1. The elements are listed in order of atomic number.



PERIODIC TABLE OF ELEMENTS

1
H
Hydrogen
1.0080

Atomic Number
Symbol
Name
Atomic Mass, u

1 H Hydrogen 1.0080																	2 He Helium 4.00260
3 Li Lithium 7.0	4 Be Beryllium 9.012183											5 B Boron 10.81	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.99840316	10 Ne Neon 20.180
11 Na Sodium 22.9897693	12 Mg Magnesium 24.305											13 Al Aluminum 26.981538	14 Si Silicon 28.085	15 P Phosphorus 30.97376200	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.9
19 K Potassium 39.098	20 Ca Calcium 40.08	21 Sc Scandium 44.95591	22 Ti Titanium 47.87	23 V Vanadium 50.941	24 Cr Chromium 51.996	25 Mn Manganese 54.93804	26 Fe Iron 55.84	27 Co Cobalt 58.93319	28 Ni Nickel 58.693	29 Cu Copper 63.55	30 Zn Zinc 65.4	31 Ga Gallium 69.72	32 Ge Germanium 72.63	33 As Arsenic 74.92159	34 Se Selenium 78.97	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.468	38 Sr Strontium 87.6	39 Y Yttrium 88.9058	40 Zr Zirconium 91.22	41 Nb Niobium 92.9064	42 Mo Molybdenum 96.0	43 Tc Technetium 97.90721	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9055	46 Pd Palladium 106.4	47 Ag Silver 107.868	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.6	53 I Iodine 126.9045	54 Xe Xenon 131.29
55 Cs Cesium 132.9054520	56 Ba Barium 137.33	*	72 Hf Hafnium 178.5	73 Ta Tantalum 180.9479	74 W Tungsten 183.8	75 Re Rhenium 186.21	76 Os Osmium 190.2	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.96657	80 Hg Mercury 200.59	81 Tl Thallium 204.383	82 Pb Lead 207	83 Bi Bismuth 208.9804	84 Po Polonium 208.98243	85 At Astatine 209.98715	86 Rn Radon 222.01758
87 Fr Francium 223.01973	88 Ra Radium 226.02541	**	104 Rf Rutherfordium 267.122	105 Db Dubnium 268.126	106 Sg Seaborgium 271.134	107 Bh Bohrium 274.144	108 Hs Hassium 277.152	109 Mt Meitnerium 278.156	110 Ds Darmstadtium 281.165	111 Rg Roentgenium 282.169	112 Cn Copernicium 285.177	113 Nh Nihonium 286.183	114 Fl Flerovium 289.191	115 Mc Moscovium 290.196	116 Lv Livermorium 293.205	117 Ts Tennessine 294.211	118 Og Oganesson 294.214
		*	57 La Lanthanum 138.9055	58 Ce Cerium 140.12	59 Pr Praseodymium 140.9077	60 Nd Neodymium 144.24	61 Pm Promethium 144.91276	62 Sm Samarium 150.4	63 Eu Europium 151.96	64 Gd Gadolinium 157.2	65 Tb Terbium 158.92535	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93033	68 Er Erbium 167.26	69 Tm Thulium 168.93422	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.967
		**	89 Ac Actinium 227.02775	90 Th Thorium 232.038	91 Pa Protactinium 231.0359	92 U Uranium 238.0289	93 Np Neptunium 237.04817	94 Pu Plutonium 244.06420	95 Am Americium 243.06138	96 Cm Curium 247.07035	97 Bk Berkelium 247.07031	98 Cf Californium 251.07959	99 Es Einsteinium 252.0830	100 Fm Fermium 257.09511	101 Md Mendelevium 258.09843	102 No Nobelium 259.10100	103 Lr Lawrencium 262.110

Figure 1.9.1: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number. (Public Domain; PubChem via NIH)

Features of the Periodic Table

Elements that have similar chemical properties are grouped in columns called groups (or families). As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements), *alkaline earth metals* (the second column of elements), *halogens* (the next-to-last column of elements), and *noble gases* (the last column of elements).

> *The word halogen comes from the Greek for “salt maker” because these elements combine with other elements to form a group of compounds called salts.*

To Your Health: Radon

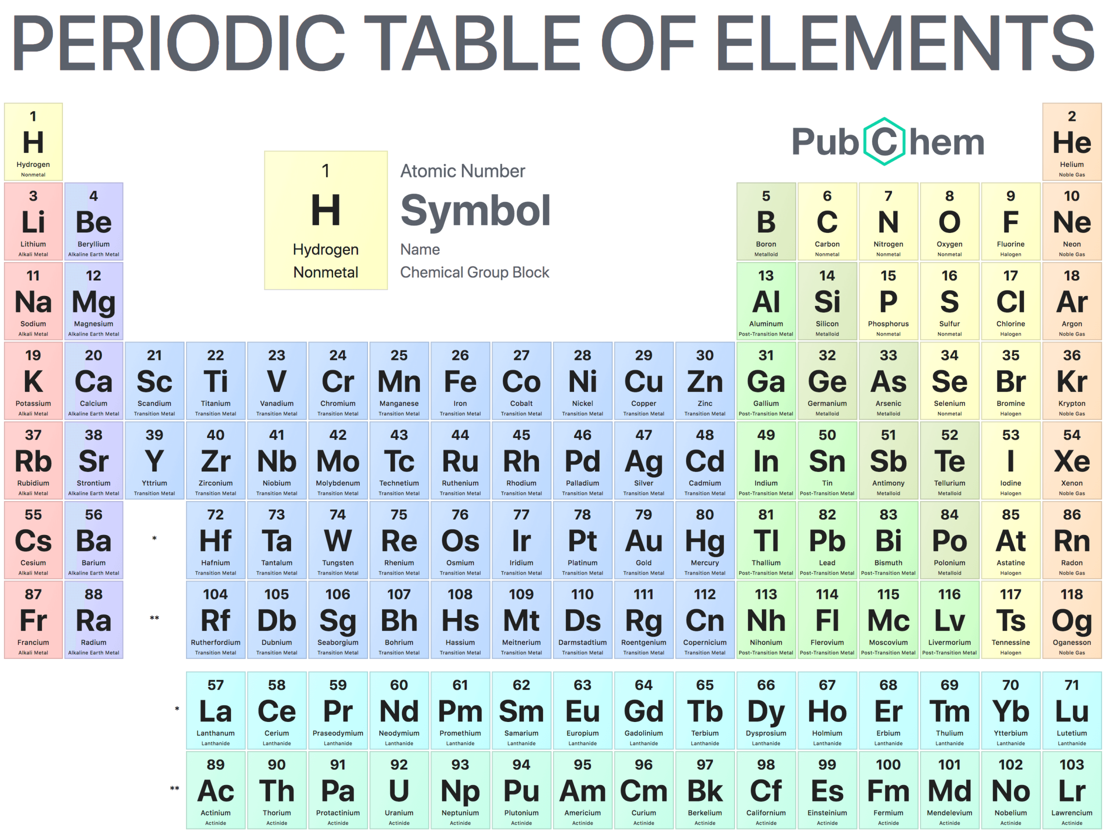
Radon is an invisible, odorless noble gas that is slowly released from the ground, particularly from rocks and soils whose uranium content is high. Because it is a noble gas, radon is not chemically reactive. Unfortunately, it is radioactive, and increased exposure to it has been correlated with an increased lung cancer risk.

Because radon comes from the ground, we cannot avoid it entirely. Moreover, because it is denser than air, radon tends to accumulate in basements, which if improperly ventilated can be hazardous to a building’s inhabitants. Fortunately, specialized ventilation minimizes the amount of radon that might collect. Special fan-and-vent systems are available that draw air from below the basement floor, before it can enter the living space, and vent it above the roof of a house.

After smoking, radon is thought to be the second-biggest *preventable* cause of lung cancer in the United States. The American Cancer Society estimates that 10% of all lung cancers are related to radon exposure. There is uncertainty regarding what levels of exposure cause cancer, as well as what the exact causal agent might be (either radon or one of its breakdown products, many of which are also radioactive and, unlike radon, not gases). The US Environmental Protection Agency recommends testing every floor below the third floor for radon levels to guard against long-term health effects.

Each row of elements on the periodic table is called a period. Periods have different lengths; the first period has only 2 elements (hydrogen and helium), while the second and third periods have 8 elements each. The fourth and fifth periods have 18 elements each, and later periods are so long that a segment from each is removed and placed beneath the main body of the table.

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure $1.9.2$. A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure $1.9.2$, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.



PERIODIC TABLE OF ELEMENTS

PubChem

1 H Hydrogen Nonmetal	Atomic Number Symbol Name Chemical Group Block

1 H Hydrogen Nonmetal																	2 He Helium Noble Gas
3 Li Lithium Alkali Metal	4 Be Beryllium Alkaline Earth Metal											5 B Boron Metalloid	6 C Carbon Nonmetal	7 N Nitrogen Nonmetal	8 O Oxygen Nonmetal	9 F Fluorine Halogen	10 Ne Neon Noble Gas
11 Na Sodium Alkali Metal	12 Mg Magnesium Alkaline Earth Metal											13 Al Aluminum Post-Transition Metal	14 Si Silicon Metalloid	15 P Phosphorus Nonmetal	16 S Sulfur Nonmetal	17 Cl Chlorine Halogen	18 Ar Argon Noble Gas
19 K Potassium Alkali Metal	20 Ca Calcium Alkaline Earth Metal	21 Sc Scandium Transition Metal	22 Ti Titanium Transition Metal	23 V Vanadium Transition Metal	24 Cr Chromium Transition Metal	25 Mn Manganese Transition Metal	26 Fe Iron Transition Metal	27 Co Cobalt Transition Metal	28 Ni Nickel Transition Metal	29 Cu Copper Transition Metal	30 Zn Zinc Transition Metal	31 Ga Gallium Post-Transition Metal	32 Ge Germanium Metalloid	33 As Arsenic Metalloid	34 Se Selenium Nonmetal	35 Br Bromine Halogen	36 Kr Krypton Noble Gas
37 Rb Rubidium Alkali Metal	38 Sr Strontium Alkaline Earth Metal	39 Y Yttrium Transition Metal	40 Zr Zirconium Transition Metal	41 Nb Niobium Transition Metal	42 Mo Molybdenum Transition Metal	43 Tc Technetium Transition Metal	44 Ru Ruthenium Transition Metal	45 Rh Rhodium Transition Metal	46 Pd Palladium Transition Metal	47 Ag Silver Transition Metal	48 Cd Cadmium Transition Metal	49 In Indium Post-Transition Metal	50 Sn Tin Post-Transition Metal	51 Sb Antimony Metalloid	52 Te Tellurium Metalloid	53 I Iodine Halogen	54 Xe Xenon Noble Gas
55 Cs Cesium Alkali Metal	56 Ba Barium Alkaline Earth Metal	*	72 Hf Hafnium Transition Metal	73 Ta Tantalum Transition Metal	74 W Tungsten Transition Metal	75 Re Rhenium Transition Metal	76 Os Osmium Transition Metal	77 Ir Iridium Transition Metal	78 Pt Platinum Transition Metal	79 Au Gold Transition Metal	80 Hg Mercury Transition Metal	81 Tl Thallium Post-Transition Metal	82 Pb Lead Post-Transition Metal	83 Bi Bismuth Post-Transition Metal	84 Po Polonium Metalloid	85 At Astatine Halogen	86 Rn Radon Noble Gas
87 Fr Francium Alkali Metal	88 Ra Radium Alkaline Earth Metal	**	104 Rf Rutherfordium Transition Metal	105 Db Dubnium Transition Metal	106 Sg Seaborgium Transition Metal	107 Bh Bohrium Transition Metal	108 Hs Hassium Transition Metal	109 Mt Meitnerium Transition Metal	110 Ds Darmstadtium Transition Metal	111 Rg Roentgenium Transition Metal	112 Cn Copernicium Transition Metal	113 Nh Nihonium Post-Transition Metal	114 Fl Flerovium Post-Transition Metal	115 Mc Moscovium Post-Transition Metal	116 Lv Livermorium Post-Transition Metal	117 Ts Tennessine Halogen	118 Og Oganesson Noble Gas
		*	57 La Lanthanum Lanthanide	58 Ce Cerium Lanthanide	59 Pr Praseodymium Lanthanide	60 Nd Neodymium Lanthanide	61 Pm Promethium Lanthanide	62 Sm Samarium Lanthanide	63 Eu Europium Lanthanide	64 Gd Gadolinium Lanthanide	65 Tb Terbium Lanthanide	66 Dy Dysprosium Lanthanide	67 Ho Holmium Lanthanide	68 Er Erbium Lanthanide	69 Tm Thulium Lanthanide	70 Yb Ytterbium Lanthanide	71 Lu Lutetium Lanthanide
		**	89 Ac Actinium Actinide	90 Th Thorium Actinide	91 Pa Protactinium Actinide	92 U Uranium Actinide	93 Np Neptunium Actinide	94 Pu Plutonium Actinide	95 Am Americium Actinide	96 Cm Curium Actinide	97 Bk Berkelium Actinide	98 Cf Californium Actinide	99 Es Einsteinium Actinide	100 Fm Fermium Actinide	101 Md Mendelevium Actinide	102 No Nobelium Actinide	103 Lr Lawrencium Actinide

Figure 1.9.2: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table. (Public Domain, PubChem via NIH)

Example 1.9.1

Based on its position in the periodic table, classify each element below as metal, a nonmetal, or a metalloid.

a. Se
b. Mg

c. Ge

SOLUTION

a. In Figure 1.9.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

Exercise 1.9.1

Based on its location in the periodic table, do you expect indium to be a nonmetal, a metal, or a metalloid?

Answer

Indium is a metal.

Another way to categorize the elements of the periodic table is shown in Figure 1.9.3. The first two columns on the left and the last six columns on the right are called the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

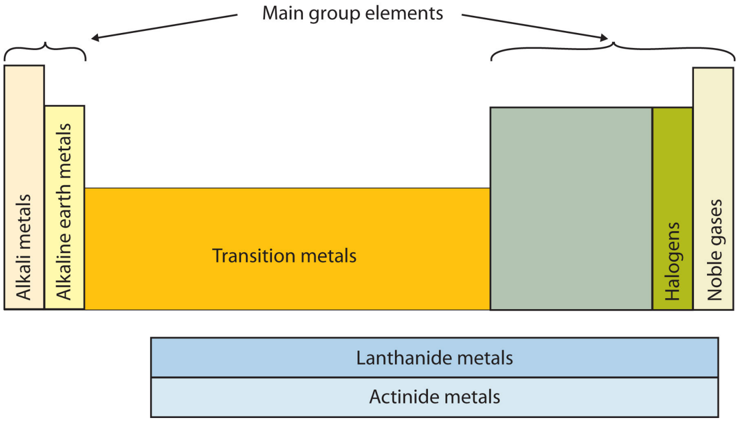


Figure 1.9.3: Special Names for Sections of the Periodic Table. (CC BY-NC-SA; Anonymous by request)

Descriptive Names

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins.

Group 1: The Alkali Metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium. Hydrogen is unique in that it is generally placed in Group 1, but it is not a metal.

The compounds of the alkali metals are common in nature and daily life. One example is table salt (sodium chloride); lithium compounds are used in greases, in batteries, and as drugs to treat patients who exhibit manic-depressive, or bipolar, behavior. Although lithium, rubidium, and cesium are relatively rare in nature, and francium is so unstable and highly radioactive that it exists in only trace amounts, sodium and potassium are the seventh and eighth most abundant elements in Earth’s crust, respectively.

Video: Alkali metals in water - Chemical elements: properties and reactions. (The Open University via https://youtu.be/6ZY6d6jrq-0)

Group 2: The Alkaline Earth Metals

The alkaline earth metals are beryllium, magnesium, calcium, strontium, barium, and radium. Beryllium, strontium, and barium are rare, and radium is unstable and highly radioactive. In contrast, calcium and magnesium are the fifth and sixth most abundant elements on Earth, respectively; they are found in huge deposits of limestone and other minerals.

Group 17: The Halogens

The halogens are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for “salt forming,” which reflects that all the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt).

Compounds that contain the fluoride ion are added to toothpaste and the water supply to prevent dental cavities. Fluorine is also found in Teflon coatings on kitchen utensils. Although chlorofluorocarbon propellants and refrigerants are believed to lead to the depletion of Earth’s ozone layer and contain both fluorine and chlorine, the latter is responsible for the adverse effect on the ozone layer. Bromine and iodine are less abundant than chlorine, and astatine is so radioactive that it exists in only negligible amounts in nature.

Group 18: The Noble Gases

The noble gases are helium, neon, argon, krypton, xenon, and radon. Because the noble gases are composed of only single atoms, they are called monatomic. At room temperature and pressure, they are unreactive gases. Because of their lack of reactivity, for many years they were called inert gases or rare gases. However, the first chemical compounds containing the noble gases were prepared in 1962. Although the noble gases are relatively minor constituents of the atmosphere, natural gas contains substantial amounts of helium. Because of its low reactivity, argon is often used as an unreactive (inert) atmosphere for welding and in light bulbs. The red light emitted by neon in a gas discharge tube is used in neon lights.

Example 1.9.2: Groups

Provide the family or group name of each element.

a. Li
b. Ar
c. Cl

SOLUTION

a. Lithium is an alkali metal (Group 1)
b. Argon is a noble gas (Group 18)
c. Chlorine is a halogen (Group 17)

Exercise 1.9.2: Groups

Provide the family or group name of each element.

a. F
b. Ca
c. Kr

Answer a:

Fluorine is a halogen (Group 17)

Answer b:

Calcium is a alkaline earth metal (Group 2)

Answer c:

Krypton is a noble gas (Group 18)

Example 1.9.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

a. Li
b. Ar
c. Am
d. Fe

SOLUTION

a. Lithium is a metal.
b. Argon is a non metal
c. Americium is an inner transition metal
d. Iron is a transition metal.

Exercise 1.9.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

a. F
b. U
c. Cu

Answer a:

Fluorine is a nonmetal.

Answer b:

Uranium is a metal (and a inner transition metal too)

Answer c:

Copper is a metal (and a transition metal too)

Summary

The periodic table is an arrangement of the elements in order of increasing atomic number. Elements that exhibit similar chemistry appear in vertical columns called groups (numbered 1–18 from left to right); the seven horizontal rows are called periods. Some of the groups have widely-used common names, including the alkali metals (Group 1) and the alkaline earth metals (Group 2) on the far left, and the halogens (Group 17) and the noble gases (Group 18) on the far right.

The elements can be broadly divided into metals, nonmetals, and semi metals. Semi metals exhibit properties intermediate between those of metals and nonmetals. Metals are located on the left of the periodic table, and nonmetals are located on the upper right. They are separated by a diagonal band of semi metals. Metals are lustrous, good conductors of electricity, and readily shaped (they are ductile and malleable). Solid nonmetals are generally brittle and poor electrical conductors. Other important groupings of elements in the periodic table are the main group elements, the transition metals, and the inner transition metals (the lanthanides, and the actinides).

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3. Petrucci, Ralph H., Carey Bissonnette, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. Custom Edition for CHEM 2. Pearson Learning Solutions, 2010.

Contributions & Attributions

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1.10: The Basic Units of Measurement

Learning Objectives

- State the different measurement systems used in chemistry.
- Describe how prefixes are used in the metric system and identify how the prefixes milli-, centi-, and kilo- compare to the base unit.

How long is a yard? It depends on whom you ask and when you asked the question. Today we have a standard definition of the yard, which you can see marked on every football field. If you move the ball ten yards, you get a first down and it doesn't matter whether you are playing in Los Angeles, Dallas, or Green Bay. But at one time that yard was arbitrarily defined as the distance from the tip of the king's nose to the end of his outstretched hand. Of course, the problem there is simple: new king, new distance (and then you have to remark all those football fields).

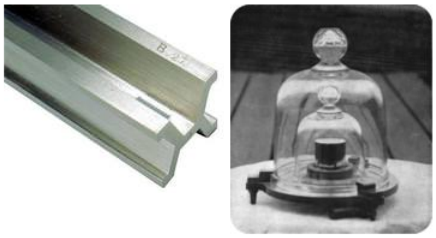


Figure 1.10.1: Meter standard (left) and Kilogram standard (right).

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units in Table 1.10.1.

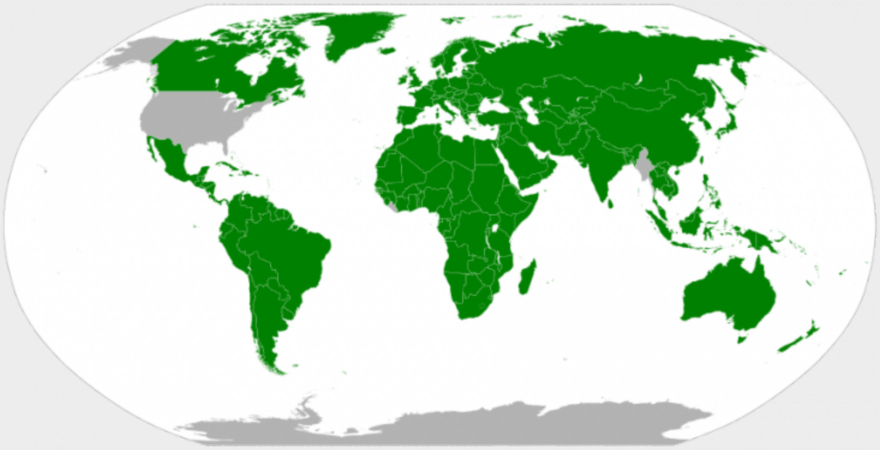
Table 1.10.1: SI Base Units of Measurement

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol
Electric Current	ampere	A
Luminous Intensity	candela	cd

The first units are frequently encountered in chemistry. All other measurement quantities, such as volume, force, and energy, can be derived from these seven base units.

Unfortunately, The Metric System is Not Ubiquitous

The map below shows the adoption of the SI units in countries around the world. It is worth noting that the US uses its own system - the US Customary system, based on the British Imperial System. According to masterclass on cooking, the US has not switched because they customary system was in place during the Industrial Revolution and manufacturing is based on it so the companies lobbied to keep the system. Only Myanmar and Liberia along with the US have not adopted the metric system as the official system of measurement. The United States uses the metric system for measurements, but does not use it in everyday practice. Great Britain and much of Canada use a combination of metric and imperial units.



Areas of world using metric system (in green). Only a few countries are slow or resistant to adoption including the United States.

Prefix Multipliers

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word $\mu\varepsilon\gamma\alpha\varsigma$, meaning "great". Table $1.10.2$ lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table 1.10.2: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Example
giga	G	1,000,000,000	1 gigameter $(\mathrm{Gm}) = 10^9$ m
mega	M	1,000,000	1 megameter $(\mathrm{Mm}) = 10^6$ m
kilo	k	1,000	1 kilometer $(\mathrm{km}) = 1{,}000$ m
hecto	h	100	1 hectometer $(\mathrm{hm}) = 100$ m
deka	da	10	1 dekameter $(\mathrm{dam}) = 10$ m
		1	1 meter (m)
deci	d	1/10	1 decimeter $(\mathrm{dm}) = 0.1$ m
centi	c	1/100	1 centimeter $(\mathrm{cm}) = 0.01$ m
milli	m	1/1,000	1 millimeter $(\mathrm{mm}) = 0.001$ m
micro	μ	1/1,000,000	1 micrometer $(\mu\mathrm{m}) = 10^{-6}$ m
nano	n	1/1,000,000,000	1 nanometer $(\mathrm{nm}) = 10^{-9}$ m
pico	p	1/1,000,000,000,000	1 picometer $(\mathrm{pm}) = 10^{-12}$ m

There are a couple of odd little practices with the use of metric abbreviations. Most abbreviations are lowercase. We use "m" for meter and not "M". However, when it comes to volume, the base unit "liter" is abbreviated as "L" and not "l". So we would write 3.5 milliliters as $3.5\ \mathrm{mL}$.

As a practical matter, whenever possible you should express the units in a small and manageable number. If you are measuring the weight of a material that weighs $6.5\ \mathrm{kg}$, this is easier than saying it weighs $6500\ \mathrm{g}$ or $0.65\ \mathrm{dag}$. All three are correct, but the kg units in this case make for a small and easily managed number. However, if a specific problem needs grams instead of kilograms, go with the grams for consistency.

Example 1.10.1 : Unit Abbreviations

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

a. kiloliter
b. microsecond
c. decimeter
d. nanogram

SOLUTIONS

	Explanation	Answer
a	The prefix kilo means “1,000 ×,” so 1 kL equals 1,000 L	kL
b	The prefix micro implies 1/1,000,000th of a unit, so 1 µs equals 0.000001 s.	µs
c	The prefix deci means 1/10th, so 1 dm equals 0.1 m.	dm
d	The prefix nano means 1/1000000000, so a nanogram is equal to 0.000000001 g	ng

Exercise 1.10.1

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

a. kilometer
b. milligram
c. nanosecond
d. centiliter

Answer a:

km

Answer b:

mg

Answer c:

ns

Answer d:

cL

Summary

- Metric prefixes derive from Latin or Greek terms. The prefixes are used to make the units manageable.
- The SI system is based on multiples of ten. There are seven basic units in the SI system. Five of these units are commonly used in chemistry.

Contributions & Attributions

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1.11: Significant Figures - Writing Numbers to Reflect Precision

Learning Objectives

- Identify the number of significant figures in a reported value.

Think about this

How many letters are on your keyboard? What is the concentration of carbon dioxide in the atmosphere? One of these is an exact number and one is a measured number. Which one is which? How do you know?

Measured numbers

A measured number is obtained with a tool like a graduated cylinder in laboratory experiments. The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit. In the ruler illustration below, the bottom ruler gave a length with 2 significant figures, while the top ruler gave a length with 3 significant figures. In a correctly reported measurement, the final digit is significant but not certain. Insignificant digits are not reported. With either ruler, it would not be possible to report the length at 2.553 cm as there is no possible way that the thousandths digit could be estimated. The 3 is not significant and would not be reported.

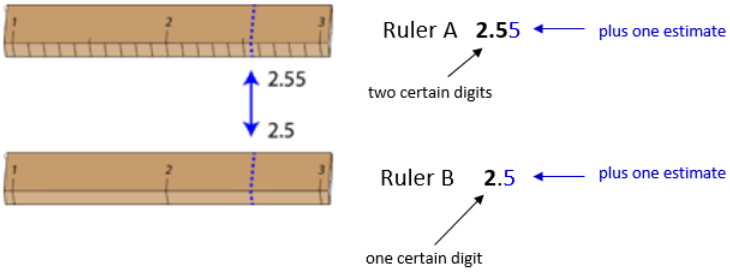


Figure 1.11.1: Measurement with two different rulers.

Measurement Uncertainty

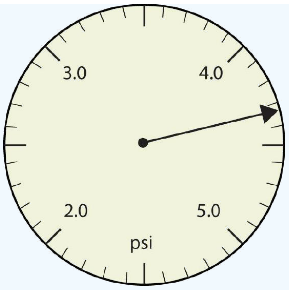
Some error or **uncertainty** always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest 0.1 g, other highly sensitive balances are capable of measuring to the nearest 0.001 g or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Figure 1.11.1 shows two rulers making the same measurement of an object (indicated by the blue arrow).

With either ruler, it is clear that the length of the object is between 2 and 3 cm. The bottom ruler contains no millimeter markings. With that ruler, the tenths digit can be estimated and the length may be reported as 2.5 cm. However, another person may judge that the measurement is 2.4 cm or perhaps 2.6 cm. While the 2 is known for certain, the value of the tenths digit is uncertain.

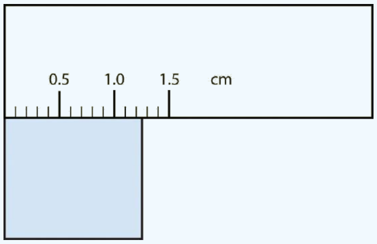
The top ruler contains marks for tenths of a centimeter (millimeters). Now the same object may be measured as 2.55 cm. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5. Again, another measurer may report the length to be 2.54 cm or 2.56 cm. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

Example 1.11.1: Reporting Measurements to the Proper number of Significant Figures

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



a.



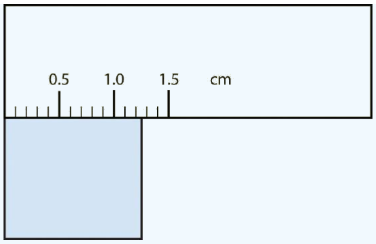
b.

Solutions

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

Exercise 1.11.1

What would be the reported width of this rectangle?



Answer

1.25 cm

When you look at a reported measurement, it is necessary to be able to count the number of significant figures. The table below details the rules for determining the number of significant figures in a reported measurement. For the examples in the table, assume that the quantities are correctly reported values of a measured quantity.

Table 1.11.1: Significant Figure Rules

Rule	Examples
1. All nonzero digits in a measurement are significant.	• 237 has three significant figures. • 1.897 has four significant figures.
2. Zeros that appear between other nonzero digits (middle zeros) are always significant.	• 39,004 has five significant figures. • 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called leading zeros. Leading zeros are never significant.	• 0.008 has one significant figure. • 0.000416 has three significant figures.
4. Zeros that appear after all nonzero digits are called trailing zeros. A number with trailing zeros that lacks a decimal point may or may not be significant. **Use scientific notation to indicate the appropriate number of significant figures.**	• 1400 is ambiguous. ∘ 1.4×10^3 has two significant figures. ∘ 1.40×10^3 three significant figures. ∘ 1.400×10^3 has four significant figures.
5. Trailing zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	• 620.0 has four significant figures. • 19.000 has five significant figures.

Exact Numbers

Integers obtained either by counting objects or from definitions are exact numbers, which are considered to have infinitely many significant figures. If we have counted four objects, for example, then the number 4 has an infinite number of significant figures (i.e., it represents 4.000…). Similarly, 1 foot (ft) is defined to contain 12 inches (in), so the number 12 in the following equation has infinitely many significant figures:

Example 1.11.2

Give the number of significant figures in each. Identify the rule for each.

a. 5.87
b. 0.031
c. 52.90
d. 00.2001
e. 500
f. 6 atoms

Solution

Explanation	Answer
a. All three numbers are significant (rule 1).	5.87 , three significant figures

Explanation	Answer
b. The leading zeros are not significant (rule 3). The 3 and the 1 are significant (rule 1)	0.031, two significant figures
c. The 5, the 2 and the 9 are significant (rule 1). The trailing zero is also significant (rule 5).	52.90, four significant figures
d. The leading zeros are not significant (rule 3). The 2 and the 1 are significant (rule 1) and the middle zeros are also significant (rule 2).	00.2001, four significant figures
e. The number is ambiguous. It could have one, two or three significant figures.	500, ambiguous
f. The 6 is a counting number. A counting number is an exact number.	6, infinite

Exercise 1.11.2

Give the number of significant figures in each.

a. 36.7 m
b. 0.006606 s
c. 2,002 kg
d. 306,490,000 people
e. 3,800 g

Answer a:

three significant figures.

Answer b:

four significant figures.

Answer c:

four significant figures.

Answer d:

Infinite (Exact number)

Answer e:

Ambiguous, could be two, three or four significant figure.

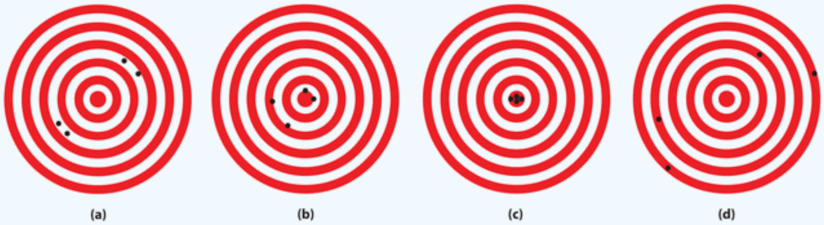
Accuracy and Precision

Measurements may be accurate, meaning that the measured value is the same as the true value; they may be precise, meaning that multiple measurements give nearly identical values (i.e., reproducible results); they may be both accurate and precise; or they may be neither accurate nor precise. The goal of scientists is to obtain measured values that are both accurate and precise. The video below demonstrate the concepts accuracy and precision.

Video 2.3.1: Difference between precision and accuracy.

Example 1.11.3

The following archery targets show marks that represent the results of four sets of measurements.



Which target shows

a. a precise, but inaccurate set of measurements?
b. a set of measurements that is both precise and accurate?
c. a set of measurements that is neither precise nor accurate?

Solution

a. Set a is a precise, but inaccurate
b. Set c is both precise and accurate
c. Set d is neither precise nor accurate

Summary

Uncertainty exists in all measurements. The degree of uncertainty is affected in part by the quality of the measuring tool. Significant figures give an indication of the certainty of a measurement. Rules allow decisions to be made about how many digits to use in any given situation.

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1.12: Significant Figures in Calculations

Learning Objectives

- Use significant figures correctly in arithmetical operations.

Rounding

Before dealing with the specifics of the rules for determining the significant figures in a calculated result, we need to be able to round numbers correctly. To **round** a number, first decide how many significant figures the number should have. Once you know that, round to that many digits, starting from the left. If the number immediately to the right of the last significant digit is less than 5, it is dropped and the value of the last significant digit remains the same. If the number immediately to the right of the last significant digit is greater than or equal to 5, the last significant digit is increased by 1.

Consider the measurement 207.518 m Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process as outlined in Table $1.12.1$.

Table 1.12.1: Rounding examples

Number of Significant Figures	Rounded Value	Reasoning
6	207.518	All digits are significant
5	207.52	8 rounds the 1 up to 2
4	207.5	2 is dropped
3	208	5 rounds the 7 up to 8
2	210	8 is replaced by a 0 and rounds the 0 up to 1
1	200	1 is replaced by a 0

Notice that the more rounding that is done, the less reliable the figure is. An approximate value may be sufficient for some purposes, but scientific work requires a much higher level of detail.

It is important to be aware of significant figures when you are mathematically manipulating numbers. For example, dividing 125 by 307 on a calculator gives 0.4071661238… to an infinite number of digits. But do the digits in this answer have any practical meaning, especially when you are starting with numbers that have only three significant figures each? When performing mathematical operations, there are two rules for limiting the number of significant figures in an answer—one rule is for addition and subtraction, and one rule is for multiplication and division.

> *In operations involving significant figures, the answer is reported in such a way that it reflects the reliability of the* ***least precise*** *operation. An answer is no more precise than the least precise number used to get the answer.*

Multiplication and Division

For multiplication or division, the rule is to count the number of significant figures in each number being multiplied or divided and then limit the significant figures in the answer to the lowest count. An example is as follows:

$$\underbrace{38.65}_{\text{4 sig figs}} \times \underbrace{105.93}_{\text{5 sig figs}} = \underbrace{4{,}094.1945}_{\text{reduce to 4 sig figs}}$$

The final answer, limited to four significant figures, is 4,094. The first digit dropped is 1, so we do not round up.

Scientific notation provides a way of communicating significant figures without ambiguity. You simply include all the significant figures in the leading number. For example, the number 450 has two significant figures and would be written in scientific notation as 4.5×10^2, whereas 450.0 has four significant figures and would be written as 4.500×10^2. In scientific notation, all significant figures are listed explicitly.

✓ Example 1.12.1

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

a. 23.096 × 90.300
b. 125 × 9.000

Solution

a

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
The calculator answer is 2,085.5688, but we need to round it to five significant figures. Because the first digit to be dropped (in the tenths place) is greater than 5, we round up to 2,085.6.	2.0856×10^3

b

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
The calculator gives 1,125 as the answer, but we limit it to three significant figures.	1.13×10^3

Addition and Subtraction

How are significant figures handled in calculations? It depends on what type of calculation is being performed. If the calculation is an addition or a subtraction, the rule is as follows: limit the reported answer to the rightmost column that all numbers have significant figures in common. For example, if you were to add 1.2 and 4.71, we note that the first number stops its significant figures in the tenths column, while the second number stops its significant figures in the hundredths column. We therefore limit our answer to the tenths column.

1.2
$\underline{4.41}$
5.61
↑ limit final answer to the tenths column: 5.6

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

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

77.2
$\underline{10.46}$
87.66
↑ limit final answer to the tenths column and round up: 87.7

✓ Example 1.12.2

a. 13.77 + 908.226
b. 1,027 + 611 + 363.06

Solution

a

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
The calculator answer is 921.996, but because 13.77 has its farthest-right significant figure in the hundredths place, we need to round the final answer to the hundredths position. Because the first digit to be dropped (in the thousandths place) is greater than 5, we round up to 922.00	$922.00 = 9.2200 \times 10^2$

b

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
The calculator gives 2,001.06 as the answer, but because 611 and 1027 has its farthest-right significant figure in the ones place, the final answer must be limited to the ones position.	$2{,}001.06 = 2.001 \times 10^3$

? Exercise 1.12.2

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

a. 217 ÷ 903
b. 13.77 + 908.226 + 515
c. 255.0 − 99
d. 0.00666 × 321

Answer a:

$$0.240 = 2.40 \times 10^{-1}$$

Answer b:

$$1,437 = 1.437 \times 10^3$$

Answer c:

$$156 = 1.56 \times 10^2$$

Answer d:

$$2.14 = 2.14 \times 10^0$$

Remember that calculators do not understand significant figures. *You* are the one who must apply the rules of significant figures to a result from your calculator.

Calculations Involving Multiplication/Division and Addition/Subtraction

In practice, chemists generally work with a calculator and carry all digits forward through subsequent calculations. When working on paper, however, we often want to minimize the number of digits we have to write out. Because successive rounding can compound inaccuracies, intermediate rounding needs to be handled correctly. When working on paper, always round an intermediate result so as to retain at least one more digit than can be justified and carry this number into the next step in the calculation. The final answer is then rounded to the correct number of significant figures at the very end.

Video 1.12.1: Significant figures in mixed operations (https://www.youtube.com/watch?v=yBntMndXQWA).

Video 1.12.2: https://www.youtube.com/watch?v=__csP0NtlGI

In the worked examples in this text, we will often show the results of intermediate steps in a calculation. In doing so, we will show the results to only the correct number of significant figures allowed for that step, in effect treating each step as a separate calculation. This procedure is intended to reinforce the rules for determining the number of significant figures, but in some cases it may give a final answer that differs in the last digit from that obtained using a calculator, where all digits are carried through to the last step.

✓ Example 1.12.3

a. 2(1.008 g) + 15.99 g
b. 137.3 s + 2(35.45 s)
c. $\frac{118.7g}{2} - 35.5g$

Solution

a.

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation for multiplication first. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer

2(1.008 g) + 15.99 g = Perform multiplication first. 2 (1.008 g *4 sig figs*) = **2.016 g** *4 sig figs* The number with the least number of significant figures is 1.008 g; *the number 2 is an exact number and therefore has an infinite number of significant figures*. **Then, perform the addition.** **2.016 g** *thousandths place* + 15.99 g *hundredths place (least precise)* = 18.006 g **Round the final answer.** Round the final answer to the hundredths place since 15.99 has its farthest right significant figure in the hundredths place (least precise).	18.01 g (rounding up)

b.

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation for multiplication first. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	Answer
137.3 s + 2(35.45 s) = **Perform multiplication first.** 2(35.45 s *4 sig figs*) = **70.90 s** *4 sig figs* The number with the least number of significant figures is 35.45; *the number 2 is an exact number and therefore has an infinite number of significant figures.* **Then, perform the addition.** 137.3 s *tenths place (least precise)* + **70.90 s** *hundredths place* = 208.20 s **Round the final answer.** Round the final answer to the tenths place based on 137.3 s.	208.2 s

c.

Table with two columns and 1 row. The first column on the left is labeled, Explanation, and underneath in the row is an explanation for division first. The second column is labeled, Answer, and underneath in the row is an answer.

Explanation	**Answer**
$\frac{118.7g}{2} - 35.5g =$ **Perform division first.** $\frac{118.7g}{2}$ *4 sig figs* = **59.35 g** *4 sig figs* The number with the least number of significant figures is 118.7 g; *the number 2 is an exact number and therefore has an infinite number of significant figures.* **Perform subtraction next.** **59.35 g** *hundredths place* − 35.5 g *tenths place (least precise)* = 23.85 g **Round the final answer.** Round the final answer to the tenths place based on 35.5 g.	23.9 g (rounding up)

? Exercise 1.12.3

Complete the calculations and report your answers using the correct number of significant figures.

a. 5(1.008s) - 10.66 s
b. 99.0 cm+ 2(5.56 cm)

Answer a

-5.62 s

Answer b

110.2 cm

Summary

- Rounding
    - If the number to be dropped is greater than or equal to 5, increase the number to its left by 1 (e.g. 2.9699 rounded to three significant figures is 2.97).
    - If the number to be dropped is less than 5, there is no change (e.g. 4.00443 rounded to four significant figures is 4.004).
- The rule in multiplication and division is that the final answer should have the same number of significant figures as there are in the number with the fewest significant figures.
- The rule in addition and subtraction is that the answer is given the same number of decimal places as the term with the fewest decimal places.

1.13: Problem Solving and Unit Conversions

Learning Objectives

- To convert a value reported in one unit to a corresponding value in a different unit using conversion factors.

During your studies of chemistry (and physics also), you will note that mathematical equations are used in many different applications. Many of these equations have a number of different variables with which you will need to work. You should also note that these equations will often require you to use measurements with their units. Algebra skills become very important here!

Converting Between Units with Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another. A simple conversion factor can convert meters into centimeters, or a more complex one can convert miles per hour into meters per second. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. Always remember that a conversion factor has to represent a fact; this fact can either be simple or more complex. For instance, you already know that 12 eggs equal 1 dozen. A more complex fact is that the speed of light is 1.86×10^5 miles/sec. Either one of these can be used as a conversion factor depending on what type of calculation you are working with (Table $1.13.1$).

Table 1.13.1: Conversion Factors from SI units to English Units

English Units	Metric Units	Quantity
1 ounce (oz)	28.35 grams (g)	*mass
1 fluid once (oz)	29.6 mL	volume
2.205 pounds (lb)	1 kilogram (kg)	*mass
1 inch (in)	2.54 centimeters (cm)	length
0.6214 miles (mi)	1 kilometer (km)	length
1 quarter (qt)	0.95 liters (L)	volume

*Pounds and ounces are technically units of force, not mass, but this fact is often ignored by the non-scientific community.

Of course, there are other ratios which are not listed in Table $1.13.1$. They may include:

- Ratios embedded in the text of the problem (using words such as *per* or *in each*, or using symbols such as / or %).
- Conversions in the metric system, as covered earlier in this chapter.
- Common knowledge ratios (such as 60 seconds $=$ 1 minute).

If you learned the SI units and prefixes described, then you know that 1 cm is 1/100th of a meter.

$$1 \text{ cm} = \frac{1}{100} \text{ m} = 10^{-2} \text{m}$$

or

$$100 \text{ cm} = 1 \text{ m}$$

Suppose we divide both sides of the equation by 1m (both the number *and* the unit):

$$\frac{100 \text{ cm}}{1 \text{ m}} = \frac{1 \text{ m}}{1 \text{ m}}$$

As long as we perform the same operation on both sides of the equals sign, the expression remains an equality. Look at the right side of the equation; it now has the same quantity in the numerator (the top) as it has in the denominator (the bottom). Any fraction that has the same quantity in the numerator and the denominator has a value of 1:

$$\frac{100 \text{ cm}}{1 \text{ m}} = \frac{1000 \text{ mm}}{1 \text{ m}} = \frac{1 \times 10^6 \mu\text{m}}{1 \text{ m}} = 1$$

We know that 100 cm *is* 1 m, so we have the same quantity on the top and the bottom of our fraction, although it is expressed in different units.

Performing Dimensional Analysis

Dimensional analysis is amongst the most valuable tools that physical scientists use. Simply put, it is the conversion between an amount in one unit to the corresponding amount in a desired unit using various conversion factors. This is valuable because certain measurements are more accurate or easier to find than others. The use of units in a calculation to ensure that we obtain the final proper units is called *dimensional analysis*.

Here is a simple example. How many centimeters are there in 3.55 m? Perhaps you can determine the answer in your head. If there are 100 cm in every meter, then 3.55 m equals 355 cm. To solve the problem more formally with a conversion factor, we first write the quantity we are given, 3.55 m. Then we multiply this quantity by a conversion factor, which is the same as multiplying it by 1. We can write 1 as $\frac{100\text{ cm}}{1\text{ m}}$ and multiply:

$$3.55\text{ m} \times \frac{100\text{ cm}}{1\text{ m}}$$

The 3.55 m can be thought of as a fraction with a 1 in the denominator. Because m, the abbreviation for meters, occurs in both the numerator *and* the denominator of our expression, they cancel out:

$$\frac{3.55\ \cancel{\text{m}}}{1} \times \frac{100\text{ cm}}{1\ \cancel{\text{m}}}$$

The final step is to perform the calculation that remains once the units have been canceled:

$$\frac{3.55}{1} \times \frac{100\text{ cm}}{1} = 355\text{ cm}$$

In the final answer, we omit the 1 in the denominator. Thus, by a more formal procedure, we find that 3.55 m equals 355 cm. A generalized description of this process is as follows:

quantity (in old units) × conversion factor = quantity (in new units)

You may be wondering why we use a seemingly complicated procedure for a straightforward conversion. In later studies, the conversion problems you encounter *will not always be so simple*. If you master the technique of applying conversion factors, you will be able to solve a large variety of problems.

In the previous example, we used the fraction $\frac{100\text{ cm}}{1\text{ m}}$ as a conversion factor. Does the conversion factor $\frac{1\text{ m}}{100\text{ cm}}$ also equal 1? Yes, it does; it has the same quantity in the numerator as in the denominator (except that they are expressed in different units). Why did we not use *that* conversion factor? If we had used the second conversion factor, the original unit would not have canceled, and the result would have been meaningless. Here is what we would have gotten:

$$3.55\text{ m} \times \frac{1\text{ m}}{100\text{ cm}} = 0.0355\frac{\text{m}^2}{\text{cm}}$$

For the answer to be meaningful, we have to *construct the conversion factor in a form that causes the original unit to cancel out*. Figure 1.13.1 shows a concept map for constructing a proper conversion.

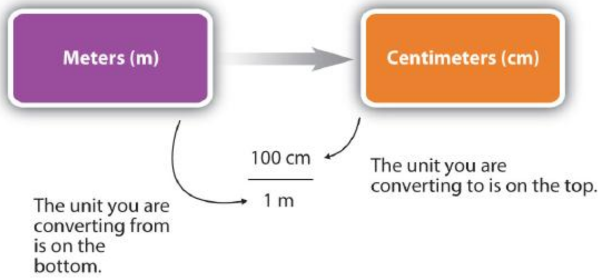


Figure 1.13.1: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another. Converting meters to centimeters: the unit you are converting is on the bottom of the fraction, and the unit to convert to is on top

General Steps in Performing Dimensional Analysis

1. Identify the "**given**" information in the problem. Look for a number with units to start this problem with.
2. What is the problem asking you to "**find**"? In other words, what unit will your answer have?
3. Use **ratios** and conversion factors to cancel out the units that aren't part of your answer, and leave you with units that are part of your answer.
4. When your units cancel out correctly, you are ready to do the **math**. You are multiplying fractions, so you multiply the top numbers and divide by the bottom numbers in the fractions.

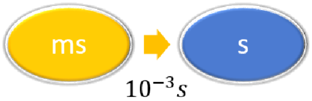
Significant Figures in Conversions

How do conversion factors affect the determination of significant figures?

- Numbers in conversion factors based on prefix changes, such as kilograms to grams, are *not* considered in the determination of significant figures in a calculation because the numbers in such conversion factors are exact.
- Exact numbers are defined or counted numbers, not measured numbers, and can be considered as having an infinite number of significant figures. (In other words, 1 kg is exactly 1,000 g, by the definition of kilo-.)
- Counted numbers are also exact. If there are 16 students in a classroom, the number 16 is exact.
- In contrast, conversion factors that come from measurements (such as density, as we will see shortly) or that are approximations have a limited number of significant figures and should be considered in determining the significant figures of the final answer.

Example 1.13.1

Steps for Problem Solving for Example 2.6.1 and 2.6.2

	Example 1.13.1	**Example 1.13.2**
Steps for Problem Solving	The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?	A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?
Identify the "given" information and what the problem is asking you to "find."	Given: 4.7 L Find: mL	Given: 18 ms Find: s
List other known quantities.	$1\,mL = 10^{-3}L$	$1\,ms = 10^{-3}s$
Prepare a concept map and use the proper conversion factor.	 $\frac{1mL}{10^{-3}L}$	 $\frac{10^{-3}s}{1ms}$
Cancel units and calculate.	$4.7\ \cancel{\text{L}} \times \frac{1\ \text{mL}}{10^{-3}\ \cancel{\text{L}}} = 4,700\ \text{mL}$ or $4.7\ \cancel{\text{L}} \times \frac{1,000\ \text{mL}}{1\ \cancel{\text{L}}} = 4,700\ \text{mL}$ or 4.7 x 10^3 2SF, not ambiguous	$18\ \cancel{\text{ms}} \times \frac{10^{-3}\ \text{s}}{1\ \cancel{\text{ms}}} = 0.018\ \text{s}$ or $18\ \cancel{\text{ms}} \times \frac{1\ \text{s}}{1,000\ \cancel{\text{ms}}} = 0.018\ \text{s}$
Think about your result.	The amount in mL should be 1000 times larger than the given amount in L.	The amount in s should be 1/1000 the given amount in ms.

Exercise 1.13.1

Perform each conversion.

a. 101,000 ns to seconds
b. 32.08 kg to grams
c. 1.53 grams to cg

Answer a:

$1.01000x10^{-4}s$

Answer b:

$3.208x10^{4}g$

Answer c:

$1.53x10^{2}cg$

Summary

- Conversion factors are used to convert one unit of measurement into another.
- Dimensional analysis (unit conversions) involves the use of conversion factors that will cancel unwanted units and produce the appropriate units.

1.14: The Nuclear Atom

Learning Objectives

- Describe the history of the atom.
- Draw a diagram of a model of the atom and label the nucleus and the electron cloud.

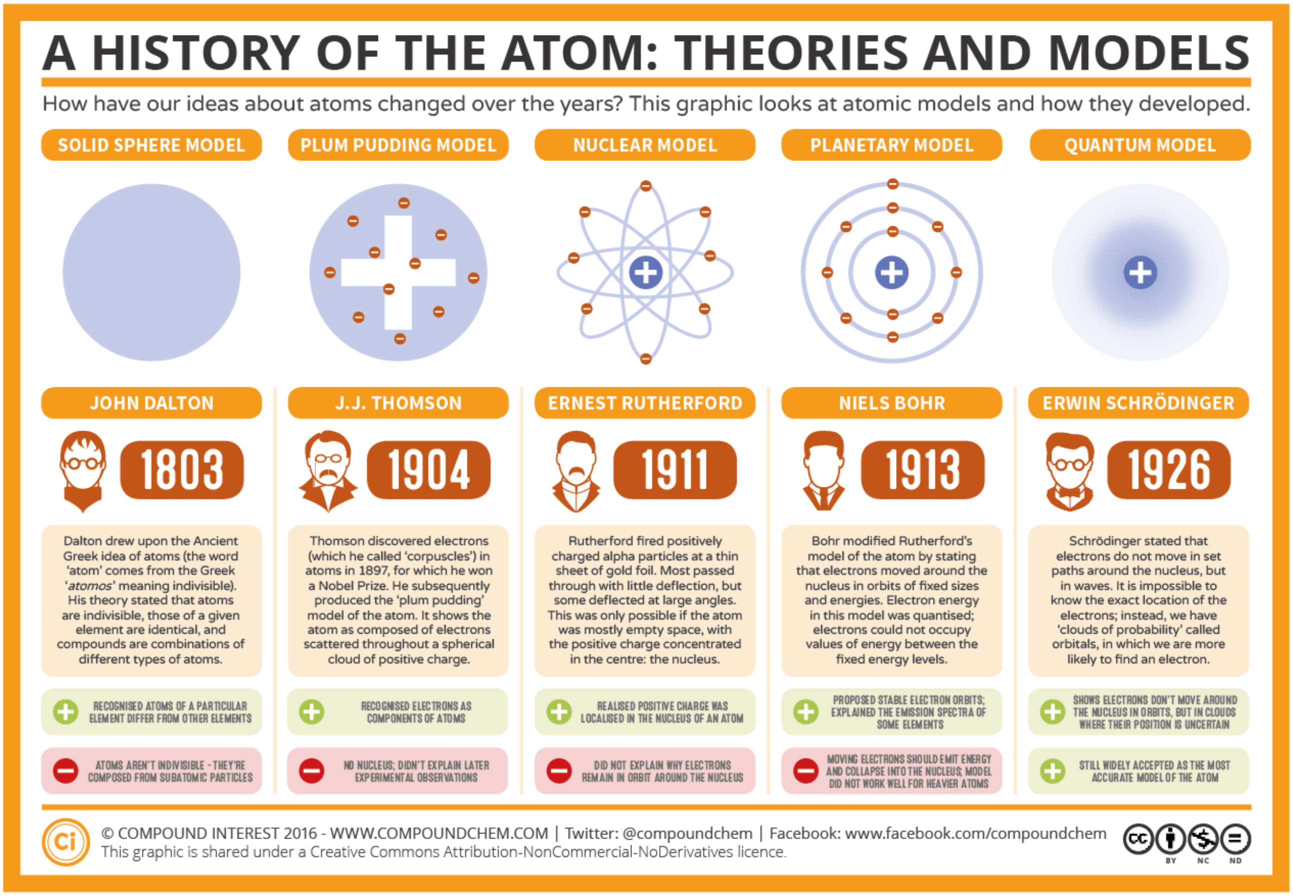
The history of the atom begins before 1800, where an atom was thought to be the smallest piece of matter. The first scientist to provide a theory about the atom was John Dalton. His atomic theory held up well to a lot of the different chemical experiments that scientists performed to test it. The following statements support Dalton's Atomic Theory.

- All matter is composed of extremely small particles called atoms.
- Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
- Atoms cannot be subdivided, created, or destroyed.
- Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
- In chemical reactions, atoms are combined, separated, or rearranged

In fact, for almost 100 years, it seemed as if Dalton's Atomic Theory was the whole truth. However, in 1897, a scientist named J. J. Thomson conducted some research that suggested that Dalton's Atomic Theory was not the entire story. He suggested that the small, negatively charged particles making up the cathode ray were actually pieces of atoms. He called these pieces "corpuscles," although today we know them as **electrons**. Thanks to his clever experiments and careful reasoning, J. J. Thomson is credited with the discovery of the electron.

History and why models are important

Below is a timeline of the understanding of the atom. If you read through the first four models, you will see they will have a reason why that specific model did not work. The only model that is acceptable today is the quantum model that was devised by Erwin Schrodinger. The nuclear contains the protons and neutrons (that is evidence in the three models (nuclear, planetary and quantum). In all three models you can see that electrons are all outside of the nucleus. However, the nuclear model and planetary (Bohr) do not show an accurate depiction of where electrons are in the atom. The quantum model is appropriate because Schrondinger shows that electrons exist in the in clouds of probability rather than exact levels. Quantum chemistry is not discussed in this course, but it helps to describe the placement of electrons. Electrons can exist as a particle or wave (discussed in Module 2), this means that they aren't sitting levels of an electron.


A HISTORY OF THE ATOM: THEORIES AND MODELS

How have our ideas about atoms changed over the years? This graphic looks at atomic models and how they developed.

SOLID SPHERE MODEL

JOHN DALTON 1803

Dalton drew upon the Ancient Greek idea of atoms (the word 'atom' comes from the Greek '*atomos*' meaning indivisible). His theory stated that atoms are indivisible, those of a given element are identical, and compounds are combinations of different types of atoms.

\+ RECOGNISED ATOMS OF A PARTICULAR ELEMENT DIFFER FROM OTHER ELEMENTS

− ATOMS AREN'T INDIVISIBLE - THEY'RE COMPOSED FROM SUBATOMIC PARTICLES

PLUM PUDDING MODEL

J.J. THOMSON 1904

Thomson discovered electrons (which he called 'corpuscles') in atoms in 1897, for which he won a Nobel Prize. He subsequently produced the 'plum pudding' model of the atom. It shows the atom as composed of electrons scattered throughout a spherical cloud of positive charge.

\+ RECOGNISED ELECTRONS AS COMPONENTS OF ATOMS

− NO NUCLEUS; DIDN'T EXPLAIN LATER EXPERIMENTAL OBSERVATIONS

NUCLEAR MODEL

ERNEST RUTHERFORD 1911

Rutherford fired positively charged alpha particles at a thin sheet of gold foil. Most passed through with little deflection, but some deflected at large angles. This was only possible if the atom was mostly empty space, with the positive charge concentrated in the centre: the nucleus.

\+ REALISED POSITIVE CHARGE WAS LOCALISED IN THE NUCLEUS OF AN ATOM

− DID NOT EXPLAIN WHY ELECTRONS REMAIN IN ORBIT AROUND THE NUCLEUS

PLANETARY MODEL

NIELS BOHR 1913

Bohr modified Rutherford's model of the atom by stating that electrons moved around the nucleus in orbits of fixed sizes and energies. Electron energy in this model was quantised; electrons could not occupy values of energy between the fixed energy levels.

\+ PROPOSED STABLE ELECTRON ORBITS; EXPLAINED THE EMISSION SPECTRA OF SOME ELEMENTS

− MOVING ELECTRONS SHOULD EMIT ENERGY AND COLLAPSE INTO THE NUCLEUS; MODEL DID NOT WORK WELL FOR HEAVIER ATOMS

QUANTUM MODEL

ERWIN SCHRÖDINGER 1926

Schrödinger stated that electrons do not move in set paths around the nucleus, but in waves. It is impossible to know the exact location of the electrons; instead, we have 'clouds of probability' called orbitals, in which we are more likely to find an electron.

\+ SHOWS ELECTRONS DON'T MOVE AROUND THE NUCLEUS IN ORBITS, BUT IN CLOUDS WHERE THEIR POSITION IS UNCERTAIN

\+ STILL WIDELY ACCEPTED AS THE MOST ACCURATE MODEL OF THE ATOM

Figure 1. The history of the atom (CC BY-NC-ND, CompoundChem.com)

Discovery of the Electron

The electron was discovered by J.J. Thomson in 1897. The existence of protons was also known, as was the fact that atoms were neutral in charge. Since the intact atom had no net charge and the electron and proton had opposite charges, the next step after the discovery of subatomic particles was to figure out how these particles were arranged in the atom. This is a difficult task because of the incredibly small size of the atom. Therefore, scientists set out to design a model of what they believed the atom could look like. The goal of each **atomic model** was to accurately represent all of the experimental evidence about atoms in the simplest way possible.

Following the discovery of the electron, J.J. Thomson developed what became known as the "**plum pudding**" model in 1904. Plum pudding is an English dessert similar to a blueberry muffin. In Thomson's plum pudding model of the atom, the electrons were embedded in a uniform sphere of positive charge like blueberries stuck into a muffin. The positive matter was thought to be jelly-like or a thick soup. The electrons were somewhat mobile. As they got closer to the outer portion of the atom, the positive charge in the region was greater than the neighboring negative charges and the electron would be pulled back more toward the center region of the atom.

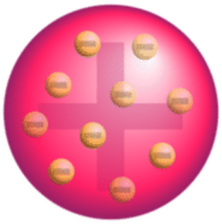


Figure 1.14.2: The "plum pudding" model.

However, this model of the atom soon gave way to a new model developed by New Zealander Ernest Rutherford (1871 - 1937) about five years later. Thomson did still receive many honors during his lifetime, including being awarded the Nobel Prize in Physics in 1906 and a knighthood in 1908.

Discovery of the Nucleus

In 1911, Rutherford and coworkers Hans Geiger and Ernest Marsden initiated a series of groundbreaking experiments that would completely change the accepted model of the atom. They bombarded very thin sheets of gold foil with fast moving **alpha particles**. Alpha particles, a type of natural radioactive particle, are positively charged particles with a mass about four times that of a hydrogen atom.

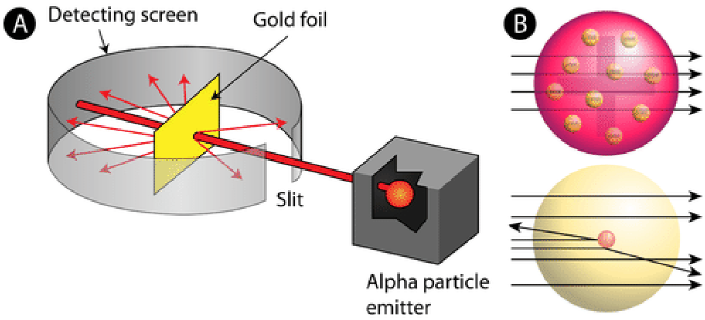


Figure 1.14.3: (A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil that was surrounded by a screen which would allow detection of the deflected particles. (B) According to the plum pudding model (top) all of the alpha particles should have passed through the gold foil with little or no deflection. Rutherford found that a small percentage of alpha particles were deflected at large angles, which could be explained by an atom with a very small, dense, positively-charged nucleus at its center (bottom).

According to the accepted atomic model (Quantum Model), in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all. Surprisingly, while most of the alpha particles were indeed undeflected, a very small percentage (about 1 in 8000 particles) bounced off the gold foil at very large angles. Some were even redirected back toward the source. No prior knowledge had prepared them for this discovery. In a famous quote, Rutherford exclaimed that it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue and it came back and hit you."

Rutherford needed to come up with an entirely new model of the atom in order to explain his results. Because the vast majority of the alpha particles had passed through the gold, he reasoned that most of the atom was empty space. In contrast, the particles that were highly deflected must have experienced a tremendously powerful force within the atom. He concluded that all of the positive charge and the majority of the mass of the atom must be concentrated in a very small space in the atom's interior, which he called the nucleus. The **nucleus** is the tiny, dense, central core of the atom and is composed of protons and neutrons.

Rutherford's atomic model became known as the nuclear model. In the nuclear atom, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in the nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom. It is worth emphasizing just how small the nucleus is compared to the rest of the atom. If we could blow up an atom to be the size of a large professional football stadium, the nucleus would be about the size of a marble.

The Best Model

Rutherford's model proved to be an important step towards a full understanding of the atom. However, it did not completely address the nature of the electrons and the way in which they occupied the vast space around the nucleus. It was not until some years later that a full understanding of the electron was achieved. Niels Bohr later modified the nuclear model to show electrons in concentric circles that showed that electrons could move between energy levels known as the Bohr Model. However, this model still did not address the full nature of electrons. Today, we understand the atom in a quantum mechanical way. The quantum model is the most appropriate description of the atom. There is a nucleus, and electrons are found in specific areas of the electron cloud outside of the nucleus. The quantum model has provided us a better understanding to the chemical properties of elements.

Atomic Nucleus

The **nucleus** (plural, nuclei) is a positively charged region at the center of the atom. It consists of two types of subatomic particles packed tightly together. The particles are protons, which have a positive electric charge, and neutrons, which are neutral in electric charge. Outside of the nucleus, an atom is mostly empty space, with orbiting negative particles called electrons whizzing through it. The figure below shows these parts of the atom.

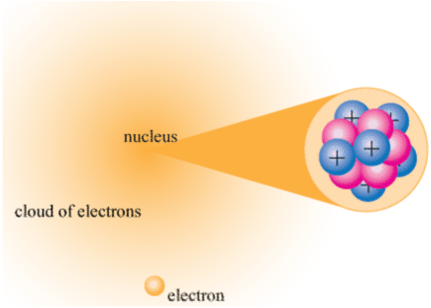


Figure 1.14.4: The nuclear atom

The nucleus of the atom is extremely small. Its radius is only about 1/100,000 of the total radius of the atom. If an atom were the size of a football stadium, the nucleus would be about the size of a pea! Electrons have virtually no mass, but protons and neutrons have a lot of mass for their size. As a result, the nucleus has virtually all the mass of an atom. Given its great mass and tiny size, the nucleus is very dense. If an object the size of a penny had the same density as the nucleus of an atom, its mass would be greater than 30 million tons!

Holding It All Together

Particles with opposite electric charges attract each other. This explains why negative electrons orbit the positive nucleus. Particles with the same electric charge repel each other. This means that the positive protons in the nucleus push apart from one another. So why doesn't the nucleus fly apart? An even stronger force - called the **strong nuclear force** - holds protons and neutrons together in the nucleus.

Summary

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.
- Bombardment of gold foil with alpha particles showed that some particles were deflected.
- The most appropriate model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons known as the quantum model.

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1.15: The Properties of Protons, Neutrons, and Electrons

Learning Objectives

- Describe the locations, charges, and masses of the three main subatomic particles.
- Determine the number of protons and electrons in an atom.
- Define atomic mass unit (amu).

The atom is composed of three subatomic particles.

Electrons

Electrons are one of three main types of particles that make up atoms. The other two types are protons and neutrons. Unlike protons and neutrons, which consist of smaller, simpler particles, electrons are fundamental particles that do not consist of smaller particles. They are a type of fundamental particles called leptons. All leptons have an electric charge of -1 or 0. Electrons are extremely small. The mass of an electron is only about 1/2000 the mass of a proton or neutron, so electrons contribute virtually nothing to the total mass of an atom. Electrons have an electric charge of -1, which is equal but opposite to the charge of a proton, which is $+1$. All atoms have the same number of electrons as protons, so the positive and negative charges "cancel out", making atoms electrically neutral.

Unlike protons and neutrons, which are located inside the nucleus at the center of the atom, electrons are found outside the nucleus. Because opposite electric charges attract each other, negative electrons are attracted to the positive nucleus. This force of attraction keeps electrons constantly moving through the otherwise empty space around the nucleus. The figure below is a common way to represent the structure of an atom. It shows the electron as a particle orbiting the nucleus, similar to the way that planets orbit the sun. This is however, an incorrect perspective, as electrons are more complicated as quantum mechanics demonstrate.

Figure 1.15.1: Electrons are much smaller than protons or neutrons. If an electrons were the mass of a penny, a proton or a neutrons would have the mass of a large bowling ball!

Protons

A proton is one of three main particles that make up the atom. Protons are found in the nucleus of the atom. This is a tiny, dense region at the center of the atom. Protons have a positive electrical charge of one $(+1)$ and a mass of 1 atomic mass unit (amu), which is about 1.67×10^{-27} kilograms. The number of protons in an element can be found on the periodic table. It is known as the atomic number.

Neutrons

Atoms of all elements - except for most atoms of hydrogen - have neutrons in their nucleus. Unlike protons and electrons, which are electrically charged, neutrons have no charge - they are electrically neutral. That's why the neutrons in the diagram above are labeled n^0. The zero stands for "zero charge". The mass of a neutron is slightly greater than the mass of a proton, which is 1 atomic mass unit (amu). (An atomic mass unit equals about 1.67×10^{-27} kilograms.) A neutron also has about the same diameter as a proton, or 1.7×10^{-15} meters.

As you might have already guessed from its name, the neutron is neutral. In other words, it has no charge whatsoever and is therefore neither attracted to nor repelled from other objects. Neutrons are in every atom (with one exception), and they're bound together with other neutrons and protons in the atomic nucleus.

Before we move on, we must discuss how the different types of subatomic particles interact with each other. When it comes to neutrons, the answer is obvious. Since neutrons are neither attracted to nor repelled from objects, they don't really interact with protons or electrons (beyond being bound into the nucleus with the protons).

Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When you compare the masses of electrons, protons, and neutrons, what you find is that electrons have an extremely small mass, compared to either protons or neutrons. On the other hand, the masses of protons and neutrons are fairly similar. Because protons and neutrons are so much more massive than electrons, almost all of the mass of any atom comes from the nucleus, which contains all of the neutrons and protons.

Table 1.15.1: Properties of Subatomic Particles

Particle	Symbol	Mass (amu)	Relative Mass (proton = 1)	Relative Charge	Location
proton	p^+	1	1	+1	inside the nucleus
electron	e^-	5.45×10^{-4}	0.00055	−1	outside nucleus
neutron	n^0	1	1	0	inside the nucleus

Table $1.15.1$ gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in "atomic mass units." An **atomic mass unit (amu)** is defined as one-twelfth the mass of a carbon-12 atom. Atomic mass units (amu) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1 in this unit system.

Negative and positive charges of equal magnitude cancel each other out. This means that the negative charge on an electron perfectly balances the positive charge on the proton. In other words, a neutral atom must have exactly one electron for every proton. If a neutral atom has 1 proton, it must have 1 electron. If a neutral atom has 2 protons, it must have 2 electrons. If a neutral atom has 10 protons, it must have 10 electrons. You get the idea. In order to be neutral, an atom must have the same number of electrons and protons.

Summary

- Electrons are a type of subatomic particle with a negative charge.
- Protons are a type of subatomic particle with a positive charge. Protons are bound together in an atom's nucleus as a result of the strong nuclear force.
- Neutrons are a type of subatomic particle with no charge (they're neutral). Like protons, neutrons are bound into the atom's nucleus as a result of the strong nuclear force.
- Protons and neutrons have approximately the same mass, but they are both much more massive than electrons (approximately 2,000 times as massive as an electron).
- The positive charge on a proton is equal in magnitude to the negative charge on an electron. As a result, a neutral atom must have an equal number of protons and electrons.
- The atomic mass unit (amu) is a unit of mass equal to one-twelfth the mass of a carbon-12 atom

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1.16: Elements- Defined by Their Number of Protons

Learning Objectives

- Define atomic number.
- Define mass number.
- Determine the number of protons, neutrons, and electrons in an atom.

It's important to be able to distinguish atoms of one element from atoms of another element. Elements are pure substances that make up all other matter, so each one is given a unique name. The names of elements are also represented by unique one- or two-letter symbols, such as H for hydrogen, C for carbon, or He for helium. However, it would more powerful if these names could be used to identify the numbers of protons and neutrons in the atoms. That's where atomic number and mass number are useful.

Figure 1.16.1: It is difficult to find qualities that differ between each element, and to distinguish one element from another. Each element, however, does have a unique number of protons. Sulfur has 16 protons, silicon has 14 protons, and gold has 79 protons. Images used with permission (public domain for sulfur and silicon, gold is licensed by CC-BY-SA-NC-ND; Alchemist-hp).

Atomic Number

Scientists distinguish between different elements by counting the number of protons in the nucleus (Table $1.16.1$). If an atom has only one proton, we know that it's a hydrogen atom. An atom with two protons is always a helium atom. If scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom . . . the list goes on.

Since an atom of one element can be distinguished from an atom of another element by the number of protons in its nucleus, scientists are always interested in this number, and how this number differs between different elements. The number of protons in an atom is called its **atomic number** (Z). This number is very important because it is unique for atoms of a given element. All atoms of an element have the same number of protons, and every element has a different number of protons in its atoms. For example, all helium atoms have two protons, and no other elements have atoms with two protons.

Table 1.16.1: Atoms of the First Six Elements

Name	Protons	Neutrons	Electrons	Atomic Number (Z)	Mass Number (A)
Hydrogen	1	0	1	1	1
Helium	2	2	2	2	4
Lithium	3	4	3	3	7
Beryllium	4	5	4	4	9
Boron	5	6	5	5	11
Carbon	6	6	6	6	12

Of course, since neutral atoms have to have one electron for every proton, an element's atomic number also tells you how many electrons are in a neutral atom of that element. For example, hydrogen has an atomic number of 1. This means that an atom of hydrogen has one proton, and, if it's neutral, one electron as well. Gold, on the other hand, has an atomic number of 79, which means that an atom of gold has 79 protons, and, if it's neutral, 79 electrons as well.

Neutral Atoms

Atoms are neutral in electrical charge because they have the same number of negative electrons as positive protons (Table $1.16.1$). Therefore, the atomic number of an atom also tells you how many electrons the atom has. This, in turn, determines many of the atom's chemical properties.

Mass Number

The **mass number** (A) of an atom is the total number of protons and neutrons in its nucleus. The mass of the atom is a unit called the atomic mass unit (amu). One atomic mass unit is the mass of a proton, or about 1.67×10^{-27} kilograms, which is an extremely small mass. A neutron has just a tiny bit more mass than a proton, but its mass is often assumed to be one atomic mass unit as well. Because electrons have virtually no mass, just about all the mass of an atom is in its protons and neutrons. Therefore, the total number of protons and neutrons in an atom determines its mass in atomic mass units (Table $1.16.1$).

Consider helium again. Most helium atoms have two neutrons in addition to two protons. Therefore the mass of most helium atoms is 4 atomic mass units ($2\ \mathrm{amu}$ for the protons + $2\ \mathrm{amu}$ for the neutrons). However, some helium atoms have more or less than two neutrons. Atoms with the same number of protons but different numbers of neutrons are called isotopes. Because the number of neutrons can vary for a given element, the mass numbers of different atoms of an element may also vary. For example, some helium atoms have three neutrons instead of two (these are called isotopes and are discussed in detail later on).

Why do you think that the "mass number" includes protons and neutrons, but not electrons? You know that most of the mass of an atom is concentrated in its nucleus. The mass of an atom depends on the number of protons and neutrons. You have already learned that the mass of an electron is very, very small compared to the mass of either a proton or a neutron (like the mass of a penny compared to the mass of a bowling ball). Counting the number of protons and neutrons tells scientists about the total mass of an atom.

$$\text{mass number } A = (\text{number of protons}) + (\text{number of neutrons})$$

An atom's mass number is very easy to calculate, provided that you know the number of protons and neutrons in an atom.

✓ Example 4.5.1

What is the mass number of an atom of helium that contains 2 neutrons?

Solution

$(\text{number of protons}) = 2$ (Remember that an atom of helium always has 2 protons.)

$(\text{number of neutrons}) = 2$

$\text{mass number} = (\text{number of protons}) + (\text{number of neutrons})$

$$\text{mass number} = \mathbf{2 + 2 = 4}$$

A **chemical symbol** is a one- or two-letter designation of an element. Some examples of chemical symbols are O for oxygen, Zn for zinc, and Fe for iron. The first letter of a symbol is always capitalized. If the symbol contains two letters, the second letter is lower case. The majority of elements have symbols that are based on their English names. However, some of the elements that have been known since ancient times have maintained symbols that are based on their Latin names, as shown in Table $1.16.2$

Table $1.16.2$: *Symbols and Latin Names for Elements*

Chemical Symbol	Name	Latin Name
Na	Sodium	Natrium
K	Potassium	Kalium
Fe	Iron	Ferrum
Cu	Copper	Cuprum
Ag	Silver	Argentum

Chemical Symbol	Name	Latin Name
Sn	Tin	Stannum
Sb	Antimony	Stibium
Au	Gold	Aurum
Pb	Lead	Plumbum

Summary

- Elements are pure substances that make up all matter, so each one is given a unique name.
- The names of elements are also represented by unique one- or two-letter symbols.
- Each element has a unique number of protons. An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
- The mass number of an atom is the sum of the protons and neutrons in the atom.
- Isotopes are atoms of the same element (same number of protons) that have different numbers of neutrons in their atomic nuclei.

1.17: Ions

As discussed in the previous section, the number of protons in an atom determines to which element an atom belongs. If the number of protons changes (as you will in see in unit 2 can happen during nuclear reactions) the identity of the atom changes. In a neutral atom, the number of protons and electrons are equal to each other. However, the number of electrons in an atom can change. If the number of electrons changes compared to the number in a neutral atom, the resulting particle is charged. Why? Consider the following examples.

Element	Protons	Electrons	Net Charge	Particle	Symbol
Potassium atom	19	19	0	Neutral atom	K
Potassium ion	19	18	+1	Cation	K^+
Sulfur atom	16	16	0	Neutral atom	S
Sulfur ion	16	18	−2	Anion	S^{2-}

Protons are positively charged particles and the number of protons is determined by the atomic number of the element. Since an electron has the same magnitude of charge as a proton but opposite sign (electrons are negative), in order for an atom to be neutral, to have no net charge, the number of protons and electrons are equal to each other. If, however, there are fewer electrons than protons as in the case of the potassium ion, the net charge is positive since there is one more positive proton than negative electrons [(+19) + (−18) = +1]. Positive ions are **cations**. When there are more electrons than protons, as in the case of the sulfur ion, the net charge is negative [(+16) + (−18) = −2]. Negative ions are **anions**.

Complete the following table and consider how you would explain your answers if you were teaching this concept to a peer.

Element	Protons	Electrons	Net Charge	Particle	Symbol
Strontium atom			0		Sr
Strontium ion			+2	Cation	
Bromine atom				Neutral atom	
Bromine ion					Br^-

1.18: Isotopes - When the Number of Neutrons Varies

Learning Objectives

- Explain what isotopes are and how an isotope affect an element's atomic mass.
- Determine the number of protons, electrons, and neutrons of an element with a given mass number.

All atoms of the same element have the same number of protons, but some may have different numbers of neutrons. For example, all carbon atoms have six protons, and most have six neutrons as well. But some carbon atoms have seven or eight neutrons instead of the usual six. Atoms of the same element that differ in their numbers of neutrons are called **isotopes**. Many isotopes occur naturally. Usually one or two isotopes of an element are the most stable and common. Different isotopes of an element generally have the same physical and chemical properties. That's because they have the same numbers of protons and electrons.

An Example: Hydrogen Isotopes

Hydrogen is an example of an element that has isotopes. Three isotopes of hydrogen are modeled in Figure 1.18.1. Most hydrogen atoms have just one proton and one electron and lack a neutron. These atoms are just called hydrogen. Some hydrogen atoms have one neutron as well. These atoms are the isotope named deuterium. Other hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

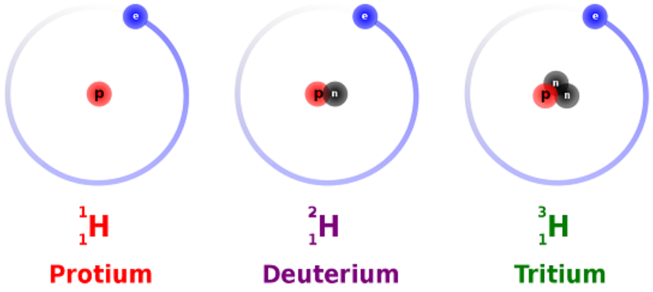


Figure 1.18.1: The three most stable isotopes of hydrogen: protium (A = 1), deuterium (A = 2), and tritium (A = 3). (CC SA-BY 3.0; Balajijagadesh).

For most elements other than hydrogen, isotopes are named for their mass number. For example, carbon atoms with the usual 6 neutrons have a mass number of 12 (6 protons + 6 neutrons = 12), so they are called carbon-12. Carbon atoms with 7 neutrons have atomic mass of 13 (6 protons + 7 neutrons = 13). These atoms are the isotope called carbon-13.

Example 1.18.1: Lithium Isotopes

a. What is the atomic number and the mass number of an isotope of lithium containing 3 neutrons.
b. What is the atomic number and the mass number of an isotope of lithium containing 4 neutrons?

Solution

A lithium atom contains 3 protons in its nucleus irrespective of the number of neutrons or electrons.

a.

$$\begin{aligned} \text{atomic number} = (\text{number of protons}) &= 3 \\ (\text{number of neutrons}) &= 3 \end{aligned}$$

$$\begin{aligned} \text{mass number } &= (\text{number of protons}) + (\text{number of neutrons}) \\ \text{mass number } &= 3 + 3 \\ &= 6 \end{aligned}$$

b.

$$\begin{aligned} \text{atomic number} = (\text{number of protons}) &= 3 \\ (\text{number of neutrons}) &= 4 \end{aligned}$$

$$\begin{aligned}\text{mass number} &= (\text{number of protons}) + (\text{number of neutrons})\\ \text{mass number} &= 3+4\\ &= 7\end{aligned}$$

Notice that because the lithium atom always has 3 protons, the atomic number for lithium is always 3. The mass number, however, is 6 in the isotope with 3 neutrons, and 7 in the isotope with 4 neutrons. In nature, only certain isotopes exist. For instance, lithium exists as an isotope with 3 neutrons, and as an isotope with 4 neutrons, but it doesn't exist as an isotope with 2 neutrons or as an isotope with 5 neutrons.

Stability of Isotopes

Atoms need a certain ratio of neutrons to protons to have a stable nucleus. Having too many or too few neutrons relative to protons results in an unstable, or radioactive, nucleus that will sooner or later break down to a more stable form. This process is called radioactive decay. Many isotopes have radioactive nuclei, and these isotopes are referred to as radioisotopes. When they decay, they release particles that may be harmful. Some we experience everyday, like Potassium-40 in bananas. However, some radioactive isotopes can be dangerous and why we need to take additional precautions when working with these radioisotopes. The isotope of carbon known as carbon-14 is an example of a radioisotope. In contrast, the carbon isotopes called carbon-12 and carbon-13 are stable.

This whole discussion of isotopes brings us back to Dalton's Atomic Theory. According to Dalton, atoms of a given element are identical. But if atoms of a given element can have different numbers of neutrons, then they can have different masses as well! How did Dalton miss this? It turns out that elements found in nature exist as constant uniform mixtures of their naturally occurring isotopes. In other words, a piece of lithium always contains both types of naturally occurring lithium (the type with 3 neutrons and the type with 4 neutrons). Moreover, it always contains the two in the same relative amounts (or "relative abundances"). In a chunk of lithium, 93% will always be lithium with 4 neutrons, while the remaining 7% will always be lithium with 3 neutrons.

Dalton always experimented with large chunks of an element - chunks that contained all of the naturally occurring isotopes of that element. As a result, when he performed his measurements, he was actually observing the averaged properties of all the different isotopes in the sample. For most of our purposes in chemistry, we will do the same thing and deal with the average mass of the atoms. Luckily, aside from having different masses, most other properties of different isotopes are similar.

There are two main ways in which scientists frequently show the mass number of an atom they are interested in. It is important to note that the mass number is *not* given on the periodic table. These two ways include writing a nuclear symbol or by giving the name of the element with the mass number written.

To write a **nuclear symbol**, the mass number is placed at the upper left (superscript) of the chemical symbol and the atomic number is placed at the lower left (subscript) of the symbol. The complete nuclear symbol for helium-4 is drawn below:

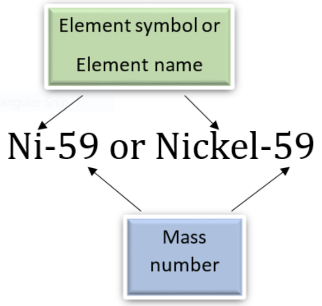
The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

$$^{59}_{28}\text{Ni} \tag{1.18.1}$$

$$^{238}_{92}\text{U} \tag{1.18.2}$$

In the nickel nucleus represented above, the atomic number 28 indicates the nucleus contains 28 protons, and therefore, it must contain 31 neutrons in order to have a mass number of 59. The uranium nucleus has 92 protons as do all uranium nuclei and this particular uranium nucleus has 146 neutrons.

Another way of representing isotopes is by adding a hyphen and the mass number to the chemical name or symbol. Thus the two nuclei would be Nickel-59 or Ni-59 and Uranium-238 or U-238, where 59 and 238 are the mass numbers of the two atoms, respectively. Note that the mass numbers (not the number of neutrons) are given to the side of the name.



Example 1.18.2: POTASSIUM-40

How many protons, electrons, and neutrons are in an atom of ${}^{40}_{19}\mathrm{K}$?

Solution

$$\text{atomic number} = (\text{number of protons}) = 19 \tag{1.18.3}$$

For all atoms with no charge, the number of electrons is equal to the number of protons.

$$\text{number of electrons} = 19 \tag{1.18.4}$$

The mass number, 40 is the sum of the protons and the neutrons.

To find the number of neutrons, subtract the number of protons from the mass number.

$$\text{number of neutrons} = 40 - 19 = 21. \tag{1.18.5}$$

Example 1.18.3: Zinc-65

How many protons, electrons, and neutrons are in an atom of zinc-65?

Solution

$$\text{number of protons} = 30 \tag{1.18.6}$$

For all atoms with no charge, the number of electrons is equal to the number of protons.

$$\text{number of electrons} = 30 \tag{1.18.7}$$

The mass number, 65 is the sum of the protons and the neutrons.

To find the number of neutrons, subtract the number of protons from the mass number.

$$\text{number of neutrons} = 65 - 30 = 35 \tag{1.18.8}$$

Exercise 1.18.3

How many protons, electrons, and neutrons are in each atom?

a. ${}^{60}_{27}\mathrm{Co}$
b. Na-24
c. ${}^{45}_{20}\mathrm{Ca}$
d. Sr-90

Answer a:

27 protons, 27 electrons, 33 neutrons

Answer b:

11 protons, 11 electrons, 13 neutrons

Answer c:

20 protons, 20 electrons, 25 neutrons

Answer d:

38 protons, 38 electrons, 52 neutrons

Summary

- The number of protons is always the same in atoms of the same element.
- The number of neutrons can be different, even in atoms of the same element.
- Atoms of the same element, containing the same number of protons, but different numbers of neutrons, are known as **isotopes**.
- Isotopes of any given element all contain the same number of protons, so they have the same atomic number (for example, the atomic number of helium is always 2).
- Isotopes of a given element contain different numbers of neutrons, therefore, different isotopes have different mass numbers.

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1.19: Atomic Mass- The Average Mass of an Element’s Atoms

Learning Objectives

- Explain what is meant by the atomic mass of an element.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

In chemistry we very rarely deal with only one isotope of an element. We use a mixture of the isotopes of an element in chemical reactions and other aspects of chemistry, because all of the isotopes of an element react in the same manner. That means that we rarely need to worry about the mass of a specific isotope, but instead we need to know the average mass of the atoms of an element. Using the masses of the different isotopes and how abundant each isotope is, we can find the average mass of the atoms of an element. The **atomic mass** of an element is the weighted average mass of the atoms in a naturally occurring sample of the element. Atomic mass is typically reported in atomic mass units.

Calculating Atomic Mass

You can calculate the atomic mass (or average mass) of an element provided you know the **relative abundance** (the fraction of an element that is a given isotope), the element's naturally occurring isotopes, and the masses of those different isotopes. We can calculate this by the following equation:

$$\text{Atomic mass} = (\%_1)(\text{mass}_1) + (\%_2)(\text{mass}_2) + \cdots$$

Look carefully to see how this equation is used in the following examples.

Example 1.19.1: Boron Isotopes

Boron has two naturally occurring isotopes. In a sample of boron, 20% of the atoms are B-10, which is an isotope of boron with 5 neutrons and mass of 10 amu. The other 80% of the atoms are B-11, which is an isotope of boron with 6 neutrons and a mass of 11 amu. What is the atomic mass of boron?

Solution

Boron has two isotopes. We will use the equation:

$$\text{Atomic mass} = (\%_1)(\text{mass}_1) + (\%_2)(\text{mass}_2) + \cdots$$

- Isotope 1: $\%_1 = 0.20$ (Write all percentages as decimals), $\text{mass}_1 = 10$
- Isotope 2: $\%_2 = 0.80$, $\text{mass}_2 = 11$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.20)(10) + (0.80)(11)$$

$$\text{Atomic mass} = 10.8 \text{ amu}$$

The mass of an average boron atom, and thus boron's atomic mass, is 10.8 amu

Example 1.19.2: Neon Isotopes

Neon has three naturally occurring isotopes. In a sample of neon, 90.92% of the atoms are Ne-20, which is an isotope of neon with 10 neutrons and a mass of 19.99 amu Another 0.3% of the atoms are Ne-21, which is an isotope of neon with 11 neutrons and a mass of 20.99 amu The final 8.85% of the atoms are Ne-22, which is an isotope of neon with 12 neutrons and a mass of 21.99 amu What is the atomic mass of neon?

Solution

Neon has three isotopes. We will use the equation:

$$\text{Atomic mass} = (\%_1)(\text{mass}_1) + (\%_2)(\text{mass}_2) + \cdots$$

- Isotope 1: $\%_1 = 0.9092$ (write all percentages as decimals), $\text{mass}_1 = 19.99$
- Isotope 2: $\%_2 = 0.003$, $\text{mass}_2 = 20.99$
- Isotope 3: $\%_3 = 0.0885$, $\text{mass}_3 = 21.99$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.9092)(19.99) + (0.003)(20.99) + (0.0885)(21.99)$$

$$\text{Atomic mass} = 20.17 \text{ amu}$$

The mass of an average neon atom is $\mathbf{20.17}$ amu

The periodic table gives the atomic mass of each element. The atomic mass is a number that usually appears below the element's symbol in each square. Notice that the atomic mass of boron (symbol B) is 10.8, which is what we calculated in Example $1.19.1$, and the atomic mass of neon (symbol Ne) is 20.8, which is what we calculated in Example $1.19.2$. Take time to notice that not all periodic tables have the atomic number above the element's symbol and the mass number below it. If you are ever confused, remember that the atomic number should always be the smaller of the two and will be a whole number, while the atomic mass should always be the larger of the two and will be a decimal number.

? Exercise 1.19.1

Chlorine has two naturally occurring isotopes. In a sample of chlorine, 75.77% of the atoms are Cl-35, with a mass of 34.97 amu Another 24.23% of the atoms are Cl-37, with a mass of 36.97 amu What is the atomic mass of chlorine?

Answer

35.45 amu

Summary

- An element's atomic mass is the weighted average of the masses of the isotopes of an element
- An element's atomic mass can be calculated provided the relative abundance of the element's naturally occurring isotopes and the masses of those isotopes are known.
- The periodic table is a convenient way to summarize information about the different elements. In addition to the element's symbol, most periodic tables will also contain the element's atomic number and the element's atomic mass.

CHAPTER OVERVIEW

2: Radiation- Pros and Cons

2.1: What is Radiation

When you spend time outside in sunlight, you may wear a hat or long sleeves or sunscreen. Why do you take those precautions? You may know that overexposure to sunlight can be harmful to your health, but did you also know that exposure to sunlight is important for your health? Sunlight consists of different types of radiation including ultraviolet (UV) radiation that activates a precursor in your skin that is converted, through a metabolic pathway in your body, to the active form of Vitamin D. Vitamin D helps your body absorb calcium and is needed for strong bones. Overexposure to this UV radiation, however, can lead to skin cancer.

Radiation is the emission of energy from some source through space as particles or electromagnetic waves. UV radiation is a form of electromagnetic radiation emitted by the sun and the energy of the sun is produced by a nuclear reaction known as fusion. We will discuss both types of radiation as we explore nuclear chemistry and continue to develop the model of the atom.

2.2: The Discovery of Radioactivity

Learning Objectives

- List the most common emissions from naturally radioactive nuclei.
- Compare the energy released per gram of matter in nuclear reactions to that in chemical reactions.
- Express the relationship between nuclear stability and the nuclei's binding energy per nucleon ratio.

No one could have known in the 1800's that the discovery of the fascinating science and art form of photography would eventually lead to the splitting of the atom. The basis of photography is the fact that visible light causes certain chemical reactions. If the chemicals are spread thinly on a surface but protected from light by a covering, no reaction occurs. When the covering is removed, however, light acting on the chemicals causes them to darken. With millions of cameras in use today we do not think of it as a strange phenomenon, but at the time of its discovery photography was a strange and wonderful thing.

Even stranger was the discovery by Roentgen, that radiation other than visible light could expose photographic film. He found that film wrapped in dark paper would react when x-rays went through the paper and struck the film.

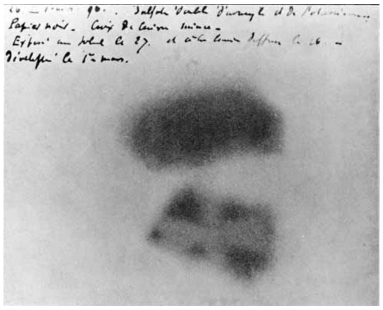

Figure 2.2.1: Image of Becquerel's photographic plate which has been fogged by exposure to radiation from a uranium salt. The shadow of a metal Maltese Cross placed between the plate and the uranium salt is clearly visible (Public Domain).

When Becquerel heard about Roentgen's discovery, he wondered if his fluorescent minerals would give the same x-rays. Becquerel placed some of his rock crystals on top of a well-covered photographic plate and sat them in the sunlight. The sunlight made the crystals glow with a bright fluorescent light, but when Becquerel developed the film he was very disappointed. He found that only one of his minerals, a uranium salt, had fogged the photographic plate. He decided to try again, and this time, to leave them out in the sun for a longer period of time. Fortunately, the weather didn't cooperate and Becquerel had to leave the crystals and film stored in a drawer for several cloudy days. Before continuing his experiments, Becquerel decided to check one of the photographic plates to make sure the chemicals were still good. To his amazement, he found that the plate had been exposed in spots where it had been near the uranium containing rocks and some of these rocks had not been exposed to sunlight at all. In later experiments, Becquerel confirmed that the radiation from the uranium had no connection with light or fluorescence, but the amount of radiation was directly proportional to the concentration of uranium in the rock. Becquerel had discovered **radioactivity**.

The Curies and Radium

One of Becquerel's assistants, a young Polish scientist named Maria Sklowdowska (to become Marie Curie after she married Pierre Curie), became interested in the phenomenon of radioactivity. With her husband, she decided tofind out if chemicals other than uranium were radioactive. The Austrian government was happy to send the Curies a ton of pitchblende from the mining region of Joachimstahl because it was waste material that had to be disposed of anyway. The Curies wanted the pitchblende because it was the residue of uranium mining. From the ton of pitchblende, the Curies separated $0.10\ g$ of a previously unknown element, radium, in the form of the compound radium chloride. This radium was many times more radioactive than uranium.

By 1902, the world was aware of a new phenomenon called radioactivity and of new elements which exhibited natural radioactivity. For this work, Becquerel and the Curies shared the 1903 Nobel Prize and for subsequent work; Marie Cure received a second Nobel Prize in 1911. She is the only person ever to receive two Nobel Prizes in science.

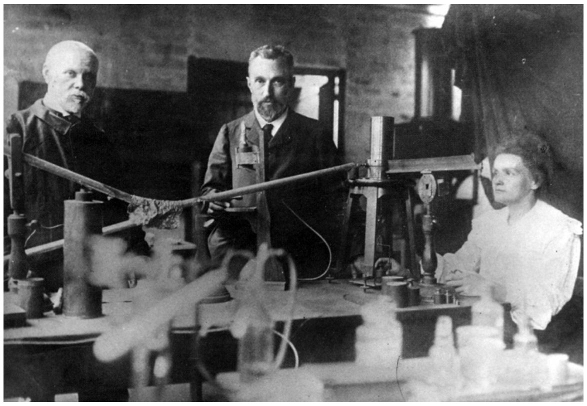

Figure 2.2.2: Marie (left) and Pierre Curie (middle) with Henri Becquerel (right) shared the 1903 Nobel Prize.

Further experiments provided information about the characteristics of the penetrating emissions from radioactive substances. It was soon discovered that there were three common types of radioactive emissions. Some of the radiation could pass easily through aluminum foil while some of the radiation was stopped by the foil. Some of the radiation could even pass through foil up to a centimeter thick. The three basic types of radiation were named alpha, beta, and gamma radiation. The actual composition of the three types of radiation was still not known.

Eventually, scientists were able to demonstrate experimentally that the alpha particle, α, was a helium nucleus (a particle containing two protons and two neutrons), a beta particle, β, was a high speed electron, and gamma rays, γ, were a very high energy form of light (even higher energy than x-rays).

Summary

- Henri Becquerel, Marie Curie, and Pierre Curie shared the discovery of radioactivity.

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2.3: Nuclear Equations

Learning Objectives

- Identify common particles and energies involved in nuclear reactions
- Write and balance nuclear equations

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.

Types of Particles in Nuclear Reactions

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 2.3.1. Protons ($^{1}_{1}\mathrm{p}$, also represented by the symbol $^{1}_{1}\mathrm{H}$) and neutrons ($^{1}_{0}\mathrm{n}$) are the constituents of atomic nuclei, and have been described previously. Alpha particles ($^{4}_{2}\mathrm{He}$, also represented by the symbol $^{4}_{2}\alpha$) are high-energy helium nuclei. Beta particles ($^{0}_{-1}\beta$, also represented by the symbol $^{0}_{-1}\mathrm{e}$) are high-energy electrons, and gamma rays are photons of very high-energy electromagnetic radiation. Positrons ($^{0}_{+1}\mathrm{e}$, also represented by the symbol $^{0}_{+1}\beta$) 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 $^{4}_{2}\mathrm{He}$. 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 2.3.1: 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 positrons 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 $E = mc^2$. For example, when a positron and an electron collide, both are annihilated and two gamma ray photons are created:

$$^{0}_{-1}\mathrm{e} + {}^{0}_{+1}\mathrm{e} \rightarrow \gamma + \gamma \tag{2.3.1}$$

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 2.3.1 shows how we can identify a nuclide by balancing the nuclear reaction.

Example 2.3.1: 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:

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

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:

$$25+4=\mathrm{A}+1$$

so

$$\mathrm{A}=28$$

Similarly, the charges must balance, so:

$$12+2=\mathrm{Z}+1$$

so

$$\mathrm{Z}=13$$

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

Exercise 2.3.1

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:

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

- 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:

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

- 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 $^{4}\mathrm{He}$:

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

- 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, ${}^{2}_{1}\text{H}$), by Emilio Segre and Carlo Perrier in 1937:

$${}^{2}_{1}\text{H} + {}^{97}_{42}\text{Mo} \longrightarrow 2\,{}^{1}_{0}\text{n} + {}^{97}_{43}\text{Tc}$$

- 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:

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

Summary

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.

Glossary

alpha particle

(α or ${}^{4}_{2}\text{He}$ or ${}^{4}_{2}\alpha$) 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

beta particle

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

gamma ray

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

nuclear reaction

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

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

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

Contributors and Attributions

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2.4: Biological Effects of Radiation

Learning Objectives

- Describe the biological impact of ionizing radiation.
- Recognize units for measuring radiation exposure.
- Explain the operation of common tools for detecting radioactivity.
- List common sources of radiation exposure in the US.

The increased use of radioisotopes has led to increased concerns over the effects of these materials on biological systems (such as humans). All radioactive nuclides emit high-energy particles or electromagnetic waves. When this radiation encounters living cells, it can cause heating, break chemical bonds, or ionize molecules. The most serious biological damage results when these radioactive emissions fragment or ionize molecules. For example, alpha and beta particles emitted from nuclear decay reactions possess much higher energies than ordinary chemical bond energies. When these particles strike and penetrate matter, they produce ions and molecular fragments that are extremely reactive. The damage this does to biomolecules in living organisms can cause serious malfunctions in normal cell processes, taxing the organism's repair mechanisms and possibly causing illness or even death (Figure $2.4.1$).

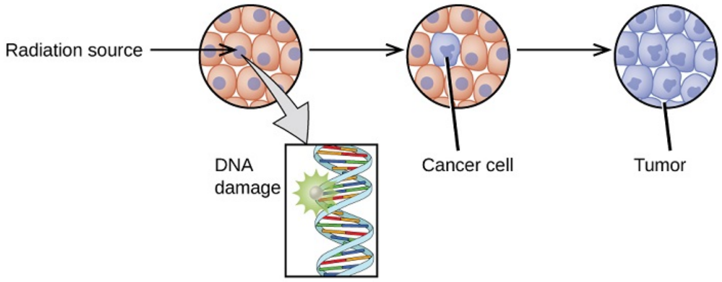

Figure $2.4.1$: Radiation can harm biological systems by damaging the DNA of cells. If this damage is not properly repaired, the cells may divide in an uncontrolled manner and cause cancer.

Ionizing vs. Nonionizing Radiation

There is a large difference in the magnitude of the biological effects of nonionizing radiation (for example, light and microwaves) and ionizing radiation, emissions energetic enough to knock electrons out of molecules (for example, α and β particles, γ rays, X-rays, and high-energy ultraviolet radiation) (Figure $2.4.2$).

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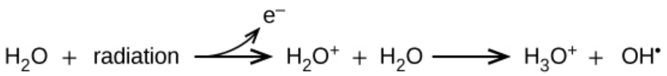
Figure $2.4.2$: Lower frequency, lower-energy electromagnetic radiation is nonionizing, and higher frequency, higher-energy electromagnetic radiation is ionizing.

Energy absorbed from nonionizing radiation speeds up the movement of atoms and molecules, which is equivalent to heating the sample. Although biological systems are sensitive to heat (as we might know from touching a hot stove or spending a day at the beach in the sun), a large amount of nonionizing radiation is necessary before dangerous levels are reached. Ionizing radiation, however, may cause much more severe damage by breaking bonds or removing electrons in biological molecules, disrupting their structure and function. The damage can also be done indirectly, by first ionizing H_2O (the most abundant molecule in living organisms), which forms a H_2O^+ ion that reacts with water, forming a hydronium ion and a hydroxyl radical:



$$\mathrm{H_2O} + \text{radiation} \xrightarrow{\mathrm{e^-}} \mathrm{H_2O^+} + \mathrm{H_2O} \longrightarrow \mathrm{H_3O^+} + \mathrm{OH^{\bullet}}$$

Figure $2.4.3$.

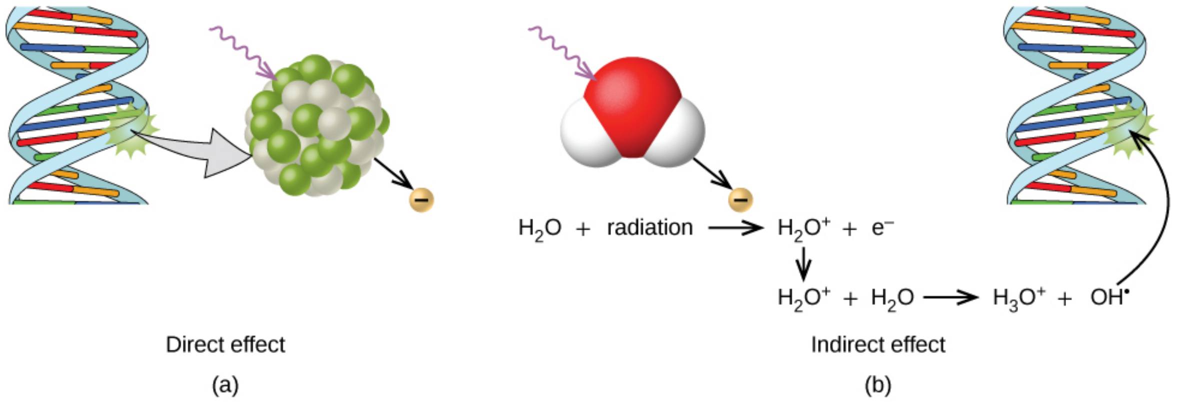


Figure 2.4.3: Ionizing radiation can (a) directly damage a biomolecule by ionizing it or breaking its bonds, or (b) create an H_2O^+ ion, which reacts with H_2O to form a hydroxyl radical, which in turn reacts with the biomolecule, causing damage indirectly.

Biological Effects of Exposure to Radiation

Radiation can harm either the whole body (somatic damage) or eggs and sperm (genetic damage). Its effects are more pronounced in cells that reproduce rapidly, such as the stomach lining, hair follicles, bone marrow, and embryos. This is why patients undergoing radiation therapy often feel nauseous or sick to their stomach, lose hair, have bone aches, and so on, and why particular care must be taken when undergoing radiation therapy during pregnancy.

Different types of radiation have differing abilities to pass through material (Figure 2.4.4). A very thin barrier, such as a sheet or two of paper, or the top layer of skin cells, usually stops alpha particles. Because of this, alpha particle sources are usually not dangerous if outside the body, but are quite hazardous if ingested or inhaled (see the Chemistry in Everyday Life feature on Radon Exposure). Beta particles will pass through a hand, or a thin layer of material like paper or wood, but are stopped by a thin layer of metal. Gamma radiation is very penetrating and can pass through a thick layer of most materials. Some high-energy gamma radiation is able to pass through a few feet of concrete. Certain dense, high atomic number elements (such as lead) can effectively attenuate gamma radiation with thinner material and are used for shielding. The ability of various kinds of emissions to cause ionization varies greatly, and some particles have almost no tendency to produce ionization. Alpha particles have about twice the ionizing power of fast-moving neutrons, about 10 times that of β particles, and about 20 times that of γ rays and X-rays.

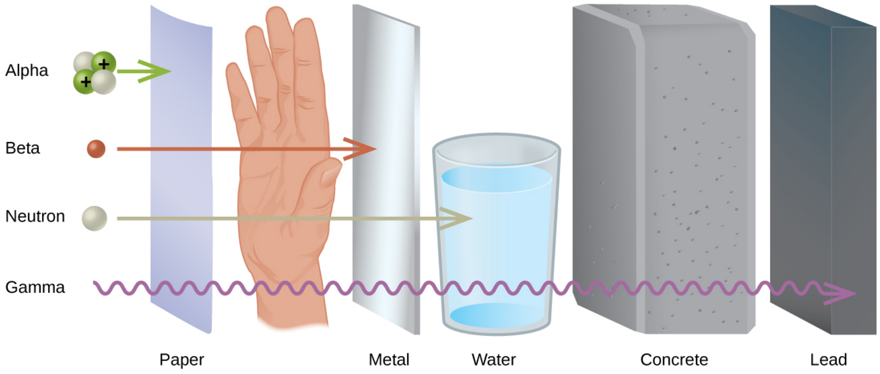


Figure 2.4.4: The ability of different types of radiation to pass through material is shown. From least to most penetrating, they are alpha < beta < neutron < gamma.

For many people, one of the largest sources of exposure to radiation is from radon gas (Rn-222). Radon-222 is an α emitter with a half–life of 3.82 days. It is one of the products of the radioactive decay series of U-238, which is found in trace amounts in soil and rocks. The radon gas that is produced slowly escapes from the ground and gradually seeps into homes and other structures above. Since it is about eight times more dense than air, radon gas accumulates in basements and lower floors, and slowly diffuses throughout buildings (Figure 2.4.5).

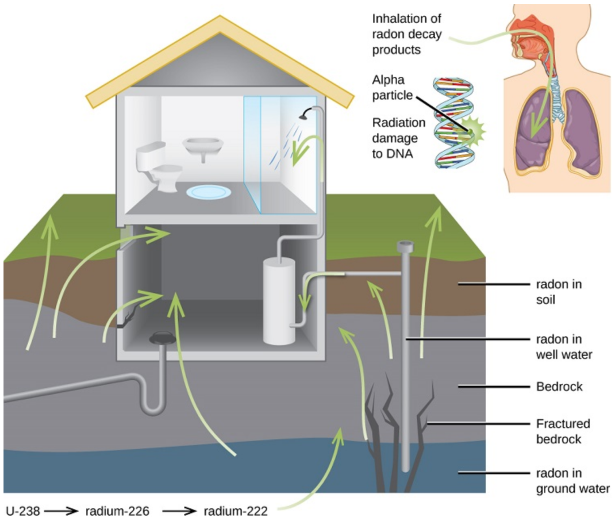


Figure 2.4.5: Radon-222 seeps into houses and other buildings from rocks that contain uranium-238, a radon emitter. The radon enters through cracks in concrete foundations and basement floors, stone or porous cinderblock foundations, and openings for water and gas pipes.

Radon is found in buildings across the country, with amounts dependent on location. The average concentration of radon inside houses in the US (1.25 pCi/L) is about three times the level found in outside air, and about one in six houses have radon levels high enough that remediation efforts to reduce the radon concentration are recommended. Exposure to radon increases one's risk of getting cancer (especially lung cancer), and high radon levels can be as bad for health as smoking a carton of cigarettes a day. Radon is the number one cause of lung cancer in nonsmokers and the second leading cause of lung cancer overall. Radon exposure is believed to cause over 20,000 deaths in the US per year.

Measuring Radiation Exposure

Several different devices are used to detect and measure radiation, including Geiger counters, scintillation counters (scintillators), and radiation dosimeters (Figure 2.4.6). Probably the best-known radiation instrument, the Geiger counter (also called the Geiger-Müller counter) detects and measures radiation. Radiation causes the ionization of the gas in a Geiger-Müller tube. The rate of ionization is proportional to the amount of radiation. A scintillation counter contains a scintillator—a material that emits light (luminesces) when excited by ionizing radiation—and a sensor that converts the light into an electric signal. Radiation dosimeters also measure ionizing radiation and are often used to determine personal radiation exposure. Commonly used types are electronic, film badge, thermoluminescent, and quartz fiber dosimeters.

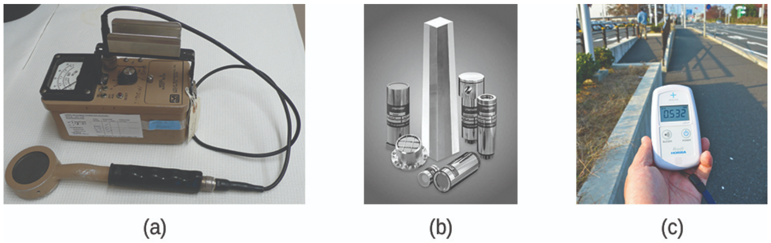


Figure 2.4.6: Devices such as (a) Geiger counters, (b) scintillators, and (c) dosimeters can be used to measure radiation. (Credit c: modification of work by "osaMu"/Wikimedia commons.)

A variety of units are used to measure various aspects of radiation (Table 2.4.1). The SI unit for rate of radioactive decay is the becquerel (Bq), with 1 Bq = 1 disintegration per second. The curie (Ci) and millicurie (mCi) are much larger units and are frequently used in medicine (1 curie = 1 Ci = 3.7×10^{10} disintegrations per second). The SI unit for measuring radiation dose is the gray (Gy), with 1 Gy = 1 J of energy absorbed per kilogram of tissue. In medical applications, the radiation absorbed dose (rad) is more often used (1 rad = 0.01 Gy; 1 rad results in the absorption of 0.01 J/kg of tissue). The SI unit measuring tissue damage caused by radiation is the sievert (Sv). This takes into account both the energy and the biological effects of the type of radiation involved in the radiation dose.

Table 2.4.1: Units Used for Measuring Radiation

Measurement Purpose	Unit	Quantity Measured	Description

Measurement Purpose	Unit	Quantity Measured	Description
activity of source	becquerel (Bq)	radioactive decays or emissions	amount of sample that undergoes 1 decay/second
	curie (Ci)		amount of sample that undergoes 3.7×10^{10} decays/second
absorbed dose	gray (Gy)	energy absorbed per kg of tissue	1 Gy = 1 J/kg tissue
	radiation absorbed dose (rad)		1 rad = 0.01 J/kg tissue
biologically effective dose	sievert (Sv)	tissue damage	Sv = RBE × Gy
	roentgen equivalent for man (rem)		Rem = RBE × rad

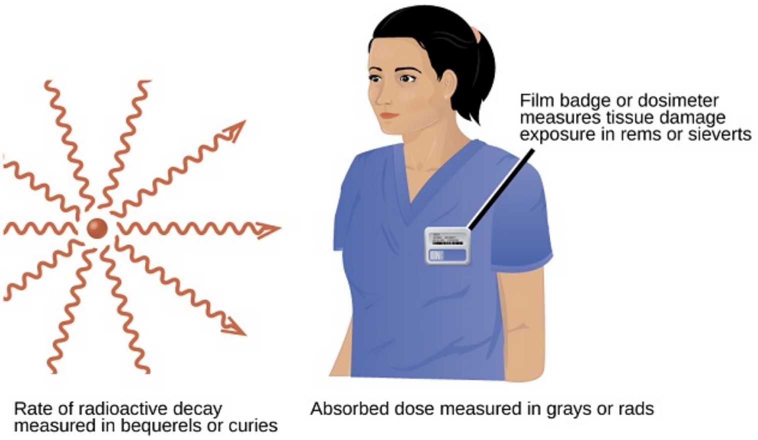


Figure 2.4.7: Different units are used to measure the rate of emission from a radioactive source, the energy that is absorbed from the source, and the amount of damage the absorbed radiation does. (CC BY 4.0; OpenStax)

Effects of Long-term Radiation Exposure on the Human Body

The effects of radiation depend on the type, energy, and location of the radiation source, and the length of exposure. As shown in Figure 2.4.8, the average person is exposed to background radiation, including cosmic rays from the sun and radon from uranium in the ground (see the Chemistry in Everyday Life feature on Radon Exposure); radiation from medical exposure, including CAT scans, radioisotope tests, X-rays, and so on; and small amounts of radiation from other human activities, such as airplane flights (which are bombarded by increased numbers of cosmic rays in the upper atmosphere), radioactivity from consumer products, and a variety of radionuclides that enter our bodies when we breathe (for example, carbon-14) or through the food chain (for example, potassium-40, strontium-90, and iodine-131).

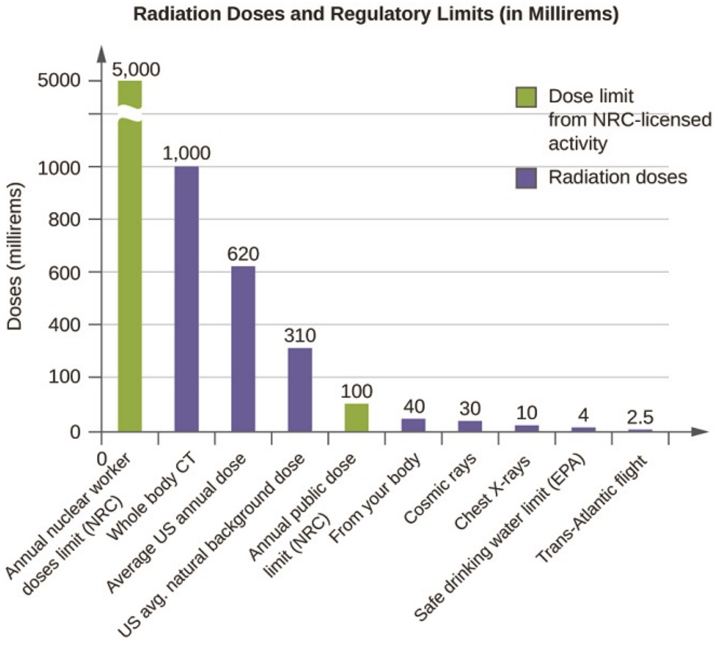


Figure 2.4.8: The total annual radiation exposure for a person in the US is about 620 mrem. The various sources and their relative amounts are shown in this bar graph. (source: U.S. Nuclear Regulatory Commission).

A short-term, sudden dose of a large amount of radiation can cause a wide range of health effects, from changes in blood chemistry to death. Short-term exposure to tens of rems of radiation will likely cause very noticeable symptoms or illness; a dose of about 500 rems is estimated to have a 50% probability of causing the death of the victim within 30 days of exposure. Exposure to radioactive emissions has a cumulative effect on the body during a person's lifetime, which is another reason why it is important to avoid any unnecessary exposure to radiation. Health effects of short-term exposure to radiation are shown in Table 2.4.2.

Table 2.4.2: Health Effects of Radiation

Exposure (rem)	Health Effect	Time to Onset (Without Treatment)
5–10	changes in blood chemistry	—
50	nausea	hours
55	fatigue	—
70	vomiting	—
75	hair loss	2–3 weeks
90	diarrhea	—
100	hemorrhage	—
400	possible death	within 2 months
1000	destruction of intestinal lining	—
	internal bleeding	—
	death	1–2 weeks
2000	damage to central nervous system	—
	loss of consciousness	minutes
	death	hours to days

It is impossible to avoid some exposure to ionizing radiation. We are constantly exposed to background radiation from a variety of natural sources, including cosmic radiation, rocks, medical procedures, consumer products, and even our own atoms. We can minimize our exposure by blocking or shielding the radiation, moving farther from the source, and limiting the time of exposure.

Summary

We are constantly exposed to radiation from a variety of naturally occurring and human-produced sources. This radiation can affect living organisms. Ionizing radiation is the most harmful because it can ionize molecules or break chemical bonds, which damages the molecule and causes malfunctions in cell processes. It can also create reactive hydroxyl radicals that damage biological molecules and disrupt physiological processes. Radiation can cause somatic or genetic damage, and is most harmful to rapidly reproducing cells. Types of radiation differ in their ability to penetrate material and damage tissue, with alpha particles the least penetrating, but potentially most damaging, and gamma rays the most penetrating.

Various devices, including Geiger counters, scintillators, and dosimeters, are used to detect and measure radiation, and monitor radiation exposure. We use several units to measure radiation: becquerels or curies for rates of radioactive decay; gray or rads for energy absorbed; and rems or sieverts for biological effects of radiation. Exposure to radiation can cause a wide range of health effects, from minor to severe, including death. We can minimize the effects of radiation by shielding with dense materials such as lead, moving away from the source of radiation, and limiting time of exposure.

Footnotes

1. 1 Source: US Environmental Protection Agency

Glossary

becquerel (Bq)

SI unit for rate of radioactive decay; 1 Bq = 1 disintegration/s.

curie (Ci)

Larger unit for rate of radioactive decay frequently used in medicine; 1 Ci = 3.7×10^{10} disintegrations/s.

Geiger counter

Instrument that detects and measures radiation via the ionization produced in a Geiger-Müller tube.

gray (Gy)

SI unit for measuring radiation dose; 1 Gy = 1 J absorbed/kg tissue.

ionizing radiation

Radiation that can cause a molecule to lose an electron and form an ion.

millicurie (mCi)

Larger unit for rate of radioactive decay frequently used in medicine; 1 Ci = 3.7×10^{10} disintegrations/s.

nonionizing radiation

Radiation that speeds up the movement of atoms and molecules; it is equivalent to heating a sample, but is not energetic enough to cause the ionization of molecules.

radiation absorbed dose (rad)

SI unit for measuring radiation dose, frequently used in medical applications; 1 rad = 0.01 Gy.

radiation dosimeter

Device that measures ionizing radiation and is used to determine personal radiation exposure.

relative biological effectiveness (RBE)

Measure of the relative damage done by radiation.

roentgen equivalent man (rem)

Unit for radiation damage, frequently used in medicine; 1 rem = 1 Sv.

scintillation counter

Instrument that uses a scintillator—a material that emits light when excited by ionizing radiation—to detect and measure radiation.

sievert (Sv)

SI unit measuring tissue damage caused by radiation; takes energy and biological effects of radiation into account.

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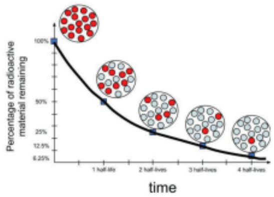
2.5: Natural Radioactivity and Half-Life

Learning Objectives

- Describe half-life and what factors affect half-life.
- Calculate the amount of radioactive material that will remain after an integral number of half-lives.
- Find the half-life of an isotope given graphical or other data.
-

Rate of Radioactive Decay

During natural radioactive decay, not all atoms of an element are instantaneously changed to atoms of another element. The decay process takes time and there is value in being able to express the rate at which a process occurs. A useful concept is **half-life**, which is the time required for half of the starting material to change or decay. Half-lives can be calculated from measurements on the change in mass of a nuclide and the time it takes to occur. The only thing we know is that in the time of that substance's half-life, half of the original nuclei will disintegrate. Although the rate of chemical reactions are impacted by factors such as temperature, concentration, etc, these factors have no effect on half-life. Each radioactive isotope will have its own unique half-life that is independent of any of these factors.



After one half-life, half of the radioactive atoms have decayed. After another half-life, half of the remaining atoms have changed. This pattern continues, with half of the atoms changing over each half-life time.
CC Tracy Poulsen

The half-lives of many radioactive isotopes have been determined and they have been found to range from extremely long half-lives of 10 billion years to extremely short half-lives of fractions of a second.

Table 17.5.1: Table of Selected Half-lives

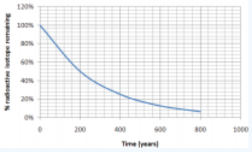
Element	Mass Number (A)	Half-life	Element	Mass Number (A)	Half Life
Uranium	238	4.5 Billion years	Californium	251	800 years
Neptunium	240	1 hour	Nobelium	254	3 seconds
Plutonium	243	5 hours	Carbon	14	5730 years
Americium	245	25 minutes	Carbon	16	740 milliseconds

The quantity of radioactive nuclei at any given time will decrease to half as much in one half-life. For example, if there were 100 g of Cf-251 in a sample at some time, after 800 years, there would be 50 g of Cf-251 remaining and after another 800 years (1600 years total), there would only be 25 g remaining.

Remember, the half-life is the time it takes for half of your sample, no matter how much you have, to remain. Each half-life will follow the same general pattern as Cf-251. The only difference is the length of time it takes for half of a sample to decay.

Example $2.5.3$

Using the graph, what is the half-life of an isotope that produces the following graph of decay over time:



Solution

We know that the half-life is the time it takes for half of a sample to change. How long did it take for half of our isotope to change? It took approximately 200 years for 100% of our sample to leave only 50% (half of the original amount) remaining. **The half-life is 200 years.**

*Notice that after another 200 years (400 years total), 25% remains (half of 50%)

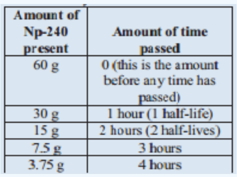
Look carefully at the graph in the previous example. All types of radioactive decay make a graph of the same general shape. The only difference is the scale and units of the x-axis, as the half-life time will be different.

Example 2.5.2

If there are 60 grams of Np-240 present, how much Np-240 will remain after 4 hours? (Np-240 has a half-life of 1 hour)

Solution

Np-240 with a half life of only 1 hour.



Amount of Np-240 present	Amount of time passed
60 g	0 (this is the amount before any time has passed)
30 g	1 hour (1 half-life)
15 g	2 hours (2 half-lives)
7.5 g	3 hours
3.75 g	4 hours

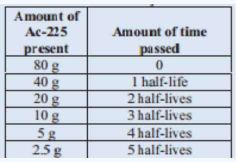
After 4 hours, only 3.75 g of our original 60 g sample would remain the radioactive isotope Np-240.

Example 2.5.3

A sample of Ac-225 originally contained 80 grams and after 50 days only 2.55 grams of the original Ac-225 remain. What is the half life of Ac-225?

Solution

We are going to tackle this problem similar to the last problem. The difference is that we are looking for the half-life time. Let's set up a similar table, though:



Amount of Ac-225 present	Amount of time passed
80 g	0
40 g	1 half-life
20 g	2 half-lives
10 g	3 half-lives
5 g	4 half-lives
2.5 g	5 half-lives

We know that 50 days is the same as 5 half-lives. Therefore, 1 half-life is 10 days. The half-life of Ac-225 is 10 days.

Radioactive Dating

An ingenious application of half-life studies established a new science of determining ages of materials by half-life calculations. For geological dating, the decay of U-238 can be used. The half-life of U-238 is 4.5×10^9 years. The end product of the decay of U-238 is Pb-206. After one half-life, a 1.00 gram sample of uranium will have decayed to 0.50 grams of U-238 and 0.43 grams of Pb-206. By comparing the amount of U-238 to the amount of Pb-206 in a sample of uranium mineral, the age of the mineral can be estimated. Present day estimates for the age of the Earth's crust from this method is at 4 billion years.

Organic material (material made from things that were once living, such as paper and fabric) is radioactively dated using the long-lived nuclide of carbon, carbon-14. This method of determining the age of organic material (or once living materials) was given the name radiocarbon dating. The carbon dioxide consumed by living systems contains a certain concentration of $^{14}CO_2$. When an organism dies, the acquisition of carbon-14 stops but the decay of the C-14 in the body continues. As time goes by, the ratio of C-14 to C-12 decreases at a rate determined by the half-life of C-14. Using half-life equations, the time since the organism died can be calculated. These procedures have been used to determine the age of organic artifacts and determine, for instance, whether art works are real or fake.

Summary and Vocabulary

The half-life of an isotope is used to describe the rate at which the isotope will decay and give off radiation. Using the half-life, it is possible to predict the amount of radioactive material that will remain after a given amount of time. C-14 dating procedures have been used to determine the age of organic artifacts. Its half-life is approximately 5700 years.

- Background radiation: Radiation that comes from environment sources including the earth's crust, the atmosphere, cosmic rays, and radioisotopes. These natural sources of radiation account for the largest amount of radiation received by most people.
- Half-life: The half-life of a radioactive substance is the time interval required for a quantity of material to decay to half its original value.

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2.6: Fission and Fusion

Learning Objectives

- Define and give examples of fission and fusion.
- Classify nuclear reactions as fission or fusion.

Nuclei that are larger than iron-56 may undergo nuclear reactions in which they break up into two or more smaller nuclei. These reactions are called **fission** reactions. Conversely, nuclei that are smaller than iron-56 become larger nuclei in order to be more stable. These nuclei undergo a nuclear reaction in which smaller nuclei join together to form a larger nucleus. Such nuclear reactions are called **fusion** reactions.

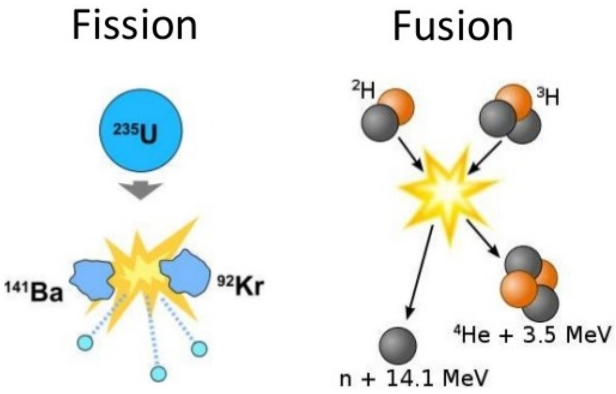

Figure 2.6.1: (left) Nuclear fission occurs with one large nuclear is split into two or more smaller nuclei. (right) Nuclear fusion happens when two small nuclei combine to make a larger nucleus.

Fission and Chain Reactions

In both fission and fusion, large amounts of energy are given off in the form of heat, light, and gamma radiation. Nuclear fission was discovered in the late 1930's when U-235 nuclides were bombarded with neutrons and were observed to split into two smaller-mass nuclei.

$$ {}^{1}_{0}\mathrm{n} + {}^{235}_{92}\mathrm{U} \rightarrow {}^{141}_{56}\mathrm{Ba} + {}^{92}_{36}\mathrm{Kr} + 3{}^{1}_{0}\mathrm{n} \tag{2.6.1} $$

The products shown are only one of many sets of products from the disintegration of a U-235 nucleus. Over 35 different elements have been observed in the fission products of U-235.

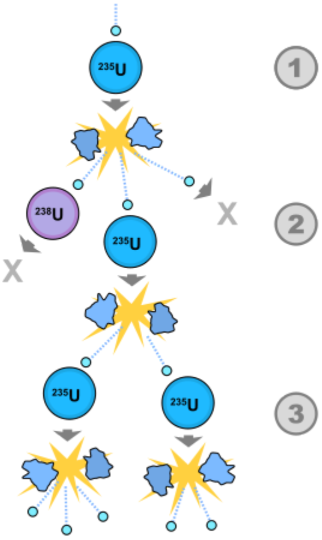

Figure 2.6.2: A possible nuclear fission chain reaction. 1. A uranium-235 atom absorbs a neutron, and fissions into two new atoms (fission fragments), releasing three new neutrons and a large amount of binding energy. 2. One of those neutrons is absorbed by an atom of uranium-238, and does not continue the reaction. Another neutron leaves the system without being absorbed. However, one neutron does collide with an atom of uranium-235, which then fissions and releases two neutrons and more binding energy. 3. Both of those neutrons collide with uranium-235 atoms, each of which fissions and releases a few neutrons, which can then continue the reaction. (Public Domain).

When a neutron strikes a U-235 nucleus and the nucleus captures a neutron, it undergoes fission producing two lighter nuclei and three free neutrons. The production of the free neutrons makes it possible to have a self-sustaining fission process - a nuclear **chain reaction**. It at least one of the neutrons goes on to cause another U-235 disintegration, the fission will be self-sustaining.

How are these reactions are used?

Fusion reactions occur in the sun and make it possible for life on earth. We are currently not able to use fusion on the earth, but we do benefit from it. Fission reactions occur in many places on our earth. At least 20% of the energy you use in your home comes from nuclear energy. This power is made in a nuclear reactor. Nuclear power is a part of our discussion on the current energy crisis. Another topic in fission reactions are nuclear weapons. There are many applications beyond bombs and the generation of electricity. Radioactivity has huge applications in scientific research, several fields of medicine both in terms of imaging and in terms of treatment, industrial processes, some very useful appliances, and even in agriculture.

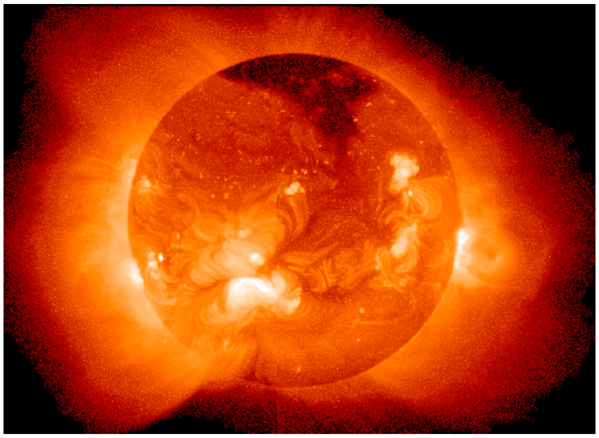

Figure 2.6.3: The energy that comes from the sun and other stars is produced by fusion. (Public Domain; NASA).

Summary and Vocabulary

Naturally radioactive elements exist in the earth and are either alpha or beta emitters. Artificial transmutation of elements can be accomplished by bombarding the nuclei of some elements with alpha or subatomic particles. Nuclear radiation also has many medical uses.

- Chain reaction: A multi-stage nuclear reaction that sustains itself in a series of fissions in which the release of neutrons from the splitting of one atom leads to the splitting of others.
- Critical mass: The smallest mass of a fissionable material that will sustain a nuclear chain reaction at a constant level.
- Fission: A nuclear reaction in which a heavy nucleus splits into two or more smaller fragments, releasing large amounts of energy.
- Fusion: A nuclear reaction in which nuclei combine to form more massive nuclei with the simultaneous release of energy.
- Control rods: Control rods are made of chemical elements capable of absorbing many neutrons and are used to control the rate of a fission chain reaction in a nuclear reactor.

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2.7: Radioactivity in Medicine and Other Applications

Learning Objectives

- List common applications of radioactive isotopes

Radioactive isotopes have the same chemical properties as stable isotopes of the same element, but they emit radiation, which can be detected. If we replace one (or more) atom(s) with radioisotope(s) in a compound, we can track them by monitoring their radioactive emissions. This type of compound is called a radioactive tracer (or radioactive label). Radioisotopes are used to follow the paths of biochemical reactions or to determine how a substance is distributed within an organism. Radioactive tracers are also used in many medical applications, including both diagnosis and treatment. They are used to measure engine wear, analyze the geological formation around oil wells, and much more.

Radioisotopes have revolutionized medical practice, where they are used extensively. Over 10 million nuclear medicine procedures and more than 100 million nuclear medicine tests are performed annually in the United States. Four typical examples of radioactive tracers used in medicine are technetium-99 $\left({}^{99}_{43}\mathrm{Tc}\right)$, thallium-201 $\left({}^{201}_{81}\mathrm{Tl}\right)$, iodine-131 $\left({}^{131}_{53}\mathrm{I}\right)$, and sodium-24 $\left({}^{24}_{11}\mathrm{Na}\right)$. Damaged tissues in the heart, liver, and lungs absorb certain compounds of technetium-99 preferentially. After it is injected, the location of the technetium compound, and hence the damaged tissue, can be determined by detecting the γ rays emitted by the Tc-99 isotope. Thallium-201 (Figure $2.7.1$) becomes concentrated in healthy heart tissue, so the two isotopes, Tc-99 and Tl-201, are used together to study heart tissue. Iodine-131 concentrates in the thyroid gland, the liver, and some parts of the brain. It can therefore be used to monitor goiter and treat thyroid conditions, such as Grave's disease, as well as liver and brain tumors. Salt solutions containing compounds of sodium-24 are injected into the bloodstream to help locate obstructions to the flow of blood.

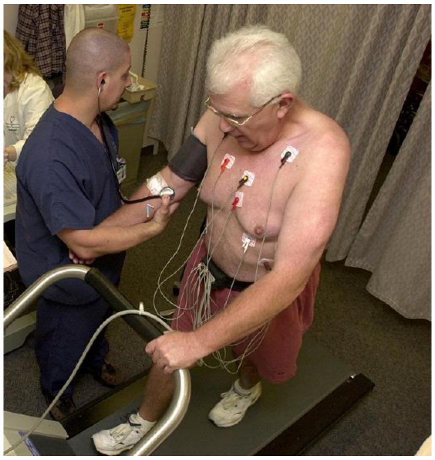

Figure $2.7.1$: Administering thallium-201 to a patient and subsequently performing a stress test offer medical professionals an opportunity to visually analyze heart function and blood flow. (credit: modification of work by "Blue0ctane"/Wikimedia Commons)

A medical professional is assisting a topless elderly man on a treadmill with sensors and electrical wires connected to his torso.

Radioisotopes used in medicine typically have short half-lives—for example, the ubiquitous Tc-99m has a half-life of 6.01 hours. This makes Tc-99m essentially impossible to store and prohibitively expensive to transport, so it is made on-site instead. Hospitals and other medical facilities use Mo-99 (which is primarily extracted from U-235 fission products) to generate Tc-99. Mo-99 undergoes β decay with a half-life of 66 hours, and the Tc-99 is then chemically extracted (Figure $2.7.2$). The parent nuclide Mo-99 is part of a molybdate ion, $\mathrm{MoO_4^{2-}}$; when it decays, it forms the pertechnetate ion, $\mathrm{TcO_4^-}$. These two water-soluble ions are separated by column chromatography, with the higher charge molybdate ion adsorbing onto the alumina in the column, and the lower charge pertechnetate ion passing through the column in the solution. A few micrograms of Mo-99 can produce enough Tc-99 to perform as many as 10,000 tests.

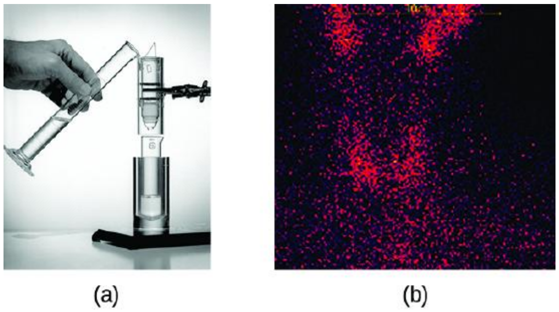

Figure $2.7.2$: (a) The first Tc-99m generator (circa 1958) is used to separate Tc-99 from Mo-99. The MoO_4^{2-} *is retained by the matrix in the column, whereas the* TcO_4^- *passes through and is collected. (b) Tc-99 was used in this scan of the neck of a patient with Grave's disease. The scan shows the location of high concentrations of Tc-99. (credit a: modification of work by the Department of Energy; credit b: modification of work by "MBq"/Wikimedia Commons)*

The first image shows a hand pouring a liquid from a measuring cylinder into a column held up by a clamp. Below the column is a glass tube. The second picture shows red dots on a dark background dispersed everywhere with four spots of heavily concentrated regions.

Radioisotopes can also be used, typically in higher doses than as a tracer, as treatment. Radiation therapy is the use of high-energy radiation to damage the DNA of cancer cells, which kills them or keeps them from dividing (Figure $2.7.3$). A cancer patient may receive external beam radiation therapy delivered by a machine outside the body, or internal radiation therapy (brachytherapy) from a radioactive substance that has been introduced into the body. Note that chemotherapy is similar to internal radiation therapy in that the cancer treatment is injected into the body, but differs in that chemotherapy uses chemical rather than radioactive substances to kill the cancer cells.

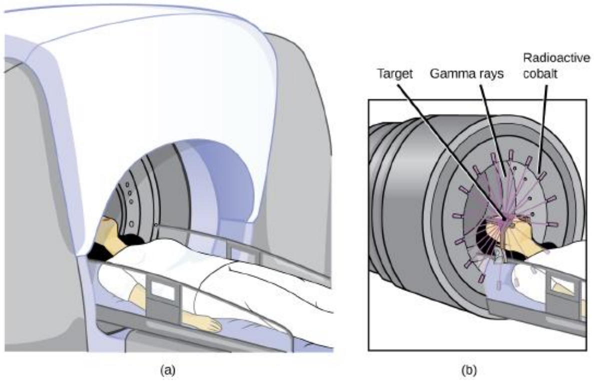

Figure $2.7.3$: The cartoon in (a) shows a cobalt-60 machine used in the treatment of cancer. The diagram in (b) shows how the gantry of the Co-60 machine swings through an arc, focusing radiation on the targeted region (tumor) and minimizing the amount of radiation that passes through nearby regions.

A. A woman lies down as she goes into a dome shaped medical machine. B. A closer view of the women's head shows gamma rays from radioactive cobalt attacks the target on the woman's head.

Cobalt-60 is a synthetic radioisotope produced by the neutron activation of Co-59, which then undergoes β decay to form Ni-60, along with the emission of γ radiation. The overall process is:

$$^{59}_{27}\mathrm{Co} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{60}_{27}\mathrm{Co} \longrightarrow {}^{60}_{28}\mathrm{Ni} + {}^{0}_{-1}\beta + 2\,{}^{0}_{0}\gamma$$

The overall decay scheme for this is shown graphically in Figure $2.7.4$.

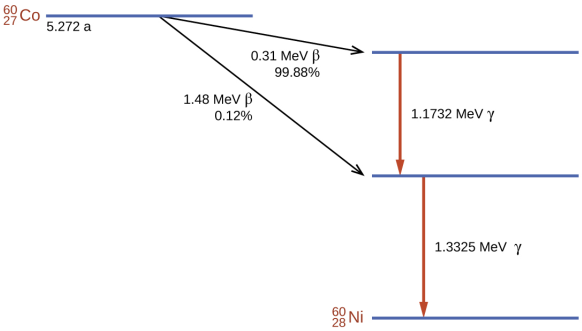

Figure 2.7.4: Co-60 undergoes a series of radioactive decays. The γ emissions are used for radiation therapy.

Radioisotopes are used in diverse ways to study the mechanisms of chemical reactions in plants and animals. These include labeling fertilizers in studies of nutrient uptake by plants and crop growth, investigations of digestive and milk-producing processes in cows, and studies on the growth and metabolism of animals and plants.

For example, the radioisotope C-14 was used to elucidate the details of how photosynthesis occurs. The overall reaction is:

$$6\,CO_2(g) + 6\,H_2O(l) \longrightarrow C_6H_{12}O_6(s) + 6\,O_2(g),$$

but the process is much more complex, proceeding through a series of steps in which various organic compounds are produced. In studies of the pathway of this reaction, plants were exposed to CO_2 containing a high concentration of $^{14}_{6}C$. At regular intervals, the plants were analyzed to determine which organic compounds contained carbon-14 and how much of each compound was present. From the time sequence in which the compounds appeared and the amount of each present at given time intervals, scientists learned more about the pathway of the reaction.

Commercial applications of radioactive materials are equally diverse (Figure 2.7.5). They include determining the thickness of films and thin metal sheets by exploiting the penetration power of various types of radiation. Flaws in metals used for structural purposes can be detected using high-energy gamma rays from cobalt-60 in a fashion similar to the way X-rays are used to examine the human body. In one form of pest control, flies are controlled by sterilizing male flies with γ radiation so that females breeding with them do not produce offspring. Many foods are preserved by radiation that kills microorganisms that cause the foods to spoil.

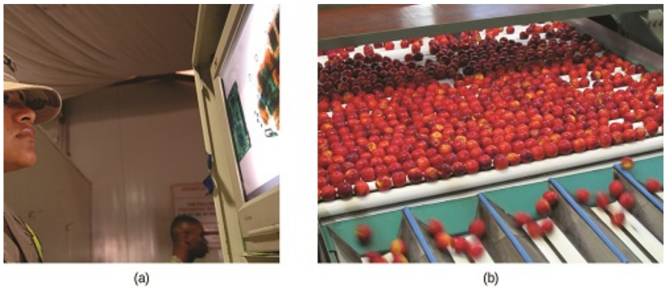

Figure 2.7.5: Common commercial uses of radiation include (a) X-ray examination of luggage at an airport and (b) preservation of food. (credit a: modification of work by the Department of the Navy; credit b: modification of work by the US Department of Agriculture)

A. A man is observing a monitor which shows the X-ray of luggages. B. Many apples on processing belts.

Americium-241, an α emitter with a half-life of 458 years, is used in tiny amounts in ionization-type smoke detectors (Figure 2.7.6). The α emissions from Am-241 ionize the air between two electrode plates in the ionizing chamber. A battery supplies a potential that causes movement of the ions, thus creating a small electric current. When smoke enters the chamber, the movement of the ions is impeded, reducing the conductivity of the air. This causes a marked drop in the current, triggering an alarm.

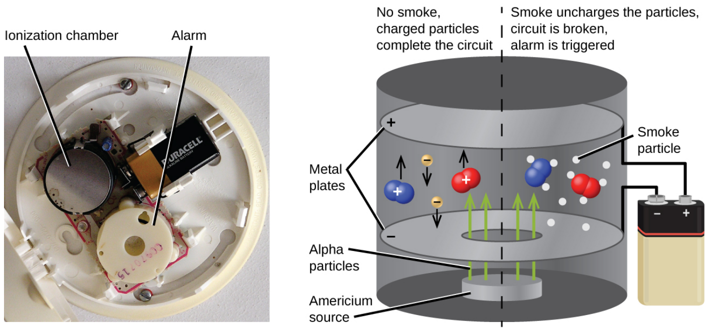

Figure 2.7.6: Inside a smoke detector, Am-241 emits α particles that ionize the air, creating a small electric current. During a fire, smoke particles impede the flow of ions, reducing the current and triggering an alarm. (credit a: modification of work by “Muffet”/Wikimedia Commons)

The inside of a smoke detector is shown with the alarm and ionization chamber labeled. In the picture beside it, a schematic shows the mechanisms of a smoke detector. The two oppositely charged metals plates in the detector are shown along with the Americium source on the bottom part emitting alpha particles. The schematic is divided into two parts, one to show the presence and the other for the absence of smoke.

Summary

Compounds known as radioactive tracers can be used to follow reactions, track the distribution of a substance, diagnose and treat medical conditions, and much more. Other radioactive substances are helpful for controlling pests, visualizing structures, providing fire warnings, and for many other applications. Hundreds of millions of nuclear medicine tests and procedures, using a wide variety of radioisotopes with relatively short half-lives, are performed every year in the US. Most of these radioisotopes have relatively short half-lives; some are short enough that the radioisotope must be made on-site at medical facilities. Radiation therapy uses high-energy radiation to kill cancer cells by damaging their DNA. The radiation used for this treatment may be delivered externally or internally.

Glossary

chemotherapy

similar to internal radiation therapy, but chemical rather than radioactive substances are introduced into the body to kill cancer cells

external beam radiation therapy

radiation delivered by a machine outside the body

internal radiation therapy

(also, brachytherapy) radiation from a radioactive substance introduced into the body to kill cancer cells

radiation therapy

use of high-energy radiation to damage the DNA of cancer cells, which kills them or keeps them from dividing

radioactive tracer

(also, radioactive label) radioisotope used to track or follow a substance by monitoring its radioactive emissions

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2.8: The Electromagnetic Spectrum

Learning Objectives

- Know the properties of different types of electromagnetic radiation.

Electromagnetic waves have an extremely wide range of wavelengths, frequencies, and energies. The highest energy form of electromagnetic waves are gamma (γ) rays and the lowest energy form are radio waves.

The figure below shows the **electromagnetic spectrum**, which is all forms of electromagnetic radiation. On the far left of Figure 2.8.1 are the highest energy electromagnetic waves. These are called **gamma rays** and can be quite dangerous, in large numbers, to living systems. The next lower energy form of electromagnetic waves are called **x-rays**. Most of you are familiar with the penetration abilities of these waves. They can also be dangerous to living systems. Humans are advised to limit as much as possible the number of medical x-rays they have per year. Next lower, in energy, are **ultraviolet rays**. These rays are part of sunlight and the upper end of the ultraviolet range can cause sunburn and perhaps skin cancer. The tiny section next in the spectrum is the **visible range of light** … this section has been greatly expanded in the bottom half of the figure so it can be discussed in more detail. The visible range of electromagnetic radiation are the frequencies to which the human eye responds. Lower in the spectrum are infrared rays and radio waves.

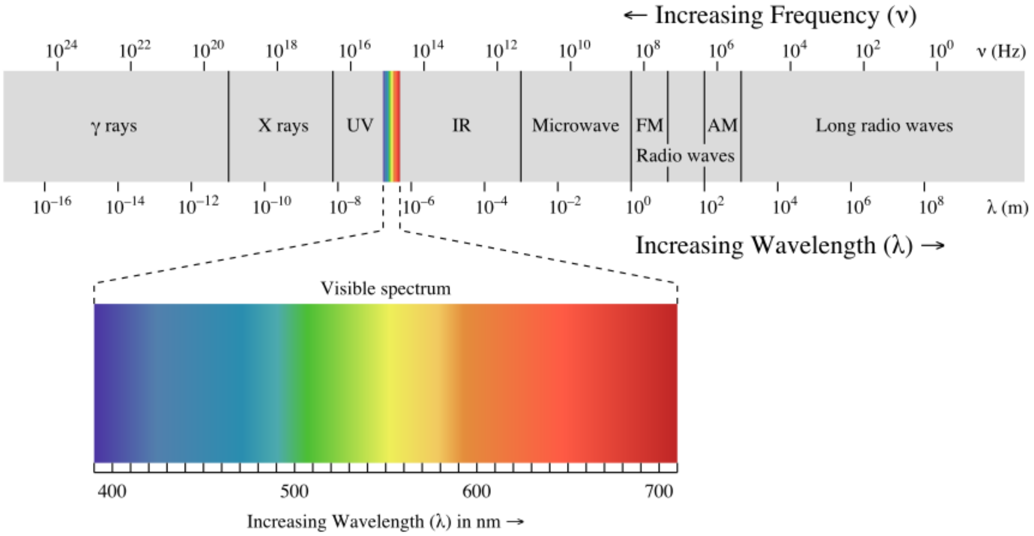

Figure 2.8.1: The electromagnetic spectrum, with its various regions labeled. The borders of each region are approximate. (CC BY-NC-SA; anonymous by request).

The light energies that are in the visible range are electromagnetic waves that cause the human eye to respond when those frequencies enter the eye. The eye sends a signal to the brain and the individual "sees" various colors. The highest energy waves in the visible region cause the brain to see violet and as the energy decreases, the colors change to blue, green, yellow, orange, and red. When the energy of the wave is above or below the visible range, the eye does not respond to them. When the eye receives several different frequencies at the same time, the colors are blended by the brain. If all frequencies of light strike the eye together, the brain sees white and if there are no visible frequencies striking the eye, the brain sees black. The objects that you see around you are light absorbers - that is, the chemicals on the surface of the object will absorb certain frequencies and not others. Your eyes detect the frequencies that strike your eye. Therefore, if your friend is wearing a red shirt, it means the dye in that shirt absorbs every frequency except red and the red frequencies are reflected. If your only light source was one exact frequency of blue light and you shined it on a shirt that was red in sunlight, the shirt would appear black because no light would be reflected. The light from fluorescent types of lights do not contain all the frequencies of sunlight and so clothes inside a store may appear to be a slightly different color than when you get them home.

Summary

- Electromagnetic radiation has a wide spectrum, including gamma rays, X-rays, UV rays, visible light, IR radiation, microwaves, and radio waves.
- The different colors of light differ in their frequencies (or wavelengths).

Contributions & Attributions

This page was constructed from content via the following contributor(s) and edited (topically or extensively) by the LibreTexts development team to meet platform style, presentation, and quality:

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2.9: The Bohr Model - Atoms with Orbits

Learning Objectives

- Define an energy level in terms of the Bohr model.
- Discuss how the Bohr model can be used to explain atomic spectra.

Electric light bulbs contain a very thin wire in them that emits light when heated. The wire is called a filament. The particular wire used in light bulbs is made of tungsten. A wire made of any metal would emit light under these circumstances, but tungsten was chosen because the light it emits contains virtually every frequency and therefore, the light emitted by tungsten appears white. A wire made of some other element would emit light of some color that was not convenient for our uses. Every element emits light when energized by heating or passing electric current through it. Elements in solid form begin to glow when they are heated sufficiently, and elements in gaseous form emit light when electricity passes through them. This is the source of light emitted by neon signs and is also the source of light in a fire.

Figure 2.9.1: Human/Need/Desire. Neon sculpture by Bruce Nauman (1983), who has been characterized as a conceptual artist.

Each Element Has a Unique Spectrum

The light frequencies emitted by atoms are mixed together by our eyes so that we see a blended color. Several physicists, including Angstrom in 1868 and Balmer in 1875, passed the light from energized atoms through glass prisms in such a way that the light was spread out so they could see the individual frequencies that made up the light. The **emission spectrum** (or **atomic spectrum**) of a chemical element is the unique pattern of light obtained when the element is subjected to heat or electricity.

Figure 2.9.2: Atomic Emission Spectrum of Hydrogen.

When hydrogen gas is placed into a tube and electric current passed through it, the color of emitted light is pink. But when the color is spread out, we see that the hydrogen spectrum is composed of four individual frequencies. The pink color of the tube is the result of our eyes blending the four colors. Every atom has its own characteristic spectrum; no two atomic spectra are alike. The image below shows the emission spectrum of iron. Because each element has a unique emission spectrum, elements can be defined using them.

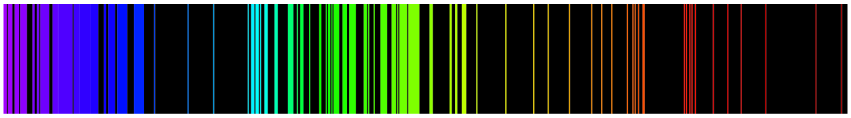

Figure 2.9.3: Atomic Emission Spectrum of Iron.

You may have heard or read about scientists discussing what elements are present in the sun or some more distant star, and after hearing that, wondered how scientists could know what elements were present in a place no one has ever been. Scientists determine what elements are present in distant stars by analyzing the light that comes from stars and finding the atomic spectrum of elements in that light. If the exact four lines that compose hydrogen's atomic spectrum are present in the light emitted by the star, that element contains hydrogen.

Bohr's Model of the Atom

By 1913, the concept of the atom had evolved from Dalton's indivisible spheres idea, to J. J. Thomson's plum pudding model, and then to Rutherford's nuclear atom theory. Rutherford, in addition to carrying out the brilliant experiment that demonstrated the presence of the atomic nucleus, also proposed that the electrons circled the nucleus in a planetary type motion. The solar system or planetary model of the atom was attractive to scientists because it was similar to something with which they were already familiar, namely the solar system.

Figure 2.9.3: Niels Bohr with Albert Einstein at Paul Ehrenfest's home in Leiden (December 1925).

Unfortunately, there was a serious flaw in the planetary model. It was already known that when a charged particle (such as an electron) moves in a curved path, it gives off some form of light and loses energy in doing so. This is, after all, how we produce TV signals. If the electron circling the nucleus in an atom loses energy, it would necessarily have to move closer to the nucleus as it loses energy, and would eventually crash into the nucleus. Furthermore, Rutherford's model was unable to describe how electrons give off light forming each element's unique atomic spectrum. These difficulties cast a shadow on the planetary model and indicated that, eventually, it would have to be replaced.

In 1913, the Danish physicist Niels Bohr proposed a model of the electron cloud of an atom in which electrons orbit the nucleus and were able to produce atomic spectra. Understanding Bohr's model requires some knowledge of electromagnetic radiation (or light).

Energy Levels

Bohr's key idea in his model of the atom is that electrons occupy definite orbitals that require the electron to have a specific amount of energy. In order for an electron to be in the electron cloud of an atom, it must be in one of the allowable orbitals and it must have the precise energy required for that orbit. Orbits closer to the nucleus would require smaller amounts of energy for an electron, and orbits farther from the nucleus would require the electron to have a greater amount of energy. The possible orbits are known as **energy levels**. One of the weaknesses of Bohr's model was that he could not offer a reason why only certain energy levels or orbits were allowed.

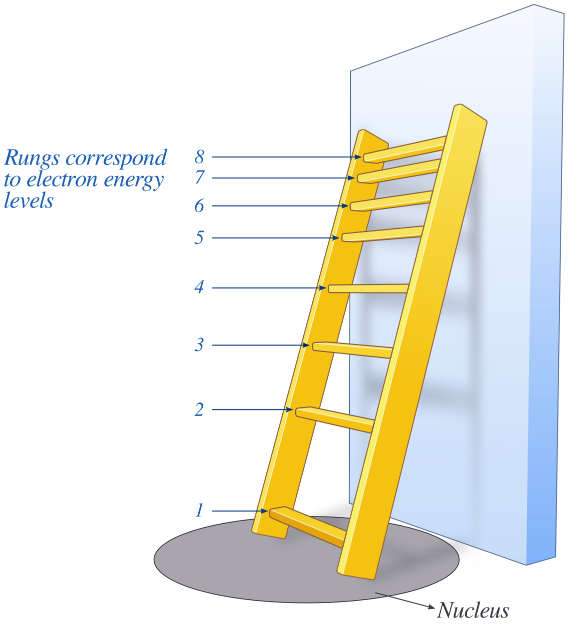

Figure 2.9.4: The energy levels of the electrons can be viewed as rungs on a ladder. Note that the spacing between rungs gets smaller at higher energies (CC BY-NC; Ümit Kaya)

Bohr hypothesized that the only way electrons could gain or lose energy would be to move from one energy level to another, thus gaining or losing precise amounts of energy. The energy levels are **quantized**, meaning that only specific amounts are possible. It would be like a ladder that had rungs only at certain heights. The only way you can be on that ladder is to be on one of the rungs, and the only way you could move up or down would be to move to one of the other rungs. Suppose we had such a ladder with 10 rungs. Other rules for the ladder are that only one person can be on a rung in the normal state, and the ladder occupants must be on the lowest rung available. If the ladder had five people on it, they would be on the lowest five rungs. In this situation, no person could move down because all of the lower rungs are full. Bohr worked out rules for the maximum number of electrons that could be in each energy level in his model, and required that an atom in its normal state (ground state) had all electrons in the lowest energy levels available. Under these circumstances, no electron could lose energy because no electron could move down to a lower energy level. In this way, Bohr's model explained why electrons circling the nucleus did not emit energy and spiral into the nucleus.

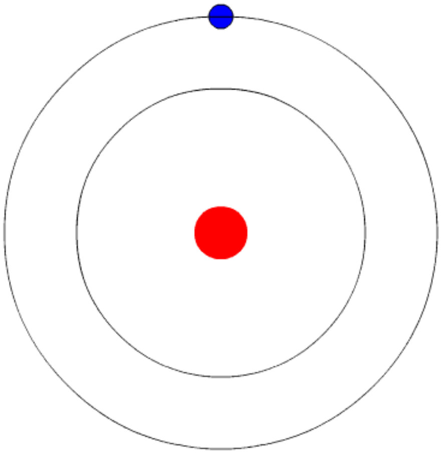

Figure 2.9.5: In Bohr's Model of the atom, electrons absorb energy to move to a higher level and release energy to move to lower levels. (CC BY-SA 3.0; Kurzon).

Bohr's Model and Atomic Spectra

The evidence used to support Bohr's model came from the atomic spectra. He suggested that an atomic spectrum is made by the electrons in an atom moving energy levels. The electrons typically have the lowest energy possible, called the **ground state**. If the electrons are given energy (through heat, electricity, light, etc.) the electrons in an atom could absorb energy by jumping to a higher energy level, or **excited state**. The electrons then give off the energy in the form of a piece of light—called a **photon**—that they had absorbed, to fall back to a lower energy level. The energy emitted by electrons dropping back to lower energy levels will always be precise amounts of energy, because the differences in energy levels are precise. This explains why you see specific lines of light when looking at an atomic spectrum—each line of light matches a specific "step down" that an electron can take in that atom. This also explains why each element produces a different atomic spectrum. Because each element has different acceptable energy levels for its electrons, the possible steps each element's electrons can take differ from all other elements.

Summary

- Bohr's model suggests each atom has a set of unchangeable energy levels, and electrons in the electron cloud of that atom must be in one of those energy levels.
- Bohr's model suggests that the atomic spectra of atoms is produced by electrons gaining energy from some source, jumping up to a higher energy level, then immediately dropping back to a lower energy level and emitting the energy difference between the two energy levels.
- The existence of the atomic spectra is support for Bohr's model of the atom.
- Bohr's model was only successful in calculating energy levels for the hydrogen atom.

Vocabulary

- **Emission spectrum** (or **atomic spectrum**) - The unique pattern of light given off by an element when it is given energy.
- **Energy levels** - Possible orbits that an electron can have in the electron cloud of an atom.
- **Ground state** - To be in the lowest energy level possible.
- **Excited state** - To be in a higher energy level.
- **Photon** - A piece of electromagnetic radiation, or light, with a specific amount of energy.
- **Quantized** - To have a specific amount.

2.10: The Quantum-Mechanical Model- Atoms with Orbitals

Learning Objectives

- Define quantum mechanics
- Differentiate between an orbit and an orbital.

How do you study something that seemingly makes no sense? We talk about electrons being in orbits and it sounds like we can tell where that electron is at any moment. We can draw pictures of electrons in orbit, but the reality is that we don't know exactly where they are. We are going to take a look at an area of science that even leaves scientists puzzled. When asked about quantum mechanics, Niels Bohr (who proposed the Bohr model of the atom) said: "Anyone who is not shocked by quantum theory has not understood it". Richard Feynman (one of the founders of modern quantum theory) stated: "I think I can safely say that nobody understands quantum theory." So, let's take a short trip into a land that challenges our everyday world.

Quantum Mechanics

The study of motion of large objects such as baseballs is called mechanics, or more specifically, classical mechanics. Because of the quantum nature of the electron and other tiny particles moving at high speeds, classical mechanics is inadequate to accurately describe their motion. **Quantum mechanics** is the study of the motion of objects that are atomic or subatomic in size and thus demonstrate wave-particle duality. In classical mechanics, the size and mass of the objects involved effectively obscures any quantum effects, so that such objects appear to gain or lose energies in any amounts. Particles whose motion is described by quantum mechanics gain or lose energy in small pieces called **quanta**.

One of the fundamental (and hardest to understand) principles of quantum mechanics is that the electron is both a particle and a wave. In the everyday macroscopic world of things we can see, something cannot be both. But this duality can exist in the quantum world of the submicroscopic on the atomic scale.

At the heart of quantum mechanics is the idea that we cannot accurately specify the location of an electron. All we can say is that there is a probability that it exists within this certain volume of space. The scientist Erwin Schrödinger developed an equation that deals with these calculations, which we will not pursue at this time.

Erwin Schrödinger.

Recall that in the Bohr model, the exact path of the electron was restricted to very well-defined circular **orbits** around the nucleus. An **orbital** is the quantum mechanical refinement of Bohr's orbit. In contrast to his concept of a simple circular orbit with a fixed radius, orbitals are mathematically derived regions of space with different *probabilities* of having an electron.

Summary

Quantum mechanics involves the study of material at the atomic level. This field deals with probabilities, since we cannot definitely locate a particle. Orbitals are mathematically derived regions of space with different *probabilities* of having an electron.

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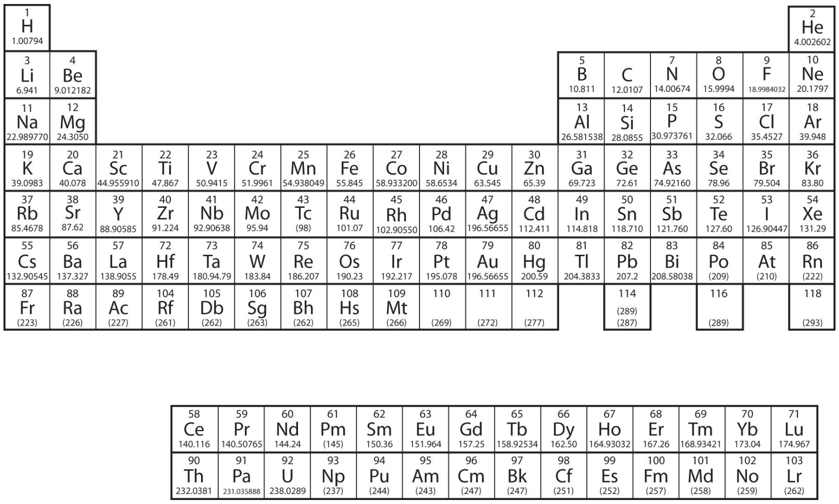
2.11: Electron Configurations and the Periodic Table

Learning Objectives

- Relate the electron configurations of the elements to the shape of the periodic table.
- Determine the expected electron configuration of an element by its place on the periodic table.

Previously, we introduced the periodic table as a tool for organizing the known chemical elements. A periodic table is shown in Figure 2.11.1. The elements are listed by atomic number (the number of protons in the nucleus), and elements with similar chemical properties are grouped together in columns.



1 H 1.00794																	2 He 4.002602
3 Li 6.941	4 Be 9.012182											5 B 10.811	C 12.0107	7 N 14.00674	8 O 15.9994	9 F 18.9984032	10 Ne 20.1797
11 Na 22.989770	12 Mg 24.3050											13 Al 26.581538	14 Si 28.0855	15 P 30.973761	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.955910	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.938049	26 Fe 55.845	27 Co 58.933200	28 Ni 58.6534	29 Cu 63.545	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.504	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.90550	46 Pd 106.42	47 Ag 196.56655	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.90447	54 Xe 131.29
55 Cs 132.90545	56 Ba 137.327	57 La 138.9055	72 Hf 178.49	73 Ta 180.94.79	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.078	79 Au 196.56655	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.58038	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (277)		114 (289) (287)		116 (289)		118 (293)

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

Figure 2.11.1: The Periodic Table

Why does the periodic table have the structure it does? The answer is rather simple, if you understand electron configurations: the shape of the periodic table mimics the filling of the subshells with electrons.

> *The shape of the periodic table mimics the filling of the subshells with electrons.*

Let us start with H and He. Their electron configurations are $1s^1$ and $1s^2$, respectively; with He, the n = 1 shell is filled. These two elements make up the first row of the periodic table (Figure 2.11.2)

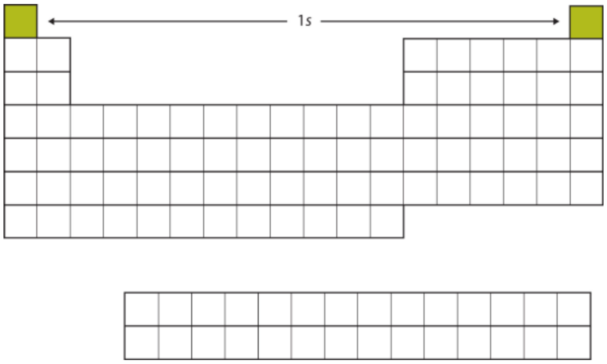


Figure 2.11.2: The 1*s* Subshell. H and He represent the filling of the 1*s* subshell.

The next two electrons, for Li and Be, would go into the 2*s* subshell. Figure 2.11.3 shows that these two elements are adjacent on the periodic table.

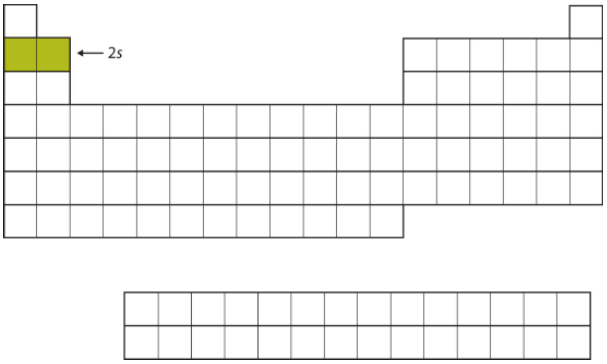


Figure 2.11.3: The 2s Subshell. In Li and Be, the 2s subshell is being filled.

For the next six elements, the 2p subshell is being occupied with electrons. On the right side of the periodic table, these six elements (B through Ne) are grouped together (Figure 2.11.4).

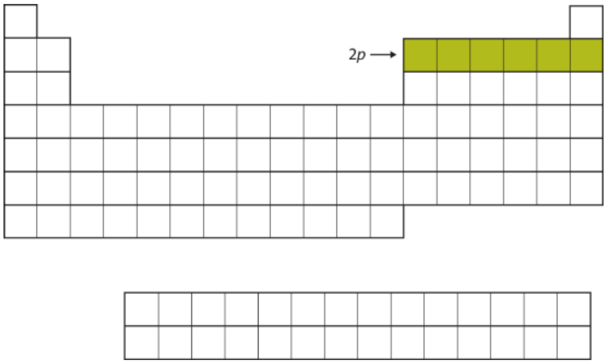


Figure 2.11.4: The 2p Subshell. For B through Ne, the 2p subshell is being occupied.

The next subshell to be filled is the 3s subshell. The elements when this subshell is being filled, Na and Mg, are back on the left side of the periodic table (Figure 2.11.5).

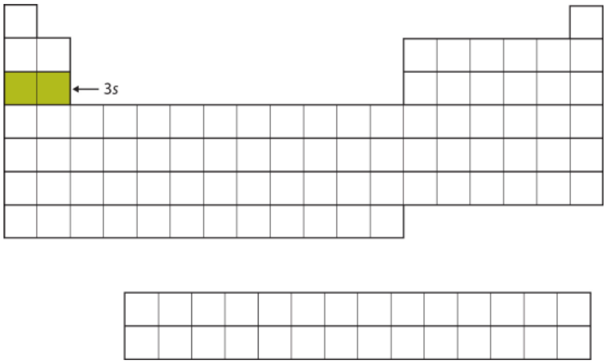


Figure 2.11.5: The 3s Subshell. Now the 3s subshell is being occupied.

Next, the 3p subshell is filled with the next six elements (Figure 2.11.6).

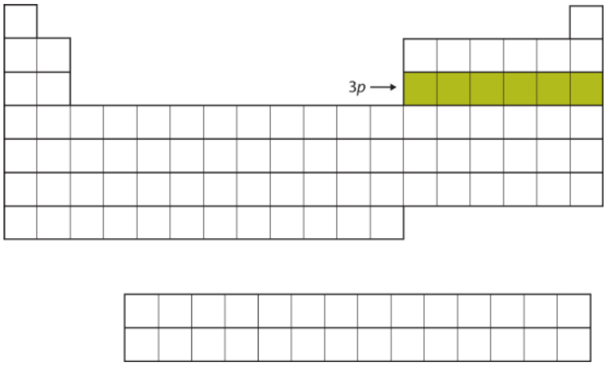


Figure 2.11.6: The 3*p* Subshell. Next, the 3*p* subshell is filled with electrons.

Instead of filling the 3*d* subshell next, electrons go into the 4*s* subshell (Figure 2.11.7).

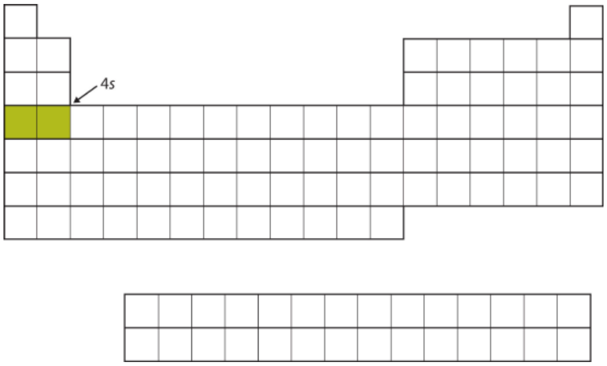


Figure 2.11.7: The 4*s* Subshell. The 4*s* subshell is filled before the 3*d* subshell. This is reflected in the structure of the periodic table.

After the 4*s* subshell is filled, the 3*d* subshell is filled with up to 10 electrons. This explains the section of 10 elements in the middle of the periodic table (Figure 2.11.8).

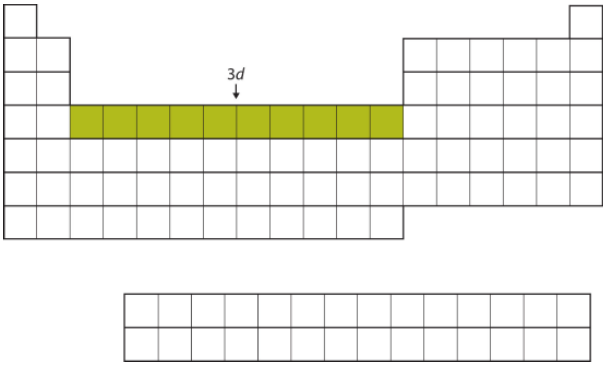


Figure 2.11.8: The 3*d* Subshell. The 3*d* subshell is filled in the middle section of the periodic table.

...And so forth. As we go across the rows of the periodic table, the overall shape of the table outlines how the electrons are occupying the shells and subshells.

The first two columns on the left side of the periodic table are where the *s* subshells are being occupied. Because of this, the first two rows of the periodic table are labeled the **s block**. Similarly, the **p block** are the right-most six columns of the periodic table, the **d block** is the middle 10 columns of the periodic table, while the **f block** is the 14-column section that is normally depicted as detached from the main body of the periodic table. It could be part of the main body, but then the periodic table would be rather long and cumbersome. Figure 2.11.9 shows the blocks of the periodic table.

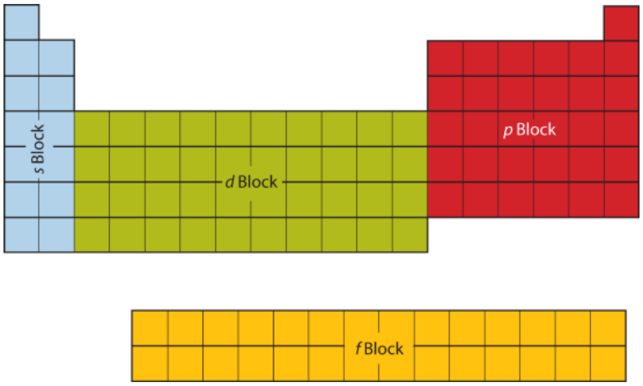


Figure 2.11.9: Blocks on the Periodic Table. The periodic table is separated into blocks depending on which subshell is being filled for the atoms that belong in that section.

The electrons in the highest-numbered shell, plus any electrons in the last unfilled subshell, are called **valence electrons**; the highest-numbered shell is called the **valence shell**. (The inner electrons are called *core electrons*.) The valence electrons largely control the chemistry of an atom. If we look at just the valence shell's electron configuration, we find that in each column, the valence shell's electron configuration is the same. For example, take the elements in the first column of the periodic table: H, Li, Na, K, Rb, and Cs. Their electron configurations (abbreviated for the larger atoms) are as follows, with the valence shell electron configuration highlighted:

Electrons, electron configurations, and the valence shell electron configuration highlighted.

H:	$1s^{\mathbf{1}}$
Li:	$1s^22s^{\mathbf{1}}$
Na:	$[Ne]3s^{\mathbf{1}}$
K:	$[Ar]4s^{\mathbf{1}}$
Rb:	$[Kr]5s^{\mathbf{1}}$
Cs:	$[Xe]6s^{\mathbf{1}}$

They all have a similar electron configuration in their valence shells: a single *s* electron. Because much of the chemistry of an element is influenced by valence electrons, we would expect that these elements would have similar chemistry—*and they do*. The organization of electrons in atoms explains not only the shape of the periodic table, but also the fact that elements in the same column of the periodic table have similar chemistry.

The same concept applies to the other columns of the periodic table. Elements in each column have the same valence shell electron configurations, and the elements have some similar chemical properties. This is strictly true for all elements in the *s* and *p* blocks. In the *d* and *f* blocks, because there are exceptions to the order of filling of subshells with electrons, similar valence shells are not absolute in these blocks. However, many similarities do exist in these blocks, so a similarity in chemical properties is expected.

Similarity of valence shell electron configuration implies that we can determine the electron configuration of an atom solely by its position on the periodic table. Consider Se, as shown in Figure 2.11.10. It is in the fourth column of the *p* block. This means that its electron configuration should end in a p^4 electron configuration. Indeed, the electron configuration of Se is $[Ar]4s^23d^{10}4p^4$, as expected.

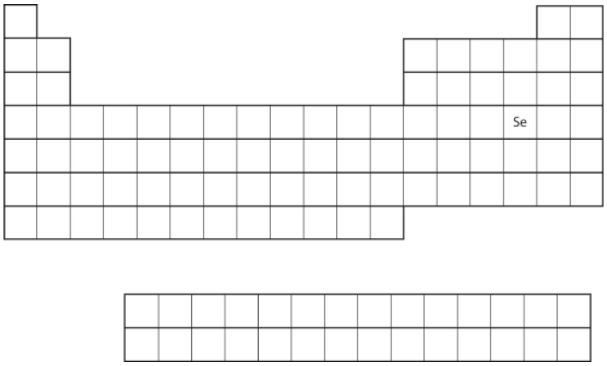


Figure 2.11.10: Selenium on the Periodic Table

Example 2.11.1: Predicting Electron Configurations

From the element’s position on the periodic table, predict the valence shell electron configuration for each atom (Figure 2.11.11).

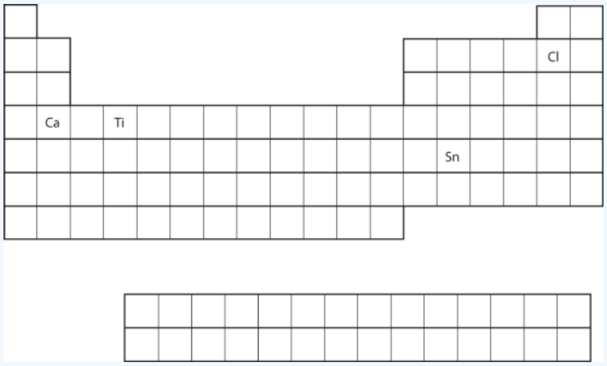


Figure 2.11.11: Various Elements on the Periodic Table

a. Ca
b. Sn

Solution

a. Ca is located in the second column of the *s* block. We expect that its electron configuration should end with s^2. Calcium’s electron configuration is $[Ar]4s^2$.
b. Sn is located in the second column of the *p* block, so we expect that its electron configuration would end in p^2. Tin’s electron configuration is $[Kr]5s^2 4d^{10} 5p^2$.

Exercise 2.11.1

From the element’s position on the periodic table, predict the valence shell electron configuration for each atom. Figure 2.11.11

a. Ti
b. Cl

Answer a

$[Ar]4s^2 3d^2$

Answer b

$[Ne]3s^2 3p^5$

Summary

The arrangement of electrons in atoms is responsible for the shape of the periodic table. Electron configurations can be predicted by the position of an atom on the periodic table.

2.12: Arrangements of Electrons

Learning Objectives

- Describe how electrons are grouped within atoms.

Although we have discussed the general arrangement of subatomic particles in atoms, we have said little about how electrons occupy the space about the nucleus. Do they move around the nucleus at random, or do they exist in some ordered arrangement?

The modern theory of electron behavior is called quantum mechanics. It makes the following statements about electrons in atoms:

- Electrons in atoms can have only certain specific energies. We say that the energies of the electrons are quantized.
- Electrons are organized according to their energies into sets called **shells** (labeled by the principle quantum number, ***n***). Generally the higher the energy of a shell, the farther it is (on average) from the nucleus. Shells do not have specific, fixed distances from the nucleus, but an electron in a higher-energy shell will spend more time farther from the nucleus than does an electron in a lower-energy shell.
- Shells are further divided into subsets of electrons called **subshells**. The first shell has only one subshell, the second shell has two subshells, the third shell has three subshells, and so on. The subshells of each shell are labeled, in order, with the letters *s*, *p*, *d*, and *f*. Thus, the first shell has only a single *s* subshell (called **1*s***), the second shell has **2*s*** and **2*p*** subshells, the third shell has **3*s***, **3*p***, and **3*d*** and so forth.

Table 2.6.1: Shells and Subshells

Shell	Number of Subshells	Names of Subshells
1	1	*1s*
2	2	*2s* and *2p*
3	3	*3s*, *3p* and *3d*
4	4	*4s*, *4p*, *4d* and *4f*

- Different subshells hold a different maximum number of electrons. **Any *s* subshell can hold up to 2 electrons; *p*, 6; *d*, 10; and *f*, 14.**

Table 2.6.2 Number of Electrons

Subshell	Maximum Number of Electrons
s	2
p	6
d	10
f	14

It is the arrangement of electrons into shells and subshells that most concerns us here, so we will focus on that.

We use numbers to indicate which shell an electron is in. As shown in Table **2.6.1**, the first shell, closest to the nucleus and with the lowest-energy electrons, is shell 1. This first shell has only one subshell, which is labeled 1*s* and can hold a maximum of 2 electrons. We combine the shell and subshell labels when referring to the organization of electrons about a nucleus and use a superscript to indicate how many electrons are in a subshell. Thus, because a hydrogen atom has its single electron in the *s* subshell of the first shell, we use $\mathbf{1s^1}$ to describe the electronic structure of hydrogen. This structure is called an electron configuration. Electron configurations are shorthand descriptions of the arrangements of electrons in atoms. The electron configuration of a hydrogen atom is spoken out loud as “one-ess-one.”

Helium atoms have 2 electrons. Both electrons fit into the 1*s* subshell because *s* subshells can hold up to 2 electrons; therefore, the electron configuration for helium atoms is $\mathbf{1s^2}$ (spoken as “one-ess-two”).

The 1*s* subshell cannot hold 3 electrons (because an *s* subshell can hold a maximum of 2 electrons), so the electron configuration for a lithium atom cannot be $1s^3$. Two of the lithium electrons can fit into the 1*s* subshell, but the third electron must go into the second shell. The second shell has two subshells, *s* and *p*, which fill with electrons in that order. The 2*s* subshell holds a maximum of 2 electrons, and the 2*p* subshell holds a maximum of 6 electrons. Because lithium's final electron goes into the 2*s* subshell, we write the electron configuration of a lithium atom as $\mathbf{1s^2 2s^1}$. The shell diagram for a lithium atom is shown below. The shell closest to the nucleus (first shell) has 2 dots representing the 2 electrons in **1*s***, while the outermost shell (**2*s***) has 1 electron.

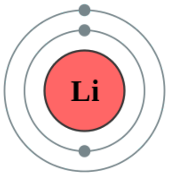


Figure 2.6.1 Shell diagram of lithium (Li) atom.

The next largest atom, beryllium, has 4 electrons, so its electron configuration is $1s^2 2s^2$. Now that the 2*s* subshell is filled, electrons in larger atoms start filling the 2*p* subshell. Thus, the electron configurations for the next six atoms are as follows:

- B: $1s^2 2s^2 2p^1$
- C: $1s^2 2s^2 2p^2$
- N: $1s^2 2s^2 2p^3$
- O: $1s^2 2s^2 2p^4$
- F: $1s^2 2s^2 2p^5$
- Ne: $1s^2 2s^2 2p^6$

With neon, the 2*p* subshell is completely filled. Because the second shell has only two subshells, atoms with more electrons now must begin the third shell. The third shell has three subshells, labeled *s*, *p*, and *d*. The *d* subshell can hold a maximum of 10 electrons. The first two subshells of the third shell are filled in order—for example, the electron configuration of aluminum, with 13 electrons, is $1s^2 2s^2 2p^6 3s^2 3p^1$. However, a curious thing happens after the 3*p* subshell is filled: the 4*s* subshell begins to fill before the 3*d* subshell does. In fact, the exact ordering of subshells becomes more complicated at this point (after argon, with its 18 electrons), so we will not consider the electron configurations of larger atoms. A fourth subshell, the *f* subshell, is needed to complete the electron configurations for all elements. An *f* subshell can hold up to 14 electrons.

Electron filling always starts with **1*s***, the subshell closest to the nucleus. Next is **2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d***, **4*p***, **5s, 4d, 5p, 6s,** etc., shown in the electron shell filling order diagram in Figure **2.6.2**. Follow each arrow in order **from top to bottom**. The subshells you reach along each arrow give the ordering of filling of subshells in larger atoms.

Example 2.12.1: Electronic Configuration of Phosphorus Atoms

Using Figure 2.6.2 as your guide, write the electron configuration of a neutral phosphorus atom. The atomic number of P is 15.

Solution

A neutral phosphorus atom has 15 electrons. Two electrons can go into the 1*s* subshell, 2 can go into the 2*s* subshell, and 6 can go into the 2*p* subshell. That leaves 5 electrons. Of those 5 electrons, 2 can go into the 3*s* subshell, and the remaining 3 electrons can go into the 3*p* subshell. Thus, the electron configuration of neutral phosphorus atoms is $1s^2 2s^2 2p^6 3s^2 3p^3$.

Exercise 2.12.1: Electronic Configuration of Chlorine Atoms

Using Figure 2.6.2 as your guide, write the electron configuration of a neutral chlorine atom. The atomic number of Cl is 17.

Answer

A neutral chlorine atom has 17 electrons. Two electrons can go into the 1*s* subshell, 2 can go into the 2*s* subshell, and 6 can go into the 2*p* subshell. That leaves 7 electrons. Of those 7 electrons, 2 can go into the 3*s* subshell, and the remaining 5 electrons can go into the 3*p* subshell. Thus, the electron configuration of neutral chlorine atoms is $1s^2 2s^2 2p^6 3s^2 3p^5$.

Since the arrangement of the periodic table is based on the electron configurations, Figure 2.6.3 provides an alternative method for determining the electron configuration. The filling order simply **begins at the top left,** with hydrogen (Z=1) and includes each subshell as you proceed in increasing atomic number (*Z)* order.

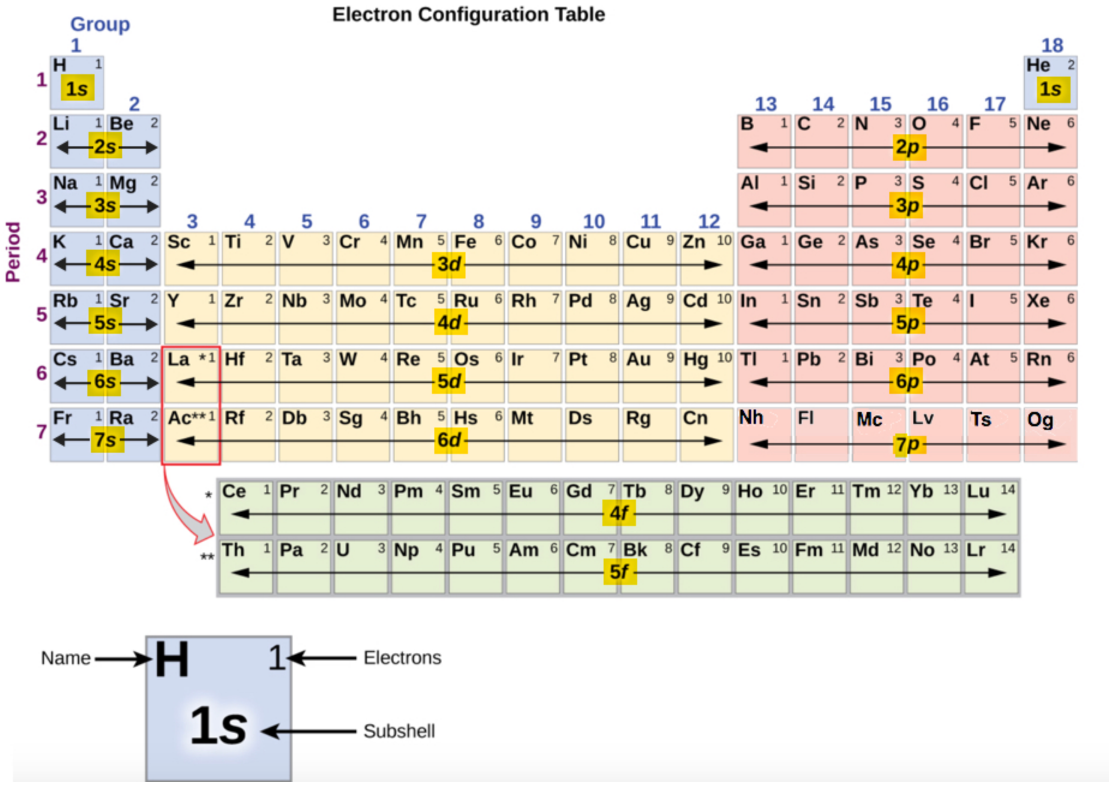


Figure 2.12.3: This periodic table shows the electron configuration for each subshell. By “building up” from hydrogen, this table can be used to determine the electron configuration for any atom on the periodic table.

For example, the first row (Period 1) contains H and He only, because only two electrons are required to fill the 1s subshell. The second row s-block, contains only two elements, Li and Be, to fill the 2s subshell. This is followed by the second row p-block, containing 6 elements (B through Ne) since six electrons are required to fill the 2p subshell. The third row is similar to the second row elements. Two electrons are needed (Na and Mg) to fill the 3s subshell and six electrons are required (Al through Ar) to complete the 3p subshell. After filling the 3*p* block up to Ar, we see the next subshell will be 4s (K, Ca), followed by the 3*d* subshell, which are filled by ten electrons (Sc through Zn). The 4p subshell is filled next by six electrons (Ga through Kr). As you can see, the periodic table shown in Figure 2.6.3 provides a simple way to remember the order of filling the subshells in determining the electron configuration. The order of filling subshells is the same: **1s, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d***, **4*p***, **5s, 4d, 5p, 6s,** etc.

Example 2.12.2

Using Figure 2.6.3 as your guide, write the electron configuration of neutral aluminum atom. The atomic number of Al is 13.

Solution

Aluminum has 13 electrons.

Start at Period 1 of the periodic table, Figure 2.6.3. Place two electrons in the 1s subshell ($\mathbf{1s^2}$).

Proceed to Period 2 (left to right direction). Place the next two electrons in the 2s subshell ($\mathbf{2s^2}$) and the next six electrons in the 2p subshell ($\mathbf{2p^6}$).

Proceed to Period 3 (left to right direction). Place the next two electrons in the 3s subshell ($\mathbf{3s^2}$) and the last one electron in the 3p subshell ($\mathbf{3p^1}$).

The electron configuration of Aluminum is **$1s^22s^22p^63s^23p^1$**

Exercise 2.12.2

Using Figure 2.6.3 as your guide, write the electron configuration of the atom that has 20 electrons

Answer

Start at Period 1 of Figure 2.6.3. Place two electrons in the 1s subshell (**$1s^2$**).

Proceed to Period 2 (left to right direction). Place the next two electrons in the 2s subshell (**$2s^2$**) and the next six electrons in the 2p subshell (**$2p^6$**).

Proceed to Period 3 (left to right direction). Place the next two electrons in the 3s subshell (**$3s^2$**) and the next six electron in the 3p subshell (**$3p^6$**).

Proceed to Period 4. Place the remaining two electrons in the 4s subshell (**$4s^2$).**

The electron configuration is **$1s^22s^22p^63s^23p^64s^2$**

Valence Electrons

When studying chemical reactivity, we find that only some electrons are involved in bonding. These electrons are the valence electrons. For elements in the s and p blocks, the valence electrons are those in the highest occupied energy level. **Valence electrons** are the electrons in the highest occupied principal energy level of an atom.

In the second period elements, the two electrons in the $1s$ sublevel are called **inner-shell electrons** and are not involved directly in the element's reactivity or in the formation of compounds. Lithium has a single electron in the second principal energy level and so we say that lithium has one valence electron. Beryllium has two valence electrons. How many valence electrons does boron have? You must recognize that the second principal energy level consists of both the $2s$ and the $2p$ sublevels and so the answer is three. In fact, the number of valence electrons goes up by one for each step across a period until the last element is reached. Neon, with its configuration ending in $2s^22p^6$, has eight valence electrons.

The alkali metal sodium (atomic number 11) has one more electron than the neon atom. This electron must go into the lowest-energy subshell available, the 3*s* orbital, giving a $1s^22s^22p^63s^1$ configuration. The electrons occupying the outermost shell orbital(s) (highest value of *n*) are called valence electrons, and those occupying the inner shell orbitals are called **core electrons** (Figure \PageIndex4). Since the core electron shells correspond to noble gas electron configurations, we can abbreviate electron configurations by writing the noble gas that matches the core electron configuration, along with the valence electrons in a condensed format. For our sodium example, the symbol [Ne] represents core electrons, ($1s^22s^22p^6$) and our abbreviated or condensed configuration is [Ne]$3s^1$.

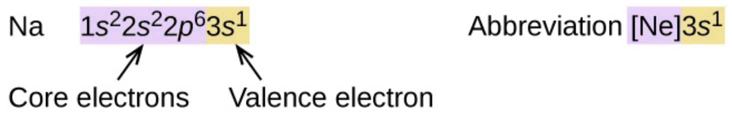


Figure 2.12.4: A core-abbreviated electron configuration (right) replaces the core electrons with the noble gas symbol whose configuration matches the core electron configuration of the other element.

Similarly, the abbreviated configuration of lithium can be represented as [He]$2s^1$, where [He] represents the configuration of the helium atom, which is identical to that of the filled inner shell of lithium. Writing the configurations in this way emphasizes the similarity of the configurations of lithium and sodium. Both atoms, which are in the alkali metal family, have only one electron in a valence *s* subshell outside a filled set of inner shells.

$$\begin{gathered}\mathrm{Li:[He]}\,2s^1\\ \mathrm{Na:[Ne]}\,3s^1\end{gathered}\tag{2.12.1}$$

A chemical reaction results from electron removal, electron addition, or electron sharing of the valence electrons of the different atoms. The path a specific element will take depends on where the electrons are in the atom and how many there are. Thus, it is convenient to separate electrons into two groups. Valence shell electrons (or, more simply, the *valence electrons*) are the electrons in the highest-numbered shell, or valence shell, while core electrons are the electrons in lower-numbered shells. We can see from the electron configuration of a carbon atom—$1s^22s^22p^2$—that it has 4 valence electrons ($2s^22p^2$) and 2 core electrons ($1s^2$). You will see in the next chapters that the chemical properties of elements are determined by the number of valence electrons.

Example 2.12.3:

Examine the electron configuration of neutral phosphorus atoms in Example 2.12.1, $1s^22s^22p^63s^23p^3$ and write the abbreviated notation.

Solution

Phosphorus has electron configuration, $1s^22s^22p^63s^23p^3$.

The highest-numbered shell is the third shell ($3s^23p^3$): **2 electrons** in the 3*s* subshell and **3 electrons** in the 3*p* subshell. That gives a total of **5 valence electrons**.

The 10 inner shell (core) electrons, $1s^22s^22p^6$ can be replaced by [Ne] (see Figure 2.6.3). Abbreviated notation is : [Ne]$3s^23p^3$

Exercise 2.12.3:

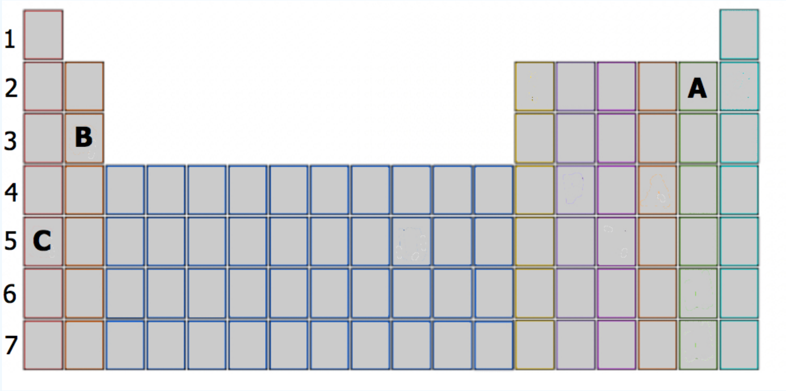
Examine the electron configuration of neutral calcium atom (Exercise 2.12.2), $1s^22s^22p^63s^23p^64s^2$, and write the abbreviated notation.

Answer

> The highest-numbered shell is the fourth shell $4s^2$, which has **2 electrons** in the *4s* subshell. Hence, Calcium has **2 valence electrons**.
>
> The 18 inner-shell (core) electrons, $1s^22s^22p^63s^23p^6$, can be replaced by [Ar], see Figure 2.6.3. The abbreviated notation is: [Ar]$4s^2$

Example 2.12.4

Based on their respective locations in the periodic table (use Figure 2.6.3), determine the number of valence electrons and the valence shell configuration of elements A, B and C.



Solution

Element A is located in Period 2, the *5th position* in *2p*-block. Before the electrons are placed in *2p* subshell, the *2s* subshell must be filled first. This means that A has ***two valence electrons*** in 2s (**$2s^2$**) and ***five valence electrons*** in 2p (**$2p^5$**). Answer: **$2s^22p^5$. It has *2 + 5 = 7* valence electrons.**

Element B is located in Period 3, the *2nd position* in *3s*-block. This means that B has ***two valence electrons*** in 3s (**$3s^2$**). Answer: **$3s^2$.**

Element C is located in Period 5, the *1st position* in *5s*-block). This means that there is only **one valence electron** in 5s (**$5s^1$**). Answer: **$5s^1$**.

Exercise 2.12.4

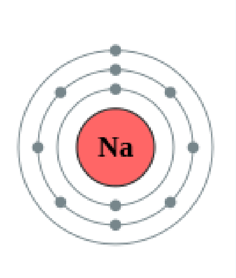
Using the location of Na is the periodic table (Figure 2.6.3), draw the shell diagram of sodium atom.

Answer

Sodium (Na) is the first element in the 3rd row (Period 3) in the periodic table. This means that the first shell and second shells of Na atom are filled to the maximum number of electrons.

The first shell **(1s)** *is filled with* ***2 electrons***. The second shell (2s and 2p) has a **total of** ***8 electrons***. And, the third (last) shell has ***1 electron***.

The shell diagram of the Na atom is shown below. The shell nearest the nucleus (first shell) has 2 electrons (2 dots), the second shell has 8 electrons and the last (outermost) shell has 1 electron. (**2.8.1**)



Concept Review Exercises

1. How are electrons organized in atoms?
2. What information does an electron configuration convey?
3. What is the difference between core electrons and valence electrons?

Answers

1. Electrons are organized into shells and subshells around nuclei.
2. The electron configuration states the arrangement of electrons in shells and subshells.
3. Valence electrons are in the highest-numbered shell; all other electrons are core electrons.

Key Takeaway

- Electrons are organized into shells and subshells about the nucleus of an atom.
- The valence electrons determine the reactivity of an atom.

Exercises

2.13: Periodic Trends

Learning Objective

- Be able to state how certain properties of atoms vary based on their relative position on the periodic table.

One of the reasons the periodic table is so useful is because its structure allows us to qualitatively determine how some properties of the elements vary versus their position on the periodic table. The variation of properties versus position on the periodic table is called **periodic trends**. There is no other tool in science that allows us to judge relative properties of a class of objects like this, which makes the periodic table a very useful tool. Many periodic trends are general. There may be a few points where an opposite trend is seen, but there is an overall trend when considered across a whole row or down a whole column of the periodic table.

The first periodic trend we will consider atomic radius. The **atomic radius** is an indication of the size of an atom. Although the concept of a definite radius of an atom is a bit fuzzy, atoms behave as if they have a certain radius. Such radii can be estimated from various experimental techniques, such as the x-ray crystallography of crystals.

As you go down a column of the periodic table, the atomic radii increase. This is because the valence electron shell is getting a larger and there is a larger principal quantum number, so the valence shell lies physically farther away from the nucleus. This trend can be summarized as follows:

$$as \downarrow PT, atomic\ radius \uparrow \tag{2.13.1}$$

where PT stands for periodic table. Going across a row on the periodic table, left to right, the trend is different. This is because although the valence shell maintains the same principal quantum number, the number of protons-and hence the nuclear charge-is increasing as you go across the row. The increasing positive charge casts a tighter grip on the valence electrons, so as you go across the periodic table, the atomic radii decrease. Again, we can summarize this trend as follows:

$$as \rightarrow PT, atomic\ radius \downarrow \tag{2.13.2}$$

Figure $2.13.1$ shows spheres representing the atoms of the *s* and *p* blocks from the periodic table to scale, showing the two trends for the atomic radius.

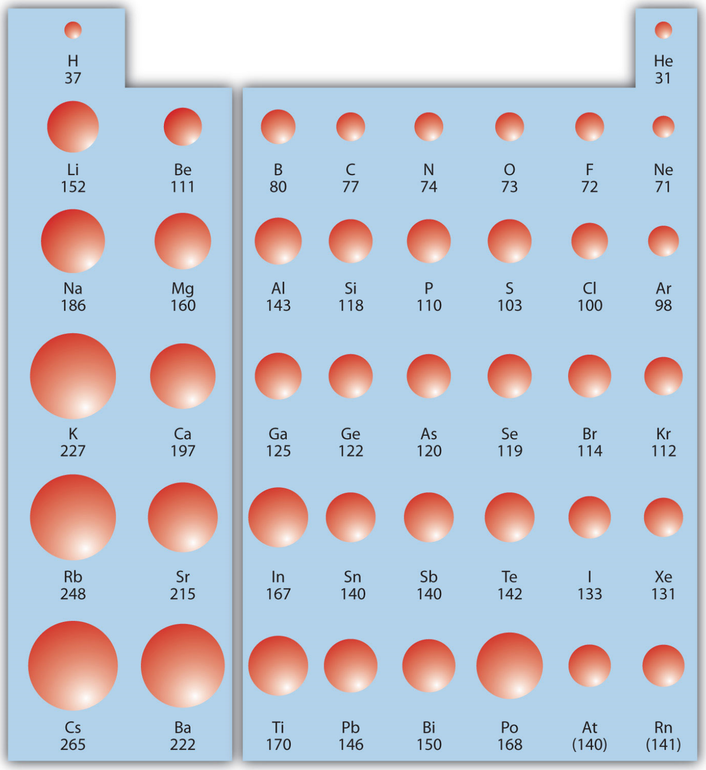


Figure $2.13.1$: Atomic Radii Trends on the Periodic Table. Although there are some reversals in the trend (e.g., see Po in the bottom row), atoms generally get smaller as you go across the periodic table and larger as you go down any one column. Numbers are the radii in pm.

Example $2.13.1$: Atomic Radii

Referring only to a periodic table and not to Figure $2.13.1$, which atom is larger in each pair?

a. Si or S
b. S or Te

Solution

a. Si is to the left of S on the periodic table, so it is larger because as you go across the row, the atoms get smaller.
b. S is above Te on the periodic table, so Te is larger because as you go down the column, the atoms get larger.

Exercise $2.13.1$: Atomic Radii

Referring only to a periodic table and not to Figure $2.13.1$, which atom is smaller, Ca or Br?

Answer

Br

Ionization energy (IE) is the amount of energy required to remove an electron from an atom in the gas phase:

$$A(g) \rightarrow A^{+}(g) + e^{-} \qquad \Delta H \equiv IE \tag{2.13.3}$$

IE is usually expressed in kJ/mol of atoms. A mole is an specified amount of a substance. It is always positive because the removal of an electron always requires that energy be put in (i.e., it is endothermic). IE also shows periodic trends. As you go down the periodic table, it becomes easier to remove an electron from an atom (i.e., IE decreases) because the valence electron is farther away from the nucleus. Thus,

$$as \downarrow PT,\ IE \downarrow \tag{2.13.4}$$

However, as you go across the periodic table and the electrons get drawn closer in, it takes more energy to remove an electron; as a result, IE increases:

$$as \rightarrow PT,\ IE \uparrow \tag{2.13.5}$$

Figure $2.13.2$ shows values of IE versus position on the periodic table. Again, the trend isn't absolute, but the general trends going across and down the periodic table should be obvious.

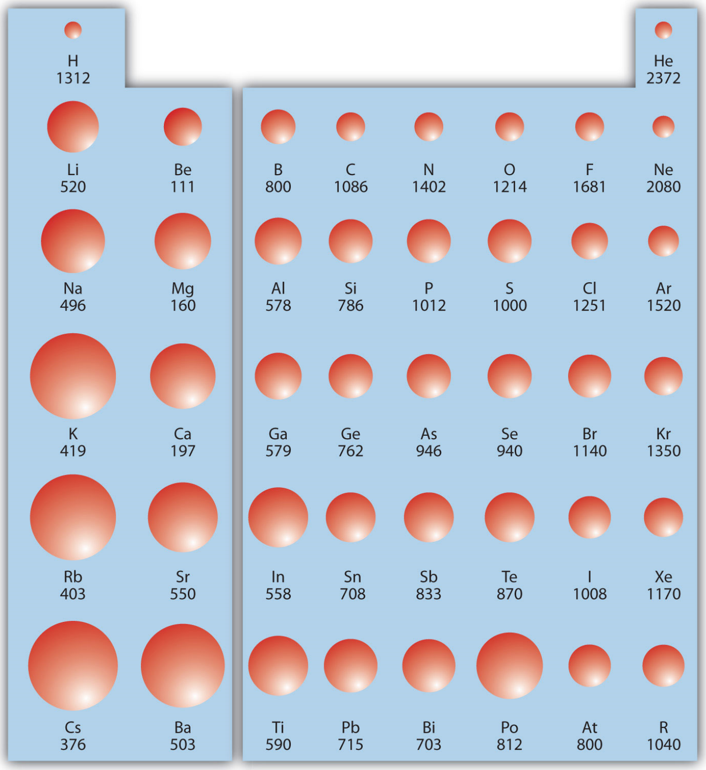


Figure 2.13.2: Ionization Energy on the Periodic Table. Values are in kJ/mol.

IE also shows an interesting trend within a given atom. This is because more than one IE can be defined by removing successive electrons (if the atom has them to begin with):

- First Ionization Energy (IE_1):

$$A(g) \rightarrow A^+(g) + e^- \tag{2.13.6}$$

- Second Ionization Energy (IE_2):

$$A^+(g) \rightarrow A^{2+}(g) + e^- \tag{2.13.7}$$

- Third Ionization Energy (IE_3):

$$A^{2+}(g) \rightarrow A^{3+}(g) + e^- \tag{2.13.8}$$

and so forth.

Each successive IE is larger than the previous because an electron is being removed from an atom with a progressively larger positive charge. However, IE takes a large jump when a successive ionization goes down into a new shell. For example, the following are the first three IEs for Mg, whose electron configuration is $1s^2 2s^2 2p^6 3s^2$:

First Ionization Energy (IE_1) = 738 kJ/mol:

$$Mg(g) \rightarrow Mg^+(g) + e^- \tag{2.13.9}$$

- Second Ionization Energy (IE_2) = 1,450 kJ/mol:

$$Mg^+(g) \rightarrow Mg^{2+}(g) + e^- \tag{2.13.10}$$

Third Ionization Energy (IE_3) = 7,734 kJ/mol:

$$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^{-} \tag{2.13.11}$$

The second IE is twice the first, which is not a surprise: the first IE involves removing an electron from a neutral atom, while the second one involves removing an electron from a positive ion. The third IE, however, is over *five times* the previous one. Why is it so much larger? Because the first two electrons are removed from the 3*s* subshell, but the third electron has to be removed from the *n* = 2 shell (specifically, the 2*p* subshell, which is lower in energy than the *n* = 3 shell). Thus, it takes much more energy than just overcoming a larger ionic charge would suggest. It is trends like this that demonstrate that electrons are organized in atoms in groups.

Example 2.13.2: Ionization Energies

Which atom in each pair has the larger first ionization energy?

a. Ca or Sr
b. K or K^+

Solution

a. Because Sr is below Ca on the periodic table, it is easier to remove an electron from it; thus, Ca has the higher IE.
b. Because K^+ has a positive charge, it will be harder to remove another electron from it, so its IE is larger than that of K. Indeed, it will be significantly larger because the next electron in K^+ to be removed comes from another shell.

Exercise 2.13.2: Ionization Energies

Which atom has the lower ionization energy, C or F?

Answer

C

The opposite of IE is described by **electron affinity (EA),** which is the energy change when a gas-phase atom accepts an electron. This is a concept you can learn about in general chemistry.

Summary

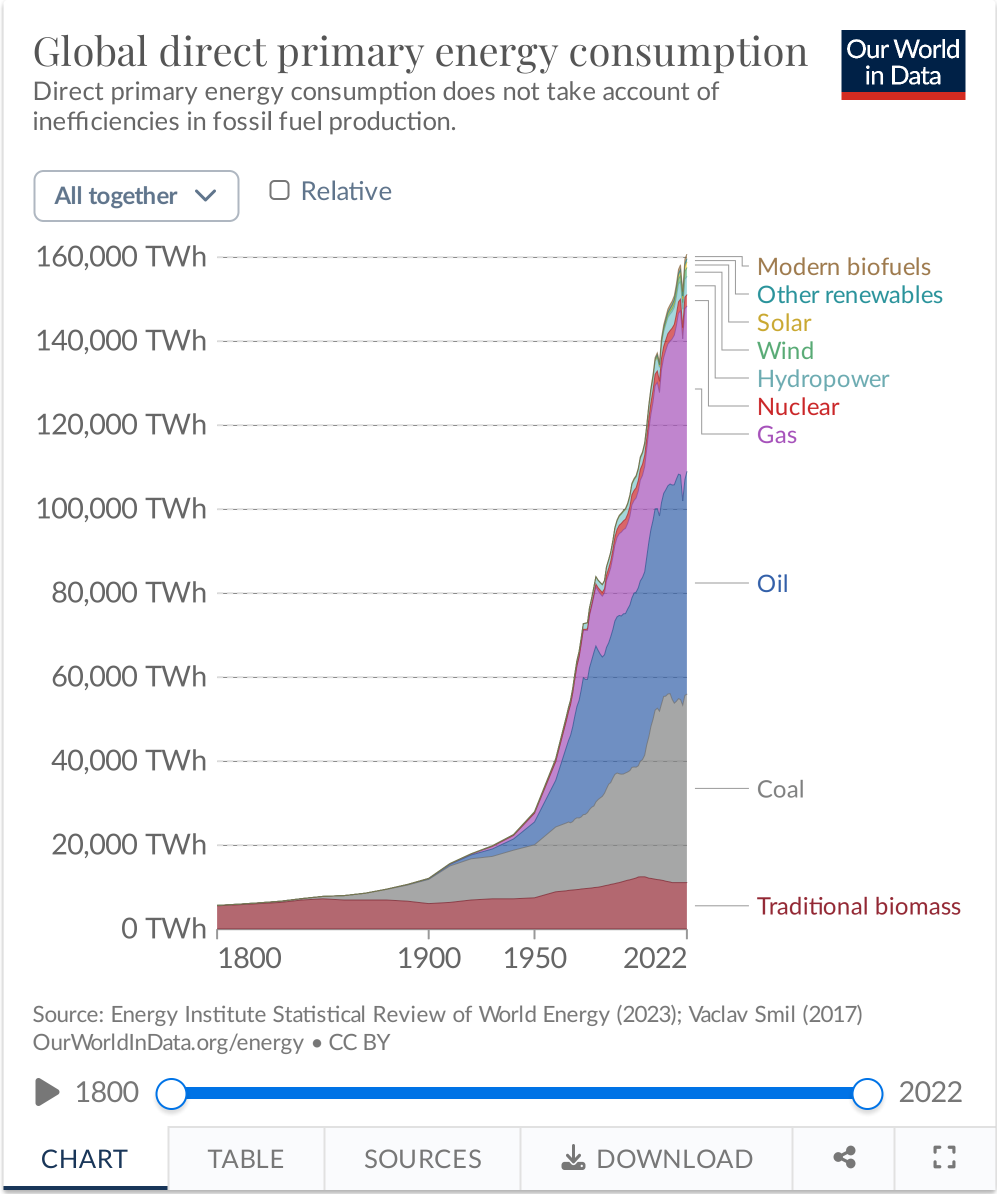
Certain properties-notably atomic radius, ionization energies, and electron affinities - can be qualitatively understood by the positions of the elements on the periodic table.

CHAPTER OVERVIEW

3: Energy Connections

3.1: Energy Production

In Module 1, you were introduced to measurements, the process of science and a little about energy. In Module 2, you learned about the atom. Through breaking apart nuclei (fission) or fusing together small nuclei (fusion), great amounts of energy can be released. However, much of what we do in chemistry does not rely only on the nuclei of atoms. In Module 3, you will learn about how atoms are combined to form molecules. By breaking apart and forming certain molecules, energy can be released. For example, in the car engines, a combustion reaction takes place between gasoline and oxygen to produce water and carbon dioxide and energy! By understanding how molecules are formed and that these molecules undergo chemical reactions, you will gain a better idea about energy production. Combustion reactions are not the only reactions that release energy. Many chemical reactions produce energy, but there is a need to look at other processes that produce energy (like wind, solar and geothermal) to help when our reactants (fuels) run low and to reduce the environmental impact of the combustion of fossil fuels.


Global direct primary energy consumption

Direct primary energy consumption does not take account of inefficiencies in fossil fuel production.

Source: Energy Institute Statistical Review of World Energy (2023); Vaclav Smil (2017)
OurWorldInData.org/energy • CC BY

Figure 1: How much energy does the world consume? (Source: Our World is Data, CC BY 4.0)

3.2: Representing Valence Electrons with Dots

Learning Objective

- Draw a Lewis electron dot diagram for an atom for s and p block elements in A group.

In almost all cases, chemical bonds are formed by interactions of valence electrons in atoms. To facilitate our understanding of how valence electrons interact, a simple way of representing those valence electrons would be useful.

A **Lewis electron dot diagram** (or electron dot diagram) is a representation of the valence electrons of an atom that uses dots around the symbol of the element. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side. (The order in which the positions are used does not matter.) For example, the Lewis electron dot diagram for hydrogen is simply

$$\mathbf{H}\cdot \tag{3.2.1}$$

Because the side is not important, the Lewis electron dot diagram could also be drawn as follows:

$$\dot{\mathbf{H}} \;\; or \;\; \cdot\mathbf{H} \;\; or \;\; \underset{\cdot}{\mathbf{H}} \tag{3.2.2}$$

The electron dot diagram for helium, with two valence electrons, is as follows:

$$\mathbf{He}\!: \tag{3.2.3}$$

By putting the two electrons together on the same side, we emphasize the fact that these two electrons are both in the 1*s* subshell; this is the common convention we will adopt, although there will be exceptions later. The next atom, lithium, has an electron configuration of $1s^2 2s^1$, so it has only one electron in its valence shell. Its electron dot diagram resembles that of hydrogen, except the symbol for lithium is used:

$$\mathbf{Li}\cdot \tag{3.2.4}$$

Beryllium has two valence electrons in its 2*s* shell, so its electron dot diagram is like that of helium:

$$\mathbf{Be}\!: \tag{3.2.5}$$

The next atom is boron. Its valence electron shell is $2s^2 2p^1$, so it has three valence electrons. The third electron will go on another side of the symbol:

$$\dot{\mathbf{B}}\!: \tag{3.2.6}$$

Again, it does not matter on which sides of the symbol the electron dots are positioned.

For carbon, there are four valence electrons, two in the 2*s* subshell and two in the 2*p* subshell. As usual, we will draw two dots together on one side, to represent the 2*s* electrons. However, conventionally, we draw the dots for the two *p* electrons on different sides. As such, the electron dot diagram for carbon is as follows:

$$\cdot\dot{\mathbf{C}}\!: \tag{3.2.7}$$

With N, which has three *p* electrons, we put a single dot on each of the three remaining sides:

$$\cdot\underset{\cdot}{\dot{\mathbf{N}}}\!: \tag{3.2.8}$$

For oxygen, which has four *p* electrons, we now have to start doubling up on the dots on one other side of the symbol. When doubling up electrons, make sure that each side has no more than two electrons.

$$\cdot\underset{\cdot}{\ddot{\mathbf{O}}}\!: \tag{3.2.9}$$

Fluorine and neon have seven and eight dots, respectively:

$$:\underset{\cdot}{\ddot{\mathbf{F}}}\!: \tag{3.2.10}$$

$$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Ne}}}: \tag{3.2.11}$$

With the next element, sodium, the process starts over with a single electron because sodium has a single electron in its highest-numbered shell, the n = 3 shell. By going through the periodic table, we see that the Lewis electron dot diagrams of atoms will never have more than eight dots around the atomic symbol.

Example 3.2.1: Lewis Dot Diagrams

What is the Lewis electron dot diagram for each element?

a. aluminum
b. selenium

Solution

a. The valence electron configuration for aluminum is $3s^2 3p^1$. So it would have three dots around the symbol for aluminum, two of them paired to represent the 3s electrons:

$$\dot{Al}\,:$$

2. The valence electron configuration for selenium is $4s^2 4p^4$. In the highest-numbered shell, the n = 4 shell, there are six electrons. Its electron dot diagram is as follows:

$$\cdot\overset{\cdot}{\underset{\cdot\cdot}{\mathbf{Se}}}:$$

Exercise 3.2.1

What is the Lewis electron dot diagram for each element?

a. phosphorus
b. argon

Answer a

$$\cdot\overset{\cdot}{\underset{\cdot}{\mathbf{P}}}:$$

Answer b

$$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Ar}}}:$$

Summary

- Lewis electron dot diagrams use dots to represent valence electrons around an atomic symbol.

Contributions & Attributions

This page was constructed from content via the following contributor(s) and edited (topically or extensively) by the LibreTexts development team to meet platform style, presentation, and quality:

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3.3: Covalent Bonds

Learning Objectives

- Define *covalent bond*.
- Illustrate covalent bond formation with Lewis electron dot diagrams.

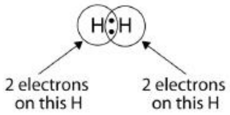
When valence electrons are shared between two atoms, they make a bond called a **covalent bond**. The driving force behind bond formation is stability - a lower potential energy state. There is a continuum of bonding between covalent bonds and ionic bonds (discussed in detail in Section 4.4) Let us illustrate a covalent bond by using H atoms, with the understanding that H atoms need only two electrons to fill the 1*s* subshell. Each H atom starts with a single electron in its valence shell:

$$\mathbf{H}\cdot \qquad \cdot\mathbf{H} \tag{3.3.1}$$

The two H atoms can share their electrons:

$$\mathbf{H:H} \tag{3.3.2}$$

We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom's valence shell:



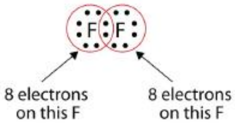
Because each H atom has a filled valence shell, this bond is stable, and we have made a diatomic hydrogen molecule. (This explains why hydrogen is one of the diatomic elements.) The shared pair of electrons is represented as a dash between the two atoms. It is possible for the same two atoms to share two pairs of electrons forming a double covalent bond or three pairs of electrons forming a triple covalent bond.

H–H

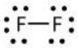
Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:

These two atoms can do the same thing that the H atoms did; they share their unpaired electrons to make a covalent bond.

Note that each F atom has a complete octet around it now:

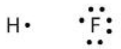


We can also write this using a dash to represent the shared electron pair:

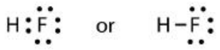


There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F atom has one bonding pair and three lone pairs of electrons.

Covalent bonds can be made between different elements as well. One example is HF. Each atom starts out with an odd number of electrons in its valence shell:



The two atoms can share their unpaired electrons to make a covalent bond:



We note that the H atom has a full valence shell with two electrons, while the F atom has a complete octet of electrons.

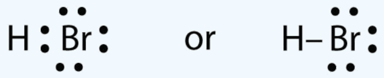
Example 3.3.1:

Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

Solution

HBr is very similar to HF, except that it has Br instead of F. The atoms are as follows:

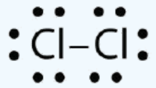
The two atoms can share their unpaired electron:



Exercise 3.3.1

Use Lewis electron dot diagrams to illustrate the covalent bond formation in Cl_2.

Answer



When working with covalent structures, it sometimes looks like you have leftover electrons. You apply the rules you learned so far and there are still some electrons hanging out there unattached. You can't just leave them there. So where do you put them?

Multiple Covalent Bonds

Some molecules are not able to satisfy the octet rule by making only single covalent bonds between the atoms. Consider the compound ethene, which has a molecular formula of C_2H_4. The carbon atoms are bonded together, with each carbon also being bonded to two hydrogen atoms.

$$\text{two C atoms} = 2 \times 4 = 8 \text{ valence electrons}$$

$$\text{four H atoms} = 4 \times 1 = 4 \text{ valence electrons}$$

$$\text{total of 12 valence electrons in the molecule}$$

If the Lewis electron dot structure was drawn with a single bond between the carbon atoms and with the octet rule followed, it would look like this:

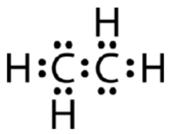


Figure 3.3.1: Incorrect dot structure of ethene.

This Lewis structure is incorrect because it contains a total of 14 electrons. However, the Lewis structure can be changed by eliminating the lone pairs on the carbon atoms and having to share two pairs instead of only one pair.

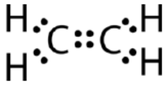


Figure 3.3.2: Correct dot structure for ethene.

A **double covalent bond** is a covalent bond formed by atoms that share two pairs of electrons. The double covalent bond that occurs between the two carbon atoms in ethane can also be represented by a structural formula and with a molecular model as shown in the figure below.

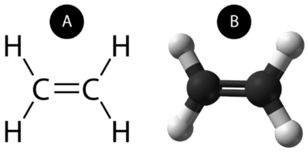


Figure 3.3.3: (A) The structural model for C_2H_4 consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of C_2H_4.

A **triple covalent bond** is a covalent bond formed by atoms that share three pairs of electrons. The element nitrogen is a gas that composes the majority of Earth's atmosphere. A nitrogen atom has five valence electrons, which can be shown as one pair and three single electrons. When combining with another nitrogen atom to form a diatomic molecule, the three single electrons on each atom combine to form three shared pairs of electrons.

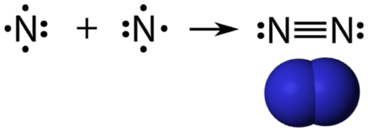


Figure 3.3.4: Triple bond in N_2.

Each nitrogen atom follows the octet rule with one lone pair of electrons and six electrons that are shared between the atoms.

Summary

- Covalent bonds are formed when atoms share electrons.
- Lewis electron dot diagrams can be drawn to illustrate covalent bond formation.
- Double bonds or triple bonds between atoms may be necessary to properly illustrate the bonding in some molecules.

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- Anonymous by request
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- Henry Agnew (UC Davis)

3.4: Drawing Lewis Structures for Covalent Compounds

Learning Objectives

- Draw Lewis structures for covalent compounds.

Lewis structures are representations of molecules that include not only what atoms are present in the molecule but also how the atoms are connected. The following procedure can be used to draw Lewis structure for simple molecules.

How-to: Constructing Lewis electron structures

1. Determine the total number of valence electrons in the molecule or ion.

- Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.)

2. Arrange the atoms to show specific bonds.

- When there is a central atom, it is usually the least electronegative element in the compound. **Electronegativity** is a measure of the relative attraction for the pair of electrons in a bond. See section 4.2 for more details and a table of electronegativity values. Chemists usually list this central atom first in the chemical formula (as in CCl_4, which both have C as the central atom), which is another clue to the compound’s structure.
- Hydrogen and the halogens (Group 7A, F, Cl, Br, I) are almost always bonded to only one other atom, so they are usually *terminal* rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O, for example, there is a bonding pair of electrons (2 electrons) between oxygen and each hydrogen.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

- These electrons will usually be lone pairs. For example with Cl, it will have three lone pairs, and one bond if the molecule CCl_4

5. If any electrons are left over, place them on the central atom as a lone pair.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and that each atom has an octet of electrons, except for hydrogen (with two electrons).
- The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

Now let’s apply this procedure to some particular compounds, beginning with one we have already discussed.

Example 3.4.1: Water

Write the Lewis Structure for H_2O.

Solution

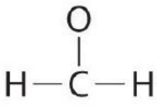
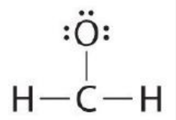
Steps for Writing Lewis Structures	Example 3.4.1
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of <u>8 valence electrons.</u>

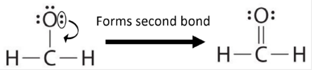
Steps for Writing Lewis Structures	Example 3.4.1
2. Arrange the atoms to show specific connections.	H O H Because H atoms are almost always terminal, the arrangement within the molecule must be HOH.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Placing one bonding pair of electrons between the O atom and each H atom gives H -O- H with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary.
7. Final check.	The Lewis structure gives oxygen an octet and each hydrogen 2 electrons.

Example 3.4.2

Write the Lewis structure for the CH_2O molecule

Solution

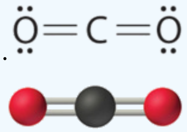
Steps for Writing Lewis Structures	Example 3.4.2
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has 1 valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = 12 valence electrons.
2. Arrange the atoms to show specific connections.	 Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.	Placing a bonding pair of electrons between each pair of bonded atoms gives the following:  6 electrons are used, and 6 are left over.
4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:  Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.

Steps for Writing Lewis Structures	Example 3.4.2
5. If any electrons are left over, place them on the central atom.	Not necessary. There are no electrons left to place on the central atom.
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon–oxygen double bond: 
7. Final check	Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

Exercise 3.4.1

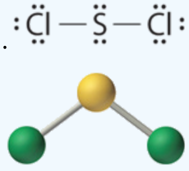
Write Lewis electron structures for CO_2 and SCl_2, a vile-smelling, unstable red liquid that is used in the manufacture of rubber.

Answer CO_2



Carbon dioxide

Answer SCl_2

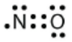


Sulfur dichloride

Exceptions to the Octet Rule

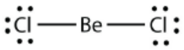
As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations. Recognize that exceptions exist.

There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO_2, and ClO_2. The Lewis electron dot diagram for NO is as follows:

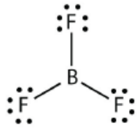


Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

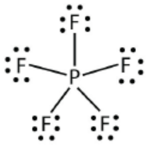
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:



Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3:



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty *d* orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell. Expanded octets are beyond the scope of this course.

Example 3.4.3: Octet Violations

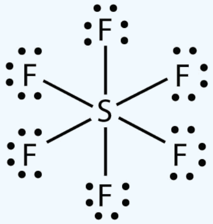
Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

a. ClO
b. SF_6

Solution

a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:

b. In SF_6, the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



Exercise 3.4.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

Answer

The Xe atom has an expanded valence shell with more than eight electrons around it.

Summary

Lewis dot symbols provide a simple rationalization of why elements form compounds with a specific number of bonds. In Lewis electron structures, we encounter **bonding pairs**, which are shared by two atoms, and **lone pairs**, which are not shared between atoms. There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules.

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3.5: Introduction to Organic Molecules

Learning Objectives

- Recognize the characteristics of the element: *carbon*
- Identify organic molecules as alkanes, alkenes, alkynes, alcohols, or carboxylic acids.
- Recognize the common types of functional groups in an organic molecule.

When methane was mentioned previously, we described it as the simplest organic compound. In this section, we introduce organic chemistry more formally. Organic chemistry is the study of the chemistry of carbon compounds. Carbon is singled out because it has a chemical diversity unrivaled by any other chemical element. Its diversity is based on the following:

- Carbon atoms bond reasonably strongly with other carbon atoms.
- Carbon atoms bond reasonably strongly with atoms of other elements.

- Curiously, elemental carbon is not particularly abundant. It does not even appear in the list of the most common elements in Earth's crust. Nevertheless, all living things consist of organic compounds. Most organic chemicals are covalent compounds, which is why we introduce organic chemistry here. By convention, compounds containing carbonate ions and bicarbonate ions, as well as carbon dioxide and carbon monoxide, are not considered part of organic chemistry, even though they contain carbon.

Alkanes

The simplest organic compounds are the **hydrocarbons**, compounds composed of carbon and hydrogen atoms only. Some hydrocarbons have only single bonds and appear as a chain (which can be a straight chain or can have branches) of carbon atoms also bonded to hydrogen atoms. These hydrocarbons are called **alkanes (saturated hydrocarbons**). Each alkane has a characteristic, systematic name depending on the number of carbon atoms in the molecule. These names consist of a stem that indicates the number of carbon atoms in the chain plus the ending -ane. The stem meth- means one carbon atom, so methane is an alkane with one carbon atom. Similarly, the stem eth- means two carbon atoms; ethane is an alkane with two carbon atoms. Continuing, the stem prop- means three carbon atoms, so propane is an alkane with three carbon atoms. The stem but- means four carbon atoms; butane is an alkane with four carbon atoms. Figure 3.5.1 gives the Lewis structures, condensed structural formulas and molecular formulas of the four simplest alkanes. In the **condensed structural formula**, the covalent bonds are understood to exist between each carbon and the hydrogens associated with it, as well as between carbon atoms.

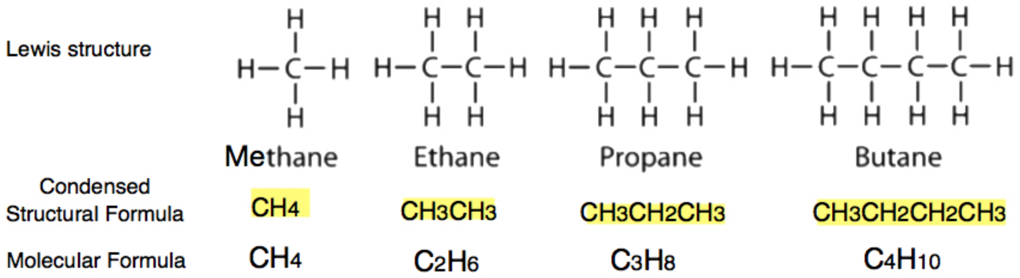


Figure 3.5.1: Formulas and Molecular Models of the Four Simplest Alkanes. The four smallest alkanes are methane, ethane, propane and butane.

Alkenes

Some hydrocarbons have one or more carbon–carbon double bonds (denoted C=C). These hydrocarbons are called alkenes. Figure 3.5.2 shows the formulas and the molecular models of the two simplest alkenes. Note that the names of alkenes have the same stem as the alkane with the same number of carbon atoms in its chain but have the ending *-ene*. Thus, ethene is an alkene with two carbon atoms per molecule, and propene is a compound with three carbon atoms and one double bond.

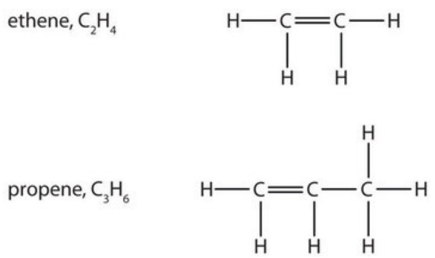


Figure 3.5.2: Formulas and Molecular Models of the Two Simplest Alkenes. Ethene is commonly called ethylene, while propene is commonly called propylene.

Alkynes

Alkynes are hydrocarbons with a carbon–carbon triple bond (denoted C≡C) as part of their carbon skeleton. Below is the formula and the molecular model of the simplest alkyne and its systematic name. Its common name is acetylene. Its chemical formula is **C_2H_2**.

$$H-C\equiv C-H$$

Ethyne

The names for alkynes have the same stems as for alkanes but with the ending *-yne*.

To Your Health: Saturated and Unsaturated Fats

Hydrocarbons are not the only compounds that can have carbon–carbon double bonds. A group of compounds called fats can have them as well, and their presence or absence in the human diet is becoming increasingly correlated with health issues.

Fats are combinations of long-chain organic compounds (fatty acids) and glycerol ($C_3H_8O_3$). The long carbon chains can have either all single bonds, in which case the fat is classified as *saturated*, or one or more double bonds, in which case it is a *monounsaturated* or a *polyunsaturated* fat, respectively. Saturated fats are typically solids at room temperature; beef fat (tallow) is one example. Mono- or polyunsaturated fats are likely to be liquids at room temperature and are often called oils. Olive oil, flaxseed oil, and many fish oils are mono- or polyunsaturated fats.

Studies have linked higher amounts of saturated fats in people's diets with a greater likelihood of developing heart disease, high cholesterol, and other diet-related diseases. In contrast, increases in unsaturated fats (either mono- or polyunsaturated) have been linked to a lower incidence of certain diseases. Thus, there have been an increasing number of recommendations by government bodies and health associations to decrease the proportion of saturated fat and increase the proportion of unsaturated fat in the diet. Most of these organizations also recommend decreasing the total amount of fat in the diet.

Recently, certain fats called trans fats have been implicated in the presence of heart disease. These are fats from animal sources and are also produced when liquid oils are exposed to partial hydrogenation, an industrial process that increases their saturation. Trans fats are used in many prepared and fried foods. Because they bring with them the health risks that naturally occurring saturated fats do, there has been some effort to better quantify the presence of trans fats in food products. US law now requires that food labels list the amount of trans fat in each serving.

Functional Groups

The carbon–carbon double and triple bonds are examples of functional groups in organic chemistry. A functional group is a specific structural arrangement of atoms or bonds that imparts a characteristic chemical reactivity to a molecule. Alkanes have no functional group. A carbon–carbon double bond is considered a functional group because carbon–carbon double bonds chemically react in specific ways that differ from reactions of alkanes (for example, under certain circumstances, alkenes react with water); a carbon–carbon triple bond also undergoes certain specific chemical reactions. We first introduced the idea of the functional group, a specific structural arrangement of atoms or bonds that imparts a characteristic chemical reactivity to the molecule. If you understand the behavior of a particular functional group, you will know a great deal about the general properties of that class of compounds. Each family is based on a common, simple functional group that contains an oxygen atom or a nitrogen atom. Some common functional groups are listed in Table 3.5.1.

Table 3.5.1: Selected Organic Functional Groups

Name of Family	General Formula	Functional Group	Suffix*
alkane	RH	none	*-ane*
alkene	$R_2C=CR_2$	C=C	*-ene*
alkyne	RC≡CR	–C≡C–	*-yne*
alcohol	ROH	–OH	*-ol*
thiol	RSH	–SH	*-thiol*
ether	ROR	–O–	ether
aldehyde	RC(O)H	-C(O)H	*-al*
ketone	RC(O)R	-C(O)-	*-one*
carboxylic acid	RC(O)OH	-C(O)OH	*-oic acid*

*Ethers do not have a suffix in their common name; all ethers end with the word *ether*.
(O) indicates the oxygen is double bonded to the carbon

A common functional group is an alcohol. If an OH group (also called a hydroxyl group) is substituted for a hydrogen atom in a hydrocarbon molecule, the compound is an alcohol. Alcohols are named using the parent hydrocarbon name but with the final *-e* dropped and the suffix *-ol* attached. The two simplest alcohols are methanol and ethanol. Figure 3.5.4 shows their formulas along with a molecular model of each.

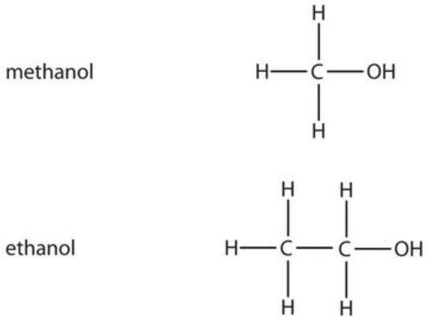


Figure 3.5.4: The Two Simplest Organic Alcohol Compounds. Alcohols have an OH functional group in the molecule.

Alcohol

Ethanol (also called ethyl alcohol) is the alcohol in alcoholic beverages. Other alcohols include methanol (or methyl alcohol), which is used as a solvent and a cleaner, and isopropyl alcohol (or rubbing alcohol), which is used as a medicinal disinfectant. Neither methanol nor isopropyl alcohol should be ingested, as they are toxic even in small quantities.

Another important family of organic compounds has a carboxyl group, in which a carbon atom is double-bonded to an oxygen atom and to an OH group. Compounds with a carboxyl functional group are called carboxylic acids, and their names end in *-oic acid*. Figure 3.5.5 shows the formulas and the molecular models of the two simplest carboxylic acids, perhaps best known by the common names formic acid and acetic acid. The carboxyl group is sometimes written in molecules as COOH.

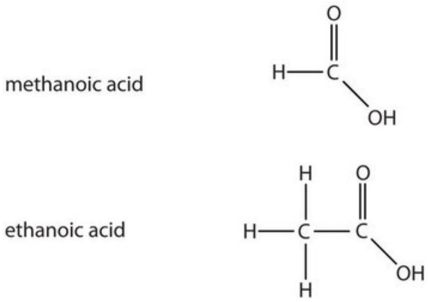


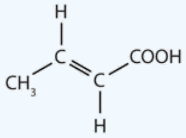
Figure 3.5.5: The Two Smallest Organic Acids. The two smallest carboxylic acids are formic acid (found in the stingers of ants) and acetic acid (found in vinegar).

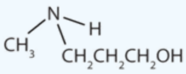
The condensed structures of methanoic acid and ethanoic acid are **HCOOH** and **CH_3COOH**, respectively.

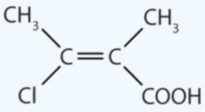
Many organic compounds are considerably more complex than the examples described here.

Example 3.5.1

Identify the functional group(s) in each molecule as a double bond, a triple bond, an alcohol, or a carboxyl.

a. 

b. 

c. 

d. $CH_3CH_2CH_2CH_2OH$

Answer a

This molecule has a double bond and a carboxyl functional group.

Answer b

This molecule has an alcohol functional group.

Answer c

This molecule has a carbon-carbon double bond and a carboxyl functional group.

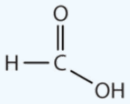
Answer d

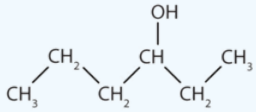
This molecule has an alcohol functional group.

Exercise 3.5.1

Identify the functional group(s) in each molecule as a double bond, a triple bond, an alcohol, or a carboxyl.

a. $CH_3C\equiv CCH_3$

b. 

c. 

d. $CH_2{=}CH{-}COOH$

Answer a:

triple bond (alkyne)

Answer b:

carboxyl group

Answer c:

alcohol group

Answer d:

double bond (alkene) and carboxyl group

Career Focus: Forensic Chemist

The main job of a forensic chemist is to identify unknown materials and their origins. Although forensic chemists are most closely associated in the public mind with crime labs, they are employed in pursuits as diverse as tracing evolutionary patterns in living organisms, identifying environmental contaminants, and determining the origin of manufactured chemicals.

In a crime lab, the forensic chemist has the job of identifying the evidence so that a crime can be solved. The unknown samples may consist of almost anything—for example, paint chips, blood, glass, cloth fibers, drugs, or human remains. The forensic chemist subjects them to a variety of chemical and instrumental tests to discover what the samples are. Sometimes these samples are extremely small, but sophisticated forensic labs have state-of-the-art equipment capable of identifying the smallest amount of unknown sample.

Another aspect of a forensic chemist's job is testifying in court. Judges and juries need to be informed about the results of forensic analyses, and it is the forensic chemist's job to explain those results. Good public-speaking skills, along with a broad background in chemistry, are necessary to be a successful forensic chemist.

Concept Review Exercises

1. What is organic chemistry?
2. What is a functional group? Give at least two examples of functional groups.

Answers

1. Organic chemistry is the study of the chemistry of carbon compounds.
2. A functional group is a specific structural arrangement of atoms or bonds that imparts a characteristic chemical reactivity to the molecule; alcohol group and carboxylic group (answers will vary).

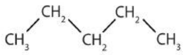
Key Takeaways

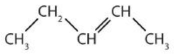
- Organic chemistry is the study of the chemistry of carbon compounds.
- Organic molecules can be classified according to the types of elements and bonds in the molecules.

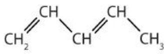
Exercises

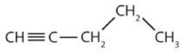
1. Give three reasons why carbon is the central element in organic chemistry.
2. Are organic compounds based more on ionic bonding or covalent bonding? Explain.

3. Identify the type of hydrocarbon in each structure.

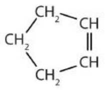
a. 

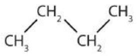
b. 

c. 

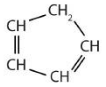
d. 

4. Identify the type of hydrocarbon in each structure.

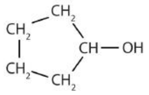
a. 

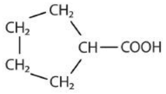
b. 

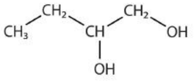
c. $CH\equiv C-CH_2-CH_3$

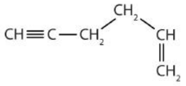
d. 

5. Identify the functional group(s) in each molecule.

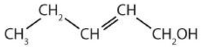
a. 

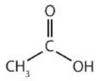
b. 

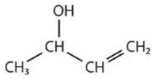
c. 

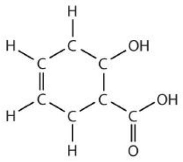
d. 

6. Identify the functional group(s) in each molecule.

a. 

b. 

c. 

d. 

7. How many functional groups described in this section contain carbon and hydrogen atoms only? Name them.
8. What is the difference in the ways the two oxygen atoms in the carboxyl group are bonded to the carbon atom?

Answers

1. Carbon atoms bond reasonably strongly with other carbon atoms. Carbon atoms bond reasonably strongly with atoms of other elements. Carbon atoms make a large number of covalent bonds (four).

2. Organic compounds are based on covalent bonding or electron sharing. The atoms C, H, O, N that make up organic compounds are all nonmetals.

3.

a. alkane
b. alkene
c. alkene
d. alkyne

4.

a. alkene
b. alkane
c. alkyne
d. alkene
a. alcohol
b. carboxyl
c. alcohol
d. alkene and alkyne

6. a. a carbon-carbon double bond and alcohol b. carboxyl group c. carbon-carbon double bond and alcohol d. carbon-carbon double bond; alcohol and carboxyl group

7. two; carbon-carbon double bonds and carbon-carbon triple bonds

8. There are two oxygen atoms in a carboxyl group: one is double-bonded while the other is OH, single bonded to the same carbon atom.

3.6: Resonance

Learning Objectives

- Define resonance
- Identify whether a molecule has resonance

Resonance

There are some cases in which more than one viable Lewis structure can be drawn for a molecule. An example is the ozone, (O_3). Ozone a gas present in different regions of the atmosphere. When ozone is present in the troposphere it is an air pollutant that is harmful to living species and a greenhouse gas that contributes to climate change. Ozone in the stratosphere helps protects the Earth from overexposure to UV radiation. There are a total of 18 valence electrons in the structure, 6 valence electrons from each oxygen atom, so the following two structures are possible in Figure $3.6.1$.

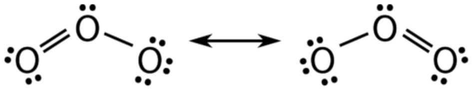


Figure $3.6.1$: Resonance forms of ozone. Note the use of the double-headed arrow.

The structure on the left ($3.6.1$) can be converted to the structure on the right by a shifting of electrons without altering the positions of the atoms.

It was once thought that the structure of a molecule such as O_3 consisted of one single bond and one double bond which then shifted back and forth as shown above. However, further studies showed that the two bonds are identical. Any double covalent bond between two given atoms is typically shorter than a single covalent bond. Studies of the O_3 and other similar molecules showed that the bonds were identical in length. Interestingly, the length of the bond is in between the lengths expected for an $O-O$ single bond and a double bond.

Resonance is the use of two or more valid (or contributing) Lewis structures to represent the covalent bonding in a molecule. The valid structures are referred to as resonance structures. It is now understood that the actual structure of a molecule which displays resonance is that of an average or a hybrid of all the valid resonance structures. In the case of the O_3 molecule, each of the covalent bonds between O atoms is best thought of as being "one and a half" bonds, as opposed to either a pure single bond or a pure double bond. This "half-bond" can be shown as a dotted line in both the Lewis structure and the molecular model (Figure $3.6.2$).

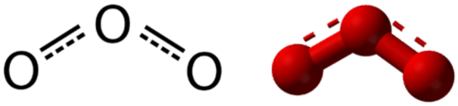


Figure $3.6.2$: "Half-bond" model of ozone molecule. This is a better description of the electornic structure of ozone than either of the resonance structures in Figure $3.6.1$.

Summary

- It is possible to draw resonance structures for molecules that are not viable structures. Additional guidelines for assessing the validity of a Lewis structure beyond the octet rule (or duet for hydrogen) are beyond the scope of this course.
- The resonance structures only differ in the placement of electrons – the connectivity of the atoms in the molecule remains the same between resonance structures.

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- Henry Agnew (UC Davis)

3.7: Air Pollutants

At the end of 2019, the coronavirus began spreading around the world, infecting humans and causing pneumonia known as COVID-19. The number of infections increased rapidly and in some places, healthcare systems were overrun. In an attempt to contain the spread, many countries limited transportation and closed down schools, colleges, universities and nonessential businesses. The negative impacts of the novel coronavirus pandemic are widespread and significant. However, one of the unintended consequences of the efforts to contain the spread of the disease was reduced air pollution. Many air pollutants are molecules produced from the combustion of fuels. NASA monitoring instruments measured decreased concentrations of NO_2 and CO_2 which resulted in a positive health benefit. (1,2)

According to the US EPA and WHO, long-term exposure to ambient air pollution increases mortality and morbidity from cardiovascular and respiratory disease and lung cancer and decreases life expectancy. (3) The two pollutants responsible for most of the disease burden are fine particulate matter ($PM_{2.5}$) and ground level ozone (O_3).

In response to poor air quality in the United States, the government established the Clean Air Act (CAA) in 1970. The law was amended in 1977 and again in 1990. This comprehensive federal law regulates air emissions from stationary sources such as factories, refineries and power plants and mobile sources such as cars, trucks and buses. CAA also authorizes the Environmental Protection Agency (EPA) to establish National Ambient Air Quality Standards (NAAQS). The goal of these standards is to protect the health and welfare of the public and regulate the emissions of hazardous air pollutants. Hazardous air pollutants are known to cause cancer and other serious health impacts. The NAAQS have been set for six common pollutants, known as the criteria air pollutants. These include ground level ozone, particulate matter carbon monoxide, lead, sulfur dioxide, and nitrogen dioxide. Four of the six criteria pollutants exist as gaseous molecules in the atmosphere. The EPA publishes the standards on their website.

These six common pollutants are not the only substances that we need to worry about in the air. Soot from fires is problematic and even hazardous pollutants from industry are in our air. An example of this was ethylene oxide was being emitted from sterilization factories in Lake County IL in 2017. Ethylene oxide is cancer causing and people in the surrounding area were inhaling it. The Illinois EPA intervened and both companies were required to reduce their emissions to protest the public health. For more information visit the Lake County Health Department. Air pollution is problematic when we breathe in these compounds, they can cause all sorts of health problems from cancer, cardiovascular diseases to asthma. You can find out more here: https://www.cdc.gov/air/pollutants.htm.

The EPA has also established the U.S. Air Quality Index (AQI) as a simple method for reporting air quality to the general public. Lead is the only criteria air pollutant not included in the AQI. The AQI ranges from 0-500. The lower the value the better the air quality. Additionally, the AQI is divided into six color-coded categories. Green (0-50) represents good air quality that poses little to no risk even to sensitive groups. Red (151-200) is considered unhealthy and some members of the public may experience health effects. Values above 300 are considered hazardous with everyone more likely to be affected. (4)

The emission of hazardous air pollutants is not limited to human-made sources nor are people only exposed outside. Natural sources include volcanic eruptions and forest fires which are of increasing concern due their greater frequency and duration. Building materials and cleaning solvents are two fairly common sources of indoor pollutants.

(1) NASA Earth Observatory. Airborne Nitrogen Dioxide Plummets Over China https://earthobservatory.nasa.gov/images/146362/airborne-nitrogen-dioxide-plummets-over-china (accessed 19 July 2020)

(2) Evans, S. Analysis: Coronavirus set to cause largest ever annual fall to CO2 emissions. https://www.carbonbrief.org/analysis-coronavirus-set-to-cause-largest-ever-annual-fall-in-co2-emissions (accessed 19 July 2020).

(3) Cohen, Aaron J et al. "Estimates and 25-year trends of the global burden of disease attributable to ambient air pollution: an analysis of data from the Global Burden of Diseases Study 2015." *Lancet (London, England)* vol. 389,10082 (2017): 1907-1918. doi:10.1016/S0140-6736(17)30505-6.

(4) AQI basics. https://www.airnow.gov/aqi/aqi-basics/ (last accessed 19 July 2020).

3.8: Naming Molecular Compounds

Learning Objectives

- Determine the name of a simple molecular compound from its chemical formula.

Molecular Compounds

Molecular compounds are inorganic compounds that take the form of discrete **molecules**. Examples include such familiar substances as water (H_2O) and carbon dioxide (CO_2). These compounds are very different from ionic compounds like sodium chloride $(NaCl)$. Ionic compounds are formed when metal atoms lose one or more of their electrons to nonmetal atoms. The resulting cations and anions are electrostatically attracted to each other.

So what holds the atoms of a molecule together? Rather than forming ions, the atoms of a molecule share their **electrons** in such a way that a **bond** forms between a pair of atoms. In a carbon dioxide molecule, there are two of these bonds, each occurring between the carbon atom and one of the two oxygen atoms.

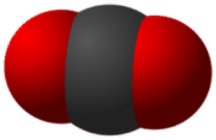


Figure 3.8.1: Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.

Larger molecules can have many, many bonds that serve to keep the molecule together. In a large sample of a given molecular compound, all of the individual molecules are identical.

Naming Binary Molecular Compounds

Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is H_2O. A molecule of octane, which is a component of gasoline, contains 8 atoms of carbon and 18 atoms of hydrogen. The molecular formula of octane is C_8H_{18}.

Figure 3.8.2: Nitrogen dioxide (NO_2) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines.

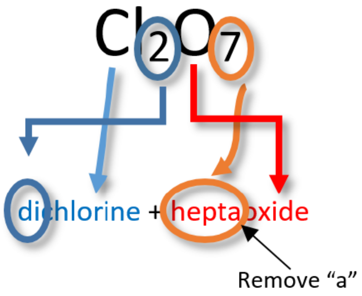
Naming *binary* (two-element) molecular compounds is similar to naming simple ionic compounds. The first element in the formula is simply listed using the name of the element. The second element is named by taking the stem of the element name and adding the suffix *-ide*. A system of numerical prefixes is used to specify the number of atoms in a molecule. Table 3.8.1 lists these numerical prefixes.

Table 3.8.1: Numerical Prefixes for Naming Binary Covalent Compounds

Number of Atoms in Compound	Prefix on the Name of the Element
1	mono-*
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-

Number of Atoms in Compound	Prefix on the Name of the Element
7	hepta-
8	octa-
9	nona-
10	deca-

*This prefix is not used for the first element’s name.



dichlorine heptoxide

Note

- Generally, the less electronegative element is written first in the formula, though there are a few exceptions. Carbon is always first in a formula and hydrogen is after nitrogen in a formula such as NH_3. The order of common nonmetals in binary compound formulas is C, P, N, H, S, I, Br, Cl, O, F.
- The *a* or *o* at the end of a prefix is usually dropped from the name when the name of the element begins with a vowel. As an example, four oxygen atoms, is tetr**o**xide instead of tetr**ao**xide.
- The prefix is "mono" is not added to the first element’s name if there is only one atom of the first element in a molecule.

Some examples of molecular compounds are listed in Table 3.8.2.

Table 3.8.2

Formula	Name
NO	nitrogen monoxide
N_2O	dinitrogen monoxide
S_2Cl_2	disulfur dichloride
Cl_2O_7	dichlorine heptoxide

Notice that the *mono*- prefix is not used with the nitrogen in the first compound, but is used with the oxygen in both of the first two examples. The S_2Cl_2 emphasizes that the formulas for molecular compounds are not reduced to their lowest ratios. The *o* of the *mono*- and the *a* of *hepta*- are dropped from the name when paired with oxide.

Exercise 3.8.1

Write the name for each compound.

a. CF_4
b. $SeCl_2$
c. SO_3

Answer a:

carbon tetrafluoride

Answer b:

selenium dichloride

Answer c:

sulfur trioxide

Simple molecular compounds with common names

For some simple covalent compounds, we use common names rather than systematic names. We have already encountered these compounds, but we list them here explicitly:

- H_2O: water
- NH_3: ammonia
- CH_4: methane
- H_2O_2: hydrogen peroxide

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system.

Some Compounds Have Both Covalent and Ionic Bonds

If you recall the introduction of polyatomic ions, you will remember that the bonds that hold the polyatomic ions together are covalent bonds. Once the polyatomic ion is constructed with covalent bonds, it reacts with other substances as an ion. The bond between a polyatomic ion and another ion will be ionic. An example of this type of situation is in the compound sodium nitrate. Sodium nitrate is composed of a sodium ion and a nitrate ion. The nitrate ion is held together by covalent bonds and the nitrate ion is attached to the sodium ion by an ionic bond.

Summary

- A molecular compound is usually composed of two or more nonmetal elements.
- Molecular compounds are named with the first element first and then the second element by using the stem of the element name plus the suffix -ide. Numerical prefixes are used to specify the number of atoms in a molecule.

3.9: Chemical Equations

Learning Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, (s), (l), (g), (aq), and $\rightarrow$ appropriately when writing a chemical reaction.

The combustion of fossil fuels results in the formation of carbon dioxide, a greenhouse gas, and other pollutants. This is an example of chemical change, in which new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

Reactants and Products

To describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $\mathrm{SO_2 + O_2}$, are reactants and sulfur trioxide, $\mathrm{SO_3}$, is the product.

$$\underbrace{2\,\mathrm{SO_2(g)} + \mathrm{O_2(g)}}_{\text{Reactants}} \rightarrow \underbrace{2\,\mathrm{SO_3(g)}}_{\text{Products}}$$

In chemical reactions, the reactants are found before the symbol "$\rightarrow$" and the products are found after the symbol "$\rightarrow$". The general equation for a reaction is:

$$\text{Reactants } \rightarrow \text{Products}$$

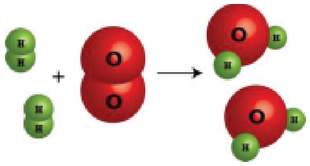
There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. Table $3.9.1$ shows a listing of symbols used in chemical equations.

Table $3.9.1$: Symbols Used in Chemical Equations

Symbol	Description	Symbol	Description
$+$	used to separate multiple reactants or products	(s)	reactant or product in the solid state
$\rightarrow$	yield sign; separates reactants from products	(l)	reactant or product in the liquid state
$\rightleftharpoons$	replaces the yield sign for reversible reactions that reach equilibrium	(g)	reactant or product in the gas state
$\xrightarrow{\text{Pt}}$	formula written above the arrow is used as a catalyst in the reaction	(aq)	reactant or product in an aqueous solution (dissolved in water)
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated		

Chemists have a choice of methods for describing a chemical reaction.

1. You can draw a picture of the chemical reaction showing the submicroscopic particles.



2. You can write a word equation for the chemical reaction:
"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."

3. You can write the equation with chemical symbols

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(g) \tag{3.9.1}$$

In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products, while symbols are used to indicate the phase of each substance.

To turn word equations into symbolic equations, we need to follow the given steps:

1. Identify the reactants and products. This will help you know which symbols go on each side of the arrow and where the + signs go.
2. Write the correct formulas for all compounds.
3. Sometimes a reactant or product is only an element (like hydrogen gas or oxygen gas). There are seven elements that are considered diatomic, meaning that they are always found in pairs in nature. They include those elements listed in the table.

Table 3.9.1*: Diatomic Elements*

Element Name	Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine
Formula	H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2

Example 3.9.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. $HCl\,(aq) + NaOH\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l)$
b. Gaseous propane, C_3H_8, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.
c. Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution

a. An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.

b. Reactants: propane (C_3H_8) and oxygen (O_2)

Product: carbon dioxide (CO_2) and water (H_2O)

$$C_3H_8\,(g) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(l)$$

c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

$$HF\,(g) + K_2CO_3\,(aq) \rightarrow KF\,(aq) + H_2O\,(l) + CO_2\,(g)$$

Exercise 3.9.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. Hydrogen gas reacts with nitrogen gas to produce gaseous ammonia.

Answer a

$H_2(g) + N_2(g) \rightarrow NH_3(g)$

Summary

- A chemical reaction is the process by which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.

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- Henry Agnew (UC Davis)

3.10: How to Write Balanced Chemical Equations

Learning Objectives

- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Explain the role of the Law of Conservation of Mass in a chemical reaction.

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. This is known as the **Law of Conservation of Mass**. The same atoms that were present in the reactants are present in the products—they are merely reorganized into different arrangements.

Coefficients and Subscripts

There are two types of numbers that appear in chemical equations. There are subscripts (CO_2) which are part of the chemical formulas of the reactants and products; and there are coefficients, ex. **2**$CO_{2,}$ that are placed in front of the formulas to indicate how many molecules of that substance is used or produced. **2**CO_2 indicates there are two molecules of carbon dioxide.

$$2H_2\,(g) + O_2\,(g) \rightarrow 2H_2O\,(g) \tag{3.10.1}$$

In the formation of water, the coefficients are **2**H_2*, and* **2**H_2O *and the subscripts for example* O_2 *,* H_2 *and* H_2O**.** *and You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties.* H_2O *is not the same as* H_2O_2 The **subscripts** are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The **coefficients** indicate the number of each substance involved in the reaction and may be changed in order to balance the equation.

Chemical reactions are done on the macroscopic scale (we can see them), and we cannot see 2 molecules of carbon dioxide, **2**CO_2 , but we can see a mole of a substance. What does this mean? A mole is a unit for an amount of something. A mole is defined as 6.02 x 10^{23} items. So in this case, in 2 moles of carbon dioxide, **2**CO_2 , there are 1.2 x 10^{24} molecules of carbon dioxide. Remember molecules are submicroscopic. We can detect moles of substance in the lab and even hold it, but we cannot hold 2 molecules.

Balancing a Chemical Equation

The simplest and most generally useful method for balancing chemical equations is "inspection," better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

Steps in Balancing a Chemical Equation

1. Identify the most complex substance.
2. Beginning with that substance, choose an element(s) that appears in only one reactant and one product, if possible. Adjust the coefficients to obtain the same number of atoms of this element(s) on both sides.
3. Balance the remaining atoms, usually ending with the least complex substance and using fractional coefficients if necessary. If a fractional coefficient has been used, multiply both sides of the equation by the denominator to obtain whole numbers for the coefficients.
4. Count the numbers of atoms of each kind on both sides of the equation to be sure that the chemical equation is balanced. This means that you have conserved mass in the chemical reaction.

Example 3.10.1: Combustion of Heptane

Balance the chemical equation for the combustion of Heptane (C_7H_{16}).

$$C_7H_{16}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

Solution

Steps	**Example**
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_7H_{16}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.**	a. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 CO_2 molecules, each of which contains 1 carbon atom, on the right side: $$\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow \underline{7}CO_2(g) + H_2O(g)}$$ - 7 carbon atoms on both reactant and product sides b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which contains 2 hydrogen atoms, on the right side: $$\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow 7\,CO_2(g) + \underline{8}H_2O(g)}$$ - 16 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, O_2, on the reactant side: $$\mathrm{C_7H_{16}(l) + \underline{11}O_2(g) \rightarrow 7\,CO_2(g) + 8\,H_2O(g)}$$ - 22 oxygen atoms on both reactant and product sides
5. **Check your work.**	The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and 22 oxygen atoms on each side. Always check to be sure that a chemical equation is balanced.

Example 3.10.2: Combustion of Isooctane

Combustion of Isooctane ($\mathrm{C_8H_{18}}$)

$$\mathrm{C_8H_{18}(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)}$$

Solution

The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start. The combustion of any hydrocarbon with oxygen produces carbon dioxide and water.

Steps	**Example**
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is $\mathrm{C_8H_{18}}$. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.

Steps	Example
2. **Adjust the coefficients.**	a. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 CO_2 molecules in the products: $C_8H_{18}(l) + O_2(g) \longrightarrow \underline{8}CO_2(g) + H_2O(g)$ • 8 carbon atoms on both reactant and product sides b. 18 hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products: $C_8H_{18}(l) + O_2(g) \longrightarrow 8\,CO_2(g) + \underline{9}H_2O(g)$ • 18 hydrogen atoms on both reactant and product sides
3. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, O_2, but because there are 2 oxygen atoms per O_2 molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms: $C_8H_{18}(l) + \underline{\frac{25}{2}}O_2(g) \rightarrow 8\,CO_2(g) + 9\,H_2O(g)$ • 25 oxygen atoms on both reactant and product sides The equation is now balanced, but we usually write equations with whole number coefficients. We can eliminate the fractional coefficient by multiplying all coefficients on both sides of the chemical equation by **2**: $\underline{2}C_8H_{18}(l) + \underline{25}O_2(g) \longrightarrow \underline{16}CO_2(g) + \underline{18}H_2O(g)$
4. **Check your work.**	The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side. Balancing equations requires some practice on your part as well as some common sense. If you find yourself using very large coefficients or if you have spent several minutes without success, go back and make sure that you have written the formulas of the reactants and products correctly.

Exercise 3.10.1

Is each chemical equation balanced?

a. $2\,Hg(\ell)^{+}O_2(g) \rightarrow Hg_2O_2(s)$
b. $C_2H_4(g) + 2\,O_2(g) \rightarrow 2\,CO_2(g) + 2\,H_2O(g)$

Answer a

yes

Answer b

no

Answer c

yes

Exercise 3.10.2

Balance the following chemical equations.

a. $N_2(g) + O_2(g) \rightarrow NO_2(g)$
b. $C_6H_{14}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Answer a

N_2 (g) + $2O_2$ (g) → $2NO_2$ (g)

Answer b

$2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

Vocabulary

- **Chemical reaction -** The process in which one or more substances are changed into one or more new substances.
- **Reactants -** The starting materials in a reaction.
- **Products -** Materials present at the end of a reaction.
- **Balanced chemical equation -** A chemical equation in which the number of each type of atom is equal on the two sides of the equation.
- **Subscripts -** Part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element.
- **Coefficient -** A small whole number that appears in front of a formula in a balanced chemical equation.

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3.11: The Mole

Learning Objectives

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

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

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

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

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

Example 3.11.1

How many molecules are present in 2.76 mol of H_2O? How many atoms is this?

Solution

The definition of a mole is an equality that can be used to construct a conversion factor. Also, because we know that there are three atoms in each molecule of H_2O, we can also determine the number of atoms in the sample.

$$2.76\ \cancel{mol\ H_2O} \times \frac{6.022 \times 10^{23} molecules\ H_2O}{\cancel{mol\ H_2O}} = 1.66 \times 10^{24} molecules\ H_2O$$

To determine the total number of atoms, we have

$$1.66 \times 10^{24}\ \cancel{molecules\ H_2O} \times \frac{3\ atoms}{1\ molecule} = 4.99 \times 10^{24}\ atoms$$

Exercise 3.11.1

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

Answer

2.78×10^{22} molecules

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

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

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

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

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

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

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

Here are some examples. The mass of a hydrogen atom is 1.0079 u; the mass of 1 mol of hydrogen atoms is 1.0079 g. Elemental hydrogen exists as a diatomic molecule, H_2. One molecule has a mass of 1.0079 + 1.0079 = 2.0158 u, while 1 mol H_2 has a mass of 2.0158 g. A molecule of H_2O has a mass of about 18.01 u; 1 mol H_2O has a mass of 18.01 g. A single unit of NaCl has a mass of 58.45 u; NaCl has a molar mass of 58.45 g. In each of these moles of substances, there are 6.022 × 10^{23} units: 6.022 × 10^{23} atoms of H, 6.022 × 10^{23} molecules of H_2 and H_2O, 6.022 × 10^{23} units of NaCl ions. These relationships give us plenty of opportunities to construct conversion factors for simple calculations.

✓ Example 3.11.2: Sugar

What is the molar mass of sugar ($C_6H_{12}O_6$)?

Solution

To determine the molar mass, we simply add the atomic masses of the atoms in the molecular formula; but express the total in grams per mole, not atomic mass units. The masses of the atoms can be taken from the periodic table.

Solutions to Example 5.3.2

6 C = 6 × 12.011	= 72.066
12 H = 12 × 1.0079	= 12.0948
6 O = 6 × 15.999	= 95.994
TOTAL	= 180.155 g/mol

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

? Exercise 3.11.2

What is the molar mass of $AgNO_3$?

Answer

169.87 g/mol

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

✓ Example 3.11.3

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

Solution

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

$$3.56 \ \cancel{mol\, HgCl_2} \times \frac{271.49\, g\, HgCl_2}{\cancel{mol\, HgCl_2}} = 967\, g\, HgCl_2$$

? Exercise 3.11.3

What is the mass of 33.7 mol of H_2O?

Answer

607 g

✓ Example 3.11.4

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

Solution

Use the molar mass of H_2O as a conversion factor from mass to moles. The molar mass of water is

$$(1.0079 + 1.0079 + 15.999) = 18.015 \text{ g/mol}.$$

However, because we want to cancel the gram unit and introduce moles, we need to take the reciprocal of this quantity, or 1 mol/18.015 g:

$$240.0 \ \cancel{g\, H_2O} \times \frac{1\, mol\, H_2O}{18.015 \ \cancel{g\, H_2O}} = 13.32\, mol\, H_2O$$

? Exercise 3.11.4

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

Answer

0.363 mol

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

✓ Example 3.11.5

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

Solution

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

$$100\ \cancel{mL}\ ethanol \times \frac{0.789\ g}{\cancel{mL}} \times \frac{1\ mol}{46.08\ \cancel{g}} = 1.71\ mol\ ethanol$$

Exercise 3.11.5

If the density of benzene, C_6H_6, is 0.879 g/mL, how many moles are present in 17.9 mL of benzene?

Answer

0.201 mol

Summary

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

3.12: Stoichiometry

Learning Objectives

- Explain the meaning of the term "stoichiometry".
- Determine the relative amounts of each substance in chemical equations.

You have learned that chemical equations provide us with information about the types of particles that react to form products. Chemical equations also provide us with the relative number of particles and moles that react to form products. In this section you will explore the quantitative relationships that exist between the quantities of reactants and products in a balanced equation. This is known as **stoichiometry**.

Stoichiometry, by definition, is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation. The word stoichiometry is actually Greek from two words: $\sigma\tau o\iota\kappa\eta\iota o\nu$, which means "element", and \(\mu \epsilon \tau \rho \omicron \nu), which means "measure".

Interpreting Chemical Equations

The mole, as you remember, is a quantitative measure that is equivalent to Avogadro's number of particles. So how does this relate to the chemical equation? Look at the chemical equation below.

$$2CuSO_4 + 4KI \rightarrow 2CuI + 2K_2SO_4 + I_2$$

The coefficients used, as we have learned, tell us the relative amounts of each substance in the equation. So for every 2 units of copper (II) sulfate ($CuSO_4$) we have, we need to have 4 units of potassium iodide (KI). For every two dozen copper (II) sulfates, we need 4 dozen potassium iodides. Because the unit "mole" is also a counting unit, we can interpret this equation in terms of moles, as well: For every two moles of copper (II) sulfate, we need 4 moles potassium iodide.

The production of ammonia $\left(NH_3\right)$ from nitrogen and hydrogen gases is an important industrial reaction called the Haber process, after German chemist Fritz Haber.

$$N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)$$

The balanced equation can be analyzed in several ways, as shown in the figure below.

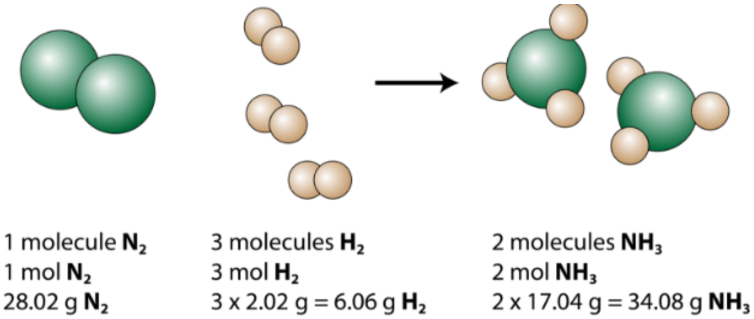


Figure 3.12.1: This representation of the production of ammonia from nitrogen and hydrogen show several ways to interpret the quantitative information of a chemical reaction.

We see that 1 molecule of nitrogen reacts with 3 molecules of hydrogen to form 2 molecules of ammonia. This is the smallest possible relative amount of the reactants and products. To consider larger relative amounts, each coefficient can be multiplied by the same number. For example, 10 molecules of nitrogen would react with 30 molecules of hydrogen to produce 20 molecules of ammonia.

The most useful quantity for counting particles is the mole. So if each coefficient is multiplied by a mole, the balanced chemical equation tells us that 1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia. This is the conventional way to interpret any balanced chemical equation.

Finally, if each mole quantity is converted to grams by using the molar mass, we can see that the law of conservation of mass is followed. 1 mol of nitrogen has a mass of 28.02 g while 3 mol of hydrogen has a mass of 6.06 g and 2 mol of ammonia has a mass of 34.08 g

$$28.02\ \text{g N}_2 + 6.06\ \text{g H}_2 \rightarrow 34.08\ \text{g NH}_3$$

Mass and the number of atoms must be conserved in any chemical reaction. The number of molecules is not necessarily conserved.

✓ Example 3.12.1

The equation for the combustion of ethane (C_2H_6) is

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

a. Indicate the number of formula units or molecules in the balanced equation.
b. Indicate the number of moles present in the balanced equation.

Solution

a. Two molecules of C_2H_6 plus seven molecules of O_2 yields four molecules of CO_2 plus six molecules of H_2O.
b. Two moles of C_2H_6 plus seven moles of O_2 yields four moles of CO_2 plus six moles of H_2O.

? Exercise 3.12.1

For the following equation below, indicate the number of formula units or molecules, and the number of moles present in the balanced equation.

$$KBrO_3 + 6KI + 6HBr \rightarrow 7KBr + 3H_2O$$

Answer

One molecules of $KBrO_3$ plus six molecules of KI plus six molecules of HBr yields seven molecules of KBr plus three molecules of I_2 and three molecules of H_2O. One mole of $KBrO_3$ plus six moles of KI plus six moles of HBr yields seven moles of KBr plus three moles of I_2 plus three moles of H_2O.

Summary

- Stoichiometry is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in a balanced chemical equation.
- The coefficients in a balanced chemical equation represent the reacting ratios of the substances in the reaction.
- The coefficients of a balanced equation can be used to determine the ratio of moles of all substances in the reaction.

Vocabulary

- **Stoichiometry -** The calculation of quantitative relationships of the reactants and products in a balanced chemical equation.
- **Formula unit -** The empirical formula of an ionic compound.
- **Mole ratio -** The ratio of the moles of one reactant or product to the moles of another reactant or product according to the coefficients in the balanced chemical equation.

3.13: Mole-to-Mole Conversions

/;Learning Objectives

- Use a balanced chemical equation to determine molar relationships between substances.

Previously, you learned to balance chemical equations by comparing the numbers of each type of atom in the reactants and products. The coefficients in front of the chemical formulas represent the numbers of molecules or formula units (depending on the type of substance). As follows, we will extend the meaning of the coefficients in a chemical equation.

Consider the simple chemical equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

The convention for writing balanced chemical equations is to use the lowest whole-number ratio for the coefficients. However, the equation is balanced as long as the coefficients are in a 2:1:2 ratio. For example, this equation is also balanced if we write it as

$$4H_2 + 2O_2 \rightarrow 4H_2O$$

The ratio of the coefficients is 4:2:4, which reduces to 2:1:2. The equation is also balanced if we were to write it as

$$22H_2 + 11O_2 \rightarrow 22H_2O$$

because 22:11:22 also reduces to 2:1:2.

Suppose we want to use larger numbers. Consider the following coefficients:

$$12.044 \times 10^{23}\ \mathrm{H_2} + 6.022 \times 10^{23}\ \mathrm{O_2} \rightarrow 12.044 \times 10^{23}\ \mathrm{H_2O}$$

These coefficients also have the ratio 2:1:2 (check it and see), so this equation is balanced. But 6.022×10^{23} is 1 mol, while 12.044×10^{23} is 2 mol (and the number is written that way to make this more obvious), so we can simplify this version of the equation by writing it as

$$2\ mol\ \mathrm{H_2} + 1\ mol\ \mathrm{O_2} \rightarrow 2\ mol\ \mathrm{H_2O}$$

We can leave out the word *mol* and not write the 1 coefficient (as is our habit), so the final form of the equation, still balanced, is

$$2\,\mathrm{H_2} + \mathrm{O_2} \rightarrow 2\,\mathrm{H_2O}$$

Now we interpret the coefficients as referring to molar amounts, not individual molecules. The lesson? *Balanced chemical equations are balanced not only at the molecular level, but also in terms of molar amounts of reactants and products.* Thus, we can read this reaction as "two moles of hydrogen react with one mole of oxygen to produce two moles of water."

By the same token, the ratios we constructed to describe a molecular reaction can also be constructed in terms of moles rather than molecules. For the reaction in which hydrogen and oxygen combine to make water, for example, we can construct the following ratios:

$$\frac{2\text{ mol H}_2}{1\text{ mol O}_2} \text{ or } \frac{1\text{ mol O}_2}{2\text{ mol H}_2}$$

$$\frac{2\text{ mol H}_2\text{O}}{1\text{ mol O}_2} \text{ or } \frac{1\text{ mol O}_2}{2\text{ mol H}_2\text{O}}$$

$$\frac{2\text{ mol H}_2}{2\text{ mol H}_2\text{O}} \text{ or } \frac{2\text{ mol H}_2\text{O}}{2\text{ mol H}_2}$$

We can use these ratios to determine what amount of a substance, in moles, will react with or produce a given number of moles of a different substance. The study of the numerical relationships between the reactants and the products in balanced chemical reactions is called *stoichiometry*.

✓ Example 3.13.1

How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O?

Solution

Solutions to Example 8.3.1

Steps for Problem Solving	How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O?
Find a balanced equation that describes the reaction.	Unbalanced: $H_2 + O_2 \rightarrow H_2O$ Balanced: $\underline{2}H_2 + O_2 \rightarrow \underline{2}H_2O$
Identify the "given" information and what the problem is asking you to "find."	Given: moles H_2O Find: moles oxygen
List other known quantities.	1 mol O_2 = 2 mol H_2O
Prepare a concept map and use the proper conversion factor.	 $\frac{1\ mol\ O_2}{2\ mol\ H_2O}$
Cancel units and calculate.	$27.6\ \cancel{\text{mol } H_2O} \times \frac{1 \text{ mol } O_2}{2\ \cancel{\text{mol } H_2O}} = 13.8 \text{ mol } O_2$ To produce 27.6 mol of H_2O, 13.8 mol of O_2 react.
Think about your result.	Since each mole of oxygen produces twice as many moles of water, it makes sense that the produced amount is greater than the reactant amount

✓ Example 3.13.2

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?

Solution

Solutions to Example 8.3.2

Steps for Problem Solving	How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?
Find a balanced equation that describes the reaction.	Unbalanced: $N_2 + H_2 \rightarrow NH_3$ Balanced: $N_2 + \underline{3}H_2 \rightarrow \underline{2}NH_3$
Identify the "given" information and what the problem is asking you to "find."	Given: $H_2 = 4.20$ mol Find: mol of NH_3
List other known quantities.	3 mol H_2 = 2 mol NH_3
Prepare a concept map and use the proper conversion factor.	 $\frac{2\ mol\ NH_3}{3\ mol\ H_2}$

Steps for Problem Solving	How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?
Cancel units and calculate.	$4.20 \text{ mol } \cancel{H_2} \times \frac{2 \text{ mol } NH_3}{3 \text{ mol } \cancel{H_2}} = 2.80 \text{ mol } NH_3$ The reaction of 4.20 mol of hydrogen with excess nitrogen produces 2.80 mol of ammonia.
Think about your result.	The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation.

? Exercise 3.13.3

a. Given the following balanced chemical equation:

$$\mathrm{C_5H_{12} + 8\,O_2 \rightarrow 5\,CO_2 + 6\,H_2O}$$

, How many moles of $\mathrm{H_2O}$ can be formed if 0.0652 mol of $\mathrm{C_5H_{12}}$ were to react?

b. Balance the following unbalanced equation and determine how many moles of $\mathrm{H_2O}$ are produced when 1.65 mol of NH_3 react:

$$\mathrm{NH_3 + O_2 \rightarrow N_2 + H_2O}$$

Answer a

0.391 mol H_2O

Answer b

$4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$; 2.48 mol H_2O

Summary

- The balanced chemical reaction can be used to determine molar relationships between substances.

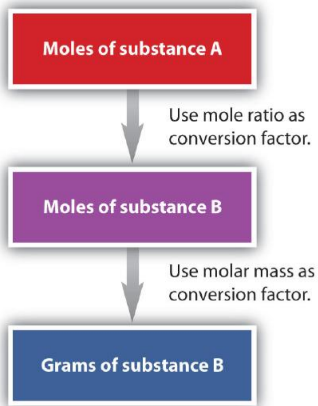
3.14: Making Molecules- Mole to Mass (or vice versa) and Mass-to-Mass Conversions

Learning Objectives

- Convert from mass or moles of one substance to mass or moles of another substance in a chemical reaction.

Mole to Mass Conversions

We have established that a balanced chemical equation is balanced in terms of moles, as well as atoms or molecules. We have used balanced equations to set up ratios, in terms of moles of materials, that we can use as conversion factors to answer stoichiometric questions—such as how many moles of substance A react with so many moles of reactant B. We can extend this technique even further. Recall that we can relate a molar amount to a mass amount using molar mass. We can use that relation to answer stoichiometry questions in terms of the masses of a particular substance, in addition to moles. We do this using the following sequence:



Collectively, these conversions are called mole-mass calculations. Flowchart of mole mass calculations: To convert from moles of substance A to moles of substance B, use the mole ratio conversion factor, and to convert from moles to grams of substance B, use molar mass conversion factor

As an example, consider the balanced chemical equation

$$\mathrm{Fe_2O_3 + 3\,SO_3 \rightarrow Fe_2(SO_4)_3} \tag{3.14.1}$$

If we have 3.59 mol of Fe_2O_3, how many grams of SO_3 can react with it? Using the mole-mass calculation sequence, we can determine the required mass of SO_3 in two steps. First, we construct the appropriate molar ratio, determined from the balanced chemical equation, to calculate the number of moles of SO_3 needed. Then, using the molar mass of SO_3 as a conversion factor, we determine the mass that this number of moles of SO_3 has.

As usual, we start with the quantity we were given:

$$3.59\ \ \mathrm{mol}\ \cancel{\mathrm{Fe_2O_3}} \times \left(\frac{3\ \mathrm{mol\ SO_3}}{1\ \ \mathrm{mol}\ \cancel{\mathrm{Fe_2O_3}}}\right) = 10.77\ \mathrm{mol\ SO_3} \tag{3.14.2}$$

The mol Fe_2O_3 units cancel, leaving mol SO_3 unit. Now, we take this answer and convert it to grams of SO_3, using the molar mass of SO_3 as the conversion factor:

$$10.77\ \ \cancel{\mathrm{mol\ SO_3}} \times \left(\frac{80.06\ \mathrm{g\ SO_3}}{1\ \ \cancel{\mathrm{mol\ SO_3}}}\right) = 862\ \mathrm{g\ SO_3} \tag{3.14.3}$$

Our final answer is expressed to three significant figures. Thus, in a two-step process, we find that 862 g of SO_3 will react with 3.59 mol of Fe_2O_3. Many problems of this type can be answered in this manner.

> The same two-step problem can also be worked out in a single line, rather than as two separate steps, as follows:
>
> $$3.59\ \ mol\ \cancel{Fe_2O_3} \times \underbrace{\left(\frac{3\ \ \cancel{mol\ SO_3}}{1\ \ mol\ \cancel{Fe_2O_3}}\right)}_{\text{converts to moles of SO}_3} \times \underbrace{\left(\frac{80.06\ g\ SO_3}{1\ \ \cancel{mol\ SO_3}}\right)}_{\text{converts to grams of SO}_3} = 862\ g\ SO_3$$
>
> We get exactly the same answer when combining all math steps together.

✓ Example 3.14.1: Generation of Aluminum Oxide

How many moles of HCl will be produced when 249 g of $AlCl_3$ are reacted according to this chemical equation?

$$2\,AlCl_3 + 3\,H_2O(\ell) \rightarrow Al_2O_3 + 6\,HCl(g)$$

Solution

Solutions to Example 8.5.1

Steps for Problem Solving	Example 3.14.1
Identify the "given" information and what the problem is asking you to "find."	Given: 249 g $AlCl_3$ Find: moles HCl
List other known quantities.	1 mol $AlCl_3$ = 133.33 g $AlCl_3$ 6 mol of HCl to 2 mol $AlCl_3$
Prepare a concept map and use the proper conversion factor.	 $\frac{1\,mol\,AlCl_3}{133.33\,g\,AlCl_3}$ $\frac{6\,mol\,HCl}{2\,mol\,AlCl_3}$ Flowchart of needed conversion factors: 1 mole AlCl3 to 133.33 grams AlCl3, and 6 moles HCl to 2 moles AlCl3
Cancel units and calculate.	$249\ \cancel{g\ AlCl_3} \times \frac{1\ \cancel{mol\ AlCl_3}}{133.33\ \cancel{g\ AlCl_3}} \times \frac{6\,mol\,HCl}{2\ \cancel{mol\ AlCl_3}} = 5.60\,mol\,HCl$
Think about your result.	Since 249 g of $AlCl_3$ is less than 266.66 g, the mass for 2 moles of $AlCl_3$ and the relationship is 6 mol of HCl to 2 mol $AlCl_3$, the answer should be less than 6 moles of HCl.

? Exercise 3.14.1: Generation of Aluminum Oxide

How many moles of Al_2O_3 will be produced when 23.9 g of H_2O are reacted according to this chemical equation?

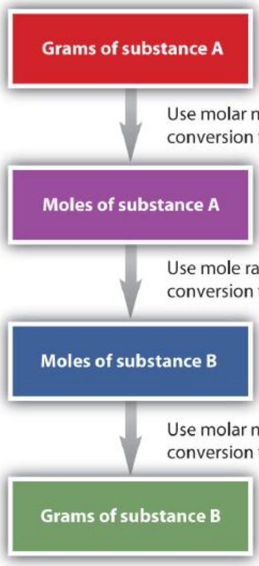
$$2\,AlCl_3 + 3\,H_2O(\ell) \rightarrow Al_2O_3 + 6\,HCl(g)$$

Answer

0.442 mol Al_2O_3

Mass to Mass Conversions

It is a small step from mole-mass calculations to mass-mass calculations. If we start with a known mass of one substance in a chemical reaction (instead of a known number of moles), we can calculate the corresponding masses of other substances in the reaction. The first step in this case is to convert the known mass into moles, using the substance’s molar mass as the conversion factor. Then—and only then—we use the balanced chemical equation to construct a conversion factor to convert that quantity to moles of another substance, which in turn can be converted to a corresponding mass. Sequentially, the process is as follows:



This three-part process can be carried out in three discrete steps or combined into a single calculation that contains three conversion factors. The following example illustrates both techniques. Flowchart of mole mass calculations: To convert from grams to moles of substance A, use molar mass conversion factor; To convert from moles of substance A to moles of substance B, use the mole ratio conversion factor, and to convert from moles to grams of substance B, use molar mass conversion factor

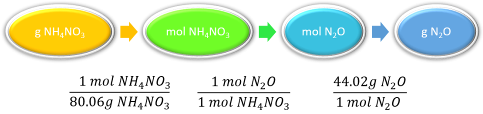
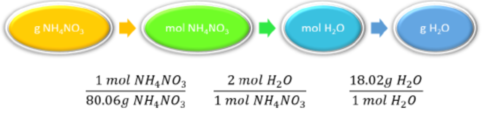
✓ Example 3.14.2: Decomposition of Ammonium Nitrate

Ammonium nitrate decomposes to dinitrogen monoxide and water according to the following equation.

$$\mathrm{NH_4NO_3}\,(s) \rightarrow \mathrm{N_2O}\,(g) + 2\mathrm{H_2O}\,(l)$$

In a certain experiment, 45.7 g of ammonium nitrate is decomposed. Find the mass of each of the products formed.

Solutions to Example 8.5.2

Steps for Problem Solving	Example $3.14.2$
Identify the "given" information and what the problem is asking you to "find."	Given: $45.7 \text{ g NH}_4\text{NO}_3$ Find: Mass $\text{N}_2\text{O} = ? \text{ g}$ Mass $\text{H}_2\text{O} = ? \text{ g}$
List other known quantities.	$1 \text{ mol NH}_4\text{NO}_3 = 80.06 \text{ g}$ $1 \text{ mol N}_2\text{O} = 44.02 \text{ g}$ $1 \text{ mol H}_2\text{O} = 18.02 \text{ g}$ 1 mol NH_4NO_3 to 1 mol N_2O to 2 mol H_2O
Prepare two concept maps and use the proper conversion factor.	 Flowchart of conversion factors: 1 mole NH4NO3 to 80.06 grams NH4NO3, 1 mole N2O to 1 mole NH4NO3, 44.02 grams N2O to 1 mole N2O  Flowchart of conversion factors: 1 mole NH4NO3 to 80.06 grams NH4NO3, 2 moles H2O to 1 mole NH4NO3, 18.02 grams H2O to 1 mole H2O
Cancel units and calculate.	$45.7 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.06 \text{ g NH}_4\text{NO}_3} \times \frac{1 \text{ mol N}_2\text{O}}{1 \text{ mol NH}_4\text{NO}_3} \times \frac{44.02 \text{ g N}_2\text{O}}{1 \text{ mol N}_2\text{O}} = 25.1$ $45.7 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.06 \text{ g NH}_4\text{NO}_3} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol NH}_4\text{NO}_3} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 20.6$

Steps for Problem Solving	Example $3.14.2$
Think about your result.	The total mass of the two products is equal to the mass of ammonium nitrate which decomposed, demonstrating the law of conservation of mass. Each answer has three significant figures.

? Exercise $3.14.2$: Carbon Tetrachloride

Methane can react with elemental chlorine to make carbon tetrachloride $(\mathrm{CCl_4})$. The balanced chemical equation is as follows:

$$\mathrm{CH_4(g) + 4\,Cl_2(g) \rightarrow CCl_2(l) + 4\,HCl(l)}$$

How many grams of HCl are produced by the reaction of 100.0 g of $\mathrm{CH_4}$?

Answer

908.7g HCl

Summary

- Calculations involving conversions between moles of a substance and the mass of that substance can be done using conversion factors.
- A balanced chemical reaction can be used to determine molar and mass relationships between substances.

3.15: Exothermic and Endothermic Processes

Learning Objectives

- Define endothermic and exothermic reactions.
- Describe how heat is transferred in endothermic and exothermic reactions.
- Determine whether a reaction is endothermic or exothermic through observations, temperature changes, or an energy diagram.

We first introduced the concept of energy in Module 1 section 8. Recall that energy is manifested in different forms and that it can be converted from one form to another and it can be transferred but must be conserved. In this section we begin to study the energy involved in chemical processes.

A campfire is an example of a combustion reaction and serves as an introduction to thermochemistry. Thermochemistry involves the study of energy changes during chemical reactions. The reaction is initiated by the application of heat from a match. The reaction converting wood to carbon dioxide and water (among other things) continues, releasing heat energy in the process. This heat energy can then be used to cook food, boil water, or just keep warm when it's cold outside.

Figure $3.15.1$: *An image of a campfire (K Hartling, CC BY-SA 3.0; Wikimedia)*

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The **law of conservation of energy** states that in any physical or chemical process, energy is neither created nor destroyed. In other words, the entire energy in the universe is conserved. In order to better understand the energy changes taking place during a reaction, we need to define two parts of the universe, called the system and the surroundings. The **system** is the specific portion of matter in a given space that is being studied during an experiment or an observation. The **surroundings** is everything in the universe that is not part of the system. In practical terms for a laboratory chemist, the system is the particular chemicals being reacted, while the surroundings is the immediate vicinity within the room. During most processes, energy is exchanged between the system and the surroundings. If the system loses a certain amount of energy, that same amount of energy is gained by the surroundings. If the system gains a certain amount of energy, that energy is supplied by the surroundings.

A chemical reaction or physical change is **endothermic** if heat is absorbed by the system from the surroundings. In the course of an endothermic process, the system gains heat from the surroundings and so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter q. The sign of q for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** if heat is released by the system into the surroundings. Because the surroundings is gaining heat from the system, the temperature of the surroundings increases. The sign of q for an exothermic process is negative because the system is losing heat.

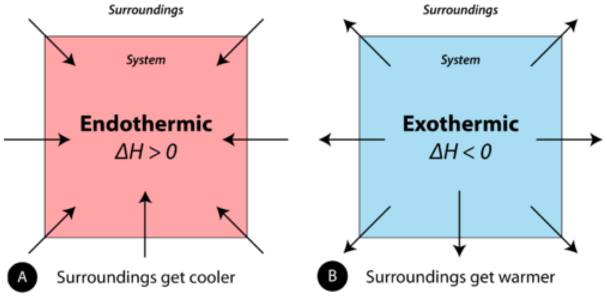


Figure 3.15.2: (A) Endothermic reaction. (B) Exothermic reaction.

The symbol ΔH, referred to the enthalpy of the reaction, is used in chemistry to represent energy changes under constant pressure conditions (a common condition for reactions that occur in a lab). Read more about enthalpy of reactions in section 3.15.

Energy Diagrams

Energy changes are frequently shown by drawing an energy diagram. Energy diagrams show the potential energy of the reactants and products. If, on an energy diagram, the products have more potential energy than the reactants started with, the reaction is endothermic, energy is added to the reaction. If, on the energy diagram, the products have less potential energy than the reactants started with, the reaction is exothermic, energy is released during the reaction. The "hill" between the reactants and products is the activation energy. This is also referred to as the activation energy barrier because the reactants must have enough energy when they collide to overcome this energy barrier and form products.

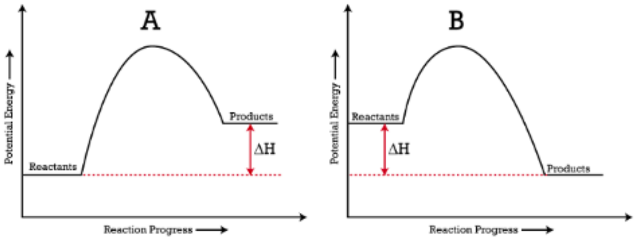


Figure 3.15.3: A potential energy diagram shows the total potential energy of a reacting system as the reaction proceeds. (A) In an endothermic reaction, the energy of the products is greater than the energy of the reactants and ΔH is positive. (B) In an exothermic reaction, the energy of the products is lower than the energy of the reactions and ΔH is negative. (CC BY-NC; CK-12)

Summary

Chemical reactions involve changes in energy. This may be a change in heat, electricity, light, or other forms of energy. Reactions that absorb energy are endothermic. Reactions that release energy are exothermic.

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3.16: Enthalpy

Enthalpy

Heat changes in chemical reactions are often measured in the laboratory under conditions in which the reacting system is open to the atmosphere. In that case, the system is at a constant pressure so the energy change is the enthalpy of the reaction, **Enthalpy (ΔH)**. The change in enthalpy of a reaction is a measure of the differences in enthalpy of the reactants and products.

Several factors influence the enthalpy of a reaction. The state of reactants and products (solid, liquid, or gas) influences the enthalpy value for a system. The direction of the reaction affects the enthalpy value. A reaction that takes place in the opposite direction has the same numerical enthalpy value, but the opposite sign. Enthalpy is an extensive property, which means it depends on the quantity of substances involved in the reaction.

Example: The combustion of hydrocarbons are exothermic reactions. For example, for every one mole of methane gas that burns, 890.3 kJ of energy are released. The energy in this case, can be treated as a product in the reaction since it is released from the system.

$$CH_4(g) + 2\ O_2(g) \text{ -> } CO_2(g) + 2H_2O(l) + 890.3\text{ kJ}$$

The enthalpy of the reaction can also be written separately from the equation for the reaction:

$$CH_4(g) + 2\ O_2(g) \text{ -> } CO_2(g) + 2H_2O(l) \quad \Delta H = -890.3\text{ kJ}$$

The negative sign on the enthalpy indicates the reaction is exothermic.

If 5.00 mol of methane are combusted, then five times the amount of energy will be released:

$$5.00\text{ mol } CH_4 \text{ x } -890.4\text{kJ}/1\text{ mol } CH_4 = -4450\text{ kJ}$$

On a fundamental level, chemical reactions involve breaking the bonds of the reactants, rearranging the atoms and forming the bonds of the products. As a result, the enthalpy of a reaction can be estimated using average bond energies. Energy is required to break bonds and energy is released when bonds form. The quantity of the energy involved in breaking and forming a particular bond is the same, but the sign is different to indicate whether the process is exothermic or endothermic. By convention, when energy is released by a system, the sign is negative and when energy is added to a system the sign is positive. The enthalpy of reaction can be estimated by comparing the energy needed to break all of the bonds of the reactants to the energy released when all of the bonds in the products form. If more energy is released when bonds form than was needed to break the bonds of the reactants, the overall reaction is exothermic. If more energy is needed to break bonds than is released when the bonds form, then the reaction is endothermic.

Table 1: Average Bond Energies (kJ/mol) (Bond Energies)

Single Bonds						**Multiple Bonds**	
H—H	432	**N—H**	391	**I—I**	149	**C = C**	614
H—F	565	**N—N**	160	**I—Cl**	208	**C ≡ C**	839
H—Cl	427	**N—F**	272	**I—Br**	175	**O = O**	495
H—Br	363	**N—Cl**	200			**C = O***	745
H—I	295	**N—Br**	243	**S—H**	347	**C ≡ O**	1072
		N—O	201	**S—F**	327	**N = O**	607

Single Bonds						Multiple Bonds	
C—H	413	**O—H**	467	**S—Cl**	253	**N = N**	418
C—C	347	**O—O**	146	**S—Br**	218	**N ≡ N**	941
C—N	305	**O—F**	190	**S—S**	266	**C ≡ N**	891
C—O	358	**O—Cl**	203			**C = N**	615
C—F	485	**O—I**	234	**Si—Si**	340		
C—Cl	339			**Si—H**	393		
C—Br	276	**F—F**	154	**Si—C**	360		
C—I	240	**F—Cl**	253	**Si—O**	452		
C—S	259	**F—Br**	237				
		Cl—Cl	239				
		Cl—Br	218				
		Br—Br	193				

*C == O(CO_2) = 799

Summary

Enthalpy changes are a measure of the energy changes in chemical reactions. The energy involved in a particular reaction depends on the quantity of the substances involved. Enthalpy changes for reactions can be estimated using average bond energies.

Contributors

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3.17: Temperature and Heat

Learning Objectives

- Identify the different between temperature and heat.
- Recognize the different scales used to measure temperature

The concept of temperature may seem familiar to you, but many people confuse temperature with heat. **Temperature** is a measurement of how hot or cold an object is relative to another object (its thermal energy content). We can measure temperature in the laboratory. H**eat** is the flow of thermal energy between objects with different temperatures. Temperature is a measure of the average kinetic energy of the particles in matter. In everyday usage, temperature indicates a measure of how hot or cold an object is. Temperature is an important parameter in chemistry. When a substance changes from solid to liquid, it is because there was in increase in the temperature of the material. Chemical reactions usually proceed faster if the temperature is increased. Many unstable materials (such as enzymes) will be viable longer at lower temperatures.

Figure 3.17.1: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy.

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K). Thermometers measure temperature by using materials that expand or contract when heated or cooled. Mercury or alcohol thermometers, for example, have a reservoir of liquid that expands when heated and contracts when cooled, so the liquid column lengthens or shortens as the temperature of the liquid changes.

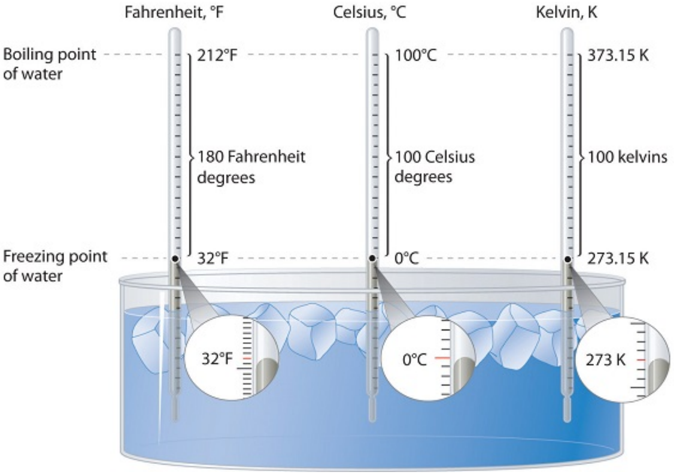

Figure 3.17.1: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit (°F). (CC BY-SA-NC 3.0; anonymous)

Summary

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K).

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3.18: Calorimetry

Learning Objectives

- Differentiate between heat capacity and specific heat capacity.
- Calculate enthalpy (heats) of reactions using calorimetry data

In the previous section we used average bond energies to estimate the change in enthalpy of a reaction (also called the heat of reaction). Heats of reactions can be experimentally measured using calorimetry. Calorimetry relies on measuring the energy change of a substance, often water, to which energy was transferred into or absorbed from during the chemical reaction. How substances respond to heat is dependent upon their heat capacity.

Heat Capacity and Specific Heat

If a glass of water and a large pot filled with water are both at the same temperature and the same amount of heat energy is added to each of them, which one will experience a greater increase in temperature? If a metal spoon and an equal mass of water, both at the same temperature initially, have the same amount of energy transferred to each of them, which would experience a greater increase in temperature?

The small glass of water and the metal will experience the greater change in temperature because the small glass of water has a lower heat capacity than the large pot of water and the metal has a lower heat capacity than the same mass of water. Heat capacity is a measure of the amount of energy needed to change the temperature of an object by 1^{o}C. These two examples show that the heat capacity of an object is dependent on its mass and on its composition.

A property related to heat capacity but independent of the mass of the object is the specific heat capacity. The specific heat capacity is defined as the amount of energy needed to change the temperature of 1 gram of a substance by 1^{o}C and is represented by the symbol c. The units on specific heat are often expressed as joules per gram per degree (J/g-^{o}C) or as calories per gram per degree (cal/g-^{o}C).

Table 3.18.1: Specific Heat Capacities

Substance	Specific Heat Capacity at 25^{o}C in J/g ^{o}C	Substance	Specific Heat Capacity at 25^{o}C in J/g ^{o}C
H_2 gas	14.267	steam @ 100^{o}C	2.010
He gas	5.300	vegetable oil	2.000
$H_2O(l)$	4.184	sodium	1.23
lithium	3.56	air	1.020
ethyl alcohol	2.460	magnesium	1.020
ethylene glycol	2.200	aluminum	0.900
ice @ 0^{o}C	2.010	concrete	0.880
steam @ 100^{o}C	2.010	glass	0.840

Water is commonly used as a coolant for machinery because it is able to absorb large quantities of heat (see table above). Coastal climates are much more moderate than inland climates because of the presence of the ocean. Water in lakes or oceans absorbs heat from the air on hot days and releases it back into the air on cool days.

Figure 3.18.1: This power plant in West Virginia, like many others, is located next to a large lake so that the water from the lake can be used as a coolant. Cool water from the lake is pumped into the plant, while warmer water is pumped out of the plant and back into the lake.

The energy transferred into or out of a substance can be calculated if the mass and temperature change is measured and the specific heat capacity is known:

$q = m \times c \times \Delta T$

where q = energy transferred

m =mass of the substance

c = specific heat of the substance

$\Delta T = T_{final} - T_{initial}$ (change in temperature of the substance)

Example:

Suppose the energy released when a fuel is burned is completely transferred into a 500.0 g sample of water initially at 18.0°C. The final temperature of the water is 28.0°C. How much energy did the water gain?

The specific heat of water is 4.184 J J/g-°C.

$$q = (500.0\text{ g})(4.184\text{ J J/g-}^{o}\text{C})(28.0^{o}\text{C} - 18.0^{o}\text{C}) = 20{,}920\text{ J}$$

Summary

- Heat capacity is the amount of heat required to raise the temperature of an object by 1°C).
- The specific heat of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by 1°C.
- You can calculate the heat (energy) of a reaction using the equation: $q = m \times c \times \Delta T$

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3.19: Climate Change - Too Much Carbon Dioxide

Carbon dioxide (CO_2) is an important heat-trapping (greenhouse) gas, which is released through human activities such as deforestation and burning fossil fuels, as well as natural processes such as respiration and volcanic eruptions. Figure 3.19.1 shows CO_2 levels during the last three glacial cycles, as reconstructed from ice cores.

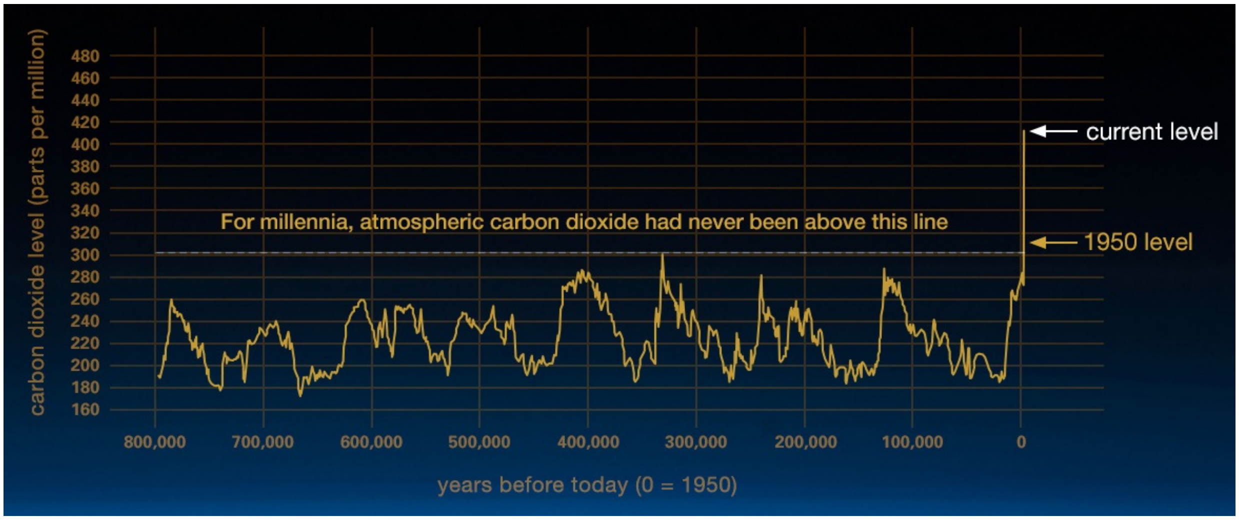


Figure 3.19.1: This graph, based on the comparison of atmospheric samples contained in ice cores and more recent direct measurements, provides evidence that atmospheric CO2 has increased since the Industrial Revolution. (Luthi, D., et al. 2008; Etheridge, D.M., et al. 2010; Vostok ice core data/J.R.R. Petit et al.; NOAA Mauna Loa CO2 record. NASA.) Line graph of carbon dioxide levels from 800 thousand before up until present day; current carbon dioxide level is higher than all historical levels.

Carbon dioxide (CO_2) is the primary greenhouse gas emitted through human activities. In 2015, CO_2 accounted for about 82.2% of all U.S. greenhouse gas emissions from human activities. Carbon dioxide is naturally present in the atmosphere as part of the Earth's carbon cycle (the natural circulation of carbon among the atmosphere, oceans, soil, plants, and animals). Human activities are altering the carbon cycle, both by adding more CO_2 to the atmosphere and by influencing the ability of natural sinks, like forests, to remove CO_2 from the atmosphere. While CO_2 emissions come from a variety of natural sources, human-related emissions are responsible for the increase that has occurred in the atmosphere since the industrial revolution.

The main human activity that emits CO_2 is the combustion of fossil fuels (coal, natural gas, and oil) for energy and transportation, although certain industrial processes and land-use changes also emit CO_2. As an example of how CO_2 can be generated, consider the combustion of octane, a component of gasoline:

$$2\,C_8H_{18}(l) + 21\,O_2(g) \rightarrow 16\,CO_2(g) + 18\,H_2O(g) \tag{3.19.1}$$

The balanced reaction in Equation 3.19.1 demonstrates that for every two molecules of octane that are burned, 16 molecules of CO_2 are generated.

Contributions & Attributions

- Earth Science Communications Team at NASA's Jet Propulsion Laboratory, California Institute of Technology
- EPA

CHAPTER OVERVIEW

4: Water

4.1: The Uniqueness of Water

Water is everywhere! It has some amazing properties which allow for life to go on and for chemistry to happen. The earth is mostly water, our bodies are mostly water, and many chemical reactions take place in water. This module introduces you to the many chemical reactions, and properties of solutions where water is the primary solvent.

Watch the following video to see some of the unique properties of water.

4.2: Predicting the Shapes of Molecules

Learning Objective

- Determine the shape of simple molecules.

Molecules have shapes. There is an abundance of experimental evidence to that effect—from their physical properties to their chemical reactivity. Small molecules—molecules with a single central atom—have shapes that can be easily predicted. The basic idea in molecular shapes is called **valence shell electron pair repulsion (VSEPR).** VSEPR says that electron pairs, being composed of negatively charged particles, repel each other to get as far away from one another as possible. VSEPR makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**: any type of bond—single, double, or triple—and lone electron pairs. When applying VSEPR to simple molecules, the first thing to do is to count the number of electron groups around the central atom. Remember that a multiple bond counts as only *one* electron group.

Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible—180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. Examples include BeH$_2$ and CO$_2$:

Figure 4.2.1: Beryllium hydride and carbon dioxide bonding.

The two molecules, shown in the figure below in a "ball and stick" model.

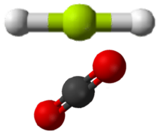


Figure 4.2.2: Beryllium hydride and carbon dioxide models. (CK12 Licence)

A molecule with three electron groups orients the three groups as far apart as possible. They adopt the positions of an equilateral triangle—120° apart and in a plane. The shape of such molecules is *trigonal planar*. An example is BF$_3$:

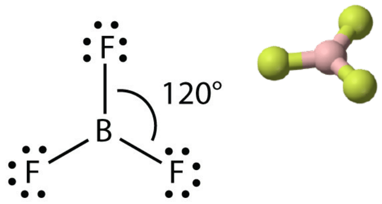


Figure 4.2.3: Boron trifluoride bonding. (CK12 Licence)

Some substances have a trigonal planar electron group distribution but have atoms bonded to only two of the three electron groups. An example is GeF$_2$:

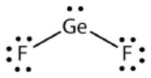


Figure 4.2.4: Germanium difluoride bonding.

From an electron group geometry perspective, GeF$_2$ has a trigonal planar shape, but its real shape is dictated by the positions of the atoms. This shape is called *bent* or *angular*.

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure 4.2.1 Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

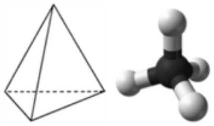


Figure 4.2.5: Tetrahedral structure of methane. (CK12 Licence)

This diagram of CH_4 illustrates the standard convention of displaying a three-dimensional molecule on a two-dimensional surface. The straight lines are in the plane of the page, the solid wedged line is coming out of the plane toward the reader, and the dashed wedged line is going out of the plane away from the reader.

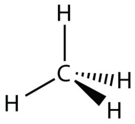


Figure 4.2.6: Methane bonding. (CK12 Licence)

NH_3 is an example of a molecule whose central atom has four electron groups, but only three of them are bonded to surrounding atoms.

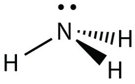


Figure 4.2.7: Ammonia bonding. (CK12 Licence)

Although the electron groups are oriented in the shape of a tetrahedron, from a molecular geometry perspective, the shape of NH_3 is *trigonal pyramidal*.

H_2O is an example of a molecule whose central atom has four electron groups, but only two of them are bonded to surrounding atoms.

Figure 4.2.8: Water bonding.

Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*. A molecule with four electron groups about the central atom, but only one electron group bonded to another atom, is linear because there are only two atoms in the molecule.

Double or triple bonds count as a single electron group. The Lewis electron dot diagram of formaldehyde (CH_2O) is shown in Figure 4.2.9.

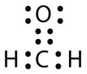


Figure 4.2.9: Lewis Electron Dot Diagram of Formaldehyde.

The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape.

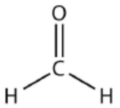


Figure 4.2.10: Formaldehyde bonding.

(The lone electron pairs on the O atom are omitted for clarity.) The molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule.

Table 4.2.1 summarizes the shapes of molecules based on the number of electron groups and surrounding atoms.

Table 4.2.1: Summary of Molecular Shapes

Number of Electron Groups on Central Atom	Number of Bonding Groups	Number of Lone Pairs	Electron Geometry	Molecular Shape
2	2	0	linear	linear
3	3	0	trigonal planar	trigonal planar
3	2	1	trigonal planar	bent
4	4	0	tetrahedral	tetrahedral
4	3	1	tetrahedral	trigonal pyramidal
4	2	2	tetrahedral	bent

✓ Example 4.2.1

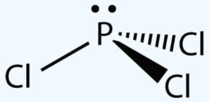
What is the approximate shape of each molecule?

a. PCl_3
b. NOF

Solution

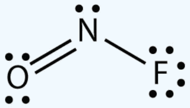
The first step is to draw the Lewis structure of the molecule.

For PCl_3, the electron dot diagram is as follows:



The lone electron pairs on the Cl atoms are omitted for clarity. The P atom has four electron groups with three of them bonded to surrounding atoms, so the molecular shape is trigonal pyramidal.

The electron dot diagram for NOF is as follows:



The N atom has three electron groups on it, two of which are bonded to other atoms. The molecular shape is bent.

? Exercise 4.2.1

What is the approximate molecular shape of CH_2Cl_2?

Answer

Tetrahedral

? Exercise 4.2.2

Ethylene (C_2H_4) has two central atoms. Determine the geometry around each central atom and the shape of the overall molecule. (Hint: hydrogen is a terminal atom.)

Answer

Trigonal planar about both central C atoms.

Summary

The approximate shape of a molecule can be predicted from the number of electron groups and the number of surrounding atoms.

4.3: Polarity of Bonds and Molecules

Learning Objectives

- Explain how polar compounds differ from nonpolar compounds.
- Determine if a molecule is polar or nonpolar.
- Given a pair of compounds, predict which would have a higher melting or boiling point.

Bond Polarity

A covalent bond forms when atoms share electrons. However, elements do not all have the same relative attraction for electrons when they are a part of a compound. This relative attraction for the electrons involved in a bond is known as electronegativity. Different scientists assigned electronegativity values based on different properties. The values in the different tables are not identical but follow the same trends across periods and down groups. Linus Pauling developed the first measurable scale for electronegativity and based the values on covalent bond energy calculations.

Differences in electronegativity values can be used to predict the primary nature of the bond ranging from pure covalent to ionic. Polar covalent bonds lie on the bonding continuum between pure covalent and ionic bonds. The larger the difference in electronegativity values, the greater the attraction for the electrons by one of the atoms, the more polar the bond.

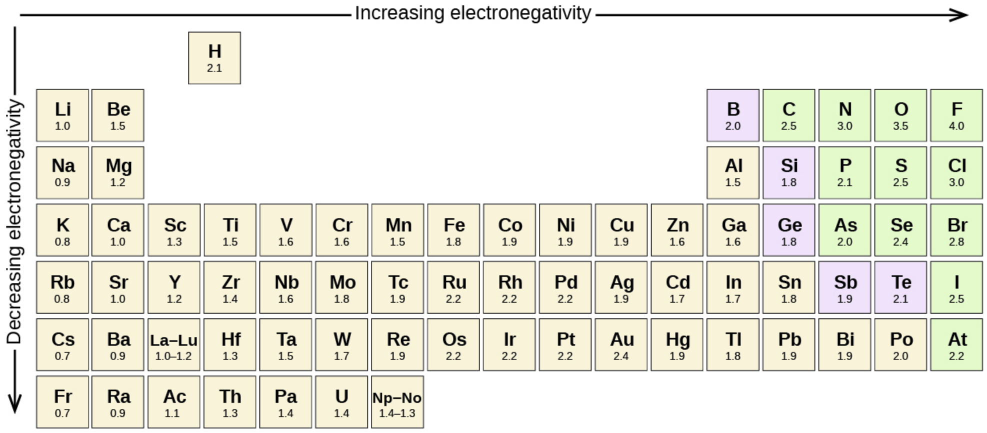


Figure 4.3.1: Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

The polarity of a covalent bond can be judged by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

Electronegativity Difference	Bond Type
0–0.4	pure covalent
0.5–2.0	polar covalent
>2.0	likely ionic

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.9 is considered to be mostly covalent in character. However, at this point we need to distinguish between two general types of covalent bonds. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons are shared equally between the two atoms. In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms.

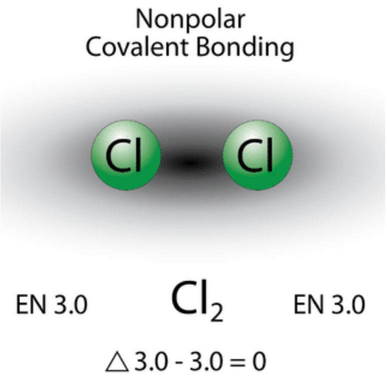


Figure 4.3.2: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal.

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the $\mathrm{Cl_2}$ molecule is symmetrical. Also note that molecules in which the electronegativity difference is very small (<0.5) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta \text{EN} = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.5 and 2.0 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical. The ΔEN difference of 2.0 as the upper limit between polar covalent and ionic is arbitrary rather than an absolute cut off and that the properties of the compound are the best indicator of the primary nature of the bond. The differences in electronegativity are most valuable when used to predict the relative polarity of covalent bonds.

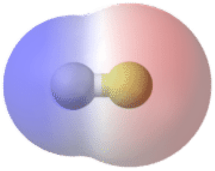


Figure 4.3.3: In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

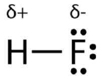


Figure 4.3.4: Use of δ to indicate partial charge.

The atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge. The delta symbol is used to indicate that the quantity of charge is less than one. A polar covalent bond results in electrical charges separated by a distance so it has a measurable dipole moment units of Coulomb-meters, C-m, or more commonly expressed in Debeye, D (1 D = 3.336 x 10^{-30} C-m). Dipole moments are vectors – quantities that have both magnitude and direction. An alternative notation to indicate a bond is polar and which end of the bond is partially positive and which is partially negative is to draw a vector (an arrow) with the arrow head pointing in the direction of the partially negative end of the bond. A vertical line is added to the end of the vector by the partially positive end of the bond so it appears to be a plus sign. The length of the arrow can be used to illustrate the relative polarity of the bond.

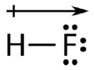


Figure 4.3.5: Use of crossed arrow to indicate polarity

Electronegativity differences in bonding using Pauling scale. Using differences in electronegativity to classify bonds as covalent, polar covalent, or ionic.

Example 4.3.1: Bond Polarity

What is the polarity of each bond?

a. C–H
b. O–H

Solution

Using Figure 4.3.1, we can calculate the difference of the electronegativities of the atoms involved in the bond.

a. For the C–H bond, the difference in the electronegativities is 2.5 − 2.1 = 0.4. Thus we predict that this bond will be non polar covalent.
b. For the O–H bond, the difference in electronegativities is 3.5 − 2.1 = 1.4, so we predict that this bond will be definitely polar covalent.

Exercise 4.3.1

What is the polarity of each bond?

a. Rb–F
b. P–Cl

Answer a

likely ionic

Answer b

polar covalent

Molecular Polarity

To determine if a molecule is polar or nonpolar, it is frequently useful to look at Lewis structures. **Nonpolar compounds** will be symmetric. For molecules with four or fewer total electron groups around the central atom, a symmetrical molecule is identical on all sides – the bonded atoms are identical and there are no unshared electrons on the central atom. **Polar molecules** are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded. This works pretty well - as long as you can visualize the molecular geometry. That's the hard part. To know how the bonds are oriented in space, you have to have a strong grasp of Lewis structures and VSEPR theory. Assuming you do, you can look at the structure of each one and decide if it is polar or not - *whether or not you know the individual atom electronegativity*. This is because you know that all bonds between dissimilar elements are polar, and in these particular examples, it doesn't matter which direction the dipole moment vectors are pointing (out or in).

A **polar molecule** is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. A molecule with two poles is called a dipole and has a net dipole moment. A diatomic molecule that consists of a polar covalent bond, such as HF, is a polar molecule. The two electrically charged regions on either end of the molecule are called poles, similar to a magnet having a north and a south pole. A molecule with two poles is called a **dipole** (see figure below). Hydrogen fluoride is a dipole.

Figure 4.3.6: A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule.

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. The figure below shows a comparison between carbon dioxide and water. Carbon dioxide $(\mathrm{CO_2})$ is a linear molecule. The oxygen atoms are more electronegative than the carbon atom, so there are two individual dipoles pointing outward from the C atom to each O atom. However, since the dipoles are of equal strength and are oriented this way, they cancel out and the overall molecular polarity of $\mathrm{CO_2}$ is zero.

Water is a bent molecule because of the two lone pairs on the central oxygen atom. The individual dipoles point from the H atoms toward the O atom. Because of the shape, the dipoles do not cancel each other out and the water molecule is polar. In the figure below, the net dipole is shown in blue and points upward.

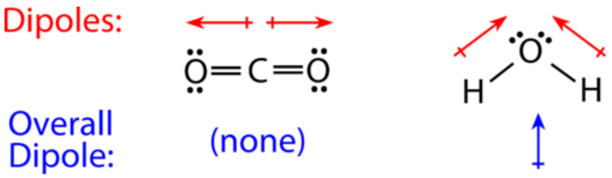


Figure 4.3.7: The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as $\mathrm{CH_4}$ is nonpolar. However, if one of the peripheral H atoms is replaced with another atom that has a different electronegativity, the molecule becomes polar. A trigonal planar molecule $(\mathrm{BF_3})$ may be nonpolar if all three peripheral atoms are the same, but a trigonal pyramidal molecule $(\mathrm{NH_3})$ is polar.

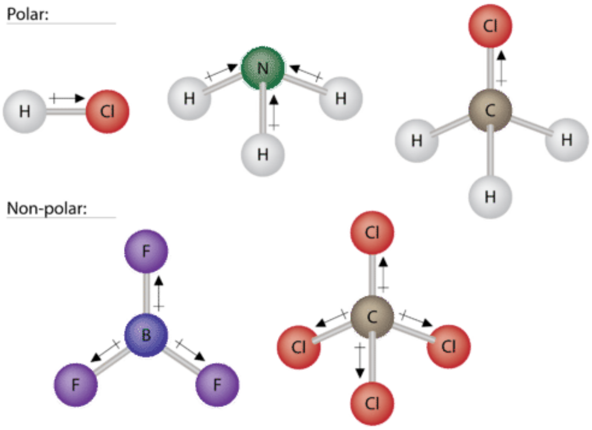


Figure 4.3.8: Some examples of polar and nonpolar molecules based on molecular geometry.

Steps to Identify Polar Molecules

1. Draw the Lewis structure
2. Figure out the geometry (using VSEPR theory)
3. Visualize or draw the geometry
4. Find the net dipole moment (you don't have to actually do calculations if you can visualize it)

5. If the net dipole moment is zero, it is non-polar. Otherwise, it is polar.

Properties of Polar Molecules

Polar molecules tend to align when placed in an electric field with the positive end of the molecule oriented toward the negative plate and the negative end toward the positive plate (Figure $4.3.14$). We can use an electrically charged object to attract polar molecules, but nonpolar molecules are not attracted. Also, polar solvents are better at dissolving polar substances, and nonpolar solvents are better at dissolving nonpolar substances.

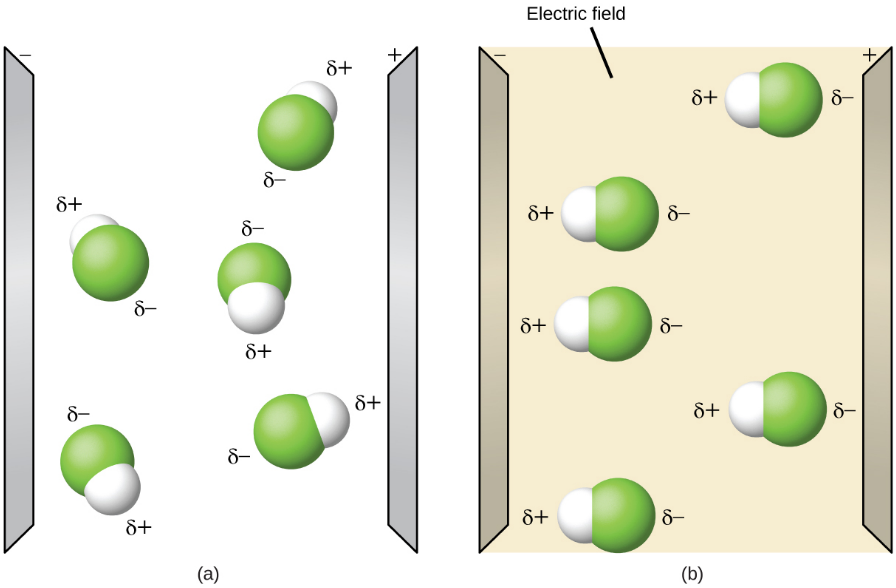


Figure $4.3.9$: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction. (OpenStax CC-BY-SA);

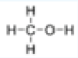
While molecules can be described as "polar covalent" or "ionic", it must be noted that this is often a relative term, with one molecule simply being *more polar* or *less polar* than another. However, the following properties are typical of such molecules. Polar molecules tend to:

- have higher melting points than nonpolar molecules
- have higher boiling points than nonpolar molecules
- be more soluble in water (dissolve better) than nonpolar molecules
- have lower vapor pressures than nonpolar molecules

Example $4.3.2$:

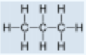
Label each of the following as polar or nonpolar.

a. Water, H_2O: 

b. Methanol, CH_3OH: 

c. Hydrogen Cyanide, HCN: 

d. Oxygen, O_2: 

e. Propane, C_3H_8: 

Solution

a. Water is polar. Any molecule with lone pairs of electrons around the central atom is polar.
b. Methanol is polar. This is not a symmetric molecule. The $-\mathrm{OH}$ side is different from the other 3 $-\mathrm{H}$ sides.
c. Hydrogen cyanide is polar. The molecule is not symmetric. The nitrogen and hydrogen have different electronegativities, creating an uneven pull on the electrons.

d. Oxygen is nonpolar. The molecule is symmetric. The two oxygen atoms pull on the electrons by exactly the same amount.
e. Propane is nonpolar, because it is symmetric, with H atoms bonded to every side around the central atoms and no unshared pairs of electrons.

Exercise $4.3.2$

Label each of the following as polar or nonpolar.

a. SO_3

b. NH_3

Answer a

non polar

Answer b

polar

Contributions & Attributions

This page was constructed from content via the following contributor(s) and edited (topically or extensively) by the LibreTexts development team to meet platform style, presentation, and quality:

-
- StackExchange (thomij).
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

4.4: What makes molecules stick together? -- Intermolecular Forces

Learning Objectives

- To describe the intermolecular forces in liquids.

The properties of liquids are intermediate between those of gases and solids, but are more similar to solids. In contrast to *intra*molecular forces, such as the covalent bonds that hold atoms together in molecules and polyatomic ions, *inter*molecular forces hold molecules together in a liquid or solid. Intermolecular forces are generally much weaker than covalent bonds. For example, it requires 927 kJ to overcome the intramolecular forces and break both O–H bonds in 1 mol of water, but it takes only about 41 kJ to overcome the intermolecular attractions and convert 1 mol of liquid water to water vapor at 100°C. (Despite this seemingly low value, the intermolecular forces in liquid water are among the strongest such forces known!) Given the large difference in the strengths of intra- and intermolecular forces, changes between the solid, liquid, and gaseous states almost invariably occur for molecular substances *without breaking covalent bonds*.

> *The properties of liquids are intermediate between those of gases and solids but are more similar to solids.*

Intermolecular forces determine bulk properties such as the melting points of solids and the boiling points of liquids. Liquids boil when the molecules have enough thermal energy to overcome the intermolecular attractive forces that hold them together, thereby forming bubbles of vapor within the liquid. Similarly, solids melt when the molecules acquire enough thermal energy to overcome the intermolecular forces that lock them into place in the solid.

Intermolecular forces are electrostatic in nature; that is, they arise from the interaction between positively and negatively charged species. Like covalent and ionic bonds, intermolecular interactions are the sum of both attractive and repulsive components. Because electrostatic interactions fall off rapidly with increasing distance between molecules, intermolecular interactions are most important for solids and liquids, where the molecules are close together.

In this section, we explicitly consider three intermolecular forces of attraction between neutral molecules: London dispersion, dipole-dipole and hydrogen bonding.

London Dispersion Forces

Thus far we have considered only interactions between polar molecules, but other factors must be considered to explain why many nonpolar molecules, such as bromine, benzene, and hexane, are liquids at room temperature, and others, such as iodine and naphthalene, are solids. Even the noble gases can be liquefied or solidified at low temperatures, high pressures, or both (Table 4.4.2).

What kind of attractive forces can exist between nonpolar molecules or atoms? This question was answered by Fritz London (1900–1954), a German physicist who later worked in the United States. In 1930, London proposed that temporary fluctuations in the electron distributions within atoms and nonpolar molecules could result in the formation of short-lived instantaneous dipole moments, which produce attractive forces called London dispersion forces between otherwise nonpolar substances.

Table 4.4.2: Normal Melting and Boiling Points of Some Elements and Nonpolar Compounds

Substance	Molar Mass (g/mol)	Melting Point (°C)	Boiling Point (°C)
Ar	40	−189.4	−185.9
Xe	131	−111.8	−108.1
N_2	28	−210	−195.8
O_2	32	−218.8	−183.0
F_2	38	−219.7	−188.1

Substance	Molar Mass (g/mol)	Melting Point (°C)	Boiling Point (°C)
I_2	254	113.7	184.4
CH_4	16	−182.5	−161.5

Consider a pair of adjacent He atoms, for example. On average, the two electrons in each He atom are uniformly distributed around the nucleus. Because the electrons are in constant motion, however, their distribution in one atom is likely to be asymmetrical at any given instant, resulting in an instantaneous dipole moment. As shown in part (a) in Figure 4.4.3, the instantaneous dipole moment on one atom can interact with the electrons in an adjacent atom, pulling them toward the positive end of the instantaneous dipole or repelling them from the negative end. The net effect is that the first atom causes the temporary formation of a dipole, called an induced dipole, in the second. Interactions between these temporary dipoles cause atoms to be attracted to one another. These attractive interactions are weak and fall off rapidly with increasing distance. London was able to show with quantum mechanics that the attractive energy between molecules due to temporary dipole–induced dipole interactions falls off as $1/r^6$. Doubling the distance therefore decreases the attractive energy by 2^6, or 64-fold.

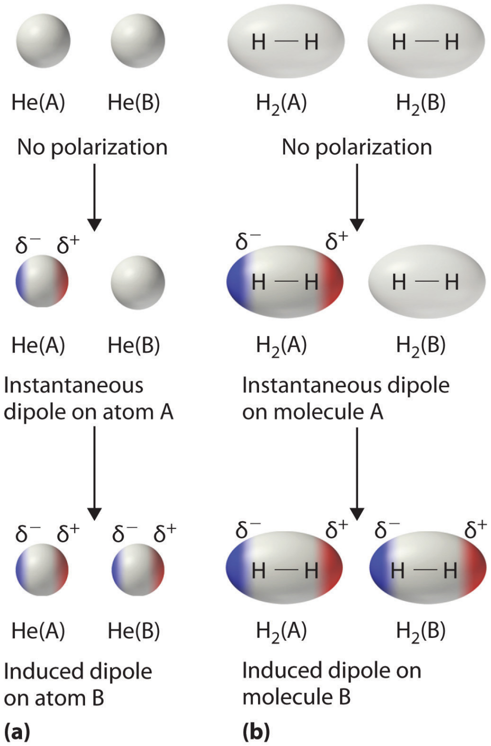


Figure 4.4.3: Instantaneous Dipole Moments. The formation of an instantaneous dipole moment on one He atom (a) or an H_2 molecule (b) results in the formation of an induced dipole on an adjacent atom or molecule.

Instantaneous dipole–induced dipole interactions between nonpolar molecules can produce intermolecular attractions just as they produce interatomic attractions in monatomic substances like Xe. This effect, illustrated for two H_2 molecules in part (b) in Figure 4.4.3, tends to become more pronounced as atomic and molecular masses increase (Table 4.4.2). For example, Xe boils at −108.1°C, whereas He boils at −269°C. The reason for this trend is that the strength of London dispersion forces is related to the ease with which the electron distribution in a given atom can be perturbed. In small atoms such as He, the two 1*s* electrons are held close to the nucleus in a very small volume, and electron–electron repulsions are strong enough to prevent significant asymmetry in their distribution. In larger atoms such as Xe, however, the outer electrons are much less strongly attracted to the nucleus because of filled intervening shells. As a result, it is relatively easy to temporarily deform the electron distribution to generate an instantaneous or induced dipole. The ease of deformation of the electron distribution in an atom or molecule is called its polarizability. Because the electron distribution is more easily perturbed in large, heavy species than in small, light species, we say that heavier substances tend to be much more *polarizable* than lighter ones.

> *For similar substances, London dispersion forces get stronger with increasing molecular size.*

The polarizability of a substance also determines how it interacts with ions and species that possess permanent dipoles. Thus London dispersion forces are responsible for the general trend toward higher boiling points with increased molecular mass and greater surface area in a homologous series of compounds, such as the alkanes (part (a) in Figure 4.4.4). The strengths of London dispersion forces also depend significantly on molecular shape because shape determines how much of one molecule can interact with its neighboring molecules at any given time. For example, part (b) in Figure 4.4.4 shows 2,2-dimethylpropane (neopentane) and *n*-pentane, both of which have the empirical formula C_5H_{12}. Neopentane is almost spherical, with a small surface area for intermolecular interactions, whereas *n*-pentane has an extended conformation that enables it to come into close contact with other *n*-pentane molecules. As a result, the boiling point of neopentane (9.5°C) is more than 25°C lower than the boiling point of *n*-pentane (36.1°C).

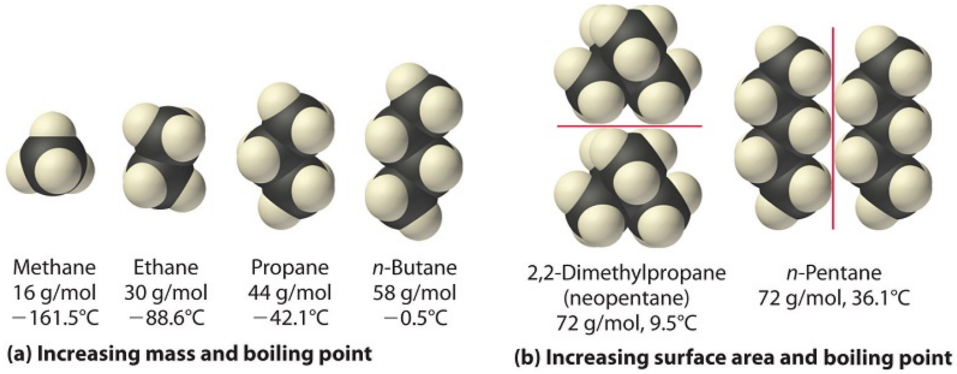


Figure 4.4.4: Mass and Surface Area Affect the Strength of London Dispersion Forces. (a) In this series of four simple alkanes, larger molecules have stronger London forces between them than smaller molecules and consequently higher boiling points. (b) Linear *n*-pentane molecules have a larger surface area and stronger intermolecular forces than spherical neopentane molecules. As a result, neopentane is a gas at room temperature, whereas *n*-pentane is a volatile liquid.

All molecules, whether polar or nonpolar, are attracted to one another by London dispersion forces in addition to any other attractive forces that may be present. In general, however, dipole–dipole interactions in small polar molecules are significantly stronger than London dispersion forces, so the former predominate.

Example 4.4.2

Arrange *n*-butane, propane, 2-methylpropane [isobutene, $(CH_3)_2CHCH_3$], and *n*-pentane in order of increasing boiling points.

Given: compounds

Asked for: order of increasing boiling points

Strategy:

Determine the intermolecular forces in the compounds and then arrange the compounds according to the strength of those forces. The substance with the weakest forces will have the lowest boiling point.

Solution:

The four compounds are alkanes and nonpolar, so London dispersion forces are the only important intermolecular forces. These forces are generally stronger with increasing molecular mass, so propane should have the lowest boiling point and *n*-pentane should have the highest, with the two butane isomers falling in between. Of the two butane isomers, 2-methylpropane is more compact, and *n*-butane has the more extended shape. Consequently, we expect intermolecular interactions for *n*-butane to be stronger due to its larger surface area, resulting in a higher boiling point. The overall order is thus as follows, with actual boiling points in parentheses: propane (−42.1°C) < 2-methylpropane (−11.7°C) < *n*-butane (−0.5°C) < *n*-pentane (36.1°C).

Exercise 4.4.2

Arrange GeH_4, $SiCl_4$, SiH_4, CH_4, and $GeCl_4$ in order of decreasing boiling points.

Answer

$GeCl_4$ (87°C) > $SiCl_4$ (57.6°C) > GeH_4 (−88.5°C) > SiH_4 (−111.8°C) > CH_4 (−161°C)

Dipole–Dipole Interactions

Polar covalent bonds behave as if the bonded atoms have localized fractional charges that are equal but opposite (i.e., the two bonded atoms generate a *dipole*). If the structure of a molecule is such that the individual bond dipoles do not cancel one another, then the molecule has a net dipole moment. Molecules with net dipole moments tend to align themselves so that the positive end of one dipole is near the negative end of another and vice versa, as shown in Figure $4.4.1a$.

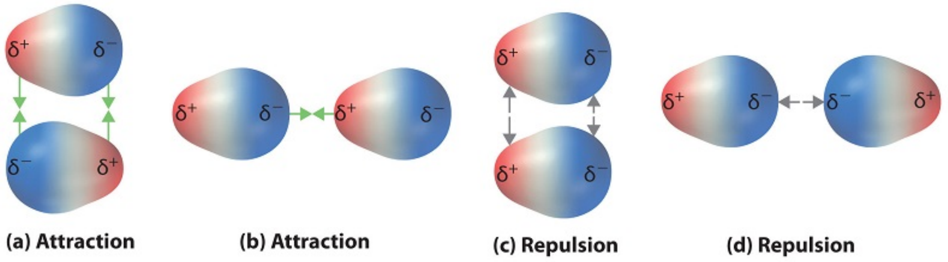


Figure $4.4.1$: Attractive and Repulsive Dipole–Dipole Interactions. (a and b) Molecular orientations in which the positive end of one dipole (δ^+) is near the negative end of another (δ^-) (and vice versa) produce attractive interactions. (c and d) Molecular orientations that juxtapose the positive or negative ends of the dipoles on adjacent molecules produce repulsive interactions.

These arrangements are more stable than arrangements in which two positive or two negative ends are adjacent (Figure $4.4.1c$). Hence dipole–dipole interactions, such as those in Figure $4.4.1b$, are *attractive intermolecular interactions*, whereas those in Figure $4.4.1d$ are *repulsive intermolecular interactions*. Because molecules in a liquid move freely and continuously, molecules always experience both attractive and repulsive dipole–dipole interactions simultaneously, as shown in Figure $4.4.2$. On average, however, the attractive interactions dominate.

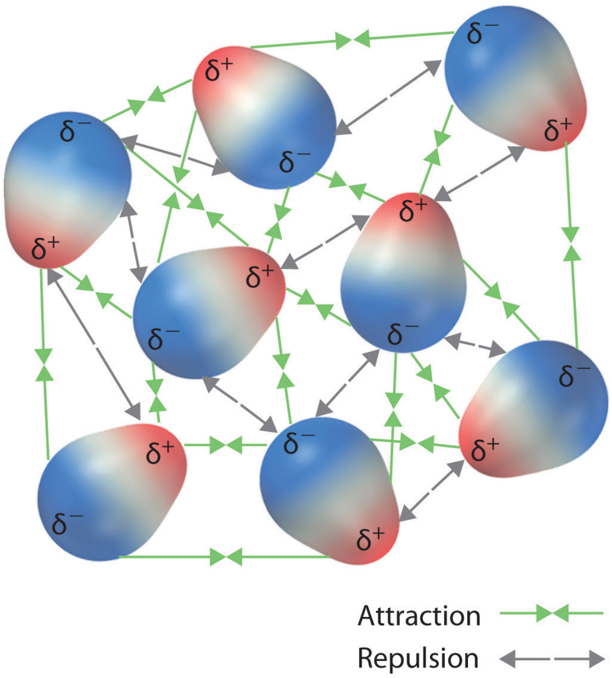


Figure $4.4.2$: Both Attractive and Repulsive Dipole–Dipole Interactions Occur in a Liquid Sample with Many Molecules

Because each end of a dipole possesses only a fraction of the charge of an electron, dipole–dipole interactions are substantially weaker than the interactions between two ions, each of which has a charge of at least ±1, or between a dipole and an ion, in which one of the species has at least a full positive or negative charge. In addition, the attractive interaction between dipoles falls off much more rapidly with increasing distance than do the ion–ion interactions. Recall that the attractive energy between two ions is proportional to $1/r$, where r is the distance between the ions. Doubling the distance ($r \rightarrow 2r$) decreases the attractive energy by one-half. In contrast, the energy of the interaction of two dipoles is proportional to $1/r^3$, so doubling the distance between the dipoles decreases the strength of the interaction by 2^3, or 8-fold. Thus a substance such as HCl, which is partially held together by dipole–dipole interactions, is a gas at room temperature and 1 atm pressure, whereas $NaCl$, which is held together by interionic interactions, is a high-melting-point solid. Within a series of compounds of similar molar mass, the strength of the intermolecular interactions increases as the dipole moment of the molecules increases, as shown in Table 4.4.1.

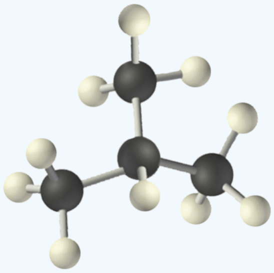
Table 4.4.1: Relationships between the Dipole Moment and the Boiling Point for Organic Compounds of Similar Molar Mass

Compound	Molar Mass (g/mol)	Dipole Moment (D)	Boiling Point (K)
C_3H_6 (cyclopropane)	42	0	240
CH_3OCH_3 (dimethyl ether)	46	1.30	248
CH_3CN (acetonitrile)	41	3.9	355

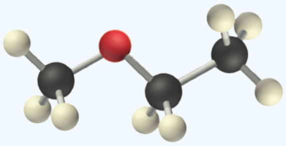
> *The attractive energy between two ions is proportional to 1/r, whereas the attractive energy between two dipoles is proportional to 1/r6.*

Example 4.4.1

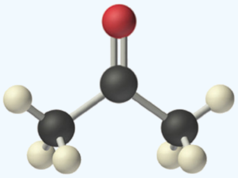
Arrange ethyl methyl ether ($CH_3OCH_2CH_3$), 2-methylpropane [isobutane, $(CH_3)_2CHCH_3$], and acetone (CH_3COCH_3) in order of increasing boiling points. Their structures are as follows:



2-Methylpropane



Ethyl methyl ether



Acetone

Given: compounds

Asked for: order of increasing boiling points

Strategy:

Compare the molar masses and the polarities of the compounds. Compounds with higher molar masses and that are polar will have the highest boiling points.

Solution:

The three compounds have essentially the same molar mass (58–60 g/mol), so we must look at differences in polarity to predict the strength of the intermolecular dipole–dipole interactions and thus the boiling points of the compounds.

The first compound, 2-methylpropane, contains only C–H bonds, which are not very polar because C and H have similar electronegativities. It should therefore have a very small (but nonzero) dipole moment and a very low boiling point.

Ethyl methyl ether has a structure similar to H_2O; it contains two polar C–O single bonds oriented at about a 109° angle to each other, in addition to relatively nonpolar C–H bonds. As a result, the C–O bond dipoles partially reinforce one another and generate a significant dipole moment that should give a moderately high boiling point.

Acetone contains a polar C=O double bond oriented at about 120° to two methyl groups with nonpolar C–H bonds. The C–O bond dipole therefore corresponds to the molecular dipole, which should result in both a rather large dipole moment and a high boiling point.

Thus we predict the following order of boiling points:

$$\text{2-methylpropane} < \text{ethyl methyl ether} < \text{acetone.}$$

This result is in good agreement with the actual data: 2-methylpropane, boiling point = −11.7°C, and the dipole moment (μ) = 0.13 D; methyl ethyl ether, boiling point = 7.4°C and μ = 1.17 D; acetone, boiling point = 56.1°C and μ = 2.88 D.

Exercise 4.4.1

Arrange carbon tetrafluoride (CF_4), ethyl methyl sulfide ($CH_3SC_2H_5$), dimethyl sulfoxide [$(CH_3)_2S{=}O$], and 2-methylbutane [isopentane, $(CH_3)_2CHCH_2CH_3$] in order of decreasing boiling points.

Answer

dimethyl sulfoxide (boiling point = 189.9°C) > ethyl methyl sulfide (boiling point = 67°C) > 2-methylbutane (boiling point = 27.8°C) > carbon tetrafluoride (boiling point = −128°C)

Hydrogen Bonds

When hydrogen is covalently bonded to a small, highly electronegative element (N, O and F) in a molecule, the molecules are capable of interacting through a particularly strong dipole-dipole interaction. The large difference in electronegativity between the hydrogen atom and N, O or F results in a highly polar bond with the hydrogen end of the bond partially positive and the N, O or F end partially negative. Additionally, since the atoms involved in this bond are so small, they can approach the comparable atoms in another molecule closely. The combination of large bond dipoles and short dipole–dipole distances results in very strong dipole–dipole interactions called hydrogen bonds, as shown for ice in Figure 4.4.6. A hydrogen bond is usually indicated by a dotted line between the hydrogen atom attached to O, N, or F (the *hydrogen bond donor*) and the atom that has the lone pair of electrons (the *hydrogen bond acceptor*). Because each water molecule contains two hydrogen atoms and two lone pairs, a tetrahedral arrangement maximizes the number of hydrogen bonds that can be formed. In the structure of ice, each oxygen atom is surrounded by a distorted tetrahedron of hydrogen atoms that form bridges to the oxygen atoms of adjacent water molecules. The bridging hydrogen atoms are *not* equidistant from the two oxygen atoms they connect, however. Instead, each hydrogen atom is 101 pm from one oxygen and 174 pm from the other. In contrast, each oxygen atom is bonded to two H atoms at the shorter distance and two at the longer distance, corresponding to two O–H covalent bonds and two O⋯H hydrogen bonds from adjacent water molecules, respectively. The resulting open, cagelike structure of ice means that the solid is actually slightly less dense than the liquid, which explains why ice floats on water rather than sinks.

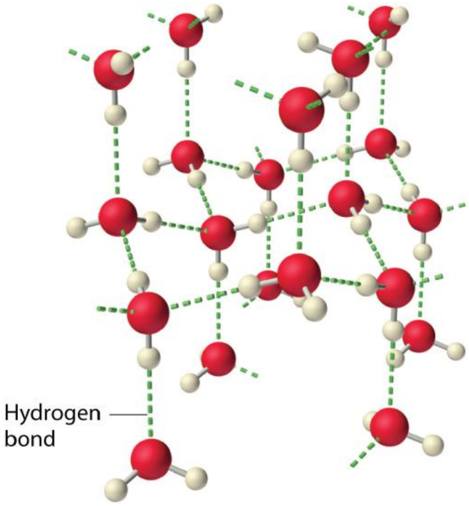


Figure 4.4.6: The Hydrogen-Bonded Structure of Ice.

Each water molecule accepts two hydrogen bonds from two other water molecules and donates two hydrogen atoms to form hydrogen bonds with two more water molecules, producing an open, cagelike structure. The structure of liquid water is very similar, but in the liquid, the hydrogen bonds are continually broken and formed because of rapid molecular motion.

> *Hydrogen bond formation requires both a hydrogen bond donor and a hydrogen bond acceptor.*

Because ice is less dense than liquid water, rivers, lakes, and oceans freeze from the top down. In fact, the ice forms a protective surface layer that insulates the rest of the water, allowing fish and other organisms to survive in the lower levels of a frozen lake or sea. If ice were denser than the liquid, the ice formed at the surface in cold weather would sink as fast as it formed. Bodies of water would freeze from the bottom up, which would be lethal for most aquatic creatures. The expansion of water when freezing also explains why automobile or boat engines must be protected by “antifreeze” and why unprotected pipes in houses break if they are allowed to freeze.

Example 4.4.3

Considering CH_3OH, C_2H_6, Xe, and $(CH_3)_3N$, which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Given: compounds

Asked for: formation of hydrogen bonds and structure

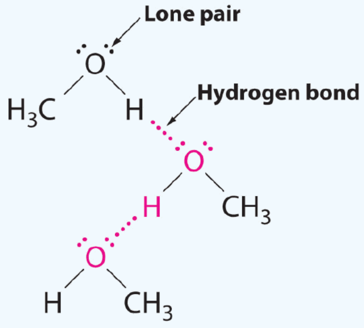
Strategy:

A. Identify the compounds with a hydrogen atom attached to O, N, or F. These are likely to be able to act as hydrogen bond donors.
B. Of the compounds that can act as hydrogen bond donors, identify those that also contain lone pairs of electrons, which allow them to be hydrogen bond acceptors. If a substance is both a hydrogen donor and a hydrogen bond acceptor, draw a structure showing the hydrogen bonding.

Solution:

A Of the species listed, xenon (Xe), ethane (C_2H_6), and trimethylamine [$(CH_3)_3N$] do not contain a hydrogen atom attached to O, N, or F; hence they cannot act as hydrogen bond donors.

B The one compound that can act as a hydrogen bond donor, methanol (CH_3OH), contains both a hydrogen atom attached to O (making it a hydrogen bond donor) and two lone pairs of electrons on O (making it a hydrogen bond acceptor); methanol can thus form hydrogen bonds by acting as either a hydrogen bond donor or a hydrogen bond acceptor. The hydrogen-bonded structure of methanol is as follows:



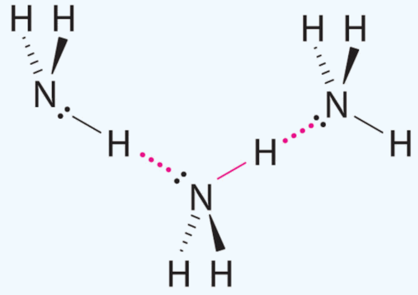
Hydrogen bonding in methanol

Exercise 4.4.3

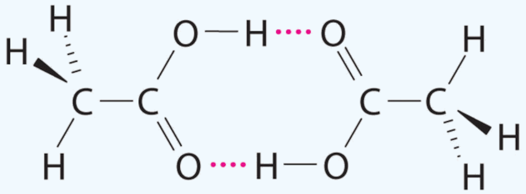
Considering CH_3CO_2H, $(CH_3)_3N$, NH_3, and CH_3F, which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Answer

CH_3CO_2H and NH_3;



Hydrogen bonding in ammonia



Hydrogen bonding in acetic acid

Although hydrogen bonds are significantly weaker than covalent bonds, with typical dissociation energies of only 15–25 kJ/mol, they have a significant influence on the physical properties of a compound. Compounds such as HF can form only two hydrogen bonds at a time as can, on average, pure liquid NH_3. Consequently, even though their molecular masses are similar to that of water, their boiling points are significantly lower than the boiling point of water, which forms *four* hydrogen bonds at a time.

Example 4.4.4: Buckyballs

Arrange C_{60} (buckminsterfullerene, which has a cage structure), NaCl, He, Ar, and N_2O in order of increasing boiling points.

Given: compounds

Asked for: order of increasing boiling points

Strategy:

Identify the intermolecular forces in each compound and then arrange the compounds according to the strength of those forces. The substance with the weakest forces will have the lowest boiling point.

Solution:

Electrostatic interactions are strongest for an ionic compound, so we expect NaCl to have the highest boiling point. To predict the relative boiling points of the other compounds, we must consider their polarity (for dipole–dipole interactions), their ability to form hydrogen bonds, and their molar mass (for London dispersion forces). Helium is nonpolar and by far the lightest, so it should have the lowest boiling point. Argon and N_2O have very similar molar masses (40 and 44 g/mol, respectively), but N_2O is polar while Ar is not. Consequently, N_2O should have a higher boiling point. A C_{60} molecule is nonpolar, but its molar mass is 720 g/mol, much greater than that of Ar or N_2O. Because the boiling points of nonpolar substances increase rapidly with molecular mass, C_{60} should boil at a higher temperature than the other nonionic substances. The predicted order is thus as follows, with actual boiling points in parentheses:

$$\text{He } (-269°\text{C}) < \text{Ar } (-185.7°\text{C}) < N_2O\ (-88.5°\text{C}) < C_{60}\ (>280°\text{C}) < \text{NaCl } (1465°\text{C}).$$

Exercise 4.4.4

Arrange 2,4-dimethylheptane, Ne, CS_2, Cl_2, and KBr in order of decreasing boiling points.

Answer

KBr (1435°C) > 2,4-dimethylheptane (132.9°C) > CS_2 (46.6°C) > Cl_2 (−34.6°C) > Ne (−246°C)

Example 4.4.5:

Identify the most significant intermolecular force in each substance.

a. C_3H_8
b. CH_3OH
c. H_2S

Solution

a. Although C–H bonds are polar, they are only minimally polar. The most significant intermolecular force for this substance would be dispersion forces.

b. This molecule has an H atom bonded to an O atom, so it will experience hydrogen bonding.

c. Although this molecule does not experience hydrogen bonding, the Lewis electron dot diagram and VSEPR indicate that it is bent, so it has a permanent dipole. The most significant force in this substance is dipole-dipole interaction.

Exercise 4.4.6

Identify the most significant intermolecular force in each substance.

a. HF
b. HCl

Answer a

hydrogen bonding

Answer b

dipole-dipole interactions

Summary

Intermolecular forces are electrostatic in nature and include van der Waals forces and hydrogen bonds. Molecules in liquids are held to other molecules by intermolecular interactions, which are weaker than the intramolecular interactions that hold the atoms together within molecules and polyatomic ions. Transitions between the solid and liquid or the liquid and gas phases are due to changes in intermolecular interactions but do not affect intramolecular interactions. The three major types of intermolecular interactions are dipole–dipole interactions, London dispersion forces (these two are often referred to collectively as **van der Waals forces**), and hydrogen bonds. **Dipole–dipole interactions** arise from the electrostatic interactions of the positive and negative ends of molecules with permanent dipole moments; their strength is proportional to the magnitude of the dipole moment and to $1/r^3$, where r is the distance between dipoles. **London dispersion forces** are due to the formation of **instantaneous dipole moments** in polar or nonpolar molecules as a result of short-lived fluctuations of electron charge distribution, which in turn cause the temporary formation of an **induced dipole** in adjacent molecules. their energy falls off as $1/r^6$. Larger atoms tend to be more **polarizable** than smaller ones because their outer electrons are less tightly bound and are therefore more easily perturbed. **Hydrogen bonds** are especially strong dipole–dipole interactions between molecules that have hydrogen bonded to a highly electronegative atom, such as O, N, or F. The resulting partially positively charged H atom on one molecule (the *hydrogen bond donor*) can interact strongly with a lone pair of electrons of a partially negatively charged O, N, or F atom on adjacent molecules (the *hydrogen bond acceptor*). Because of strong O⋯H hydrogen bonding between water molecules, water has an unusually high boiling point, and ice has an open, cagelike structure that is less dense than liquid water.

4.5: Electron Transfer - Ionic Bonds

Learning Objectives

- State the octet rule.
- Define *ionic bond*.
- Demonstrate electron transfer between atoms to form ionic bonds.

In Section 1.17, we saw how ions are formed by losing electrons to make cations or by gaining electrons to form anions. The astute reader may have noticed something: Many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the *lower* shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged. For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of electrons. The trend that atoms like to have eight electrons in their valence shell is called the **octet rule**. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

It is not impossible to violate the octet rule. Consider sodium: in its elemental form, it has one valence electron and is stable. It is rather reactive, however, and does not require a lot of energy to remove that electron to make the Na^+ ion. We *could* remove another electron by adding even more energy to the ion, to make the Na^{2+} ion. However, that requires much more energy than is normally available in chemical reactions, so sodium stops at a 1+ charge after losing a single electron. It turns out that the Na^+ ion has a complete octet in its new valence shell, the n = 2 shell, which satisfies the octet rule. The octet rule is a result of trends in energies and is useful in explaining why atoms form the ions that they do.

Now consider an Na atom in the presence of a Cl atom. The two atoms have these Lewis electron dot diagrams and electron configurations:

$$\mathbf{Na}\cdot \qquad \cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}:$$

$$[Ne]\,3s^1 \qquad [Ne]\,3s^23p^5$$

For the Na atom to obtain an octet, it must lose an electron; for the Cl atom to gain an octet, it must gain an electron. An electron transfers from the Na atom to the Cl atom:

$$\mathbf{Na}\cdot \curvearrowright \cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}:$$

resulting in two ions-the Na^+ ion and the Cl^- ion:

$$\mathbf{Na}\cdot^{+} \qquad :\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}:^{-}$$

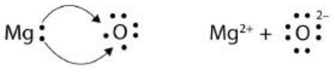
$$[Ne] \qquad [Ne]\,3s^23p^6$$

Both species now have complete octets, and the electron shells are energetically stable. From basic physics, we know that opposite charges attract. This is what happens to the Na^+ and Cl^- ions:

$$\mathbf{Na}\cdot^{+} + :\overset{\cdot\cdot}{\underset{\cdot\cdot}{\mathbf{Cl}}}:^{-} \rightarrow Na^{+}Cl^{-} \quad or \quad NaCl$$

where we have written the final formula (the formula for sodium chloride) as per the convention for ionic compounds, without listing the charges explicitly. The attraction between oppositely charged ions is called an **ionic bond,** and it is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by electrons *transferring* from one atom to another.

In electron transfer, the number of electrons lost must equal the number of electrons gained. We saw this in the formation of NaCl. A similar process occurs between Mg atoms and O atoms, except in this case two electrons are transferred:



The two ions each have octets as their valence shell, and the two oppositely charged particles attract, making an ionic bond:

$$\mathbf{Mg}^{2+} + :\ddot{\underset{\cdot\cdot}{\mathbf{O}}}:^{2-} \qquad Mg^{2+}O^{2-} \text{ or } MgO$$

Remember, in the final formula for the ionic compound, we do not write the charges on the ions.

What about when an Na atom interacts with an O atom? The O atom needs two electrons to complete its valence octet, but the Na atom supplies only one electron:

$$\mathbf{Na}\cdot \curvearrowright \cdot\ddot{\underset{\cdot}{\mathbf{O}}}:$$

The O atom still does not have an octet of electrons. What we need is a second Na atom to donate a second electron to the O atom:

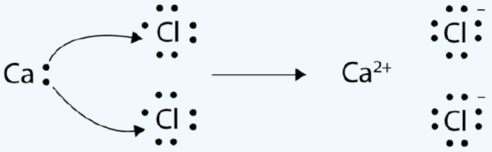
These three ions attract each other to give an overall neutral-charged ionic compound, which we write as Na_2O. The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the law of conservation of matter as well.

Example 4.5.1

With arrows, illustrate the transfer of electrons to form calcium chloride from Ca atoms and Cl atoms.

Solution

A Ca atom has two valence electrons, while a Cl atom has seven electrons. A Cl atom needs only one more to complete its octet, while Ca atoms have two electrons to lose. Thus we need two Cl atoms to accept the two electrons from one Ca atom. The transfer process looks as follows:

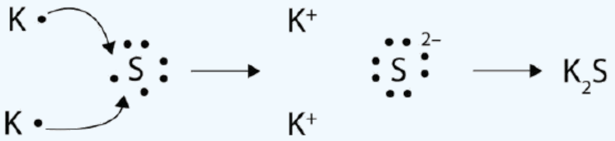


The oppositely charged ions attract each other to make $CaCl_2$.

Exercise 4.5.1

With arrows, illustrate the transfer of electrons to form potassium sulfide from K atoms and S atoms.

Answer



The strength of ionic bonding depends on two major characteristics: the magnitude of the charges and the size of the ion. The greater the magnitude of the charge, the stronger the ionic bond. The smaller the ion, the stronger the ionic bond (because a smaller ion size allows the ions to get closer together). The measured strength of ionic bonding is called the **lattice energy**. Some lattice energies are given in Table 4.5.1 - Lattice Energies of Some Ionic Compounds.

Table 4.5.1: Lattice Energies of Some Ionic Compounds

Compound	Lattice Energy (kJ/mol)
LiF	1,036

Compound	Lattice Energy (kJ/mol)
LiCl	853
NaCl	786
NaBr	747
MgF_2	2,957
Na_2O	2,481
MgO	3,791

Chemistry Is Everywhere: Salt

The element sodium (part [a] in the accompanying figure) is a very reactive metal; given the opportunity, it will react with the sweat on your hands and form sodium hydroxide, which is a very corrosive substance. The element chlorine (part [b] in the accompanying figure) is a pale yellow, corrosive gas that should not be inhaled due to its poisonous nature. Bring these two hazardous substances together, however, and they react to make the ionic compound sodium chloride (part [c] in the accompanying figure), known simply as salt.

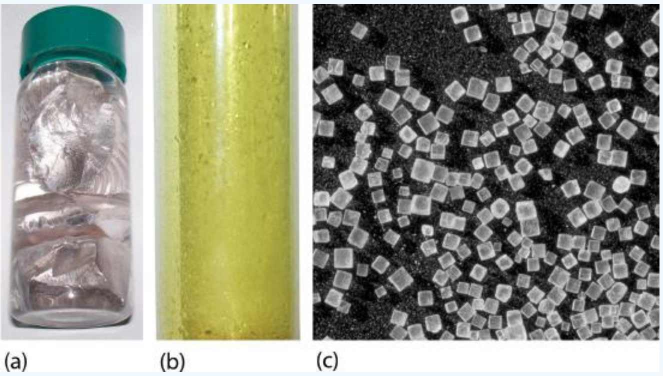


Figure 4.5.1: Sodium + Chlorine = Sodium Chloride (a) Sodium is a very reactive metal. (b) Chlorine is a pale yellow, noxious gas. (c) Together, sodium and chlorine make sodium chloride-salt-which is necessary for our survival. Source: Photo on the left courtesy of Greenhorn1. Photo in the center courtesy of Benjah-bmm27, . Photo on the right © Thinkstock.

Salt is necessary for life. Na^+ ions are one of the main ions in the human body and are necessary to regulate the fluid balance in the body. $Cl^−$ ions are necessary for proper nerve function and respiration. Both of these ions are supplied by salt. The taste of salt is one of the fundamental tastes; salt is probably the most ancient flavoring known, and one of the few rocks we eat.

The health effects of too much salt are still under debate, although a 2010 report by the US Department of Agriculture concluded that "excessive sodium intake…raises blood pressure, a well-accepted and extraordinarily common risk factor for stroke, coronary heart disease, and kidney disease."US Department of Agriculture Committee for Nutrition Policy and Promotion, Report of the Dietary Guidelines Advisory Committee on the Dietary Guidelines for Americans, accessed January 5, 2010. It is clear that most people ingest more salt than their bodies need, and most nutritionists recommend curbing salt intake. Curiously, people who suffer from low salt (called *hyponatria*) do so not because they ingest too little salt but because they drink too much water. Endurance athletes and others involved in extended strenuous exercise need to watch their water intake so their body's salt content is not diluted to dangerous levels.

- The tendency to form species that have eight electrons in the valence shell is called the octet rule.
- The attraction of oppositely charged ions caused by electron transfer is called an ionic bond.
- The strength of ionic bonding depends on the magnitude of the charges and the sizes of the ions.

4.6: Writing Formulas for Ionic Compounds

Learning Objectives

- Write the correct formula for an ionic compound.
- Recognize polyatomic ions in chemical formulas.

Ionic compounds do not exist as molecules. In the solid state, ionic compounds are in crystal lattice containing many ions each of the cation and anion. An ionic formula, like $NaCl$, is an empirical formula. This formula merely indicates that sodium chloride is made of an equal number of sodium and chloride ions. Sodium sulfide, another ionic compound, has the formula Na_2S. This formula indicates that this compound is made up of twice as many sodium ions as sulfide ions. This section will teach you how to find the correct ratio of ions, so that you can write a correct formula.

If you know the name of a binary ionic compound, you can write its **chemical formula**. Start by writing the metal ion with its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many of each ion is needed in order for the positive and negative charge to cancel each other out.

Example 4.6.1: Aluminum Nitride and Lithium Oxide

Write the formula for aluminum nitride and lithium oxide.

Solution

	Write the formula for aluminum nitride	Write the formula for lithium oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Al^{3+} N^{3-}	Li^{+} O^{2-}
2. Use a multiplier to make the total charge of the cations and anions equal to each other.	total charge of cations = total charge of anions 1(3+) = 1(3-) +3 = -3	total charge of cations = total charge of anions 2(1+) = 1(2-) +2 = -2
3. Use the multipliers as subscript for each ion.	Al_1N_1	Li_2O_1
4. Write the final formula. Leave out all charges and all subscripts that are 1.	AlN	Li_2O

Exercise 4.6.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the calcium ion and the oxygen ion
b. the 2+ copper ion and the sulfur ion
c. the 1+ copper ion and the sulfur ion

Answer a:

CaO

Answer b:

CuS

Answer c:

Cu_2S

> *Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions.*

Example 4.6.3: Sulfur Compound

Write the formula for sodium combined with sulfur.

Exercise 4.6.3

Write the formula for each ionic compound.

a. sodium bromide
b. lithium chloride
c. magnesium oxide

Answer a:

NaBr

Answer b:

LiCl

Answer c:

MgO

Polyatomic Ions

Some ions consist of groups of atoms bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called polyatomic ions. Polyatomic ions have characteristic formulas, names, and charges that should be memorized. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1− charge. Table 4.6.1 lists the most common polyatomic ions.

Table 4.6.1: Some Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$C_2H_3O_2^-$ (also written $CH_3CO_2^-$)
carbonate ion	CO_3^{2-}
chromate ion	CrO_4^{2-}
dichromate ion	$Cr_2O_7^{2-}$
hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
permanganate ion	MnO_4^-
phosphate ion	PO_4^{3-}
hydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	$H_2PO_4^-$
sulfate ion	SO_4^{2-}
hydrogen sulfate ion (bisulfate ion)	HSO_4^-
sulfite ion	SO_3^{2-}

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. An example is $Ba(NO_3)_2$.

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

Writing a formula for ionic compounds containing polyatomic ions also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion.

Example 4.6.5

Write the chemical formula for an ionic compound composed of the potassium ion and the sulfate ion

Solution

Explanation	Answer
Potassium ions have a charge of 1+, while sulfate ions have a charge of 2−. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.	K_2SO_4

Exercise 4.6.5

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the magnesium ion and the carbonate ion
b. the aluminum ion and the acetate ion

Answer a:

$MgCO_3$

Answer b:

$Al(CH_3COO)_3$

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds. First, compounds between metal and nonmetal elements are usually ionic. For example, $CaBr_2$ contains a metallic element (calcium, a group 2 (or 2A) metal) and a nonmetallic element (bromine, a group 17 (or 7A) nonmetal). Therefore, it is most likely an ionic compound. (In fact, it *is* ionic.) In contrast, the compound NO_2 contains two elements that are both nonmetals (nitrogen, from group 15 (or 5A), and oxygen, from group 16 (or 6A). It is not an ionic compound; it belongs to the category of covalent compounds discussed elsewhere. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Second, if you recognize the formula of a polyatomic ion in a compound, the compound is ionic. For example, if you see the formula $Ba(NO_3)_2$, you may recognize the "NO_3" part as the nitrate ion, NO_3^-. (Remember that the convention for writing formulas for ionic compounds is not to include the ionic charge.) This is a clue that the other part of the formula, Ba, is actually the Ba^{2+} ion, with the 2+ charge balancing the overall 2− charge from the two nitrate ions. Thus, this compound is also ionic.

Example 4.6.6

Identify each compound as ionic or not ionic.

a. Na_2O
b. PCl_3
c. NH_4Cl

d. OF_2

Solution

Explanation	Answer
a. Sodium is a metal, and oxygen is a nonmetal. Therefore, Na_2O is expected to be ionic.	Na_2O, ionic
b. Both phosphorus and chlorine are nonmetals. Therefore, PCl_3 is not ionic.	PCl_3 , not ionic
c. The NH_4 in the formula represents the ammonium ion, NH_4^+, which indicates that this compound is ionic.	NH_4Cl, ionic
d. Both oxygen and fluorine are nonmetals. Therefore, OF_2 is not ionic.	OF_2, ionic

Exercise 4.6.6

Identify each compound as ionic or not ionic.

a. N_2O
b. $FeCl_3$
c. $(NH_4)_3PO_4$
d. $SOCl_2$

Answer a:

not ionic

Answer b:

ionic

Answer c:

ionic

Answer d:

not ionic

Summary

Formulas for ionic compounds contain the symbols and number of each atom present in a compound in the lowest whole number ratio.

Contributions & Attributions

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4.7: Naming Ionic Compounds

Learning Objectives

- To use the rules for naming ionic compounds.

After learning a few more details about the names of individual ions, you will be a step away from knowing how to name ionic compounds. This section begins the formal study of nomenclature, the systematic naming of chemical compounds.

Naming Ions

The name of a monatomic cation is simply the name of the element followed by the word *ion*. Thus, Na^+ is the sodium ion, Al^{3+} is the aluminum ion, Ca^{2+} is the calcium ion, and so forth.

A small number of metals form one stable ion, other metals form more than one stable ion (Figure 3.3). Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}. The same issue arises for other ions with more than one possible charge.

There are two ways to make this distinction. In the simpler, more modern approach, called the **Stock system**, an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word *ion*. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name *sodium ion*.

Table 4.7.1: The Modern and Common System of Cation Names

Element	Stem	Charge	Modern Name	Common Name
iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion
tin	stann-	2+	tin(II) ion	stannous ion
		4+	tin(IV) ion	stannic ion
lead	plumb-	2+	lead(II) ion	plumbous ion
		4+	lead(IV) ion	plumbic ion
chromium	chrom-	2+	chromium(II) ion	chromous ion
		3+	chromium(III) ion	chromic ion
gold	aur-	1+	gold(I) ion	aurous ion
		3+	gold(III) ion	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table 4.7.1 lists the elements that use the common system, along with their respective cation names.

Table 4.7.2: Some Monatomic Anions

Ion	Name
F^-	fluoride ion
Cl^-	chloride ion

Ion	Name
Br^-	bromide ion
I^-	iodide ion
O^{2-}	oxide ion
S^{2-}	sulfide ion
P^{3-}	phosphide ion
N^{3-}	nitride ion

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, Cl^- is “chlor-” + “-ide ion,” or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table 4.7.2 lists the names of some common monatomic ions. The polyatomic ions have their own characteristic names, as discussed earlier.

Example 4.7.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. SO_3^{2-}
d. NH_4^+
e. Cu^+

Solution

a. the calcium ion
b. the sulfide ion
c. the sulfite ion
d. the ammonium ion
e. the copper(I) ion or the cuprous ion

Exercise 4.7.1

Name each ion.

a. Fe^{2+}
b. Fe^{3+}
c. SO_4^{2-}
d. Ba^{2+}
e. HCO_3^-

Answer a:

iron(II) ion

Answer b:

iron(III) ion

Answer c:

sulfate ion

Answer d:

barium ion

Answer e:

hydrogen carbonate ion or bicarbonate ion

Example 4.7.2

Write the formula for each ion.

a. the bromide ion
b. the phosphate ion
c. the cupric ion
d. the magnesium ion

Solution

a. Br^{-}
b. PO_4^{3-}
c. Cu^{2+}
d. Mg^{2+}

Exercise 4.7.2

Write the formula for each ion.

a. the fluoride ion
b. the carbonate ion
c. the stannous ion
d. the potassium ion

Answer a:

F^{-}

Answer b:

CO_3^{2-}

Answer c:

Sn^{2+}

Answer d:

K^{+}

Naming Binary Ionic Compounds with a Metal that Forms Only One Type of Cation

Now that we know how to name ions, we are ready to name ionic compounds. A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**. The metal cation is named first, followed by the nonmetal anion as illustrated in Figure 4.7.1 for the compound $BaCl_2$. The word *ion* is dropped from both parts.

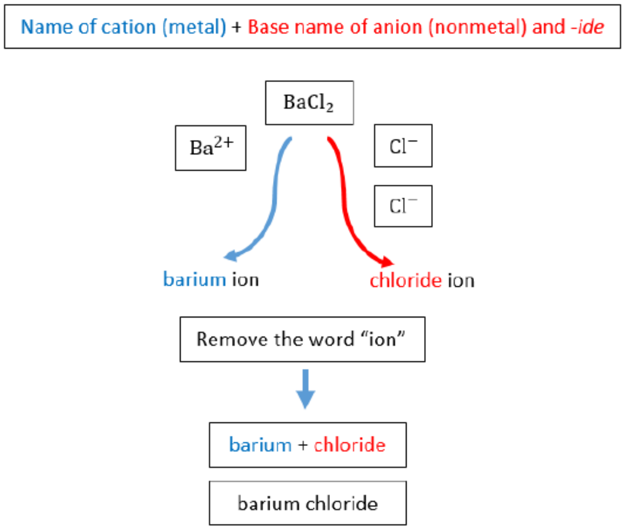


Figure 4.7.1: Naming $BaCl_2$

Subscripts in the formula do not affect the name.

Example 4.7.3: Naming Ionic Compounds

Name each ionic compound.

a. $CaCl_2$
b. AlF_3
c. KCl

Solution

a. Using the names of the ions, this ionic compound is named calcium chloride.
b. The name of this ionic compound is aluminum fluoride.
c. The name of this ionic compound is potassium chloride

Exercise 4.7.3

Name each ionic compound.

a. AgI
b. MgO
c. Ca_3P_2

Answer a:

silver iodide

Answer b:

magnesium oxide

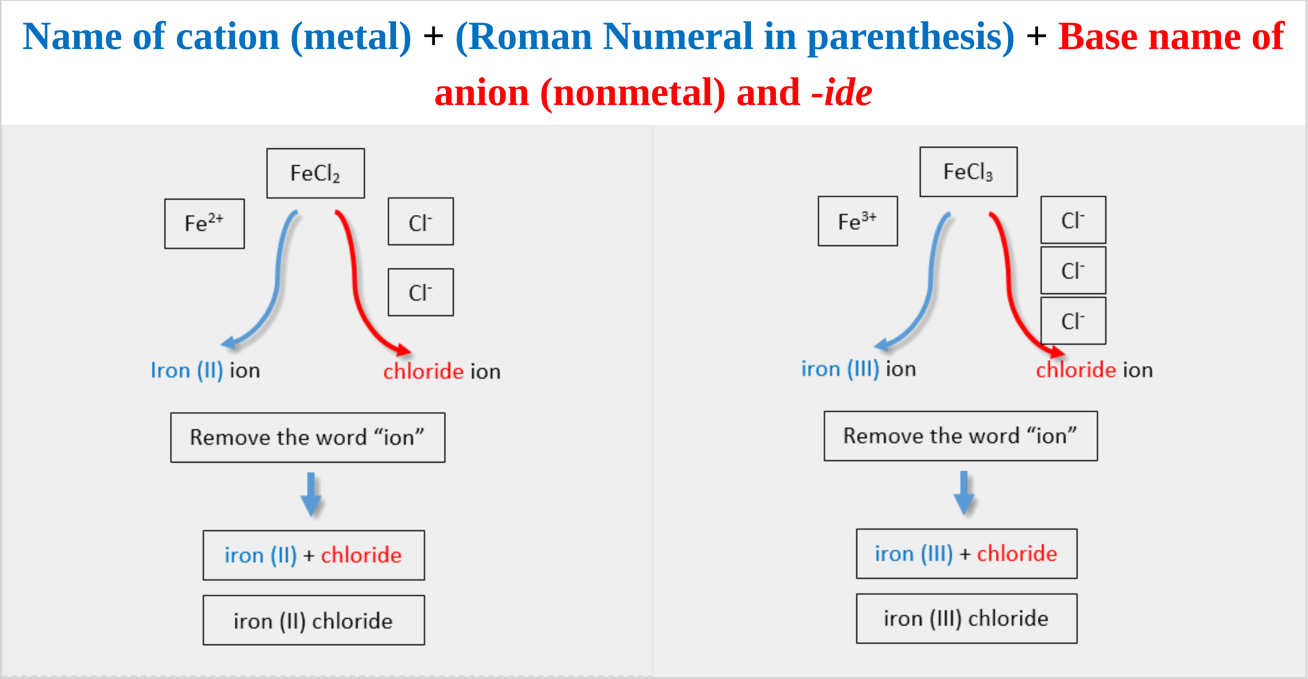
Answer c:

calcium phosphide

Naming Binary Ionic Compounds with a Metal That Forms More Than One Type of Cation

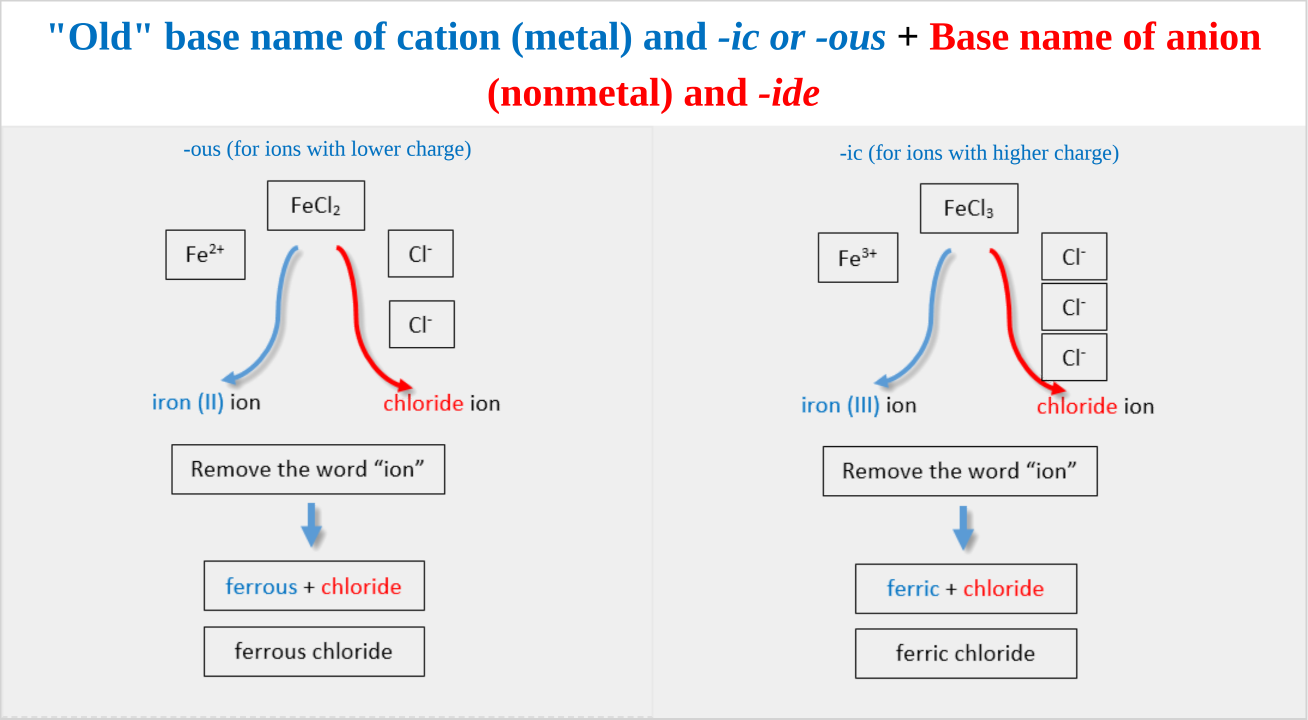
If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the iron ion has a 2+ charge because there are two Cl^- ions in the formula (1− charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride (Figure 4.7.2).

Table 4.7.3: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Modern/Stock System.



If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively (Figure 4.7.3) .

Table 4.7.4: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Old/Common System.



Example 4.7.4:

Name each ionic compound.

a. Co_2O_3
b. FeCl2

Solution

	Explanation	Answer
a	We know that cobalt can have more than one possible charge; we just need to determine what it is. • Oxide always has a 2− charge, so with three oxide ions, we have a total negative charge of 6−. • This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. • Therefore, the proper name for this ionic compound is cobalt(III) oxide.	cobalt(III) oxide
b	Iron can also have more than one possible charge. • Chloriide always has a 1− charge, so with two chloride ions, we have a total negative charge of 2−. • This means that the one iron ion must have a 2+charge. • Therefore, the proper name for this ionic compound is iron(II) chloride.	iron(II) chloride

Exercise 4.7.4

Name each ionic compound.

a. $AuCl_3$
b. PbO_2
c. CuO

Answer a:

gold(III) chloride

Answer b:

lead(IV) oxide

Answer c:

copper(II) oxide

Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium Sulfate compound in Figure 4.7.6.

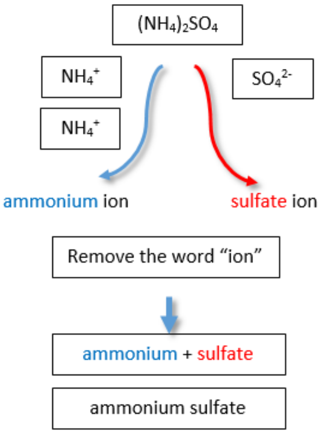


Figure 4.7.2: Naming Ionic Compounds with Polyatomic Ions

Example 4.7.5: Naming Ionic Compounds

Write the proper name for each ionic compound.

a. $(NH_4)_2S$
b. $AlPO_4$,
c. $Fe_3(PO_4)_2$

Solution

Explanation	Answer
a. The ammonium ion has a 1+ charge and the sulfide ion has a 2− charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound's name is ammonium sulfide.	ammonium sulfide
b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate
c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6−, two phosphate ions are needed . The compound's name is iron(II) phosphate.	iron(II) phosphate

Exercise $4.7.5A$

Write the proper name for each ionic compound.

a. $(NH_4)_3PO_4$
b. $Co(NO_2)_3$

Answer a:

ammonium phosphate

Answer b:

cobalt(III) nitrite

Figure 4.7.1 is a synopsis of how to name simple ionic compounds.

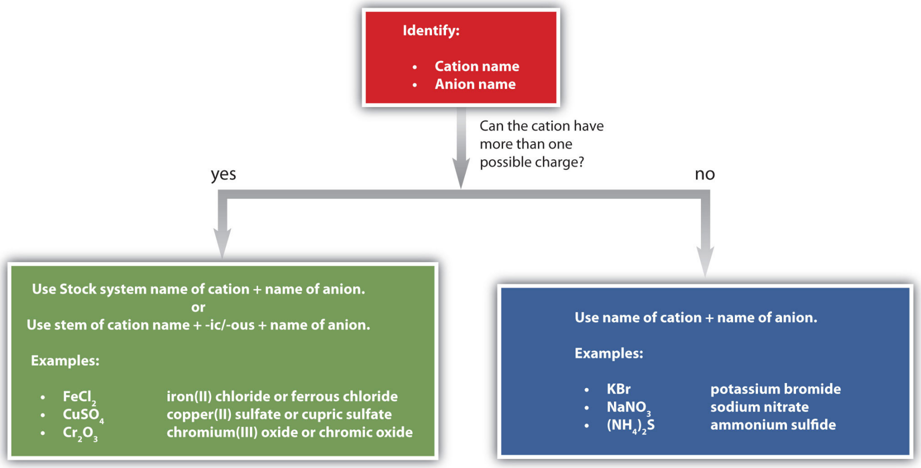


Figure $4.7.3$: A Guide to Naming Simple Ionic Compounds.

Exercise $4.7.5B$

Name each ionic compound.

a. $ZnBr_2$
b. Al_2O_3
c. $(NH_4)_3PO_4$
d. AuF_3
e. AgF

Answer a:

zinc bromide

Answer b:

aluminum oxide

Answer c:

ammonium phosphate

Answer d:

gold(III) fluoride or auric fluoride

Answer e:

silver fluoride

Summary

- Ionic compounds are named by stating the cation first, followed by the anion.
- Positive and negative charges must balance.
- Some anions have multiple forms and are named accordingly with the use of roman numerals in parenthes.
- Ternary compounds are composed of three or more elements.

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4.8: Defining terms associated with solutions

Learning Objective

- Learn some terminology involving solutions.
- Recognize which terminology is qualitative and which terminology is quantitative.
- Explain why certain substances dissolve in other substances.

The major component of a solution is called the **solvent.** The minor component of a solution is called the **solute**. "Major" and "minor" indicate which component has the greater or lesser presence by mass or moles, respectively. Sometimes this becomes confusing, especially when considering substances with very different molar masses. We will confine the discussion here to solutions for which the major component and the minor component are obvious.

Solutions exist for every possible phase of the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water; soda water is a solution of gaseous CO_2 in liquid water, while air is a solution of a gaseous solute (O_2) in a gaseous solvent (N_2). In all cases, however, the overall phase of the solution is the same phase as the solvent.

Example 4.8.1: Sugar Water

A solution is made by dissolving 1.00 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Solution

Either by mass or by moles, the obvious minor component is sucrose, so it is the solute. Water—the majority component—is the solvent. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

Exercise 4.8.1

A solution is made by dissolving 3.33 g of HCl(g) in 40.0 g of liquid methyl alcohol (CH_3OH). Identify the solvent and solute in the resulting solution.

Answer

solute: HCl(g); solvent: CH_3OH

One important concept of solutions is in defining how much solute is dissolved in a given amount of solvent. This concept is called **concentration**. Various words are used to describe the relative amounts of solute. **Dilute** describes a solution that has very little solute, while **concentrated** describes a solution that has a lot of solute. One problem is that these terms are qualitative; they describe more or less but not exactly how much.

In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is called the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. Table 4.8.1 lists the solubilities of some simple ionic compounds. These solubilities vary widely: NaCl can dissolve up to 31.6 g per 100 g of H_2O, while AgCl can dissolve only 0.00019 g per 100 g of H_2O.

Table 4.8.1: Solubilities of Some Ionic Compounds

Solute	Solubility (g per 100 g of H_2O at 25°C)
AgCl	0.00019
$CaCO_3$	0.0006
KBr	70.7
NaCl	36.1
$NaNO_3$	94.6

When the maximum amount of solute has been dissolved in a given amount of solvent, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solute, the solution is **unsaturated**. These terms are also qualitative terms because each solute has its own solubility. A solution of 0.00019 g of AgCl per 100 g of H_2O may be saturated, but with so little solute dissolved, it is also rather dilute. A solution of 36.1 g of NaCl in 100 g of H_2O is also saturated but rather concentrated. Ideally, we need more precise ways of specifying the amount of solute in a solution.

In some circumstances, it is possible to dissolve more than the maximum amount of a solute in a solution. Usually, this happens by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, and letting the solution cool down slowly and carefully. Such solutions are called **supersaturated** solutions and are not stable; given an opportunity (such as dropping a crystal of solute in the solution), the excess solute will precipitate from the solution.

It should be obvious that some solutes dissolve in certain solvents but not others. NaCl, for example, dissolves in water but not in vegetable oil. Beeswax dissolves in liquid hexane but not water. What is it that makes a solute soluble in some solvents but not others?

The answer is intermolecular interactions. The intermolecular interactions include London dispersion forces, dipole-dipole interactions, and hydrogen bonding (as described in Section 4.4). From experimental studies, it has been determined that if molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl—a very polar substance because it is composed of ions—dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane but not in polar water. This concept leads to the general rule that "like dissolves like" for predicting whether a solute is soluble in a given solvent. However, this is a general rule, not an absolute statement, so it must be applied with care.

Example 4.8.2: Polar and Nonpolar Solvents

Would I_2 be more soluble in CCl_4 or H_2O? Explain your answer.

Solution

I_2 is nonpolar. Of the two solvents, CCl_4 is nonpolar and H_2O is polar, so I_2 would be expected to be more soluble in CCl_4.

Exercise 4.8.2

Would C_3H_7OH be more soluble in CCl_4 or H_2O? Explain your answer.

Answer

H_2O because both experience hydrogen bonding

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- Concentration is the expression of the amount of solute in a given amount of solvent and can be described by several qualitative terms.
- Solubility is a specific amount of solute that can dissolve in a given amount of solvent.
- "Like dissolves like" is a useful rule for deciding if a solute will be soluble in a solvent.

4.9: What is a solution?

More than 70% of the earth's surface is covered by a very important solution-seawater. It is likely that without seawater, no life would exist on Earth.

At its simplest, seawater is mostly H_2O. But about 3.5% of seawater is dissolved solids, mostly NaCl but other ions as well. Table 4.9.1 lists the percentage by mass of the various ions in seawater. Because it is highly likely that life on Earth originated in the oceans, it should not be surprising that many bodily fluids resemble seawater-especially blood. Table 4.9.1 also lists the percentage by mass of ions in a typical sample of blood.

Table 4.9.1: Percentage by Mass of Ions in Seawater and Blood

Ion	Percentage in Seawater	Percentage in Blood
Na^+	2.36	0.322
Cl^-	1.94	0.366
Mg^{2+}	0.13	0.002
SO_4^{2-}	0.09	-
K^+	0.04	0.016
Ca^{2+}	0.04	0.0096
HCO_3^-	0.002	0.165
HPO_4^{2-}, $H_2PO_4^-$	-	0.01

Most ions are more abundant in seawater than they are in blood, with some notable exceptions. There is far more hydrogen carbonate ion (HCO_3^-) in blood than in seawater; indeed, it is the third most common ion in blood. This difference is significant because the HCO_3^- ion and some related species [CO_3^{2-}, $CO_2(aq)$] have an important role in controlling the acid-base properties of blood. Although there is a negligible amount of the two hydrogen phosphate ions (HPO_4^{2-} and $H_2PO_4^-$) in seawater, there is a small amount in blood, where these ions affect acid-base properties. Another notable difference is that blood has a negligible amount of the sulfate ion (SO_4^{2-}), but this ion is present in seawater. Gold is present in seawater-but only a tiny amount. A current estimate of the amount of gold is about 1 part per every 1×10^{13} parts of seawater, which makes the extraction of gold from seawater unfeasible. However, it does mean that there are about 1.4×10^{14} g of gold in the world's oceans!

Figure 4.9.1: Gold in the Ocean. There are approximately 1.4×10^{14} g of gold in the oceans, but extracting it effectively is beyond current technologies. (CC SA-BY; Evgeni Dinev, Flickr).

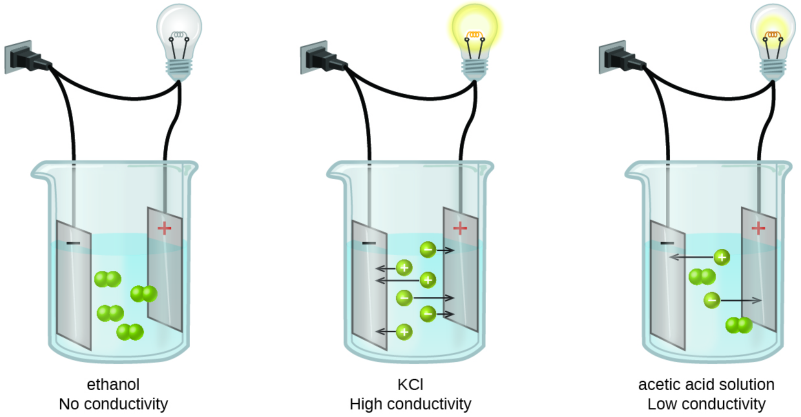
A solution is a *homogeneous mixture*-a mixture of two or more substances that are so intimately mixed that the mixture behaves in many ways like a single substance. Many chemical reactions occur when the reactants are dissolved in solution. In this module, we will introduce concepts that are applicable to solutions and the chemical reactions that occur in them.

4.10: Aqueous Solutions and Solubility - Compounds Dissolved in Water

Learning Objectives

- Identify compounds as electrolytes and non-electrolytes

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 4.10.1). Solutions of strong electrolytes conduct well because the compound exists almost 100% as ions in solution. Solutions of weak electrolytes still contain ions, put the extent to which they form ions is solution is much lower than for strong electrolytes, so they do not conduct as well. Solutions of nonelectrolytes do not contain any ions and therefore do not conduct electricity.



ethanol
No conductivity

KCl
High conductivity

acetic acid solution
Low conductivity

Figure 4.10.1: 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 4.10.2 The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water.

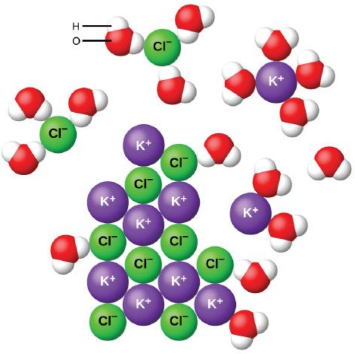


Figure 4.10.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown.

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. Ion-dipole 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 $4.10.2$ 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, or the ion-dipole attractive forces between the ions and water molecules are so weak, 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).

Summary

Substances that dissolve in water to yield ions are called electrolytes. Nonelectrolytes are substances that do not produce ions when dissolved in water. Solubility rules allow prediction of what products will be insoluble in water.

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4.11: Quantitative Units of Concentration

Learning Objective

- Determine specific concentrations with several common units.

Rather than qualitative terms (Section 11.2 - Definitions) we need quantitative ways to express the amount of solute in a solution; that is, we need specific units of concentration. In this section, we will introduce several common and useful units of concentration.

Molarity (M) is defined as the number of moles of solute divided by the number of liters of solution:

$$molarity\ =\frac{moles\ of\ solute}{liters\ of\ solution}$$

which can be simplified as

$$M\ =\frac{mol}{L},\ or\ mol/L$$

As with any mathematical equation, if you know any two quantities, you can calculate the third, unknown, quantity.

For example, suppose you have 0.500 L of solution that has 0.24 mol of NaOH dissolved in it. The concentration of the solution can be calculated as follows:

$$molarity\ =\frac{0.24\,mol\ NaOH}{0.500L}=0.48\ M\ NaOH$$

The concentration of the solution is 0.48 M, which is spoken as "zero point forty-eight molarity" or "zero point forty-eight molar." If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L? First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4\ g\cancel{HCl}\times\frac{1\ mol\ HCl}{36.5\ g\cancel{HCl}}=0.614\,mol\ HCl$$

Now we can use the definition of molarity to determine a concentration:

$$M\ =\frac{0.614\,mol\ HCl}{1.56L}=0.394\,M$$

Example 4.11.1:

What is the molarity of a solution made when 32.7 g of NaOH are dissolved to make 445 mL of solution?

Solution

To use the definition of molarity, both quantities must be converted to the proper units. First, convert the volume units from milliliters to liters:

$$445\ m\cancel{L}\times\frac{1\ L}{1000\ m\cancel{L}}=0.445\ L$$

Now we convert the amount of solute to moles, using the molar mass of NaOH, which is 40.0 g/mol:

$$32.7\ gNa\cancel{OH}\times\frac{1\ mol\ NaOH}{40.0\ gNa\cancel{OH}}=0.818\,mol\ NaOH$$

Now we can use the definition of molarity to determine the molar concentration:

$$M\ =\frac{0.818\,mol\ NaOH}{0.445L}=1.84\,M\ NaOH$$

Exercise 4.11.1

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer

1.57 M

The definition of molarity can be used to determine the amount of solute or the volume of solution, if the other information is given. Example 4 illustrates this situation.

Example 4.11.1:

How many moles of solute are present in 0.108 L of a 0.887 M NaCl solution?

Solution

We know the volume and the molarity; we can use the definition of molarity to mathematically solve for the amount in moles. Substituting the quantities into the definition of molarity:

$$0.887\,M = \frac{mol\ NaCl}{0.108 L}$$

We multiply the 0.108 L over to the other side of the equation and multiply the units together; "molarity × liters" equals moles, according to the definition of molarity. So

mol NaCl = (0.887 M)(0.108 L) = 0.0958 mol

Exercise 4.11.1

How many moles of solute are present in 225 mL of a 1.44 M $CaCl_2$ solution?

Answer

0.324 mol

If you need to determine volume, remember the rule that the unknown quantity must be by itself and in the numerator to determine the correct answer. Thus rearrangement of the definition of molarity is required.

Example 4.11.1:

What volume of a 2.33 M $NaNO_3$ solution is needed to obtain 0.222 mol of solute?

Solution

Using the definition of molarity, we have

$$2.33\,M = \frac{0.222 mol}{L}$$

To solve for the number of liters, we bring the 2.33 M over to the right into the denominator, and the number of liters over to the left in the numerator. We now have

$$L = \frac{0.222 mol}{2.33\,M}$$

Dividing, **the volume is 0.0953 L = 95.3 mL.**

Exercise 4.11.1

What volume of a 0.570 M K_2SO_4 solution is needed to obtain 0.872 mol of solute?

Answer

1.53 L

A similar unit of concentration is **molality** (*m*), which is defined as the number of moles of solute per kilogram of solvent, not per liter of solution:

$$molality\ =\frac{moles\ solute}{kilograms\ solvent}$$

Mathematical manipulation of molality is the same as with molarity.

Another way to specify an amount is **percentage composition by mass** (or *mass percentage*, % m/m). It is defined as follows:

$$\%m/m\ =\frac{mass\ of\ solute}{mass\ of\ entire\ sample}\times 100\%$$

It is not uncommon to see this unit used on commercial products (Figure 4.11.1- Concentration in Commercial Applications)

Figure 4.11.1 Concentration in Commercial Applications © Thinkstock

The percentage of urea in this package is 5% m/m, meaning that there are 5 g of urea per 100 g of product.

Example 4.11.1:

What is the mass percentage of Fe in a piece of metal with 87.9 g of Fe in a 113 g sample?

Solution

Using the definition of mass percentage, we have

$$\%m/m\ =\frac{87.9\,g\,Fe}{113\,g\,sample}\times 100\%=77.8\%\,Fe$$

Exercise 4.11.1

What is the mass percentage of H_2O_2 in a solution with 1.67 g of H_2O_2 in a 55.5 g sample?

Answer

3.01%

Related concentration units are **parts per thousand (ppth), parts per million (ppm)** and **parts per billion (ppb).** Parts per thousand is defined as follows:

$$ppth\ =\frac{mass\ of\ solute}{mass\ of\ sample}\times 1000$$

There are similar definitions for parts per million and parts per billion:

$$ppm\ =\frac{mass\ of\ solute}{mass\ of\ sample}\times 1,000,000$$

and

$$ppb\ =\frac{mass\ of\ solute}{mass\ of\ sample}\times 1,000,000,000$$

Each unit is used for progressively lower and lower concentrations. The two masses must be expressed in the same unit of mass, so conversions may be necessary.

Example 4.11.1:

If there is 0.6 g of Pb present in 277 g of solution, what is the Pb concentration in parts per thousand?

Solution

Use the definition of parts per thousand to determine the concentration. Substituting

$$\frac{0.6 g\, Pb}{277 g\, solution}\times 1000 = 2.17\, ppth$$

Exercise 4.11.1

If there is 0.551 mg of As in 348 g of solution, what is the As concentration in ppm?

Answer

1.58 ppm

As with molarity and molality, algebraic rearrangements may be necessary to answer certain questions.

Example 4.11.1:

The concentration of Cl^{-} ion in a sample of H_2O is 15.0 ppm. What mass of Cl^{-} ion is present in 240.0 mL of H_2O, which has a density of 1.00 g/mL?

Solution

First, use the density of H_2O to determine the mass of the sample:

$$240.0\ \cancel{mL}\times\frac{1.00\ g}{\cancel{mL}} = 240.0\ g$$

Now we can use the definition of ppm:

$$15.0\,ppm\;=\frac{mass\;of\;solute}{240.0\;g\;solution}\times 1,000,000$$

Rearranging to solve for the mass of solute,

$$mass\;solute=\frac{(15.0\,ppm)(240.0\;g\;solution)}{1,000,000}=0.0036g=3.6\,mg$$

Exercise 4.11.1

The concentration of Fe^{3+} ion in a sample of H_2O is 335.0 ppm. What mass of Fe^{3+} ion is present in 3,450 mL of H_2O, which has a density of 1.00 g/mL?

Answer

1.16 g

For ionic solutions, we need to differentiate between the concentration of the salt versus the concentration of each individual ion. Because the ions in ionic compounds go their own way when a compound is dissolved in a solution, the resulting concentration of the ion may be different from the concentration of the complete salt. For example, if 1 M NaCl were prepared, the solution could also be described as a solution of 1 M Na^+(aq) and 1 M $Cl^−$(aq) because there is one Na^+ ion and one $Cl^−$ ion per formula unit of the salt. However, if the solution were 1 M $CaCl_2$, there are two $Cl^−$(aq) ions for every formula unit dissolved, so the concentration of $Cl^−$(aq) would be 2 M, not 1 M.

In addition, the total ion concentration is the sum of the individual ion concentrations. Thus for the 1 M NaCl, the total ion concentration is 2 M; for the 1 M $CaCl_2$, the total ion concentration is 3 M.

Key Takeaway

- Quantitative units of concentration include molarity, molality, mass percentage, parts per thousand, parts per million, and parts per billion.

Exercise 4.11.1

1. Differentiate between molarity and molality.
2. Differentiate between mass percentage and parts per thousand.
3. What is the molarity of a solution made by dissolving 13.4 g of $NaNO_3$ in 345 mL of solution?
4. What is the molarity of a solution made by dissolving 332 g of $C_6H_{12}O_6$ in 4.66 L of solution?
5. How many moles of $MgCl_2$ are present in 0.0331 L of a 2.55 M solution?
6. How many moles of NH_4Br are present in 88.9 mL of a 0.228 M solution?
7. What volume of 0.556 M NaCl is needed to obtain 0.882 mol of NaCl?
8. What volume of 3.99 M H_2SO_4 is needed to obtain 4.61 mol of H_2SO_4?
9. What volume of 0.333 M $Al(NO_3)_3$ is needed to obtain 26.7 g of $Al(NO_3)_3$?
10. What volume of 1.772 M $BaCl_2$ is needed to obtain 123 g of $BaCl_2$?
11. What are the individual ion concentrations and the total ion concentration in 0.66 M $Mg(NO_3)_2$?
12. What are the individual ion concentrations and the total ion concentration in 1.04 M $Al_2(SO_4)_3$?
13. If the $C_2H_3O_2{}^−$ ion concentration in a solution is 0.554 M, what is the concentration of $Ca(C_2H_3O_2)_2$?
14. If the $Cl^−$ ion concentration in a solution is 2.61 M, what is the concentration of $FeCl_3$?

Answers

1. Molarity is moles per liter, whereas molality is moles per kilogram of solvent.

2.
3. 0.457 M
4.
5. 0.0844 mol
6.
7. 1.59 L
8.
9. 0.376 L
10.
11. Mg^{2+} = 0.66 M; NO_3^- = 1.32 M; total: 1.98 M
12.
13. 0.277 M

4.12: Concentrations as Conversion Factors

Learning Objective

- Apply concentration units as conversion factors.

Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition. For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

$$0.108\,L\,\cancel{NaCl} \times \frac{0.887\,mol\,NaCl}{L\,\cancel{NaCl}} = 0.0958\,mol\,NaCl$$

(There is an understood 1 in the denominator of the conversion factor.) If we used the definition approach, we get the same answer, but now we are using conversion factor skills. Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

Example 4.12.1

Using concentration as a conversion factor, how many liters of 2.35 M $CuSO_4$ are needed to obtain 4.88 mol of $CuSO_4$?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:

$$4.88\,mol\,\cancel{CuSO_4} \times \frac{1\,L}{2.35\,\cancel{mol}} = 2.08\,L\,of\,solution$$

Exercise 4.12.1

Using concentration as a conversion factor, how many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?

Answer:

1.74 L

Of course, once quantities in moles are available, another conversion can give the mass of the substance, using molar mass as a conversion factor.

Example 4.12.2

What mass of solute is present in 0.765 L of 1.93 M NaOH?

Solution

This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of NaOH (40.0 g/mol) to convert to mass:

$$0.765\,\cancel{L} \times \frac{1.93\,mol\,\cancel{NaOH}}{L\,\cancel{solution}} \times \frac{40.0g\,NaOH}{1\,mol\,\cancel{NaOH}} = 59.1\,g\,NaOH$$

Exercise 4.12.2

What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?

Answer

6.12 g

More complex stoichiometry problems using balanced chemical reactions can also use concentrations as conversion factors. For example, suppose the following equation represents a chemical reaction:

$$2\,AgNO_3(aq) + CaCl_2(aq) \rightarrow 2\,AgCl(s) + Ca(NO_3)_2(aq)$$

If we wanted to know what volume of 0.555 M $CaCl_2$ would react with 1.25 mol of $AgNO_3$, we first use the balanced chemical equation to determine the number of moles of $CaCl_2$ that would react and then use concentration to convert to liters of solution:

$$1.25\ mol\ \cancel{AgNO_3} \times \frac{1\ mol\ \cancel{CaCl_2}}{2\ mol\ \cancel{AgNO_3}} \times \frac{1 L\ solution}{0.555\ mol\ \cancel{CaCl_2}} = 1.13\ L\ CaCl_2$$

This can be extended by starting with the mass of one reactant, instead of moles of a reactant.

Example 4.12.3

What volume of 0.0995 M $Al(NO3)3$ will react with 3.66 g of Ag according to the following chemical equation?

$$3\,Ag(s) + Al(NO_3)_3(aq) \rightarrow 3\,AgNO_3 + Al(s)$$

Solution

Here, we first must convert the mass of Ag to moles before using the balanced chemical equation and then the definition of molarity as a conversion factor:

$$3.66\ g\ \cancel{Ag} \times \frac{1\ mol\ \cancel{Ag}}{107.97\ g\ \cancel{Ag}} \times \frac{1\ mol\ \cancel{Al(NO_3)_3}}{3\ mol\ \cancel{Ag}} \times \frac{1 L\ solution}{0.0995\ mol\ \cancel{Al(NO_3)_3}} = 0.114\ L$$

The strikeouts show how the units cancel.

Exercise 4.12.3

What volume of 0.512 M NaOH will react with 17.9 g of $H2C2O4(s)$ according to the following chemical equation?

$$H_2C_2O_4(s) + 2\,NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\,H_2O(\ell)$$

Answer:

0.777 L

We can extend our skills even further by recognizing that we can relate quantities of one solution to quantities of another solution. Knowing the volume and concentration of a solution containing one reactant, we can determine how much of another solution of another reactant will be needed using the balanced chemical equation.

Example 4.12.4

A student takes a precisely measured sample, called an *aliquot*, of 10.00 mL of a solution of $FeCl_3$. The student carefully adds 0.1074 M $Na_2C_2O_4$ until all the $Fe^{3+}(aq)$ has precipitated as $Fe_2(C_2O_4)_3(s)$. Using a precisely measured tube called a burette, the student finds that 9.04 mL of the $Na_2C_2O_4$ solution was added to completely precipitate the $Fe^{3+}(aq)$. What was the concentration of the $FeCl_3$ in the original solution? (A precisely measured experiment like this, which is meant to determine the amount of a substance in a sample, is called a *titration*.) The balanced chemical equation is as follows:

$$2\,FeCl_3(aq) + 3\,Na_2C_2O_4(aq) \rightarrow Fe_2(C_2O_4)_3(s) + 6\,NaCl(aq)$$

Solution

First we need to determine the number of moles of $Na_2C_2O_4$ that reacted. We will convert the volume to liters and then use the concentration of the solution as a conversion factor:

$$9.04\ m\cancel{L} \times \frac{1\ \cancel{L}}{1000\ m\cancel{L}} \times \frac{0.1074 mol\ Na_2C_2O_4}{\cancel{L}} = 0.000971\ mol\ Na_2C_2O_4$$

Now we will use the balanced chemical equation to determine the number of moles of Fe^{3+}(aq) that were present in the initial aliquot:

$$0.000971\ mol\ \cancel{Na_2C_2O_4} \times \frac{2 mol\ FeCl_3}{3\ mol \cancel{Na_2C_2O_4}} = 0.000647 mol\ FeCl_3$$

Then we determine the concentration of $FeCl_3$ in the original solution. Converting 10.00 mL into liters (0.01000 L), we use the definition of molarity directly:

$$M = \frac{mol}{L} = \frac{0.000647 mol\ FeCl_3}{0.01000 L} = 0.0647 M\ FeCl_3$$

Exercise 4.12.4

A student titrates 25.00 mL of H_3PO_4 with 0.0987 M KOH. She uses 54.06 mL to complete the chemical reaction. What is the concentration of H_3PO_4?

$$H_3PO_4(aq) + 3\,KOH(aq) \rightarrow K_3PO_4(aq) + 3\,H_2O$$

Answer:

0.0711 M

Figure 4.12.1 Titration © Thinkstock When a student performs a titration, a measured amount of one solution is added to another reactant.

We have used molarity exclusively as the concentration of interest, but that will not always be the case. The next example shows a different concentration unit being used.

Example 4.12.5:

H_2O_2 is used to determine the amount of Mn according to this balanced chemical equation:

$$2\,\mathrm{MnO_4^-(aq)} + 5\,\mathrm{H_2O_2(aq)} + 6\,\mathrm{H^+(aq)} \rightarrow 2\,\mathrm{Mn_2^+(aq)} + 5\,\mathrm{O_2(g)} + 8\,\mathrm{H_2O}(\ell)$$

What mass of 3.00% m/m H_2O_2 solution is needed to react with 0.355 mol of $MnO_4^-(aq)$?

Solution

Because we are given an initial amount in moles, all we need to do is use the balanced chemical equation to determine the number of moles of H_2O_2 and then convert to find the mass of H_2O_2. Knowing that the H_2O_2 solution is 3.00% by mass, we can determine the mass of solution needed:

$$3.55\ mol\ \cancel{MnO_4^-} \times \frac{5\ mol\ \cancel{H_2O_2}}{2\ mol \cancel{MnO_4^-}} \times \frac{34.02\ g\, \cancel{H_2O_2}}{mol\ \cancel{H_2O_2}} \times \frac{100 g\ solution}{3\ g\, \cancel{H_2O_2}} = 1006 g\ sol$$

The first conversion factor comes from the balanced chemical equation, the second conversion factor is the molar mass of H_2O_2, and the third conversion factor comes from the definition of percentage concentration by mass.

Exercise 4.12.5

Use the balanced chemical reaction for MnO_4^- and H_2O_2 to determine what mass of O_2 is produced if 258 g of 3.00% m/m H_2O_2 is reacted with MnO_4^-.

Answer

7.28 g

Summary

Know how to apply concentration units as conversion factors.

4.13: Dilutions and Concentrations

Learning Objective

- Learn how to dilute solutions.

Solutions can be prepared by diluting a more concentrated solution rather than starting with the pure solute. **Dilution** is the addition of solvent, which decreases the concentration of the solute in the solution.

In both dilution and concentration, the amount of solute stays the same. This gives us a way to calculate what the new solution volume must be for the desired concentration of solute. From the definition of molarity,

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

we can solve for the number of moles of solute:

moles of solute = (molarity)(liters of solution)

A simpler way of writing this is to use *M* to represent molarity and *V* to represent volume. So the equation becomes

moles of solute = *MV*

Because this quantity does not change before and after the change in concentration, the product *MV* must be the same before and after the concentration change. Using numbers to represent the initial and final conditions, we have

$$M_1V_1 = M_2V_2$$

as the **dilution equation**. The volumes must be expressed in the same units. Note that this equation gives only the initial and final conditions, not the amount of the change. The amount of change is determined by subtraction.

Example 4.13.1

If 25.0 mL of a 2.19 M solution are diluted to 72.8 mL, what is the final concentration?

Solution

It does not matter which set of conditions is labeled 1 or 2, as long as the conditions are paired together properly. Using the dilution equation, we have

$$(2.19 \text{ M})(25.0 \text{ mL}) = M_2(72.8 \text{ mL})$$

Solving for the second concentration (noting that the milliliter units cancel),

$$M_2 = 0.752 \text{ M}$$

The concentration of the solution has decreased. In going from 25.0 mL to 72.8 mL, 72.8 − 25.0 = 47.8 mL of solvent must be added.

Exercise 4.13.1

A 0.885 M solution of KBr whose initial volume is 76.5 mL has more water added until its concentration is 0.500 M. What is the new volume of the solution?

Answer

135.4 mL

Concentrating solutions involves removing solvent. Usually this is done by evaporating or boiling, assuming that the heat of boiling does not affect the solute. The dilution equation is used in these circumstances as well.

Chemistry Is Everywhere: Preparing IV Solutions

In a hospital emergency room, a physician orders an intravenous (IV) delivery of 100 mL of 0.5% KCl for a patient suffering from hypokalemia (low potassium levels). Does an aide run to a supply cabinet and take out an IV bag containing this concentration of KCl?

Not likely. It is more probable that the aide must make the proper solution from an IV bag of sterile solution and a more concentrated, sterile solution, called a *stock solution*, of KCl. The aide is expected to use a syringe to draw up some stock solution and inject it into the waiting IV bag and dilute it to the proper concentration. Thus the aide must perform a dilution calculation.

Figure 4.13.1 Preparing IV Solution © Thinkstock. Medical personnel commonly must perform dilutions for IV solutions.

If the stock solution is 10.0% KCl and the final volume and concentration need to be 100 mL and 0.50%, respectively, then it is an easy calculation to determine how much stock solution to use:

$$(10\%)V_1 = (0.50\%)(100\text{ mL})V_1 = 5\text{ mL}$$

Of course, the addition of the stock solution affects the total volume of the diluted solution, but the final concentration is likely close enough even for medical purposes.

Medical and pharmaceutical personnel are constantly dealing with dosages that require concentration measurements and dilutions. It is an important responsibility: calculating the *wrong* dose can be useless, harmful, or even fatal!

Summary

- Calculate the new concentration or volume for a dilution or concentration of a solution.

4.14: Types of Chemical Reactions - Single and Double Replacement Reactions

Learning Objectives

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

You were introduced to chemical reactions in module 3. Many chemical reactions can be classified as belonging to a particular type of reaction which enables you to predict products given the reactants. You have already learned how to predict the products and write the balanced equations for the complete combustion of hydrocarbons. In this module you will be introduced to two other classes of reactions that take place in aqueous solutions.

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

$$2\,\mathrm{HCl(aq)} + \mathrm{Zn(s)} \rightarrow \mathrm{ZnCl_2(aq)} + \mathrm{H_2(g)}$$

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

$$2\,\mathrm{NaCl(aq)} + \mathrm{F_2(g)} \rightarrow 2\,\mathrm{NaF(s)} + \mathrm{Cl_2(g)}$$

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

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

$$\mathrm{CuCl_2(aq)} + 2\,\mathrm{AgNO_3(aq)} \rightarrow \mathrm{Cu(NO_3)_2(aq)} + 2\,\mathrm{AgCl(s)}$$

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

Example 4.14.3

Predict the products of this double-replacement equation:

$$\mathrm{BaCl_2} + \mathrm{Na_2SO_4} \rightarrow$$

Solution

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

Exercise 4.14.3

Predict the products of this double-replacement equation:

$$\mathrm{KBr} + \mathrm{AgNO_3} \rightarrow$$

Answer

KNO3 and AgBr

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

To judge whether double-replacement reactions will occur, we need to know what kinds of ionic compounds form precipitates. For this, we use **solubility rules**, which are general statements that predict which ionic compounds dissolve (are soluble) and which do not (are not soluble or insoluble). Table 4.14.1 lists some general solubility rules. We need to consider each ionic compound (both the reactants and the possible products) in light of the solubility rules. If a compound is soluble, we use the (aq) label with it, indicating it dissolves. If a compound is not soluble, we use the (s) label with it and assume that it will precipitate out of solution. If everything is soluble, then no reaction will be expected.

Table 4.14.1: Some Useful Solubility Rules (soluble)

These compounds generally dissolve in water (are soluble):	Exceptions:
All compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, and NH_4^+	None
All compounds of NO_3^- and $C_2H_3O_2^-$	None
Compounds of Cl^-, Br^-, I^-	Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds of SO_4^2	Hg_2^{2+}, Pb^{2+}, Sr^{2+}, Ba^{2+}

Table 4.14.2: Some Useful Solubility Rules (insoluble)

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

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

$$Na_2SO_4(aq) + SrCl_2(aq) \rightarrow 2\,NaCl(aq) + SrSO_4(s)$$

You would expect to see a visual change corresponding to $SrSO_4$ precipitating out of solution (Figure 4.14.2).

Figure 4.14.2: Double-Replacement Reactions. Some double-replacement reactions are obvious because you can see a solid precipitate coming out of solution. Source: Photo courtesy of Choij, http://commons.wikimedia.org/wiki/File:Copper_solution.jpg.

Example 4.14.4:

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

1. $Ca(NO_3)_2$ + KBr → ?
2. NaOH + $FeCl_2$ → ?

Solution

1. According to the solubility rules, both $Ca(NO_3)_2$ and KBr are soluble. Now we consider what the double-replacement products would be by switching the cations (or the anions)-namely, $CaBr_2$ and KNO_3. However, the solubility rules predict that these two substances would also be soluble, so no precipitate would form. Thus, we predict no reaction in this case.
2. According to the solubility rules, both NaOH and $FeCl_2$ are expected to be soluble. If we assume that a double-replacement reaction may occur, we need to consider the possible products, which would be NaCl and $Fe(OH)_2$. NaCl is soluble, but, according to the solubility rules, $Fe(OH)_2$ is not. Therefore, a reaction would occur, and $Fe(OH)_2(s)$ would precipitate out of solution. The balanced chemical equation is

$$2\,\mathrm{NaOH(aq)} + \mathrm{FeCl_2(aq)} \rightarrow 2\,\mathrm{NaCl(aq)} + \mathrm{Fe(OH)_2(s)}$$

Exercise 4.14.4

$$\mathrm{Sr(NO_3)_2} + \mathrm{KCl} \rightarrow$$

Answer

No reaction; all possible products are soluble.

Key Takeaways

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

4.15: Precipitation Reactions

Learning Objectives

- To identify a precipitation reaction and predict solubility.

A precipitation reaction is a reaction that yields an insoluble product—a precipitate—when two solutions are mixed. When a colorless solution of silver nitrate is mixed with a yellow-orange solution of potassium dichromate a reddish precipitate of silver dichromate is produced.

$$\mathrm{AgNO_3(aq) + K_2Cr_2O_7(aq) \rightarrow Ag_2Cr_2O_7(s) + KNO_3(aq)} \tag{4.15.1}$$

This unbalanced equation has the general form of an exchange reaction:

$$\mathrm{AC + BD \rightarrow \underset{\textit{insoluble}}{AD} + BC} \tag{4.15.2}$$

Thus precipitation reactions are a subclass of exchange reactions that occur between ionic compounds when one of the products is insoluble. Because both components of each compound change partners, such reactions are sometimes called *double-displacement reactions*. Two important uses of precipitation reactions are to isolate metals that have been extracted from their ores and to recover precious metals for recycling.

***Video**: Mixing Potassium dicromate and Silver Nitrate together to initiate a precipitation reaction (Equation 4.15.1).*

Just as important as predicting the product of a reaction is knowing when a chemical reaction will *not* occur. Simply mixing solutions of two different chemical substances does *not* guarantee that a reaction will take place. For example, if 500 mL of aqueous $NaCl$ solution is mixed with 500 mL of aqueous KBr solution, the final solution has a volume of 1.00 L and contains $\mathrm{Na^+(aq)}$., $\mathrm{Cl^-(aq)}$., $\mathrm{K^+(aq)}$., and $\mathrm{Br^-(aq)}$. As you will see in (Figure 4.15.1), none of these species reacts with any of the others. When these solutions are mixed, the only effect is to dilute each solution with the other.

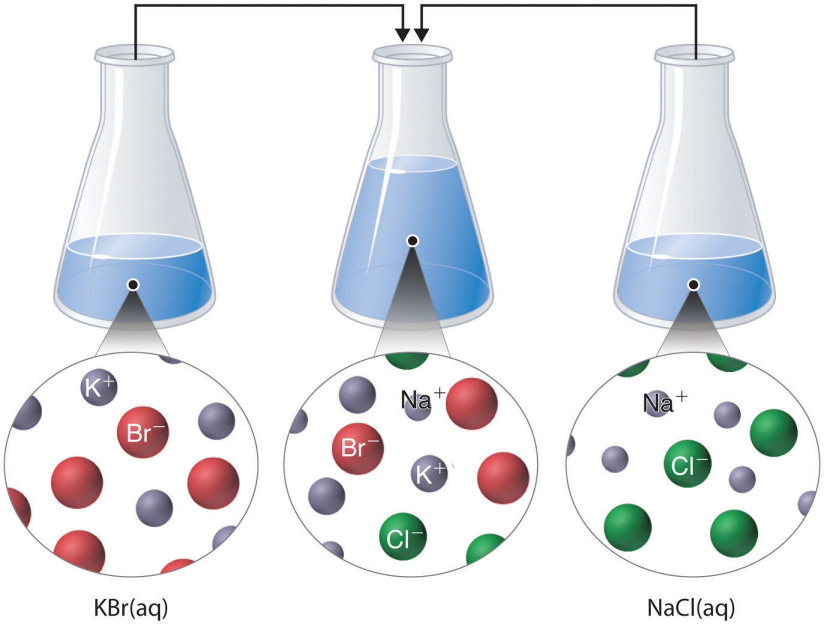


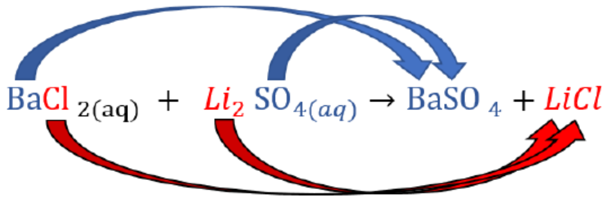
Figure 4.15.1: The Effect of Mixing Aqueous KBr and NaCl Solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

Predicting Precipitation Reactions

A precipitation reaction occurs when a solid precipitate forms after mixing two strong electrolyte solutions. As stated previously, if none of the species in the solution reacts then no net reaction occurred.

Predict what will happen when aqueous solutions of barium chloride and lithium sulfate are mixed.

Change the partners of the anions and cations on the reactant side to form new compounds (products).



Because barium chloride and lithium sulfate are strong electrolytes, each dissociates completely in water to give a solution that contains the constituent anions and cations. Mixing the two solutions *initially* gives an aqueous solution that contains Ba^{2+}, Cl$^-$, Li$^+$, and SO$_4$$^{2-}$ ions. The only possible exchange reaction is to form LiCl and BaSO$_{4.}$

Correct the formulas of the products based on the charges of the ions.

No need to correct the formula as both compounds already have their charges balanced.

$$BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4 + LiCl$$

Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.

$$BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + LiCl(aq)$$

Table 7.5.1 from the previous section shows that LiCl is soluble in water, but \(\ce{BaSO$_{4\}\backslash)}$ is not soluble in water.

Balance the equation

$$BaCl_2(aq) + Li_2SO_4(aq) \rightarrow BaSO_4(s) + 2\,LiCl(aq)$$

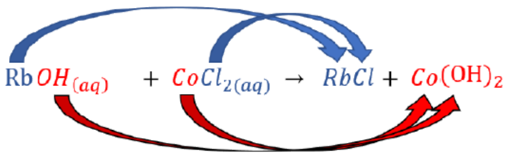
Although soluble barium salts are toxic, BaSO$_4$ is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a “barium milkshake” or a “barium enema”—a suspension of very fine $BaSO_4$ particles in water.

An x-ray of the digestive organs of a patient who has swallowed a “barium milkshake.” A barium milkshake is a suspension of very fine $BaSO_4$ particles in water; the high atomic mass of barium makes it opaque to x-rays. from Wikipedia.

Example 4.15.1

Predict what will happen if aqueous solutions of rubidium hydroxide and cobalt(II) chloride are mixed.

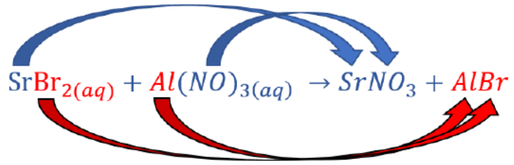
Solution

Steps	**Example**
Change the partners of the anions and cations on the reactant side to form new compounds (products).	
Correct the formulas of the products based on the charges of the ions.	$Rb(OH)_2(aq) + CoCl_2(aq) \rightarrow RbCl_2 + Co(OH)_2$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$Rb(OH)_2(aq) + CoCl_2(aq) \rightarrow RbCl_2(aq) + Co(OH)_2(s)$
Balance the equation	Coefficients already balanced $Rb(OH)_2(aq) + CoCl_2(aq) \rightarrow RbCl_2 + Co(OH)_2$

Example 4.15.2

Predict what will happen if aqueous solutions of strontium bromide and aluminum nitrate are mixed.

Solution

Steps	**Example**
Change the partners of the anions and cations on the reactant side to form new compounds (products).	
Correct the formulas of the products based on the charges of the ions.	$SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow Sr(NO_3)_2 + AlBr_3$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow Sr(NO_3)_2(aq) + AlBr_3(aq)$ According to Table 7.5.1 from the previous section, both $AlBr_3$ (rule 4) and $Sr(NO_3)_2$ (rule 2) are soluble.
If all possible products are soluble then there is no net reaction will occur	$SrBr_2(aq) + Al(NO_3)_3(aq) \rightarrow$ **NO REACTION**

Exercise 4.15.2

Using the information in Table 7.5.1 from the previous section, predict what will happen in each case involving strong electrolytes.

a. An aqueous solution of strontium hydroxide is added to an aqueous solution of iron(II) chloride.
b. Solid potassium phosphate is added to an aqueous solution of mercury(II) perchlorate.
c. Solid sodium fluoride is added to an aqueous solution of ammonium formate.
d. Aqueous solutions of calcium bromide and cesium carbonate are mixed.

Answer a

$Fe(OH)_2$ precipitate is formed.

Answer b

$Hg_3(PO_4)_2$ precipitate is formed.

Answer c

No Reaction

Answer d

$CaCO_3$ is precipitate formed.

-

> Precipitation reactions are a subclass of double displacement reactions.

Summary

In a **precipitation reaction**, a subclass of exchange reactions, an insoluble material (a **precipitate**) forms when two electrolyte solutions are mixed. To predict the product of a precipitation reaction, all species initially present in the solutions are identified, as are any combinations likely to produce an insoluble salt.

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4.16: Arrhenius Acids and Bases

Learning Objective

- Identify an Arrhenius acid and an Arrhenius base.
- Write the chemical reaction between an Arrhenius acid and an Arrhenius base.

Historically, the first chemical definition of an acid and a base was put forward by Svante Arrhenius, a Swedish chemist, in 1884. An **Arrhenius acid** is a compound that increases the H^+ ion concentration in aqueous solution. The H^+ ion is just a bare proton, and it is rather clear that bare protons are not floating around in an aqueous solution. Instead, chemistry has defined the **hydronium ion** (H_3O^+) as the actual chemical species that represents an H^+ ion. H^+ ions and H_3O^+ ions are often considered interchangeable when writing chemical equations (although a properly balanced chemical equation should also include the additional H_2O). Classic Arrhenius acids can be considered ionic compounds in which H^+ is the cation. Table 4.16.1 lists some Arrhenius acids and their names.

Table 4.16.1 Some Arrhenius Acids

Formula	Name
$HC_2H_3O_2$ (also written CH_3COOH)	acetic acid
$HClO_3$	chloric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydriodic acid
HF	hydrofluoric acid
HNO_3	nitric acid
$H_2C_2O_4$	oxalic acid
$HClO_4$	perchloric acid
H_3PO_4	phosphoric acid
H_2SO_4	sulfuric acid
H_2SO_3	sulfurous acid

An **Arrhenius base** is a compound that increases the OH^- ion concentration in aqueous solution. Ionic compounds of the OH^- ion are classic Arrhenius bases.

Example 4.16.1:

Identify each compound as an Arrhenius acid, an Arrhenius base, or neither.

a. HNO_3
b. CH_3OH
c. $Mg(OH)_2$

Solution

a. This compound is an ionic compound between H^+ ions and NO_3^- ions, so it is an Arrhenius acid.
b. Although this formula has an OH in it, we do not recognize the remaining part of the molecule as a cation. It is neither an acid nor a base. (In fact, it is the formula for methanol, an organic compound.)
c. This formula also has an OH in it, but this time we recognize that the magnesium is present as Mg^{2+} cations. As such, this is an ionic compound of the OH^- ion and is an Arrhenius base.

Exercise 4.16.1

Identify each compound as an Arrhenius acid, an Arrhenius base, or neither.

1. KOH
2. H_2SO_4
3. C_2H_6

Answer

1. Arrhenius base
2. Arrhenius acid
3. neither

Acids have some properties in common. They turn litmus, a plant extract, red. They react with some metals to give off H_2 gas. They react with carbonate and hydrogen carbonate salts to give off CO_2 gas. Acids that are ingested typically have a sour, sharp taste. (The name *acid* comes from the Latin word *acidus*, meaning "sour.") Bases also have some properties in common. They are slippery to the touch, turn litmus blue, and have a bitter flavor if ingested.

Acids and bases have another property: they react with each other to make water and an ionic compound called a salt. A **salt**, in chemistry, is any ionic compound made by combining an acid with a base. A reaction between an acid and a base is called a **neutralization reaction** and can be represented as follows:acid + base → H_2O + salt

The stoichiometry of the balanced chemical equation depends on the number of H^+ ions in the acid and the number of OH^- ions in the base.

Example 4.16.1:

Write the balanced chemical equation for the neutralization reaction between H_2SO_4 and KOH. What is the name of the salt that is formed?

Solution

The general reaction is as follows:

$$H_2SO_4 + KOH \rightarrow H_2O + \text{salt}$$

Because the acid has two H^+ ions in its formula, we need two OH^- ions to react with it, making two H_2O molecules as product. The remaining ions, K^+ and SO_4^{2-}, make the salt potassium sulfate (K_2SO_4). The balanced chemical reaction is as follows:

$H_2SO_4 + 2KOH \rightarrow 2H_2O + K_2SO_4$

Exercise 4.16.1

Write the balanced chemical equation for the neutralization reaction between HCl and $Mg(OH)_2$. What is the name of the salt that is formed?

Answer

$2HCl + Mg(OH)_2 \rightarrow 2H_2O + MgCl_2$; magnesium chloride

Key Takeaways

- An Arrhenius acid is a compound that increases the H^+ ion concentration in aqueous solution.
- An Arrhenius base is a compound that increases the OH^- ion concentration in aqueous solution.
- The reaction between an Arrhenius acid and an Arrhenius base is called neutralization and results in the formation of water and a salt.

4.17: Reactions of Acids and Bases

Learning Objectives

- Write acid-base neutralization reactions.
- Write reactions of acids with metals.
- Write reactions of bases with metals.

Neutralization Reactions

The reaction that happens when an acid, such as HCl, is mixed with a base, such as NaOH:

$$\mathrm{HCl(aq)+NaOH(aq)\rightarrow NaCl(aq)+H_2O(l)}$$

When an acid and a base are combined, water and a salt are the products. Salts are ionic compounds containing a positive ion other than $\mathrm{H^+}$ and a negative ion other than the hydroxide ion, $\mathrm{OH^-}$. Double displacement reactions of this type are called **neutralization reactions**. We can write an expanded version of this equation, with aqueous substances written in their longer form:

$$\mathrm{H^+(aq)+Cl^-(aq)+Na^+(aq)+OH^-(aq)\rightarrow Na^+(aq)+Cl^-(aq)+H_2O(l)}$$

After removing the spectator ions, we get the net ionic equation:

$$\mathrm{H^+(aq)+OH^-(aq)\rightarrow H_2O(l)}$$

When a strong acid and a strong base are combined in the proper amounts—when $[\mathrm{H^+}]$ equals $[\mathrm{OH^-}]$\)—a neutral solution results in which pH = 7. The acid and base have neutralized each other, and the acidic and basic properties are no longer present.

Salt solutions do not always have a pH of 7, however. Through a process known as **hydrolysis**, the ions produced when an acid and base combine may react with the water molecules to produce a solution that is slightly acidic or basic. As a general concept, if a strong acid is mixed with a weak base, the resulting solution will be slightly acidic. If a strong base is mixed with a weak acid, the solution will be slightly basic.

Video: Equimolar (~0.01 M) and equivolume solutions of HCl and NaOH are combined to make salt water. https://youtu.be/TS-I9KrUjB0

Example 4.17.1: Propionic Acid + Calcium Hydroxide

Calcium propionate is used to inhibit the growth of molds in foods, tobacco, and some medicines. Write a balanced chemical equation for the reaction of aqueous propionic acid ($CH_3CH_2CO_2H$) with aqueous calcium hydroxide [$Ca(OH)_2$].

Solution

Solutions to Example 14.5.1

Steps	Reaction

Steps	Reaction
Write the unbalanced equation. This is a double displacement reaction, so the cations and anions swap to create the water and the salt.	$CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + H_2O(l)$
Balance the equation. Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of propionic acid, $CH_3CH_2CO_2H$, to provide H^+ ions.	$\underline{2}CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + \underline{2}H_2O(l)$

? Exercise 4.17.1

Write a balanced chemical equation for the reaction of solid barium hydroxide with dilute acetic acid.

Answer

$$\mathrm{Ba(OH)_2(s) + 2\,CH_3CO_2H(aq) \rightarrow Ba(CH_3CO_2)_2(aq) + 2\,H_2O(l)}$$

Acids and Bases React with Metals

Acids react with most metals to form a salt and hydrogen gas. As discussed previously, metals that are more active than acids can undergo a **single displacement reaction**. For example, zinc metal reacts with hydrochloric acid, producing zinc chloride and hydrogen gas.

$$\mathrm{Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)}$$

Bases also react with certain metals, like zinc or aluminum, to produce hydrogen gas. For example, sodium hydroxide reacts with zinc and water to form sodium zincate and hydrogen gas.

$$\mathrm{Zn(s) + 2\,NaOH(aq) + 2\,H_2O(l) \rightarrow Na_2Zn(OH)_4(aq) + H_2(g).}$$

4.18: The pH Scale

Learning Objectives

- Define *pH*.
- Determine the pH of acidic and basic solutions.

As we have seen, $[H^+]$ and $[OH^-]$ values can be markedly different from one aqueous solution to another. So chemists defined a new scale that succinctly indicates the concentrations of either of these two ions.

pH is a logarithmic function of $[H^+]$:

$$pH = -\log[H^+]$$

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on $[H^+]$, we can summarize as follows:

- If $pH < 7$, then the solution is acidic.
- If $pH = 7$, then the solution is neutral.
- If $pH > 7$, then the solution is basic.

This is known as the pH scale and is the range of values from 0 to 14 that describes the acidity or basicity of a solution. You can use pH to quickly determine whether a given aqueous solution is acidic, basic, or neutral.

✓ Example 4.18.1

Label each solution as acidic, basic, or neutral based only on the stated pH.

a. milk of magnesia, pH = 10.5
b. pure water, pH = 7
c. wine, pH = 3.0

Solution

a. With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
b. Pure water, with a pH of 7, is neutral.
c. With a pH of less than 7, wine is acidic.

? Exercise 4.18.1

Identify each substance as acidic, basic, or neutral based only on the stated pH.

1. human blood, pH = 7.4
2. household ammonia, pH = 11.0
3. cherries, pH = 3.6

Answers

1. basic
2. basic
3. acidic

Table 4.18.1 gives the typical pH values of some common substances. Note that several food items are on the list, and most of them are acidic.

Table 4.18.1 Typical pH Values of Various Substances*

Substance	pH
*Actual values may vary depending on conditions	

Substance	pH
stomach acid	1.7
lemon juice	2.2
vinegar	2.9
soda	3.0
wine	3.5
coffee, black	5.0
milk	6.9
pure water	7.0
blood	7.4
seawater	8.5
milk of magnesia	10.5
ammonia solution	12.5
1.0 M NaOH	14.0
*Actual values may vary depending on conditions	

pH is a *logarithmic* scale. A solution that has a pH of 1.0 has 10 times the $[H^+]$ as a solution with a pH of 2.0, which in turn has 10 times the $[H^+]$ as a solution with a pH of 3.0 and so forth.

Using the definition of pH, it is also possible to calculate $[H^+]$ (and $[OH^-]$) from pH and vice versa. The general formula for determining $[H^+]$ from pH is as follows:

$$[H^+] = 10^{-pH}$$

Key Takeaways

- pH is a logarithmic function of $[H^+]$.
- $[H^+]$ can be calculated directly from pH.
- pOH is related to pH and can be easily calculated from pH.

4.E: Solutions (Exercises)

Additional Exercises

1. One brand of ethyl alcohol (Everclear) is 95% ethyl alcohol, with the remaining 5% being water. What is the solvent and what is the solute of this solution?
2. Give an example of each type of solution from your own experience.
   a. A solution composed of a gas solute in a liquid solvent.
   b. A solution composed of a solid solute in a liquid solvent.
   c. A solution composed of a liquid solute in a liquid solvent.
   d. A solution composed of a solid solute in a solid solvent. (Hint: usually such solutions are made as liquids and then solidified.)
3. Differentiate between the terms *saturated* and *concentrated*.

-

Differentiate between the terms *unsaturated* and *dilute*.

-

What mass of $FeCl_2$ is present in 445 mL of 0.0812 M $FeCl_2$ solution?

-

What mass of SO_2 is present in 26.8 L of 1.22 M SO_2 solution?

-

What volume of 0.225 M $Ca(OH)_2$ solution is needed to deliver 100.0 g of $Ca(OH)_2$?

-

What volume of 12.0 M HCl solution is needed to obtain exactly 1.000 kg of HCl?

-

The World Health Organization recommends that the maximum fluoride ion concentration in drinking water is 1.0 ppm. Assuming water has the maximum concentration, if an average person drinks 1,920 mL of water per day, how many milligrams of fluoride ion are being ingested?

-

For sanitary reasons, water in pools should be chlorinated to a maximum level of 3.0 ppm. In a typical 5,000 gal pool that contains 21,200 kg of water, what mass of chlorine must be added to obtain this concentration?

-

Given its notoriety, you might think that uranium is very rare, but it is present at about 2–4 ppm of the earth's crust, which is more abundant than silver or mercury. If the earth's crust is estimated to have a mass of 8.50×10^{20} kg, what range of mass is thought to be uranium in the crust?

-

Chromium is thought to be an ultratrace element, with about 8.9 ng present in a human body. If the average body mass is 75.0 kg, what is the concentration of chromium in the body in pptr?

-

What mass of 3.00% H_2O_2 solution is needed to produce 35.7 g of $O_2(g)$ at 295 K at 1.05 atm pressure?

$2H_2O_2(aq) \rightarrow 2H_2O(\ell) + O_2(g)$

-

What volume of pool water is needed to generate 1.000 L of $Cl_2(g)$ at standard temperature and pressure if the pool contains 4.0 ppm HOCl and the water is slightly acidic? The chemical reaction is as follows:

$HOCl(aq) + HCl(aq) \rightarrow H_2O(\ell) + Cl_2(g)$

Assume the pool water has a density of 1.00 g/mL.

- A 0.500 *m* solution of $MgCl_2$ has a freezing point of −2.60°C. What is the true van 't Hoff factor of this ionic compound? Why is it less than the ideal value?

- The osmotic pressure of a 0.050 M LiCl solution at 25.0°C is 2.26 atm. What is the true van 't Hoff factor of this ionic compound? Why is it less than the ideal value?

- Order these solutions in order of increasing boiling point, assuming an ideal van 't Hoff factor for each: 0.10 *m* $C_6H_{12}O_6$, 0.06 *m* NaCl, 0.4 *m* $Au(NO_3)_3$, and 0.4 *m* $Al_2(SO_4)_3$.

- Order these solutions in order of decreasing osmotic pressure, assuming an ideal van 't Hoff factor: 0.1 M HCl, 0.1 M $CaCl_2$, 0.05 M $MgBr_2$, and 0.07 M $Ga(C_2H_3O_2)_3$

Answers

1. solvent: ethyl alcohol; solute: water
2.
3. Saturated means all the possible solute that can dissolve is dissolved, whereas concentrated implies that a lot of solute is dissolved.
4.
5. 4.58 g
6.
7. 6.00 L
8.
9. 1.92 mg
10.
11. 1.7×10^{15} to 3.4×10^{15} kg
12.
13. 2,530 g
14.
15. 2.80; it is less than 3 because not all ions behave as independent particles.
16.
17. 0.10 *m* $C_6H_{12}O_6$ < 0.06 *m* NaCl < 0.4 *m* $Au(NO_3)_3$ < 0.4 *m* $Al_2(SO_4)_3$

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